

MIDDLE SCHOOL CHEMISTRY

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Chapter 1: Matter—Solids, Liquids, and Gases

Students are introduced to the idea that matter is composed of atoms and molecules that are attracted to each other and in constant motion. Students explore the attractions and motion of atoms and molecules as they experiment with and observe the heating and cooling of a solid, liquid, and gas.

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Chapter 1, Lesson 1: Molecules Matter

Key Concepts

- Chemistry is the study of matter.
- Matter is made up of extremely tiny particles called atoms and molecules.
- Atoms and molecules make up the three common states of matter on Earth—solids, liquids, and gases.
- The particles of a liquid are attracted to one another, are in motion, and are able to move past one another.
- Being a solid, liquid, or gas is a property of a substance.

Summary

Students discuss the meaning of “chemistry” and “matter.” Students investigate a drop of water hanging from a dropper and drops of water beading up on wax paper. They also look at a molecular animation that models the motion of water molecules. Students are introduced to the idea that matter is made up of extremely tiny particles that are attracted to one another.

Objective

Students will describe their observations about water on the molecular level using the idea that water is composed of tiny molecules that are attracted to one another.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- Water in small cup
- Dropper
- 2 popsicle
- Wax paper
- 2 large index cards (5 × 8”)
- Tape

Materials for the Demonstration

- Tall clear plastic cup
- Water (room temperature)
- White sheet of paper
- Food coloring (red, blue, or green)

About this Lesson

You may be able to complete this lesson in less than 1 class period. If you think this will be the case, you can move on to Lesson 2, which is an application of the concepts covered in Lesson 1.

Note: Some solids, liquids, and gases are made of atoms, and some are made of molecules. Since the concepts covered in Chapter 1 apply to both atoms and molecules, the term “particle” is used as a generic term to include both. At this point, it is enough to give students simple working definitions of “atom” and “molecule.” You can tell students that an atom is the smallest building block of matter and that a molecule is two or more atoms connected together. Let students know that even though atoms and molecules are different, in Chapters 1 and 2, they will both be represented by circles or spheres. Chapters 3 and 4 will go into more depth about the structure of atoms and molecules and will use more detailed models to represent them.

ENGAGE

1. Have a discussion about chemistry and matter.

You could begin the first class by leading a short discussion. Ask students what they think the study of chemistry might be about. You can get a sense of student prior knowledge, identify some misconceptions, and just try to get students on the “same page.”

Tell students that chemistry is the study of matter and what matter does. You could go so far as to say that chemistry is the study of stuff and what stuff does on a very small scale. Ask students for the three common types of matter on Earth (solid, liquid, and gas).

Ask students questions such as the following to guide their thinking:

- **What are some examples of matter?**

Tell students that matter is often defined as anything that has mass and takes up space. Continue the discussion by using water as an example.

- **Does water have mass, and does it take up space?**

A bucket of water is pretty heavy to lift. It definitely has mass. It also takes up space in the bucket. Since it has mass and takes up space, water is matter. But that’s just the very beginning. In chemistry, we want to look deeper and find out more about what matter is made of and how it acts.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Do an activity to explore the attractions water molecules have for each other.

In this activity, students look closely at a drop of water and move drops of water on wax paper. They see that the water holds together well and is not so easy to separate. The goal is for students to begin thinking about water, or any substance, on the molecular level and to conclude that water molecules must be attracted to one another. The reason for these attractions will be dealt with in later chapters.

Question to investigate

Does water hold together well or come apart easily?



Materials for each group

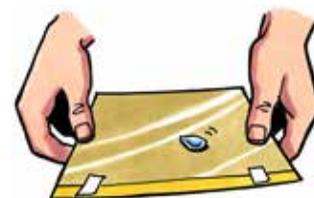
- Water in small cup
- Dropper
- 2 popsicle sticks
- Wax paper
- 2 large index cards (5 × 8")
- Tape

Teacher preparation

Cover a large index card with a piece of wax paper so that the wax paper completely covers the card. Tape the wax paper in place. Prepare two cards for each group.

Procedure

1. Use the dropper to gently squeeze out a drop of water but try not to let the drop fall completely out of the dropper. See how far you can make the drop hang off the end of the dropper without the drop falling.
2. Place 4 or 5 drops of water together on a piece of wax paper to make one medium-size drop.
3. Gently tilt the wax paper in different directions so that the drop moves.
4. Use a popsicle stick to slowly drag the drop around the wax paper a bit. Try using your popsicle stick to separate your drop into two.
5. Use your popsicle stick to move the drops near each other. Then move one drop so that the two drops touch.



3. Record and discuss student observations.

Give students time after the activity to record their observations by answering the following questions on their activity sheet. Once they have answered the questions, discuss their observations as a whole group.

- When you squeezed the drop of water out of the dropper, did the water break apart or did it hold together?
- When you tilted the wax paper, did the drop split apart or stay together?
- When you were pulling the drop around the wax paper, did the water seem to hold together or come apart easily?
- When you tried to split your drop, did the drop separate easily?
- What happened when the two small drops touched?

Expected results

The water beads up on the wax paper and stays together when the wax paper is tilted and when the drop is moved around with a straw. It is difficult to separate the drop into two drops. When the drops touch, they combine quickly and easily.

4. Do a demonstration to show that water molecules are in motion.

Materials

- Tall clear plastic cup
- Water (room temperature)
- White sheet of paper
- Food coloring (red, blue, or green)

Procedure

1. Add water to the cup until it is about $\frac{3}{4}$ filled.
2. Ask students to watch closely as you add one or two drops of food coloring to the water. Do not stir. Instead, allow the color to slowly mix into the water on its own.
3. Hold the cup up with a sheet of white paper behind it so it is easier for students to see the color moving and mixing in the water.

Expected results

The drops of food coloring will slowly move and mix into the water. Eventually all the water in the cup will be evenly colored.

Ask students:

How do your observations support the idea that water molecules are moving?

Help students understand that the drop of coloring mixes into the water because the water

molecules move and push the color in all directions. The molecules of the food coloring themselves are also in motion.

Note: In chapter 5, students will learn that water molecules and coloring molecules are attracted to each other. These attractions also help explain the mixing of the color in the water.

EXPLAIN

5. Show an animation of the molecules in liquid water.



Show the molecular model animation *Particles of a Liquid*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson1#particles_of_a_liquid

Explain that the little balls represent the particles of a liquid, in this case water molecules. Let students know that for now, they will use circles or spheres to represent atoms and molecules, but eventually they will use a more detailed model. For now, students should focus on the motion of the molecules, how they interact, and their distance from one another.

Point out that the molecules of a liquid are in motion but they are attracted to each other. That's why they move past each other but don't get very far apart from one another.

6. Have students draw their own model of water on the molecular level and complete the activity sheet.

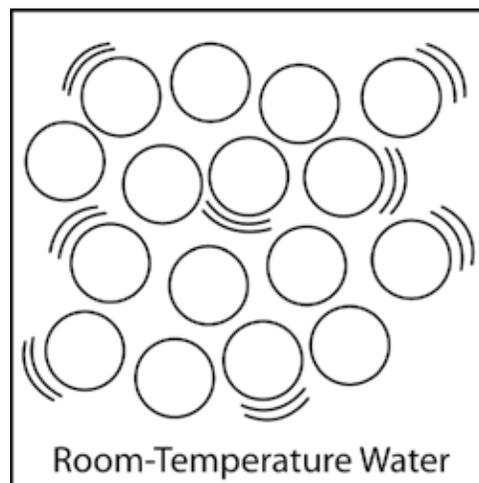


Draw or project the illustration *Water Molecules*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson1#water_molecules

Explain to students that this is a model of water molecules. Point out that the molecules are not in any exact order but are near each other. They have little curved “motion lines” to show that the molecules are moving.

Have students draw a model of water on the molecular level on their activity sheet. They should use the model you have shown them to guide their own drawing.



Students' drawings should show that the molecules are:

- Randomly arranged
- Close together
- Moving

Be sure students realize that this model shows water molecules enormously bigger than they actually are. Not only are water molecules much smaller, they are also much more numerous. A single drop of water is made up of more than a billion trillion extremely tiny water molecules.



To give students an idea of how small and numerous water molecules are, you could tell students the following: In about 1 tablespoon of water, there are about 600 billion trillion water molecules. If you could count 1 million water molecules every second, it would take about 200 million centuries to count all the molecules in that tablespoon of water. Atoms and molecules are huge in number and incredibly small in size.

EXTEND

7. Show a video so that students can see an example that water molecules are attracted to one another.

Show a video of a water balloon popping in slow motion.

www.middleschoolchemistry.com/multimedia/chapter1/lesson1#water_balloon

Ask students:

- **Why do you think the water keeps its shape the moment the balloon is popped?**
Students should realize that water holds together pretty well because the water molecules are attracted to each other.

- **Imagine a drop of water hanging from your finger. How is this similar to the water staying together after the balloon is popped?**

This can also be explained by the fact that water molecules are very attracted to each other.

EXTRA EXTEND

4. If you have time, give students the opportunity to play games with drops of water.

Water Drops Unite!

Teacher Preparation

Print 2 “Water Drops Unite” sheets for each group.

Procedure

1. Tape a piece of wax paper over the “Water Drops Unite” sheet.
2. Place about 5 drops of water in each of the small circles around the outside.
3. As fast as you can, use your straw to drag each drop of water to the center. When all the drops are united in the center, you are done.
4. Challenge your partner to see who can unite all their water drops the fastest.



Race Drop Raceway

Teacher Preparation

Print 2 “Race Drop Raceway” sheets for each group. You and a partner can follow the directions below to race each other.

Procedure

1. Tape the “Race Drop Raceway” sheet onto a piece of cardboard to give it support.
2. Tape a piece of wax paper over the “Race Drop Raceway” sheet.
3. Place 2–4 drops of water together to make one larger drop at the “Start.”
4. As fast as you can, tilt the cardboard and guide your race drop around the track to the “Finish.” Try not to touch the edge of the track. The first to finish is the winner.

Activity Sheet
Chapter 1, Lesson 1
Molecules Matter

Name: _____

Date: _____

In the activity below, you will investigate some of the characteristics of water. You will also begin to model and explain, on the molecular level, why water acts the way it does.

ACTIVITY

Question to investigate

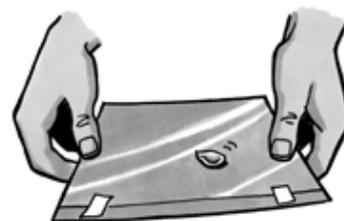
Does water hold together well or come apart easily?

Materials for each group

- Water in small cup
- Dropper
- 2 popsicle sticks
- 2 index cards covered with wax paper

Procedure

1. Use a dropper to gently squeeze out one drop of water but try not to let the drop fall completely out of the dropper. See how far you can make the drop hang off the end of the dropper without the drop falling.
2. Place 4 or 5 drops of water together on the wax paper to make a medium-size drop.
3. Gently tilt the wax paper in different directions so that the drop moves.
4. Use a popsicle stick to slowly move your drop around the wax paper. Try using your popsicle stick to separate your drop into two.
5. Use your popsicle stick to move the two drops near each other. Then move one drop so that the two drops touch.



WHAT DID YOU OBSERVE?

1. When you squeezed the drop of water out of the dropper, did the water break apart did it hold together?

2. When you were pulling the drop around the wax paper, did the water seem to hold together or come apart easily?
3. When you tried to split your drop, did the drop separate easily?
4. Was it easy or difficult to make the drops come together?

DEMONSTRATION

5. Your teacher placed a drop of food coloring in a cup of water. The color slowly mixed into the water without being stirred. What does this tell you about water molecules?

EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animated molecular model of water. Now you will draw your own molecular model.

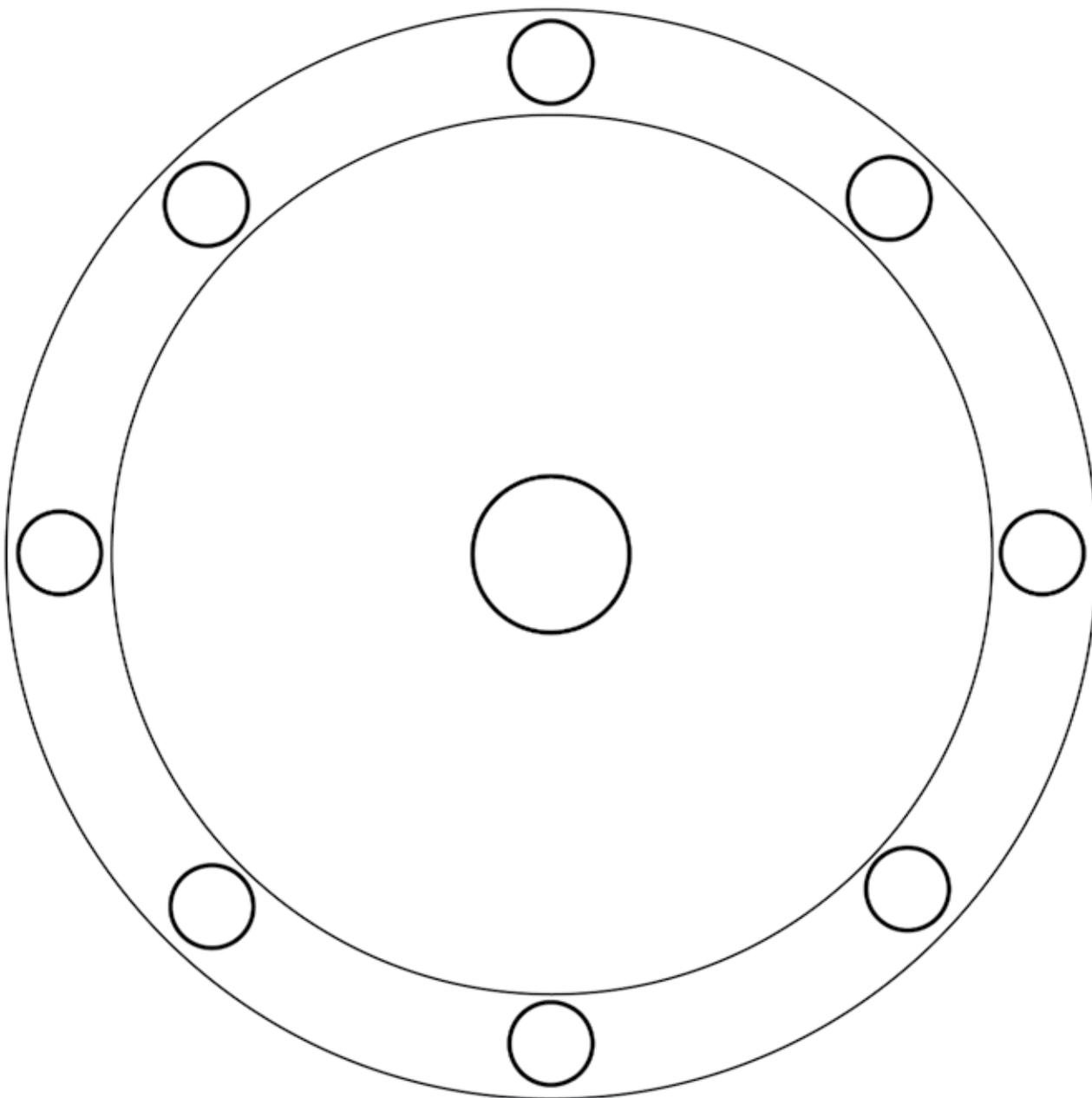
6. Using circles and motion lines to represent water molecules, draw a model of water on the molecular level. Be sure to show that water molecules are:

- Randomly arranged.
- Close together because they attract each other.
- Moving



7. What is it about water molecules that helps explain why the water drops were difficult to split apart but easy to join together?

WATER DROPS **UNITE!**



Chapter 1, Lesson 2: Molecules in Motion

Key Concepts

- Heating a liquid increases the speed of the molecules.
- An increase in the speed of the molecules competes with the attraction between molecules and causes molecules to move a little further apart.
- Cooling a liquid decreases the speed of the molecules.
- A decrease in the speed of the molecules allows the attractions between molecules to bring them a little closer together.

Summary

Students add food coloring to hot and cold water to see whether heating or cooling affects the speed of water molecules. Students watch molecular model animations to see the effect of heating and cooling on the molecules of a liquid. Students will also draw their own molecular model.

Objective

Students will be able to explain, on the molecular level, that heating and cooling affect molecular motion.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- Hot water (about 50 °C) in a clear plastic cup
- Cold water in a clear plastic cup
- Yellow food coloring in a small cup
- Blue food coloring in a small cup
- 4 droppers

ENGAGE

- 1. Ask students to help you design an experiment to see if the speed of water molecules is different in hot water compared to cold water.**

Ask students questions such as the following:

- **Is the speed of water molecules different in hot and cold water? What can we do to find out?**

Students may guess that molecules in hot water move faster. There are several possible experiments that students might suggest to find out if this is true. One of the more obvious ones is to heat water a lot so that it boils. Then you can see the water moving. You could do that but it requires a hot plate, takes a fair amount of time, and may have to be done as a demonstration instead of being an activity the students can do.

Tell students that one possible method is to use hot water and cold water and add food coloring to the water. If the water molecules move faster at one temperature than another, the food coloring should move faster too and make the movement easy to see.

Ask students:

- **Should we use the same amount of hot and cold water in our experiment?** Yes
- **Should we use the same type of cup for the hot and cold water?** Yes
- **Should we use the same number of drops of food coloring in each cup?** Yes
- **Should we put the coloring in at the same time?** Yes

Explain that the different things like the amount of water, type of cup, and number of drops of food coloring are called *variables*. It is important to keep all the variables the same except for the one you are testing. Because we are trying to find out if *temperature* affects the motion of water molecules, we should keep everything else about the experiment the same. Temperature should be the only variable. This way, if we notice something different between the two samples of water, we will know that the difference in temperature is causing it.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. **Do an activity to compare the speed of water molecules in hot and cold water.**



Question to investigate

Is the speed of water molecules different in hot and cold water?

Teacher preparation

This activity works best if there is a big difference between the temperatures of the hot and cold water.

1. Squirt 4–5 drops of blue food coloring into a small cup for each group.
2. Squirt 4–5 drops of yellow food coloring into another small cup for each group.
3. Add ice to about 6 cups of tap water to make it sufficiently cold.
4. Pour about $\frac{3}{4}$ cup of cold water (no ice) into a cup for each group.
5. Pour about $\frac{3}{4}$ cup of hot water into a cup for each group.

Materials for each group

- Hot water (about 50 °C) in a clear plastic cup
- Cold water in a clear plastic cup
- Yellow food coloring in a small cup
- Blue food coloring in a small cup
- 4 droppers



Procedure

1. With the help of your partners, use droppers to carefully place 1 drop of yellow and 1 drop of blue food coloring into the hot and cold water at the same time.
2. Allow the colors to mix on their own as you watch them for a couple of minutes.

3. Record and discuss student observations.

Give students time after the activity to record their observations by answering the following questions on their activity sheet. Once they have answered the questions, discuss their observations as a whole group.



- Describe what the colors looked like and how they moved and mixed in the cold water.
- Describe what the colors looked like and how they moved and mixed in the hot water.
- What does the speed of the mixing colors tell you about the speed of the molecules in hot and cold water?

Expected results

The yellow and blue food coloring will spread faster in hot water than in cold. The colors will combine and turn green in the hot water while the colors will remain separate longer in the cold water. Students should agree that the colors mix faster in the hot water because the molecules of both the water and food coloring move faster in hot water than they do in cold water.

EXPLAIN

4. Show an animation of water molecules at different temperatures.



Show the molecular model animation *Heating and Cooling a Liquid*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson2#heating_and_cooling

Move the slider at the bottom of the window all the way to the right to show that the water molecules are moving faster and are a little farther apart in hot water.

Explain that the little balls represent the particles of a liquid, in this case water molecules. Let students know that for now, they will use circles or spheres to represent atoms and molecules, but eventually they will use a more detailed model. For now, students should focus on the motion of the molecules, how they interact, and their distance from one another.

Ask students:

- **Are the molecules moving faster in cold or hot water?**

Students should realize that the molecules of hot water are moving faster. The molecules of cold water are moving slower.

- **How does this match with your observations with the food coloring?**

The food coloring in the hot water mixed faster than the coloring in the cold water did.

- **Look closely at the space between the molecules in cold and hot water. Is there more space in between the molecules in hot water or in cold water? Is it a lot of space?**

Point out to students that molecules of hot water are moving faster and are slightly further apart. The molecules of cold water are moving slower and are a little closer together. If students do not notice a difference, move the slider all the way to the left again and then quickly to the right. Show the animation a few times to give students a chance to notice the differences.

5. Have students answer questions about the animation and draw a model of water molecules on their activity sheet.



Have students fill in the blank with the word *increases* or *decreases* on their activity sheet as you read each sentence.

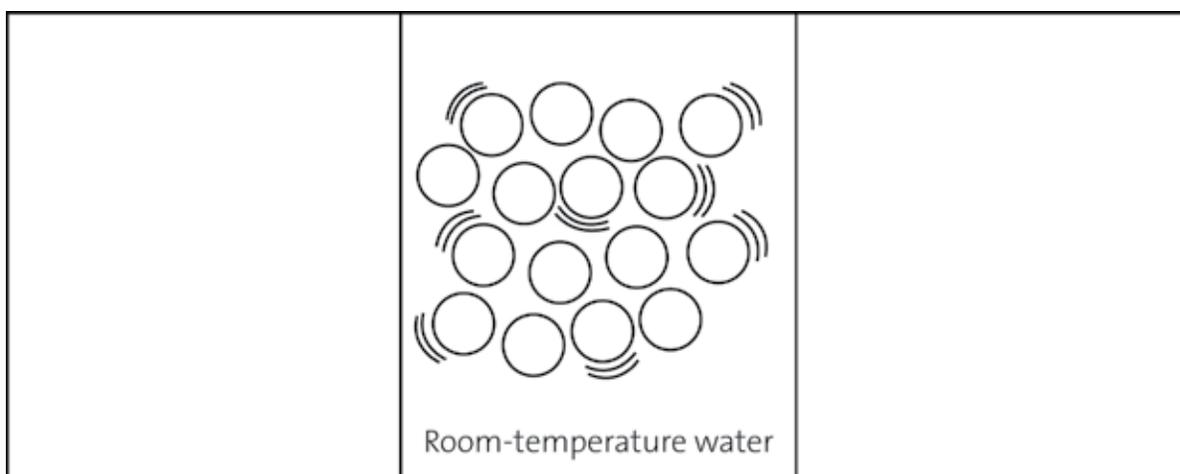
- Heating a substance *increases* molecular motion.
- Cooling a substance *decreases* molecular motion.
- As molecular motion increases, the space between molecules *increases*.
- As molecular motion decreases, the space between molecules *decreases*.

Project the image *Water Molecules at Different Temperatures*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson2#water_molecules_at_different_temperatures



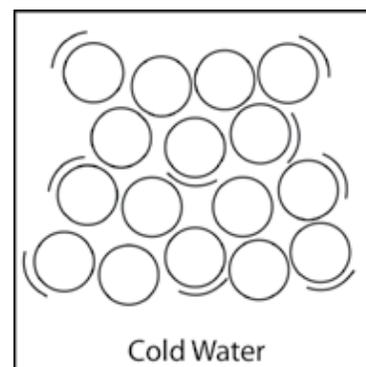
Have students refer to the drawing of room-temperature water on their activity sheet and discuss how they should represent the molecules in cold and hot water.



Cold water

Ask students:

- **Would the water molecules be closer together or further apart?** Students should draw the circles a little closer together than the circles in the room-temperature water. The water molecules are closer together because the slower motion allows the attractions to bring the molecules a little closer together.
- **Would there be more or fewer motion lines?** Students should realize that since the molecules in the cold water are moving slower, they should have fewer motion lines than the molecules in room-temperature water.



Hot water

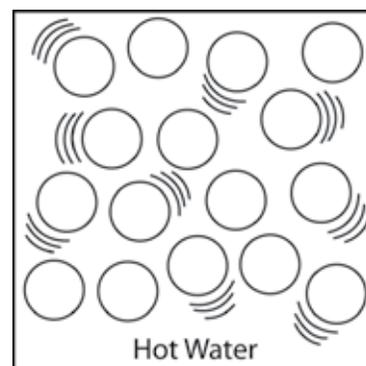
Ask students:

- **Would the water molecules be closer together or further apart?**

Students should draw the circles a little further apart than the circles in the room-temperature water. The faster motion competes with the attractions water molecules have for each other and causes the molecules to move a little further apart.

- **Would there be more or fewer motion lines?**

Students should realize that since the molecules in hot water are moving faster than in cold or room-temperature water, they should draw more motion lines.



EXTEND

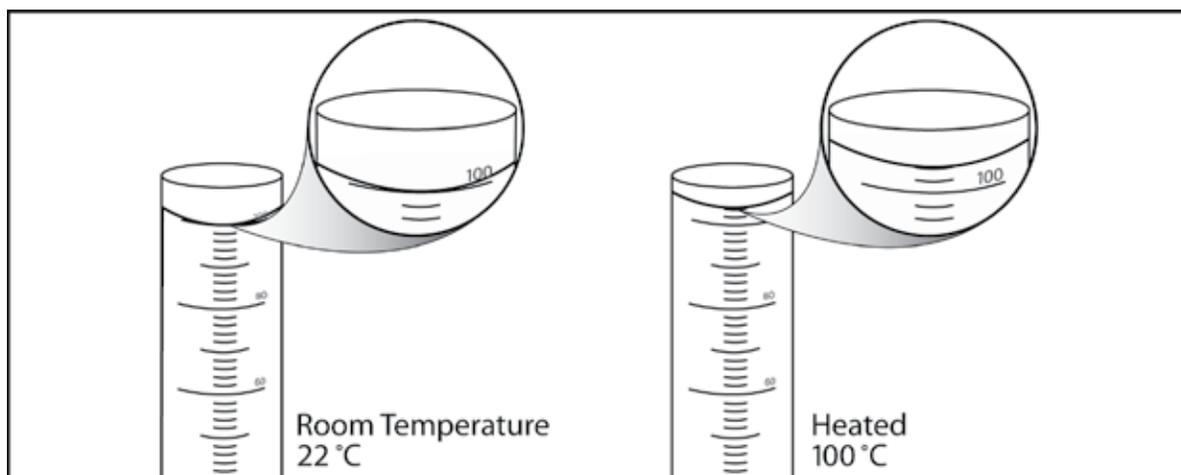
6. **Have students explain why hot water takes up more space than room-temperature water.**

Have students read and discuss the *Take It Further* question on the activity sheet. After the class discussion, have students write their own response to the following question in the space provided on the activity sheet.



- **Let's say that you measure exactly 100 milliliters of water in a graduated cylinder. You heat the water to 100 °C and notice that the volume increases to 104 milliliters. Using what you know about the attractions between water molecules and the way heating affects molecular motion, explain why the volume of water in the cylinder increases when it is heated.**

Students should realize that the molecules in hot water move slightly further apart, accounting for the increased volume.



Name _____

Name _____

ACTIVITY

Question to investigate

Is the speed of water molecules different in hot and cold water?

Materials for each group

- Hot water in a clear plastic cup
- Cold water in a clear plastic cup
- Food coloring (yellow and blue)
- 4 droppers



Procedure

1. With the help of your partners, use droppers to carefully place 1 drop of yellow and 1 drop of blue food coloring into the hot and cold water at the same time.
2. Allow the colors to mix on their own as you watch them for a couple of minutes.

WHAT DID YOU OBSERVE?

1. Describe what the colors looked like and how they moved and mixed in the cold water.
2. Describe what the colors looked like and how they moved and mixed in the hot water.
3. What does the speed of the mixing colors tell you about the speed of the molecules in hot and cold water?

4. There were several variables in this experiment:

- Amount of water in each cup
- Type of cup used
- Number of drops of food coloring
- When the coloring was added to the water

Pick one of these variables and explain why you made sure it was kept the same in the two cups.

EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animation of water molecules being heated and cooled. Now you can draw your own molecular model.

5. Based on your observations and the animations, fill in the blanks with the words *increases* or *decreases*.

Heating a substance _____ molecular motion.

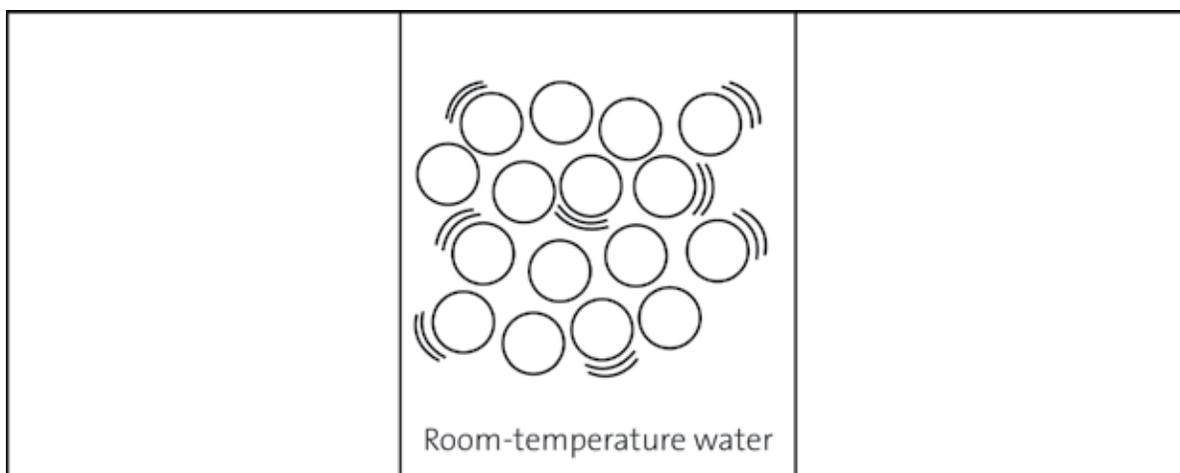
Cooling a substance _____ molecular motion.

As molecular motion increases, the space between molecules _____.

As molecular motion decreases, the space between molecules _____.

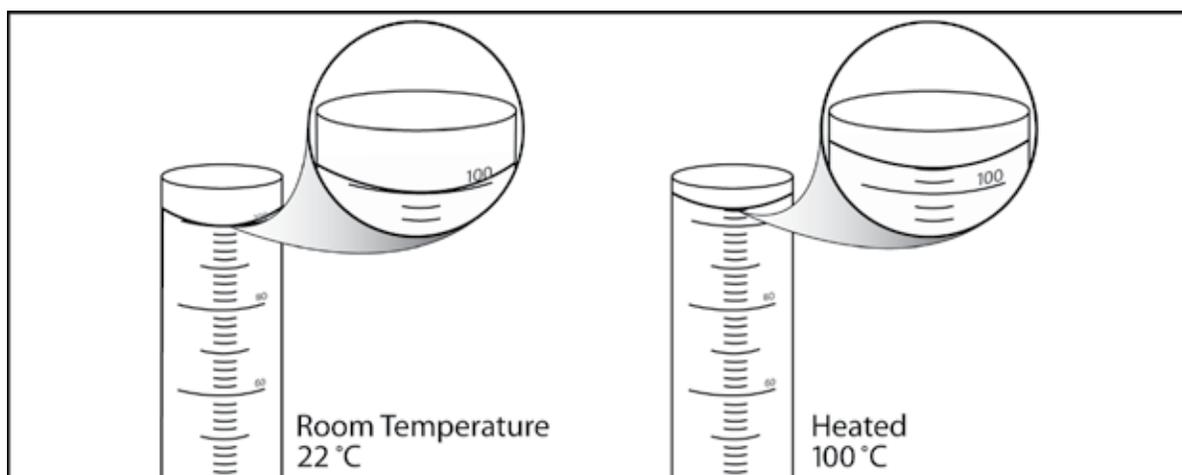
6. Using circles to represent water molecules, draw a model of the molecules in cold and hot water.

- Use motion lines to show the speed of the molecules.
- Consider the space between molecules in each temperature of water.



TAKE IT FURTHER

Let's say that you measure exactly 100 milliliters of water in a graduated cylinder. You heat the water to 100 °C and notice that the volume increases to 104 milliliters.



- Using what you know about the attractions between water molecules and the way heat affects molecular motion, explain why the volume of water in the cylinder increases when it is heated.

Chapter 1, Lesson 3: The Ups and Downs of Thermometers

Key Concepts

- The way a thermometer works is an example of heating and cooling a liquid.
- When heated, the molecules of the liquid in the thermometer move faster, causing them to get a little farther apart. This results in movement up the thermometer.
- When cooled, the molecules of the liquid in the thermometer move slower, causing them to get a little closer together. This results in movement down the thermometer.

Summary

Students will look closely at the parts of a thermometer. After placing a thermometer in hot and cold water, students will look at molecular model animations of the liquid in a thermometer. Students will then draw a model of the molecules of a thermometer after it has been placed in hot and then cold water.

Objective

Students will be able to describe, on the molecular level, why the liquid in a thermometer goes up when it is heated and down when it is cooled.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

- Be sure you and the students wear properly fitting goggles.
- Students should use care when handling hot tap water.
- When using isopropyl alcohol, read and follow all warnings on the label. Isopropyl alcohol is flammable. Keep it away from any flames or spark sources.

Materials for Each Group

- Student thermometer
- Magnifier
- Cold water
- Hot water (about 50 °C)

Notes about the materials

Student thermometers are available from Sargent Welch (WL5679), Flinn Scientific (AP5406) and other suppliers.

ENGAGE

1. Find out what students know about thermometers.

Hold up an alcohol thermometer and ask students:

- **Why do you think the liquid in a thermometer moves up and down when it is heated and cooled?**

Students should realize that the movement of the liquid in a thermometer is related to the motion of the molecules of the liquid when they are heated and cooled.

Remind students that molecules move faster and a little further apart when they are heated. Molecules also move slower and a little closer together when they are cooled.

Tell students that they will apply their understanding of what happens when liquids are heated and cooled to explain how a thermometer works.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Do an activity to investigate what makes the liquid in a thermometer go up and down.

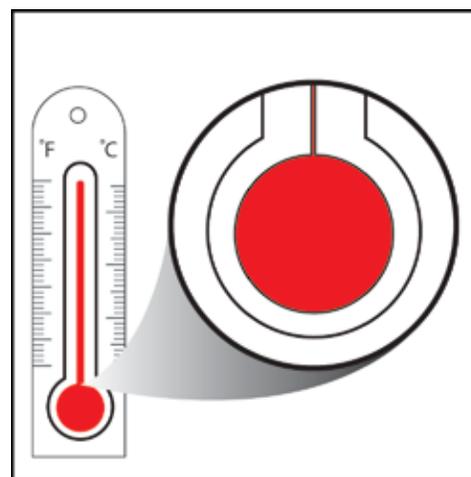


Question to investigate

What makes the liquid in a thermometer go up and down?

Materials for each group

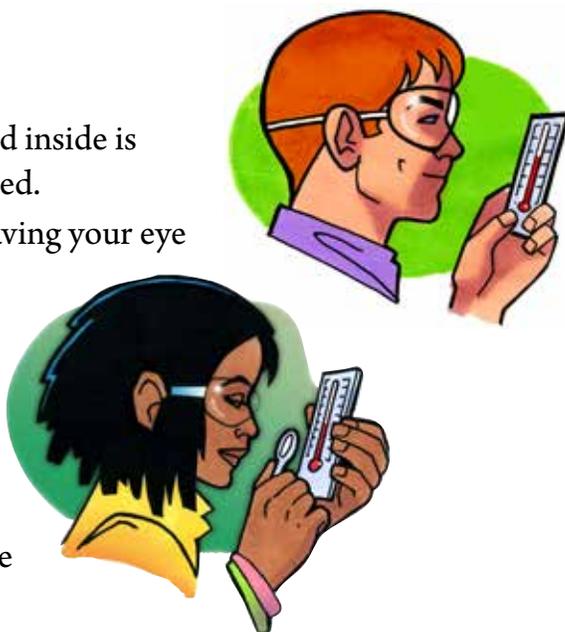
- Student thermometer
- Magnifier
- Cold water
- Hot water (about 50 °C)



Procedure

A. Look closely at the parts of a thermometer.

1. Look closely at your thermometer. The liquid inside is probably a type of alcohol that's been dyed red.
2. Practice reading the temperature in $^{\circ}\text{C}$ by having your eye at the same level as the top of the red liquid. What is the temperature?
3. Use a magnifier to look closely at the thermometer from the front and from the side. Look at the bulb and the thin tube that contains the red liquid.
4. Put your thumb or finger on the bulb and see if the red liquid moves in the thin tube.



B. Observe the red liquid in the thermometer when it is heated and cooled.

5. Place the thermometer in hot water and watch the red liquid. Keep it in the hot water until the liquid stops moving. Record the temperature in $^{\circ}\text{C}$.
6. Now put the thermometer in cold water. Keep it in the cold water until the liquid stops moving. Record the temperature in $^{\circ}\text{C}$.



Expected results

The red liquid goes up in hot water and down in cold water. Students will have an opportunity to relate these observations to an explanation, on the molecular level, of why the liquid moves the way it does.

If you have time, you can have students pick a temperature somewhere between the temperature of cold water and hot water and then attempt to combine an amount of hot and cold water to achieve that temperature in one try. They can see how close they can get.

3. Record and discuss student observations.

Give students time after the activity to record their observations by answering the following questions on their activity sheet. Once they have answered the questions, discuss their observations as a whole group.



1. Based on what you know about the way molecules move in hot liquids, explain why the liquid in the thermometer goes up when heated.
2. Based on what you know about the way molecules move in cold liquids, explain why the liquid in the thermometer goes down when cooled.
3. Why do you think the tube that contains the red liquid is so thin?
4. What do you think is the purpose of the larger outer tube?

When heated, the molecules of the red liquid inside the thermometer move faster. This movement competes with the attractions the molecules have for each other and causes the molecules to spread a little further apart. They have nowhere to go other than up the tube. When the thermometer is placed in cold water, the molecules slow down and their attractions bring them a little closer together bringing them down the tube. The red liquid is contained in a very thin tube so that a small difference in the volume of the liquid will be noticeable. The large outer tube has two purposes—to protect the fragile inner tube and act as a magnifier to help you better see the red liquid.

EXPLAIN

4. Show an animation of the molecules of liquid in a thermometer as they are heated and cooled.



Note: Alcohol molecules are composed of different atoms, but in the model shown in the animation the molecules are represented as simple red spheres.

Show the molecular model animation *Heating and Cooling a Thermometer*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson3#heating_and_cooling_a_thermometer

Point out that when the thermometer is heated, the molecules move faster, get slightly further apart, and move up the tube. When the thermometer is cooled, the molecules move more slowly, get closer together, and move down the tube. Help students realize that the attractions the molecules in the thermometer have for each other remain the same whether the thermometer is heated or cooled. The difference is that when heated, the molecules are moving so fast that the movement competes with the attractions, causing the molecules to move further apart and up the tube. When cooled, the movement of the molecules is slower and does not compete as much with the attractions the molecules have for one another. This is why the molecules in the thermometer move closer together and down the tube.

Ask students:

- **The animation shows that the molecules spread out slightly when heated. Do you think the thermometer would work as well if the tube of red liquid were wider?**



The molecules spread out in all directions when heated. If the tube were wide, the molecules would be free to spread out sideways as well as up. In the thin tube, the molecules can't move very far sideways, so they go up. This causes a greater difference in the height of the liquid, which is easier to see.

5. Have students draw a molecular model to represent the molecules of the liquid in a thermometer.

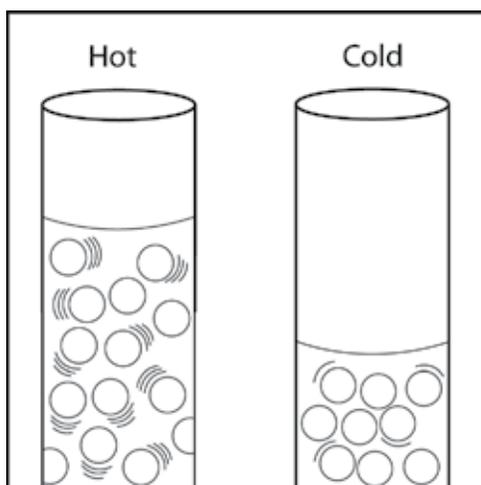
Project the image *Molecules in a Thermometer*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson1#molecules_in_a_thermometer

In the drawing, lines have been added to indicate the level of the liquid in each tube. In reality, there is no line. The “line” is made up of molecules. Students should draw circles representing molecules all the way up to the line drawn in each tube.



Have students use the projected illustration as a guide as they draw a model of the molecules in a hot and cold thermometer on their activity sheet.



The *Hot Thermometer* illustration should show random circles with more motion lines. The circles should be a little further apart than in the cold thermometer.

The *Cold Thermometer* should show random circles with fewer motion lines. The circles should be a little closer together than the circles in the hot thermometer.

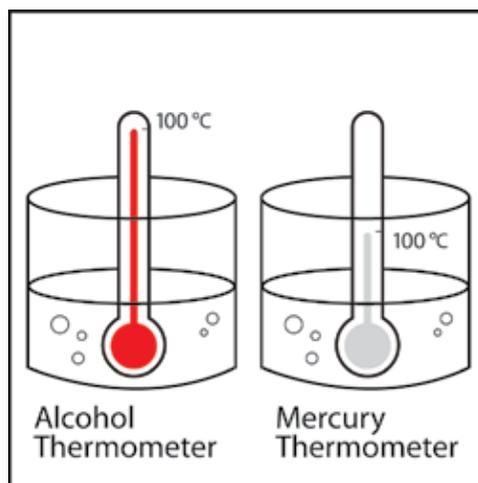
EXTEND

6. Discuss with students why thermometers with different liquids in them rise to different heights even at the same temperature.

Project the image *Different Thermometers Same Temperature*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson3#different_thermometers_same_temperature

Tell students that this picture shows two thermometers that are identical in every way, except one has alcohol and the other has mercury inside. Point out that both thermometers are placed in hot water that is 100 °C. The levels of the alcohol and mercury are shown.



Ask students:

- How can the liquids in the thermometers be at different levels even though they are in water that is the same temperature?

Hint: Alcohol and mercury are both liquids but are made of different atoms and molecules. Use what you know about the motion and attractions the particles in a liquid have for one another to explain why the levels of alcohol and mercury in the thermometers are different.



The main reason why the level of liquid in each thermometer is different is that they are different substances with different properties. The molecules that make up the alcohol have different attractions for each other than the atoms that make up the mercury. Therefore, heating and cooling them are going to make them move different distances up or down the tube.

After the class discussion, have students write their own response to the question about the two different thermometers on the activity sheet.

EXTRA EXTEND

7. If you have time and the materials, give students the opportunity to make and test their own thermometers.

Materials for each group

- Test tube, about 20 mL
- Isopropyl rubbing alcohol, 91%

- Small cup
- 2 small disposable water bottles (8 oz.)
- Food coloring
- Parafilm
- Thin transparent straw (either cut from a dropper or from a Capri Sun® drink)
- Hot water (about 50 °C)
- Ice water

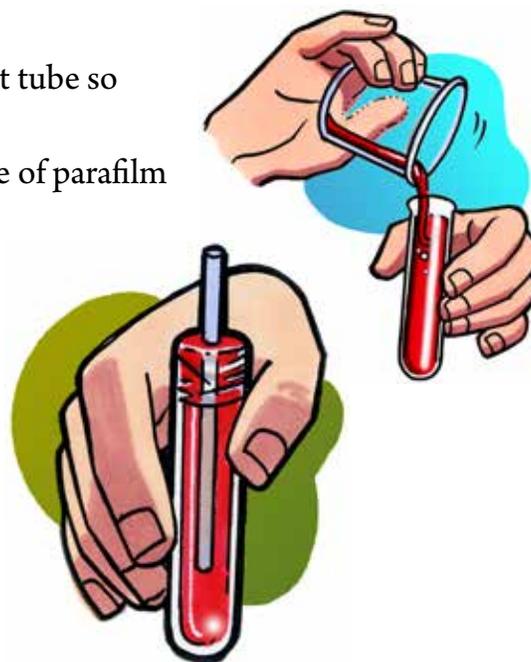
Teacher preparation

- Place an amount of isopropyl alcohol that will completely fill a test tube into a small cup for each group. Add 1 drop of red food coloring to the alcohol. Gently swirl to mix.
- Use scissors to cut the bulb end off of a dropper. Then cut diagonally on one end to give the “straw” a sharp point. Instead, you may use a thin straw from a Capri Sun® drink.



Procedure

1. Carefully pour the colored alcohol into the test tube so that the tube is almost completely full.
2. With the help of a partner, stretch a small piece of parafilm over the top of the test tube so that it is well-sealed.
3. While holding the parafilm tight with one hand and the straw in the other, carefully pierce through the plastic so that the straw goes into the alcohol. About 3–4 centimeters of straw should stick out of the top of the plastic. This is your thermometer.
4. Pour ice water into one water bottle until it is about $\frac{2}{3}$ full.
5. Ask your teacher to put hot water in another water bottle so it is about $\frac{2}{3}$ full.
6. Carefully lower the test tube into the hot water. Watch the straw above the plastic for about 1 minute.
7. Ask students to predict what will happen when the warm test tube is placed in cold water.
8. Carefully lower the warm test tube into the cold water. Watch the straw above the plastic for about 1 minute.



Student research about thermometers

Students have seen that adding energy to a liquid increases the motion of the molecules and causes them to spread a little further apart. They have also seen that this movement is the basis for how a thermometer works. Introduce the idea that measuring temperature with a thermometer gives us an idea of the average energy of a substance. This is an opportunity for students to research different types of thermometers, the history of thermometers, and how thermometers are made and calibrated.

ACTIVITY

Question to investigate

What makes the liquid in a thermometer go up and down?

Materials for each group

Student thermometer

Magnifier

Cold water

Hot water (about 50 °C)

Procedure

A. *Look closely at the parts of the thermometer.*

1. Look closely at your thermometer. The liquid inside is probably a type of alcohol that's been dyed red.
2. Read the temperature in °C by having your eye on the same level as the top of the red liquid. What is the temperature?
3. Use a magnifier to look closely at the thermometer from the front and from the side. Look at the bulb and the thin tube that contain the red liquid.
4. Put your thumb or finger on the red bulb and see if the red liquid moves in the thin tube.



B. *Observe the red liquid in the thermometer when it is heated and cooled.*

1. Place the thermometer in hot water and watch the red liquid. Keep it in the hot water until the liquid stops moving.

Record the temperature in °C. _____

2. Now put the thermometer in cold water. Keep it in the cold water until the liquid stops moving.

Record the temperature in °C. _____



WHAT DID YOU OBSERVE?

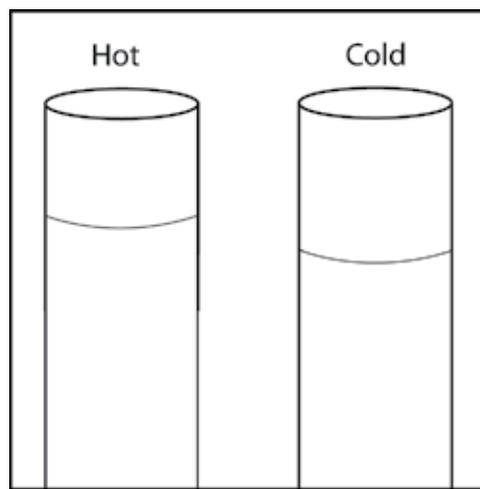
1. Based on what you know about the way molecules move in hot liquids, explain why the liquid in the thermometer goes up when heated.
2. Based on what you know about the way molecules move in cold liquids, explain why the liquid in the thermometer goes down when cooled.
3. Why do you think the tube that contains the red liquid is so thin?
4. What do you think is the purpose of the larger outer tube?

EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animated molecular model of a thermometer at different temperatures. Now you will draw your own model.

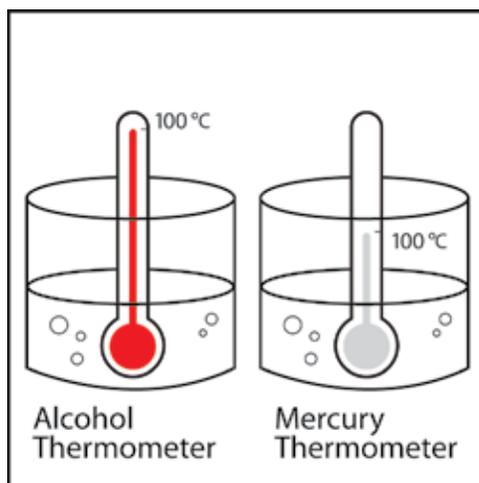
The drawing shows two close-ups of a thin tube in a thermometer like the one you used. One picture represents the thermometer in hot water, while the other is the thermometer in cold water.

5. Based on what you know about the motion of molecules in a liquid and what you saw in the animations, draw circles to represent alcohol molecules in the liquid in the thermometer. Try to show the difference in distance between the molecules when the liquid is hot and cold. Use motion lines to represent their movement (fast or slow).



TAKE IT FURTHER

6. Imagine that you have two thermometers that are identical in every way, except one has alcohol and the other has mercury inside. Each thermometer is placed in hot water that is 100 °C. The levels of the alcohol and mercury are shown in the picture.



Why do you think the liquids in the thermometers are at different levels even though they are in water that is the same temperature?

Hint: Alcohol and mercury are both liquids but are made of different atoms and molecules. Use what you know about the motion and attractions the particles in a liquid have for one another to explain why the levels of alcohol and mercury in the thermometers are different.

Chapter 1, Lesson 4: Moving Molecules in a Solid

Key Concepts

- In a solid, the atoms are very attracted to one another. The atoms vibrate but stay in fixed positions because of their strong attractions for one another.
- Heating a solid increases the motion of the atoms.
- An increase in the motion of the atoms competes with the attraction between atoms and causes them to move a little further apart.
- Cooling a solid decreases the motion of the atoms.
- A decrease in the motion of the atoms allows the attractions between atoms to bring them a little close together.

Summary

Students will see a demonstration with a metal ball and ring showing that heat causes atoms to spread a little further apart. They will also see that cooling a solid causes the atoms to get a little closer together. The same rules they discovered about liquids also apply to solids.

Objective

Based on their observations students will describe, on the molecular level, how heating and cooling affect the motion of atoms in a solid.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for the Demonstration

- Ball and ring designed specifically for this demonstration]
- Bunsen burner for heating the ball
- Room-temperature water (to cool the ball)

Notes about the materials

The metal ball and ring is available from Sargent Welch (WL1661-10) or Flynn Scientific (AP9031) or other suppliers.

About this Lesson

The solid explored in this lesson is a metal. Metal is composed of individual atoms instead of molecules like in the water and alcohol students learned about in Lessons 1–3. Although atoms and molecules are different, we will represent atoms the same way we represented molecules, using a circle or sphere. This simple representation will help students focus on the motion and position of the particles when they are heated and cooled.

ENGAGE

1. Review what students have discovered about molecules in a liquid and discuss whether these same ideas might apply to solids, too.

Ask students:

- **What do you know about the molecules in a liquid?**

Be sure students understand that the molecules in a liquid are attracted to each other but are able to move past each other.

- **How does heating or cooling affect the speed of the molecules and the distance between them?**

Heating speeds up the motion of molecules and cooling slows them down. We've also seen that speeding the molecules up makes them move a little further apart and slowing them down allows them to move a little closer together.

Ask students if these statements also apply to solids:

- **Do you think the atoms in a solid are attracted to each other?**

Students will probably realize that the atoms of a solid are attracted to each other. Explain that this is how a solid stays together.

- **Do you think heating or cooling a solid might affect the motion of the atoms?**

Students should realize that if you heat a solid, the atoms or molecules move faster and move further apart. If you cool a solid, the molecules move more slowly and move a little closer together.

2. Show an animation to help students compare atoms and molecules in solids and liquids.



Explain that the little balls represent the particles of a solid, in this case the atoms in a metal. Although atoms and molecules are different, this same simple model of balls is used for both. Let students know that for now, they will use circles or spheres to represent atoms and molecules, but eventually they will use a more detailed model. Tell students that they should focus on the motion of the molecules, how they interact, and their distance from one another.

Show the molecular model animation *Particles of a Solid*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson4#particles_of_a_solid

Point out the following about solids:

- The particles (atoms or molecules) are attracted to each other.
- The particles (atoms or molecules) vibrate but do not move past one another.
- The solid retains its shape.

Show the molecular model animation *Comparing Solid and Liquid*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson4#comparing_solid_and_liquid

Click on both tabs and make sure students notice the differences in the movement of the atoms and molecules.

- The atoms in a solid are so attracted to each other that they vibrate and don't move past each other.
- The molecules of a liquid are attracted to each other, but move more freely and past one another.



Give each student an activity sheet.

Students will record their observations and answer questions about the animation on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

EXPLORE

3. Do a demonstration to show that solid metal expands when it is heated and contracts when cooled.

It is harder to show that the particles of a solid move faster when heated than it is to show the same thing with a liquid like in Lesson 2. But you can do it if you have a special ball and ring apparatus that shows the expansion of a metal when heated. This inexpensive device, available through science education equipment companies, consists of a rod with a metal ball on the end and another rod with a metal ring. At room-temperature, the ball just barely fits through the ring. But when the ball is heated sufficiently, it will not pass through the ring. If you do not have this equipment, you can show students a video of this demonstration titled *Heating and Cooling a Metal Ball*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson4#heating_cooling_metal_ball

Question to investigate

How do heating and cooling affect a solid?



Materials for the presenter

- Ball and ring designed specifically for this demonstration
- Bunsen burner for heating the ball
- Room-temperature water (to cool the ball)

Procedure

A. Heating the metal ball

1. Hold the ball in one hand and the ring in the other. Show students how the ball fits through the ring.
2. Place the metal ball in the flame of a Bunsen burner for about 1–2 minutes.
3. Try to push the ball through the metal ring again.



Expected results

The ball will not fit through the ring.

Ask students:

- **Why won't the ball fit through the ring?**

Students should infer that the speed of the atoms in the metal ball has increased. This increased motion competes with the attractions the atoms have for each other, causing the atoms to move slightly further apart. This is why the heated ball is too big to fit through the ring.



When students see that the ball expands, they may wonder if the atoms themselves expanded. Tell students that the atoms do not expand. Instead, the atoms in a solid follow the same rules as the molecules in a liquid. Heating increases molecular motion, causing the atoms to spread a little further apart.

B. Cooling the metal ball

Ask students:

- **What could we do to the metal ball to make it fit through the ring again?**

Students should suggest cooling the ball.

4. Dip the ball in room-temperature water.
5. Push the ball through the metal ring.



Expected results

The ball will fit through the ring.

Ask students:

- **Why does the ball fit through the ring now?**

Students should infer that the atoms slow down enough so that their attractions pull them closer together, making the ball smaller so that it can fit through the ring.



EXPLAIN

4. Show an animation and explain what happened to the atoms in the metal ball as it was heated and cooled.



Show the molecular model animation *Heating and Cooling a Solid*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson4#heating_and_cooling_a_solid

Point out that when metal is heated, the atoms move faster and move slightly further apart. This makes the heated ball expand, which prevents it from passing through the ring.

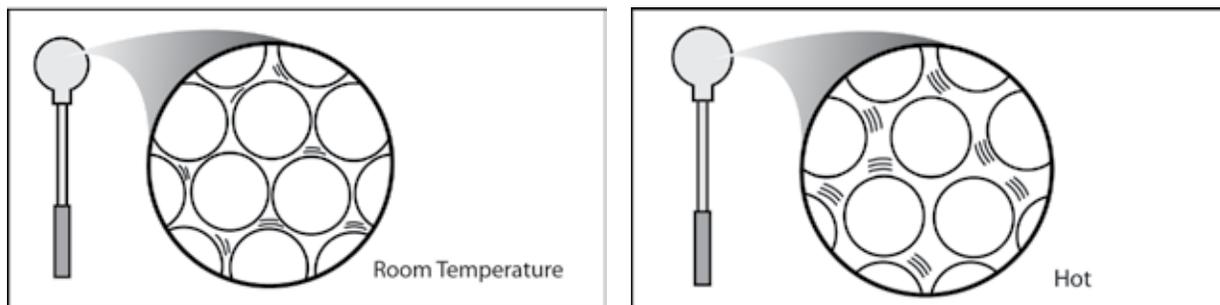
Point out that when the metal is cooled, the atoms move slower and move slightly closer together. This makes the cooled metal ball get slightly smaller so that it fits through the ring again.



Give students time to complete the questions and drawings on the activity sheet about heating and cooling the metal ball.

Project the image *Molecules in a Room-Temperature and Hot Metal Ball*.

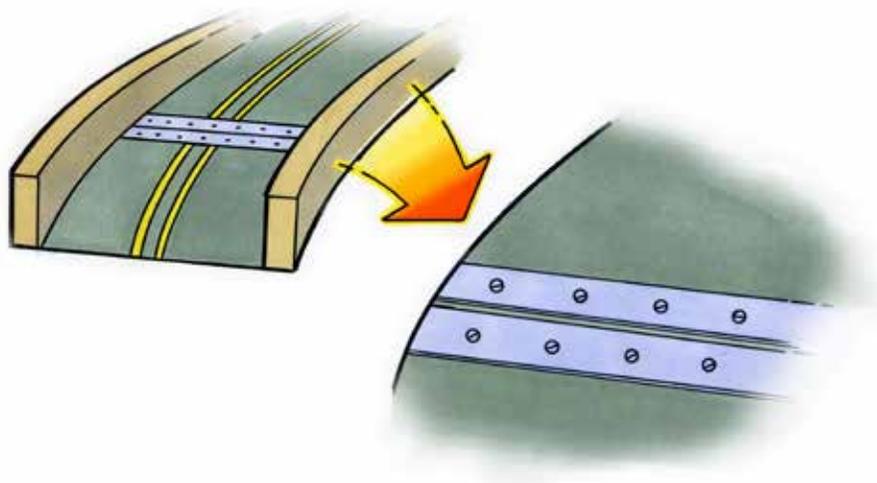
Help students draw circles to represent the atoms in the ball at room-temperature and after it is heated. Have students write captions describing the speed and distance of the atoms in each picture.



EXTEND

5. Have students apply what they have learned about heating and cooling solids to explain why bridges have flexible connections.

Show students the picture of the flexible connection in the road on a bridge. Explain that the surface of the bridge gets colder in winter and hotter in summer than the road on either end of the bridge. This is because the bridge is completely surrounded by cold air in the winter and by hot air in the summer. It is not insulated by the ground beneath it.



Ask students:

- **Knowing what you do about how solids act when they are heated and cooled, why do you think they put flexible connections in the surface of a bridge?**



Students should realize that if the bridge is hotter than the land around it, it should be able to expand a bit without breaking. If it is colder than the land around it, it should be able to contract a bit without breaking.

After the class discussion, have students write their own response to the question about flexible bridge connections on the activity sheet.

EXPLAIN IT WITH ATOMS & MOLECULES

After you watch the molecular model animations of liquids and solids, answer the questions below.

1. How is the motion of the atoms in solid metal different from the motion of the molecules in liquid water?
2. What is it about atoms and molecules in liquids and solids that keep them close to one another even though they are moving?

DEMONSTRATION

3. At room-temperature the metal ball fit through the ring. What happened when your teacher tried to push the heated ball through the ring?
4. What happened to the atoms in the heated metal ball so that it didn't fit through the ring?



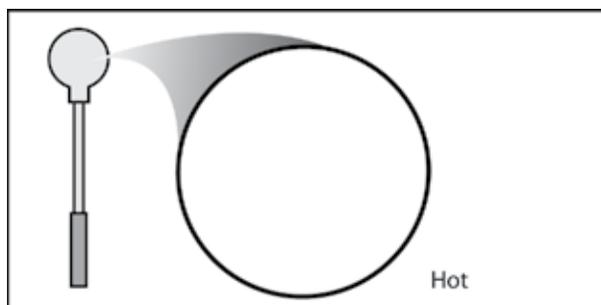
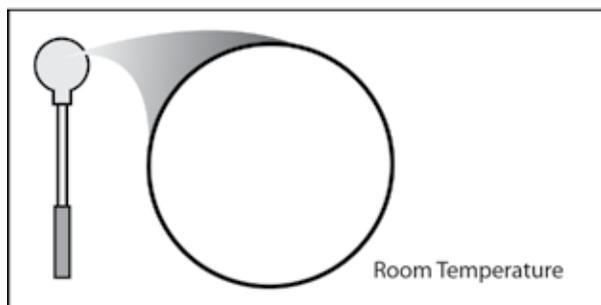
5. After the ball was cooled by putting it in the water, why do you think it fit through the ring again?



EXPLAIN IT WITH ATOMS & MOLECULES

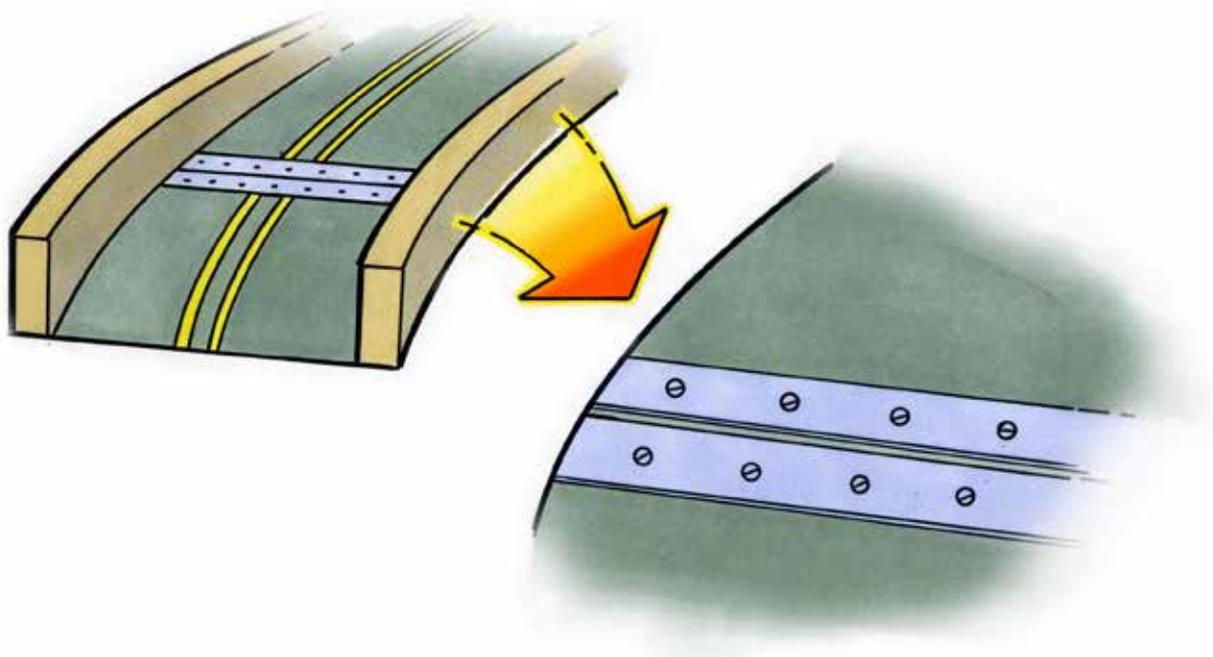
You saw in the animation that atoms in a solid move faster and get slightly further apart when heated. You also saw that they slow down and get slightly closer together when cooled. Use this information to make your own drawing on the molecular level of the metal ball.

6. Draw a model of the atoms in the metal ball at room-temperature and after it has been heated. Use circles and motion lines to show the speed and spacing of the atoms in the room-temperature ball. Include captions like “atoms faster and further apart” or “atoms slower and closer together” to describe your drawings



TAKE IT FURTHER

Look at the picture of the road of a bridge. The road on a bridge gets colder in the winter and hotter in the summer than the road leading to it and away from it. Many bridges have a flexible connection like the one shown in the picture.



7. Knowing what you do about how solids act when they are heated and cooled, why do you think they put flexible connections in the road on a bridge?

Chapter 1, Lesson 5: Air, It's Really There

Key Concepts

- In a gas, the particles (atoms and molecules) have weak attractions for one another. They are able to move freely past each other with little interaction between them.
- The particles of a gas are much more spread out and move more independently compared to the particles of liquids and solids.
- Whether a substance is a solid, liquid, or gas at a certain temperature depends on the balance between the motion of the particles at that temperature and how strong their attractions are for one another.
- Heating a gas increases the speed of its atoms or molecules.
- Cooling a gas decreases the speed of its atoms or molecules.

Summary

Students compare the mass of a basketball when it is deflated and after it has been inflated. The inflated ball has the greater mass, so students can conclude that gas is matter because it has mass and takes up space. Then students consider how heating and cooling affect molecular motion in gases. They dip the mouth of a bottle in detergent solution and observe a bubble growing and shrinking when the bottle is warmed and cooled. Students will learn that the attractions between the particles of gases are weak compared to the attractions between the particles of liquids or solids.

Objective

Students will be able to describe gas as matter. They will also be able to describe, on the molecular level, the effect of heating and cooling on the motion of molecules of a gas.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

- Be sure you and the students wear properly fitting goggles.
- Students should use care when handling hot tap water.

Materials for Each Group

- 2 clear plastic cups
- 8-oz plastic bottle
- Detergent solution in a cup
- Hot water (about 50 °C)
- Cold water

Materials for the Demonstration

- Basketball, very deflated
- Balance that measures in grams
- Pump
- Can of compressed gas (available at any office supply store)

ENGAGE

1. Discuss with students whether they think gas is matter.

Ask students about gases:

- **Are gases, such as the gases in air, matter?**

Students may have questions about whether or not gases are matter. They may also have only a very vague sense of what gases are at all. After students reply, explain that the air around them is made up of some different gases—nitrogen, oxygen, carbon dioxide, water vapor, and very small amounts of some others. Tell students that gases are made of molecules but that the molecules are much further apart than the molecules in liquids or solids. Since the molecules of a gas have mass and take up space, gas is matter.

If students have trouble accepting or appreciating that a gas is made up of molecules, you could try helping them by giving them some numbers to think about. Although these numbers are huge and may be difficult to comprehend, at least students will get the idea that a gas is definitely made of something, takes up space, and has mass. Tell students that in an amount of air about the size of a standard beach ball, there are about 6×10^{23} gas molecules. This is about 600 billion trillion molecules.

Students may have difficulty imagining that gases have mass. It seems like balloons and beach balls, for example, get lighter when we inflate them. When you add air to a balloon or beach ball it actually gets a little heavier. The reason why it seems lighter is not because it has less mass, but because its volume increases so much when it is inflated. This big increase in volume with a small increase in mass makes the balloon or beach ball less dense. That is why it seems lighter when it is inflated. (We'll get to this when we study density in Chapter 3.)

2. Do a demonstration to show that gas has mass.

You will need a balance that measures in grams for either demonstration. If you don't have this type of balance, you can show videos of each demonstration.

www.middleschoolchemistry.com/multimedia/chapter1/lesson5#air_has_mass_ball
www.middleschoolchemistry.com/multimedia/chapter1/lesson5#air_has_mass_can



Materials for the demonstration

- Basketball, very deflated
- Balance that measures in grams
- Pump
- Can of compressed gas (available at any office supply store)

Procedure

A. Basketball

1. Place the deflated ball on the balance to get the initial mass.
2. Ask students if they think the ball will weigh more or less after you pump air into it.
3. Pump as much air into the basketball as you can and then put it back on the balance.



B. Can of compressed gas

4. Place a can of compressed gas on a scale and check its mass.
5. Ask students whether it will weigh more, less, or the same if you squeeze the trigger and let some gas out.
6. Shoot gas out of the can for a few seconds and then place the can back on the scale.



Expected results

The basketball should weigh 2–4 grams more than when it was deflated. The can will weigh a few grams less than it did initially.

3. Show an animation of the molecules of a gas.

Show the molecular model animation *Particles of a Gas*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson5#particles_of_a_gas

Explain to students that the molecules of a gas have very little attraction for one another and barely interact with each other. They just collide and bounce off. It may be hard for students to accept, but in the space between the gas molecules there is nothing.



Note: An inquisitive student might ask: If gas molecules aren't attracted to each other and can just float around, why don't they all just float away? That is a very good question. In fact very light gases like hydrogen and helium have floated away and there are very little of these gases in our atmosphere. Different heavier gases, such as nitrogen, oxygen, water vapor, and carbon dioxide, surround the Earth. In the big picture, gravity holds these gases near the Earth as our atmosphere.

Give each student an activity sheet.

Students will answer questions about the demonstration on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

4. Have students do an activity to find out how heating and cooling affect gases.



Question to investigate

How does heating and cooling affect a gas?

Materials for each group

- 2 clear plastic cups
- 8-oz plastic bottle
- Detergent solution in a cup
- Hot water (about 50 °C)
- Cold water

Teacher preparation

Make the detergent solution for the entire class by adding 4 teaspoons of dishwashing liquid and 4 teaspoons of sugar to $\frac{1}{2}$ cup of water. Gently stir until the detergent and sugar are dissolved. Place about 1 tablespoon of detergent solution in a wide clear plastic cup for each group.



Procedure

A. Warming the air inside the bottle

1. Pour hot water into an empty cup until it is about $\frac{1}{2}$ -full.
2. Turn the bottle over and dip the opening of the bottle into the detergent to get a film of detergent covering the rim.



3. While holding the bottle, slowly push the bottom of the bottle down into the hot water.

Ask students:

- **What can you do to make the bubble go down?**

If students have trouble thinking of an answer, remind them that heating the gas increased the speed of the molecules, which made the bubble grow. Students should suggest that they should cool the gas in the bottle. This can be done by putting the base of the bottle into cold water.

B. Cooling the air inside the bottle

4. Pour cold water into another cup until it is about $\frac{1}{2}$ -full.
5. If there is still a bubble on the bottle, slowly push the bottom of the bottle down into the cold water.
6. If a bubble is not still on the bottle, make another bubble by dipping the opening into detergent and then pushing the bottom of the bottle into hot water again.
7. While holding the bottle, slowly push the bottom of the bottle down into the cold water.



Expected results

When the bottle is placed in hot water, a bubble forms at the top of the bottle. When the bottle is placed in cold water, the bubble gets smaller. It may actually be pushed down into the bottle.

5. Record and discuss student observations.

Give students time after the activity to record their observations by answering the following questions on their activity sheet. Once they have answered the questions, discuss their observations as a whole group.



- **What happened to the film of detergent solution when you placed the bottle in hot water?**
It formed a bubble.
- **What happened to the bubble when you placed the bottle in cold water?**
It shrunk and went into the bottle.

Tell students that you will show them an animation to help explain what caused the bubble to grow and shrink when the air in the bottle was heated and cooled.

EXPLAIN

6. Show an animation of a bubble as it is heated and cooled.



Show the animation *Heating and Cooling Gas in a Bottle*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson5#heating_cooling_gas_in_bottle

Tell students that the red arrows in the animation represent the outside air pushing down on the bubble film. Explain that heating the air inside the bottle makes the molecules move faster. These faster-moving molecules hit the inside of the bottle and the bubble film harder and more frequently. These molecules push against the inside of the bubble film harder than the surrounding air pushes from the outside. This pushes the bubble film up and out, forming a bubble.

Cooling the gas makes the molecules move more slowly. These slower-moving molecules hit the inside of the bottle and the bubble film less often and with less force. The molecules in the surrounding air are moving faster and push against the bubble from the outside. Since these outside molecules are pushing harder, the bubble gets pushed down and gets smaller.

7. Have students answer the questions about the growing and shrinking bubble on the activity sheet.



Give students time to complete the following questions. They should refer to the drawing included on the next page and on the activity sheet.

Once they have answered the questions, discuss their explanations as a whole group.

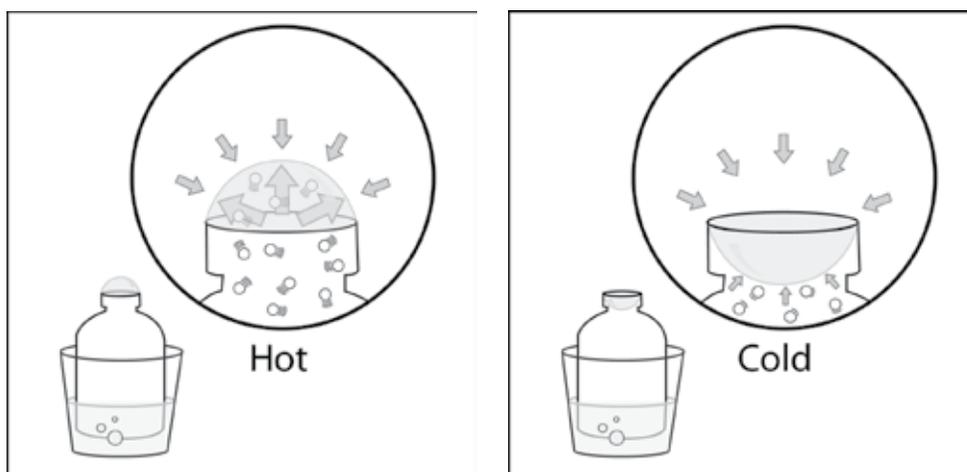
- **What caused the bubble to form when you placed the bottle in hot water? Be sure to write about the speed of the molecules inside the bottle and the pressure from the outside air.**

Point out that the molecules of air inside the bottle move faster when they are heated and push harder against the outside air. This makes the bubble form.

- **Why did the bubble get smaller when you placed the bottle in cold water? Be sure to write about the speed of the molecules inside the bottle and the pressure from the outside air.**

When the air inside the bottle is cooled, the molecules move slower and do not push as hard against the outside air. The outside air pushes against the bubble, making it go down.

Read more about the contraction and expansion of gases in the additional teacher background section at the end of this lesson.



8. Have students compare the molecules in solids, liquids, and gases.

Show the molecular model animation *Comparing Solids, Liquids, and Gases*.

www.middle-school-chemistry.com/multimedia/chapter1/lesson5#comparing_solids_liquids_and_gases



Be sure students realize that the molecules shown are from three different substances all at room-temperature. The solid is not melting to become a liquid and the liquid is not evaporating to become a gas. The model is not trying to show state changes but instead show three different substances (such as metal, water, and air), which are solid, liquid, and gas at room-temperature.

Explain the following differences to students:

- **Solid**—Particles (atoms or molecules) are very attracted to one another. They vibrate but do not move past one another. The atoms or molecules stay in fixed positions because of their strong attractions for one another. A solid has a definite volume and a definite shape.
- **Liquid**—Particles (atoms or molecules) are attracted to one another. They vibrate but are also able to move past one another. A liquid has a definite volume but does not have a definite shape.
- **Gas**—Particles (atoms or molecules) are not attracted to each other much at all and move freely. A gas does not have a definite shape or volume. The atoms or molecules of a gas will spread out evenly to fill any container.

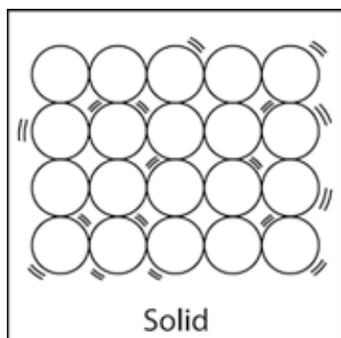
You could use the following example to help students appreciate how far apart the molecules of a gas are compared to the molecules in a liquid or solid:

- The molecules that make up a gas are about 100 to 1000 times further apart than the molecules of a solid or liquid. Imagine what a tablespoon of water looks like. If that same number of molecules was a gas, they would be spread out enough to fill up a whole beach ball. At room-temperature they are moving at about 1000 miles per hour, but over very short distances.

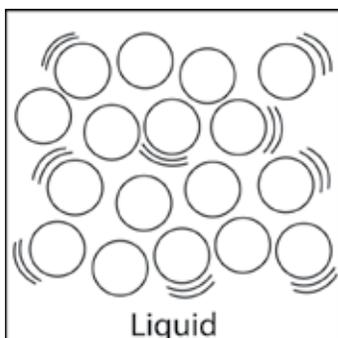
Draw or project the illustration *Solid, Liquid, and Gas*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson5#solid_liquid_and_gas

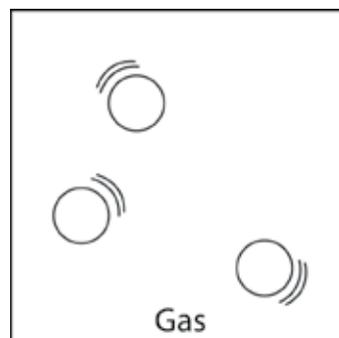
Have students use the projected illustration as a reference as they draw a model of solids, liquids, and gases on their activity sheet. Point out that the number of motion lines is the same for the solid, the liquid, and the gas. This indicates that the different substances are at the same temperature. Have students write captions like those listed below to describe the molecules in solids, liquids, and gases.



- Attractions strong enough to keep atoms in orderly arrangement
- Vibrate in fixed positions
- Definite volume and shape



- Attractions keep particles together but they can slide past each other
- Random arrangement
- Definite volume, not definite shape



- Attractions too weak to keep particles together
- Particles move independently
- No definite volume or shape

EXTEND

9. Have students apply what they have learned to explain why a balloon grows when it is heated.

Tell students to consider the following scenario:

Imagine that you work at a party store during the summer. You are going to ride home with the owner of the store whose car has been sitting in the hot sun all day long. The owner tells you that you can take home a big bunch of balloons, but advises you to not blow the balloons up all of the way. Knowing what you do about heating the molecules of a gas, explain why the owner's advice is wise.

You may choose to show the animation *Heating Molecules of a Gas* if you would like to give students a hint.

www.middleschoolchemistry.com/chapter1/lesson5/#heating_and_cooling_gas

DEMONSTRATION—BASKETBALL AND COMPRESSED GAS

The demonstrations with the basketball and the can of compressed air were meant to show something about gases and matter. Matter is anything that has mass and takes up space.

1. Think about the demonstration with the deflated and inflated basketball. The basketball weighed more after it was inflated with air than when it was deflated. How does this show that gas is matter?



2. Think about the demonstration with the can of compressed gas. The can weighed less after some gas was shot out of the can. How does this show that gas is matter?



EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animation of gas molecules inside a balloon.

3. What did you notice about the molecules of a gas:
 - Do the molecules of a gas have strong or weak attractions?
 - Are the molecules of a gas randomly or orderly arranged?
 - When the molecules of a gas hit each other, do they normally stick together or bounce off?

ACTIVITY

Question to investigate

How do heating and cooling affect a gas?



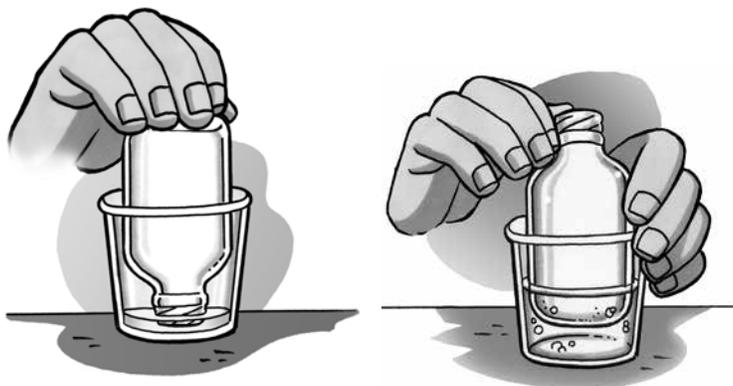
Materials for each group

- 2 clear plastic cups
- 8-oz plastic bottle
- Detergent solution in cup
- Hot water
- Cold water

Procedure

A. Warming the air inside the bottle

1. Pour hot water into an empty cup until it is about $\frac{1}{2}$ -full.
2. Turn the bottle over and dip the opening of the bottle into the detergent to get a film of detergent covering the rim.
3. While holding the bottle, slowly push the bottom of the bottle down into the hot water.



B. Cooling the air inside the bottle

4. Pour cold water into another cup until it is about $\frac{1}{2}$ -full.
5. If there is still a bubble on the bottle, slowly push the bottom of the bottle down into the cold water.
6. If a bubble is not still on the bottle, make another bubble by dipping the opening into detergent and then pushing the bottom of the bottle into hot water again.
7. While holding the bottle, slowly push the bottom of the bottle down into the cold water.

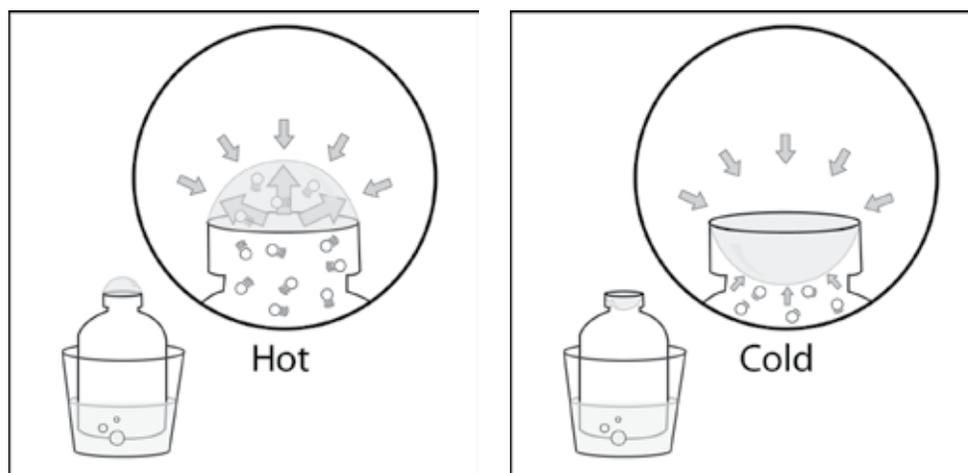


WHAT DID YOU OBSERVE?

4. What happened to the film of detergent solution when you placed the bottle in hot water?
5. What happened to the bubble when you placed the bottle in cold water?

EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animation showing the air molecules inside a bottle when it is placed in hot and cold water. Think of the animation and use the drawing below as a reference to answer the questions at the top of the next page.

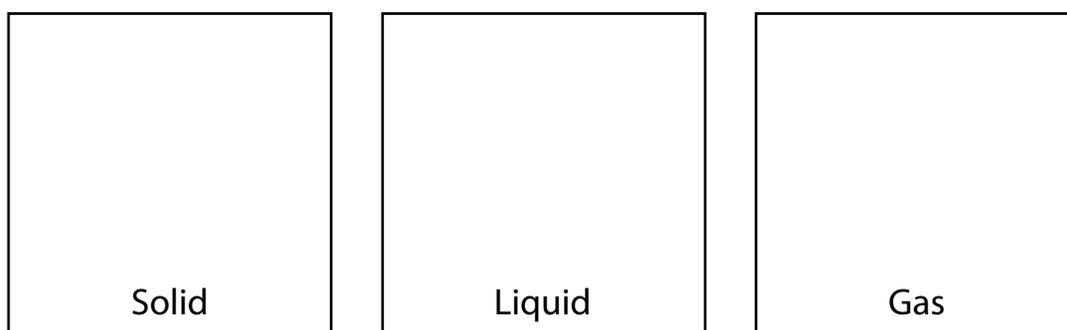


6. What caused the bubble to form when you placed the bottle in hot water? Be sure to write about the speed of the molecules inside the bubble and the force on the bubble from the outside air.

7. Why did the bubble get smaller when you placed the bottle in cold water? Be sure to write about the speed of the molecules inside the bubble and the force on the bubble from the outside air.

You saw an animation about the molecules in solids, liquids, and gases.

8. Draw circles to represent the molecules in a solid, liquid, and gas. Because all three different substances are all at the same temperature, draw the same number of motion lines near the circles for each substance. Under each box, write about the arrangement and motion of the molecules and the attractions the molecules have for one another.



TAKE IT FURTHER

9. Imagine that you work at a party store during the summer. You are going to ride home with the owner of the store whose car has been sitting in the hot sun all day long. The owner tells you that you can take home a big bunch of balloons, but advises you to not blow the balloons up all of the way. Explain why the owner's advice is wise. Be sure to discuss how heating affects the motion of the molecules in a gas.

Chapter 1—Student Reading

Chemistry is the study of matter

You could say that chemistry is the science that studies all the stuff in the entire world. A more scientific term for “stuff” is “**matter**.” So chemistry is the study of matter. Matter is all the physical things in the universe. All the stars in the galaxies, the sun and planets in our solar system, the Earth, and everything on it and in it are matter.



All human-made objects, all organisms, the gases in the atmosphere, and anything else that has mass and takes up space, including you, are examples of matter.

Chemistry is special because it looks at matter all the way down to its smallest parts: the **atoms** and **molecules** that matter is made of. To give you an idea about how small atoms and molecules are, use a metric ruler to look at the length of one millimeter. It is about the size of a dash like this one -. Try drawing a tiny line or dot that is about $1/10$ as long as the dash. It might be about the size of a period like the one at the end of this sentence. A hydrogen atom is about 1 ten millionth of the size of the period. So it would take about 10 million hydrogen atoms lined up next to each other to go from one side of the period to the other.

Here is another way to imagine how small atoms and molecules are. In about 1 tablespoon of water, there are about 600 billion trillion water molecules. That's 600,000,000,000,000,000,000,000. This number is so huge that even if you could count one million molecules every second, it would take you about 200 million centuries or about 20 billion years to count all the molecules in a tablespoon of water.

Studying chemistry can help make sense of many of the different things you see and do every day. What you eat and drink, the weather outside, the soap and water you wash with, and the clothes you wear, are all a result of chemistry. The sports equipment you use, the materials your house is made of, the way you get to school, and the electronic equipment you use are all a result of the interactions of atoms and molecules.



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Having a better idea of what atoms and molecules are and how they interact can help you better understand the world around you.

Matter is made of atoms and molecules

We have already used the term *atom* and *molecule* a couple of times. You will learn a lot more about atoms and molecules in later chapters. For now, let's say that atoms and molecules are the extremely tiny particles that make up all the matter on Earth. An atom is the basic building block of all matter. A molecule is made of two or more atoms connected or bonded together.

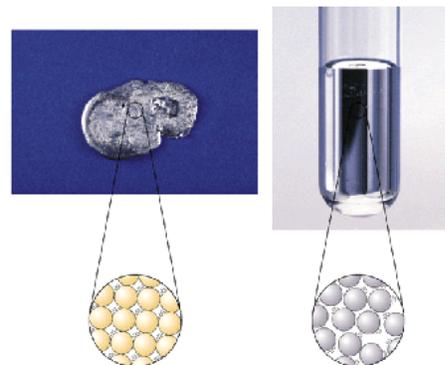
Even though atoms and molecules are not the same, the model we are using in Chapter 1 shows both atoms and molecules as little circles or spheres. This model makes it easier to show some of the basic characteristics of the different states of matter on Earth.



Matter—Solid, Liquid, Gas

On Earth, matter is either found as a **solid**, **liquid**, or **gas**. A particular solid, liquid, or gas might be made up of individual atoms or molecules.

Here is a simplified model of three different substances. One is a solid, another is a liquid and the other is a gas



In the picture, the little motion lines show that the particles (atoms or molecules) that make up the solid, liquid, and gas are moving. In later chapters, the models of atoms and molecules will be shown with more detail.

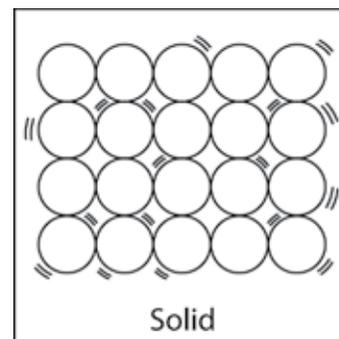
As you look at these pictures, think about these two big ideas which are always true when talking about matter:

- Matter (solid, liquid, and gas) is made up of tiny particles called atoms and molecules.
- The atoms or molecules that make up matter are always in motion.
- These first two ideas make up a very important theory called the Kinetic-molecular theory of matter.

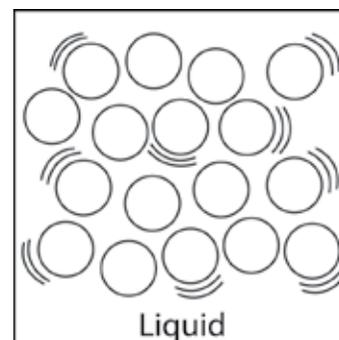
Another big idea is that:

The **atoms** or molecules that make up a solid, liquid or gas are attracted to one another.

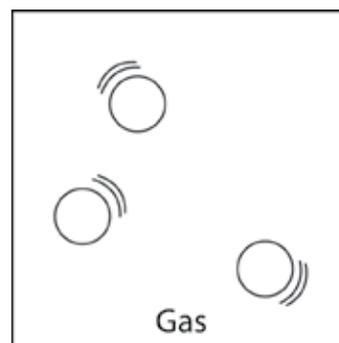
In a **solid**, the atoms are very attracted to one another. Because of this strong attraction, the atoms are held tightly together. The attractions are strong enough that the atoms can only vibrate where they are. They cannot move past one another. This is why a solid keeps its shape.



In a **liquid**, the molecules are also in motion. The attractions between the molecules in liquids are strong enough to keep the molecules close to each other but not in fixed positions. Although the molecules stay very near one another, the attractions allow the molecules of a liquid to move past one another. This is why a liquid can easily change its shape.



In a **gas**, the molecules are also moving. The attractions between the molecules of a gas are too weak to bring the molecules together. This is why gas molecules barely interact with one another and are very far apart compared to the molecules of liquids and solids. A gas will spread out evenly to fill any container.

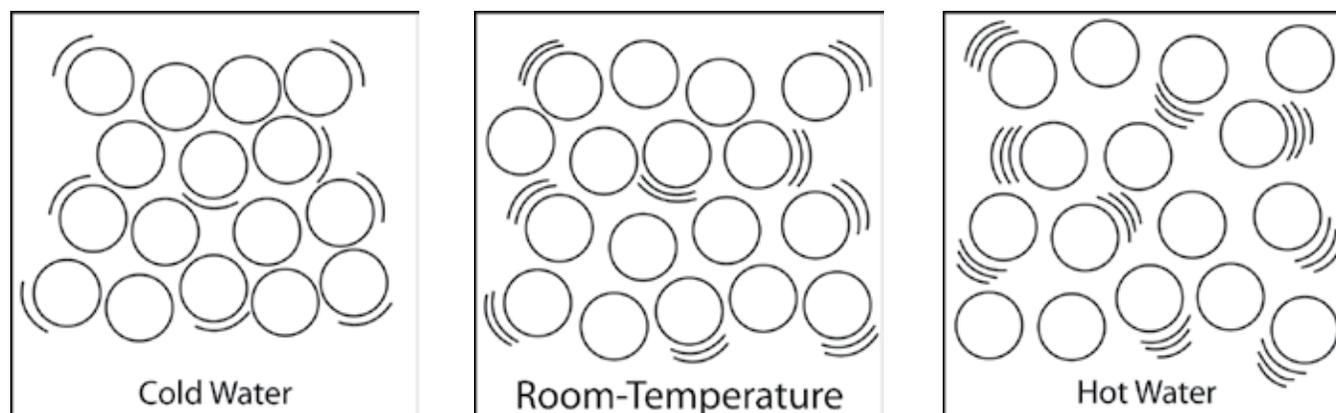


When looking at the different states of matter, it's kind of like a competition between the attractions the molecules have for each other compared to the motion of their molecules. The attractions tend to keep the atoms or molecules together while the motion tends to make the atoms or molecules come apart.

Comparing Matter			
	Solids	Liquids	Gases
Attractions	Atoms or molecules are very attracted to one another.	Atoms or molecules are attracted to one another.	Atoms or molecules are barely attracted to one another.
Movement	Vibrate but do not move past one another.	Vibrate but are able to move past one another.	Vibrate and move freely past each other.
Volume and Shape	Have a definite volume and a definite shape.	Have a definite volume, but does not have a definite shape.	Does not have a definite volume or a definite shape.

Heating and cooling liquids

Heating and cooling a liquid can affect how far apart or close together the molecules are.



One example is the red alcohol inside the thin tube of a thermometer. When the thermometer is heated, the molecules of alcohol move faster. This faster motion competes with the attraction between the molecules which causes them to spread out a little. They have nowhere else to go so they move up the tube.

When the thermometer is cooled, the molecules of alcohol slow down and the attractions bring the molecules closer together. This attraction between the molecules brings the alcohol down in the tube.

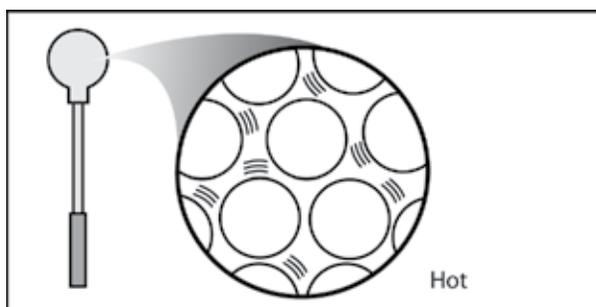
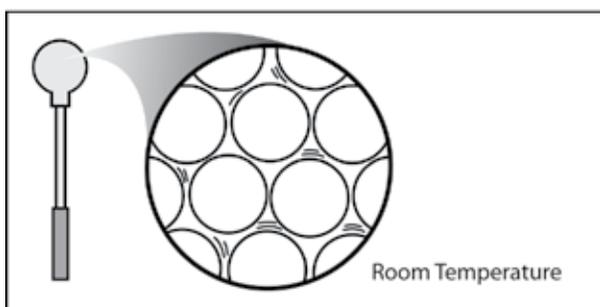
Heating and cooling solids

There is a device made out of a metal ball and ring that lets you see the effect of heating and cooling a solid. At room temperature, the ball just barely fits through the ring.

When the ball is heated sufficiently, it will not fit through the ring.

This is because heating the metal ball increases the motion of its atoms. This motion competes with the attractions between the atoms and makes the atoms move slightly further apart. The slightly larger ball no longer fits through the ring.

When the metal ball is cooled, the atoms slow down and their attractions bring the atoms closer together. This allows the metal ball to fit through the ring again.

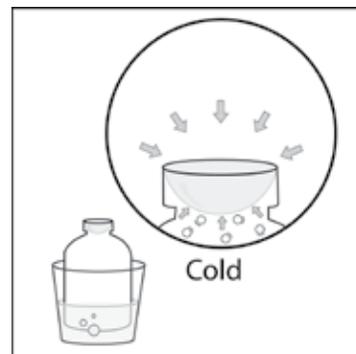
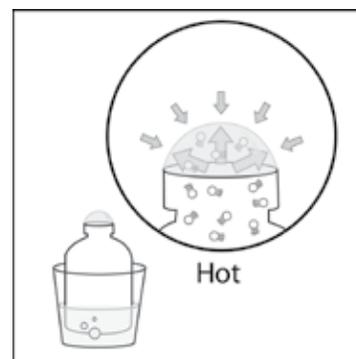


Heating and cooling gases

The molecules of a gas are not very attracted to each other and are much further apart than in liquids and solids. This is why heating a gas easily increases the motion of the gas.

For example, if you dip the opening of a bottle in a detergent solution and then heat the bottle, a bubble will form on the bottle. This happens because heating the bottle increases the motion of the gas molecules inside the bottle. Since molecules of the gas are not very attracted to each other, they spread out quickly and easily. The molecules hit the inside of the bottle and the bubble film harder and more often. The molecules push against the inside of the film harder than the surrounding air pushes from the outside. This pushes the bubble film out and forms a bubble.

If you cool the bottle while the bubble is still on top, the bubble will shrink and may go inside the bottle. This happens because cooling the gas causes its molecules to slow down. These slower-moving molecules hit the inside of the bubble film less often and with less force. The molecules in the outside air are moving faster and push against the bubble from the outside. Since the outside molecules are pushing harder, the bubble gets pushed down and in and gets smaller.



Additional Teacher Background

Chapter 1, Lesson 5, p.49

Contracting and expanding happens differently for a gas than it does for liquids and solids.

Cooling a liquid or solid

When a liquid or solid is cooled, the particles (atoms or molecules) slow down. The slower motion allows the attractions between the particles to pull themselves closer together causing the liquid or solid to contract.

Cooling a gas

But it works differently for gases. When a gas in a flexible container like a balloon or bubble is cooled, its atoms or molecules also slow down. But the attractions between the particles of a gas are so weak that they cannot pull themselves closer together like in a liquid or solid. Instead, the slower-moving molecules hit the inside of the balloon or bubble less frequently and with less force. This results in a lower pressure inside the bubble than outside. The higher pressure on the outside pushes on the bubble, making it contract.

Heating a liquid or solid

When a liquid or solid is heated, the particles it is made of move faster. This increased motion competes with the attractions the particles have for each other causing them to move slightly further apart.

Heating a gas

When a gas is heated, the particles also move faster but they do not have to compete with significant attractions. The particles hit the inside of the bubble more frequently and with more force. This results in a higher pressure inside the bubble than outside. The higher pressure on the inside pushes on the bubble, making it expand.

Chapter 2, Lesson 1: Heat, Temperature, and Conduction

Key Concepts

- Adding energy (heating) atoms and molecules increases their motion, resulting in an increase in temperature.
- Removing energy (cooling) atoms and molecules decreases their motion, resulting in a decrease in temperature.
- Energy can be added or removed from a substance through a process called conduction.
- In conduction, faster-moving molecules contact slower-moving molecules and transfer energy to them.
- During conduction the slower-moving molecules speed up and the faster-moving molecules slow down.
- Temperature is a measure of the average kinetic energy of the atoms or molecules of a substance.
- Heat is the transfer of energy from a substance at a higher temperature to a substance at a lower temperature.
- Some materials are better conductors of heat than others.

Summary

Students will do an activity in which heat is transferred from hot water to metal washers and then from hot metal washers to water. Students will view a molecular animation to better understand the process of conduction at the molecular level. Students will also draw their own model of the process of conduction.

Objective

Students will be able to describe and draw a model, on the molecular level, showing how energy is transferred from one substance to another through conduction.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles. Use caution when handling hot water.

Materials for Each Group

- 2 sets of large metal washers on a string
- Styrofoam cup filled with hot water
- Room-temperature water
- 2 thermometers
- Graduated cylinder or beaker

Materials for the Teacher

- 1 Styrofoam cup
- Thermometer
- Hot plate or coffee maker
- Large beaker or coffee pot

Note: Energy can also be transferred through radiation and convection, but this chapter only deals with heat transfer through conduction.

ENGAGE

1. Discuss what happens when a spoon is placed in a hot liquid like soup or hot chocolate.

Ask students:

- Did you ever put a metal spoon in hot soup or hot chocolate and then touch the spoon to your mouth? What do you think might be happening, between the molecules in the soup and the atoms in the spoon, to make the spoon get hot? It's not necessary for students to answer these questions completely at this time. It is more important that they begin to think that something is going on at the molecular level that causes one substance to be able to make another hotter.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Have students explore what happens when room-temperature metal is placed in hot water.

If you cannot get the materials for all groups to do this activity, you can do the activity as a demonstration or show students the videos:

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#heating_washers

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#cooling_washers

Question to investigate

Why does the temperature of an object change when it is placed in hot water?

Materials for each group

- 2 sets of large metal washers on a string
- Styrofoam cup filled with hot water
- Room-temperature water
- 2 thermometers
- Graduated cylinder or beaker



Materials for the teacher

- 1 Styrofoam cup
- Thermometer
- Hot plate or coffee maker
- Large beaker or coffee pot

Teacher preparation

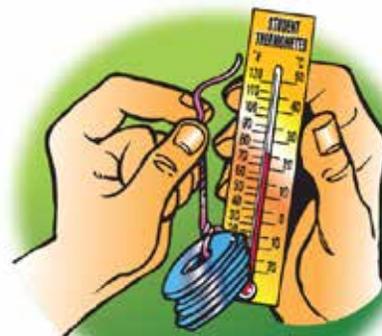
- Use a string to tie 5 or 6 metal washers together as shown. Each group of students will need two sets of washers, each tied with a string.
- Hang one set of washers for each group in hot water on a hot plate or in water in a coffee maker so that the washers can get hot. These washers will need to remain hot until the second half of the activity.
- The other set should be left at room-temperature and may be distributed to students along with the materials for the activity.
- Immediately before the activity, pour about 30 milliliters (2 tablespoons) of hot water (about 50 °C) into a Styrofoam cup for each group. Be sure to pour one cup of hot water for you to use as a control.

Tell students that they are going to see if the temperature of hot water changes as a result of placing room-temperature metal washers in the water. The only way to tell if the washers cause the temperature to change is to have a cup of hot water without washers. Explain that you will have this cup of hot water, which will be the control.

You will need to place your thermometer in the cup of hot water at the same time the students do. Have students record the initial temperature of the control in their charts on the activity sheet, along with the initial temperature of their own cup of hot water. The temperature of the two samples should be about the same.

Procedure

1. Place a thermometer in your cup to measure the initial temperature of the water. Record the temperature of the water in the “Before” column in the chart on the activity sheet. Be sure to also record the initial temperature of the water in the control cup.
2. Use another thermometer to measure the temperature of the washers. Record this in the “Before” column.



Note: It is a little awkward to take the temperature of the washers with a regular thermometer because there is such a small point of contact between the bulb of the thermometer and the surface of the washers. The washers should be about room-temperature.

Ask students to make a prediction:

- **What will happen to the temperature of the water and the washers if you place the washers into the hot water?**

3. With the thermometer still in the water, hold the string and lower the metal washers all the way into the hot water.
4. Observe any change in the temperature of the water. Leave the washers in the water until the temperature stops changing. Record the temperature of the water in each cup in the “After” column.



Room-temperature washers placed in hot water		
Temperature of...	Before	After
Water in your cup		
Water in the control cup		
Metal washers		

5. Remove the washers from the water. Then take and record the temperature of the washers in the “After” column.
6. Empty the cup in a waste container or sink.

Expected results

The temperature of the water will decrease a bit and the temperature of the washers will increase a bit. The *amount* of temperature decrease and increase is really not that important. What is important is that there is a temperature decrease in the water and a temperature increase in the washers.

Note: Eventually two objects at different temperatures that are in contact will come to the same temperature. In the activity, the washers and water will most likely be different temperatures. For the purposes of this activity, the washers and water are only in contact for a short time, so most likely will not come to the same temperature.

Read more about energy and temperature in the additional teacher background section at the end of this lesson.

Students may ask why the temperature of the water went down by a different amount than the temperature of the washers went up. The same amount of energy left the water as went into the washers, but it takes a different amount of energy to change the temperature of different substances.

3. Have students explore what happens when hot metal is placed in room-temperature water.



Ask students:

- **How do you think the temperature will change if you place hot washers into room-temperature water?**

Pour about 30 milliliters of room-temperature water into the control cup. Place a thermometer in the cup and tell students the temperature of the water.

1. Pour about 30 milliliters of room-temperature water into your Styrofoam cup.
2. Place a thermometer into the water and record its temperature in the “Before” column in the chart on the activity sheet. Be sure to also record the initial temperature of the water in the control cup.
3. Remove the washers from the hot water where they have been heating and quickly use a thermometer to measure the temperature of the washers. Record this in the “Before” column on your activity sheet.
4. With the thermometer still in the water, hold the string and lower the hot metal washers all the way into the water.
5. Observe any change in the temperature of the water. Leave the washers in the water until the temperature stops changing. Record the temperature of the water in your cup in the “After” column in the chart below. Also record the temperature of the water in the control cup.
6. Remove the washers from the water. Take and record the temperature of the washers.

Hot washers placed in room-temperature water		
Temperature of...	Before	After
Water in your cup		
Water in the control cup		
Metal washers		

Expected results

The temperature of the water increases and the temperature of the washers decreases.

4. Discuss student observations and what may have caused the temperature of the metal washers and water to change.

Ask students:

- **How did the temperature of the washers and water change in both parts of the activity?**
Based on their data, students should realize that the temperature of both the washers and water changed.
- **Knowing what you do about heating and cooling atoms and molecules, why do you think the temperature changed?**
If necessary, guide students' thinking about why the temperature of each changed by asking them which were probably moving faster, the atoms in the metal washers or the molecules in the water. Tell students that the molecular model animation you will show next will show them why the temperature of both changed.

EXPLAIN

5. Show two animations to help students understand how energy is transferred from one substance to another.



Show the molecular model animation *Heated Spoon*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#heated_spoon

Point out to students that the water molecules in the hot water are moving faster than the atoms in the spoon. The water molecules strike the atoms of the spoon and transfer some of their energy to these atoms. This is how the energy from the water is transferred to the spoon. This increases the motion of the atoms in the spoon. Since the motion of the atoms in the spoon increases, the temperature of the spoon increases.

It is not easy to notice, but when the fast-moving water molecules hit the spoon and speed up the atoms in the spoon, the water molecules slow down a little. So when energy is

transferred from the water to the spoon, the spoon gets warmer and the water gets cooler.

Explain to students that when fast-moving atoms or molecules hit slower-moving atoms or molecules and increase their speed, energy is transferred. The energy that is transferred is called *heat*. This energy transfer process is called *conduction*.

Show the molecular model animation *Cooled Spoon*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#cooled_spoon



Point out to students that in this case, the atoms in the spoon are moving faster than the water molecules in the cold water. The faster-moving atoms in the spoon transfer some of their energy to the water molecules. This causes the water molecules to move a little faster and the temperature of the water to increase. Since the atoms in the spoon transfer some of their energy to the water molecules, the atoms in the spoon slow down a little. This causes the temperature of the spoon to decrease.

Ask students:

- **Describe how the process of conduction caused the temperature of the washers and water to change in the activity.**

Room-temperature washers in hot water

When the room-temperature washers are placed in hot water, the faster-moving water molecules hit the slower-moving metal atoms and make the atoms in the washers move a little faster. This causes the temperature of the washers to increase. Since some of the energy from the water was transferred to the metal to speed them up, the motion of the water molecules decreases. This causes the temperature of the water to decrease.

Hot washers in room-temperature water

When the hot metal washers are placed in the room temperature water, the faster-moving metal atoms hit the slower-moving water molecules and make the water molecules move a little faster. This causes the temperature of the water to increase. Since some of the energy from the metal atoms was transferred to the water molecules to speed them up, the motion of the metal atoms decreases. This causes the temperature of the washers to decrease.

6. Discuss the connection between molecular motion, temperature, and conduction.

Ask students:

- **How does the motion of the atoms or molecules of a substance affect the temperature of the substance?**

If the atoms or molecules of a substance are moving faster, the substance has a higher temperature. If its atoms or molecules are moving slower, then it has a lower temperature.

- **What is conduction?**

Conduction occurs when two substances at different temperatures are in contact. Energy is always transferred from the substance with the higher temperature to the one at lower temperature. As energy is transferred from the hotter substance to the colder one, the colder substance gets warmer and the hotter substance gets cooler. Eventually the two substances become the same temperature.

Students tend to understand heating but often have a misconception about how things are cooled. Just like heating a substance, cooling a substance also works by conduction. But instead of focusing on the slower-moving molecules speeding up, you focus on the faster-moving molecules slowing down. The faster-moving atoms or molecules of the hotter substance contact slower-moving atoms or molecules of the cooler substance. The faster-moving atoms and molecules transfer some of their energy to the slower-moving atoms and molecules. The atoms and molecules of the hotter substance slow down, and its temperature decreases. An object or substance can't get colder by adding "coldness" to it. Something can only get colder by having its atoms and molecules transfer their energy to something that is colder.

7. Have students draw molecular models to show conduction between a spoon and water.

Note: In the model you will show students, the change in speed of both the water molecules and the atoms in the spoon is represented with different numbers of motion lines. Students may remember that when atoms or molecules move faster, they get further apart, and when they move slower, they get closer together. For this activity, the change in distance between water molecules or between atoms in the spoon is not the focus, and therefore it is not shown in the model. You could tell students that models can emphasize one feature over another, in order to help focus on the main point being represented.



Room-temperature spoon placed in hot water

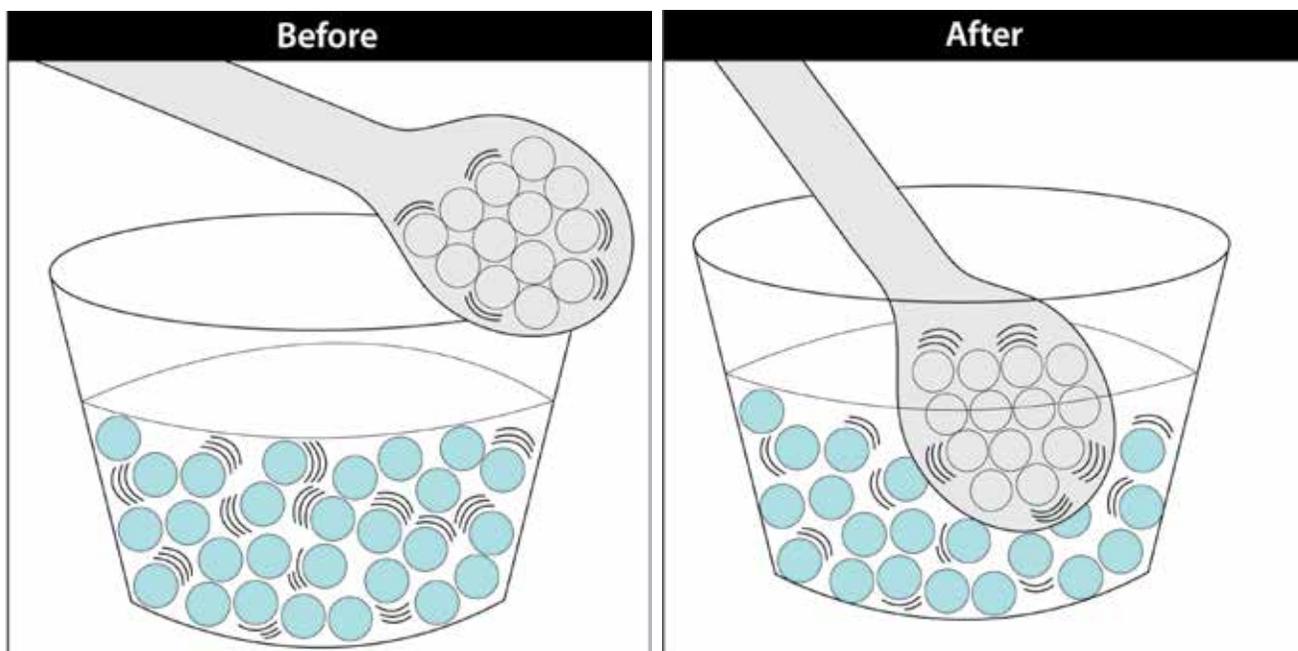
Project the illustrations *Spoon in Hot Water Before & After* from the activity sheet.

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#spoon_in_hot_water

Have students look at the motion lines in the “Before” picture on their activity sheet. Then ask students how the motion of the atoms and molecules would change in the “After” picture. The activity sheet, along with the image you are projecting, does not have motion lines drawn in the “After” picture. Putting these in correctly is the students’ task.



Tell students to add motion lines to the “After” illustration and add descriptive words like “warmer” or “cooler” to describe the change in temperature of the water and the spoon.

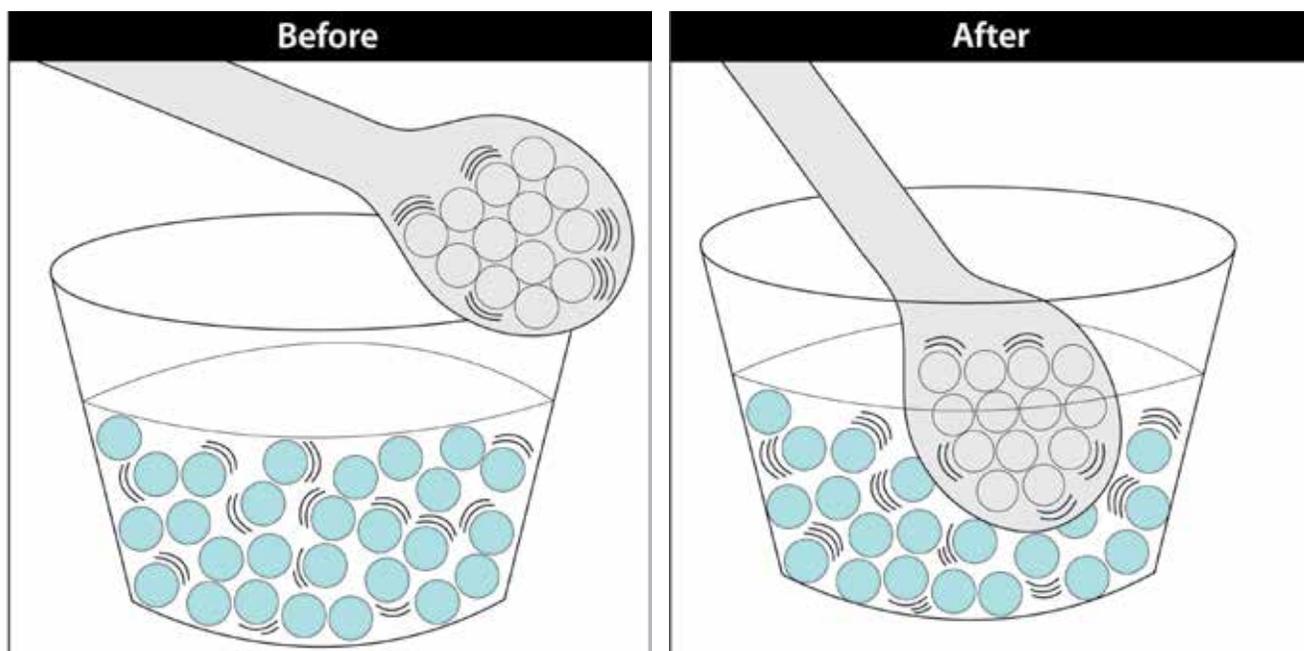


Hot spoon placed in room-temperature water

Project the illustrations *Hot spoon in room-temperature water before & after* from the activity sheet

www.middleschoolchemistry.com/chapter2/lesson1#spoon_in_room_temperature_water

Have students look at the second set of “Before” and “After” pictures. Tell students to add motion lines to the “After” illustration and add descriptive words like “warmer” or “cooler” to describe the change in temperature of the water and the spoon.



8. Show a simulation to illustrate that temperature is the average kinetic energy of atoms or molecules.

The following simulation shows that at any temperature, the atoms or molecules of a substance are moving at a variety of speeds. Some molecules are moving faster than others, some slower, but most are in-between.

Note: After pressing “Start”, the simulation works best if you cycle through all the buttons before using it for instruction with students..

Show the simulation Temperature.

www.middleschoolchemistry.com/multimedia/temperature.php

- After cycling through the “Cold”, “Medium”, and “Hot” buttons, choose “Medium” to begin the discussion with students. Tell students that this simulation shows the relationship between energy, molecular motion, and temperature.

Tell students that anything that has mass and is moving, no matter how big or small, has a certain amount of energy, called kinetic energy. The temperature of a substance gives you information about the kinetic energy of its molecules. The faster the molecules of a substance move, the higher the kinetic energy, and the higher the temperature. The slower the molecules move, the lower the kinetic energy, and the lower the temperature. But at any temperature, the molecules don't all move at the same speed so temperature is actually a measure of the average kinetic energy of the molecules of a substance.

- These ideas apply to solids, liquids, and gases. The little balls in the simulation represent molecules and change color to help visualize their speed and kinetic energy. The slow ones are blue, the faster ones are purple or pink, and the fastest are red. Explain also that individual molecules change speed based on their collisions with other molecules. Molecules transfer their kinetic energy to other molecules through conduction. When a fast-moving molecule hits a slower-moving molecule, the slower molecule speeds up (and turns more red) and the faster molecule slows down (and turns more blue).
- Explain that at any temperature, most of the molecules are moving at about the same speed and have about the same kinetic energy, but there are always some that are moving slower and some that are moving faster. The temperature is actually a combination, or average, of the kinetic energy of the molecules. If you could place a thermometer in this simulation, it would be struck by molecules going at different speeds so it would register the average kinetic energy of the molecules.

To add energy, start with “Cold” and then press “Medium” and then “Hot”.

Ask students:

- **What do you notice about the molecules as energy is added?**

As energy is added, more molecules are moving faster. There are more pink and red molecules but there are still some slower-moving blue ones..

To remove energy, start with “Hot” and then press “Medium” and then “Cold”.

Ask students:

- **What do you notice about the molecules as energy is removed?**

As energy is removed, more molecules are moving slower. There are more purple and blue molecules, but a few still change to pink.

EXTEND

9. **Have students try one or more extensions and use conduction to explain these common phenomena.**

Compare the actual temperature and how the temperature feels for different objects in the room.

Ask students:

- **Touch the metal part of your chair or desk leg and then touch the cover of a text-book. Do these surfaces feel like they are the same or a different temperature?**
They should feel different.
- **Why does the metal feel colder even though it is the same temperature as the cardboard?**

Tell students that even though the metal feels colder, the metal and the cardboard are actually the same temperature. If students don't believe this, they can use a thermometer to take the temperature of metal and cardboard in the room. After being in the same room with the same air temperature, both surfaces should be at the same temperature.

Show the animation *Conducting Energy* to help answer the question about why metal feels colder than cardboard.

www.middleschoolchemistry.com/multimedia/chapter2/lesson1#conducting_energy

Tell students to watch the motion of the molecules in the metal, cardboard, and in the finger.

Explain that the molecules in your finger are moving faster than the molecules in the room-temperature metal. Therefore the energy from your finger is transferred to the metal. Because metal is a good conductor, the energy is transferred away from the surface through the metal. The molecules in your skin slow down as your finger continues to lose energy to the metal, so your finger feels cooler.

Like the metal, the molecules in your finger are moving faster than the molecules in the room-temperature cardboard. Energy is transferred from your finger to the surface of the cardboard. But because cardboard is a poor conductor, the energy is not easily transferred away from the surface through the cardboard. The molecules in your skin move at about the same speed. Because your finger does not lose much energy to the cardboard, your finger stays warm.

Compare the actual temperature and how the temperature feels for water and air.

Have students use two thermometers to compare the temperature of room-temperature water and the temperature of the air. They should be about the same.

Ask students:

- **Put your finger in room-temperature water and another finger in the air. Do the water and the air feel like they are the same or a different temperature?**
The finger in the water should feel colder.
- **Why does the water feel cooler even though it is the same temperature as the air?**

Remind students that even though the water feels colder, the water and the air are actually about the same temperature. Students should realize that water is better than air at conducting energy. As energy is drawn more rapidly away from your finger, your skin feels colder.

Consider why cups of cold and hot water both come to room-temperature.

Have students think about and explain the following situation:

- **Let's say that you put a cup of cold water in one room and a cup of hot water in another room. Both rooms are at the same room-temperature. Why does the cold water get warmer and the hot water get cooler?**

In both cases, energy will move from an area of higher temperature to an area of lower temperature. So, the energy from room-temperature air will move into the cold water, which warms the water. And the energy from the hot water will move into the cooler air, which cools the water.

Activity Sheet
Chapter 2, Lesson 1
Heat, Temperature, and Conduction

Name _____

Date _____

In this activity, you will place a room-temperature set of washers in hot water and then place a set of hot washers in room-temperature water. Find out what happens to the temperature of each.

ACTIVITY

Question to investigate

Why does the temperature of an object change when it is placed in hot water?



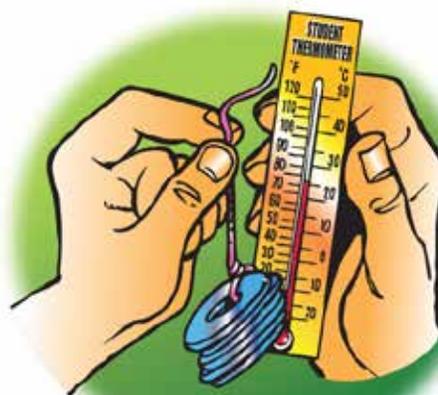
Materials for each group

- 2 sets of large metal washers on a string
- Styrofoam cup filled with hot water
- Room-temperature water
- 2 thermometers
- Graduated cylinder or beaker

Procedure

Room-temperature washers placed in hot water

1. Place a thermometer in your cup to measure the initial temperature of the water. Record the temperature of the water in the “Before” column in the chart on the activity sheet. Be sure to also record the initial temperature of the water in the control cup.
2. Use another thermometer to measure the temperature of the washers. Record this in the “Before” column.
3. With the thermometer still in the water, hold the string and lower the metal washers all the way into the water.
4. Observe any change in the temperature of the water. Leave the washers in the water until the temperature stops changing. Record the temperature of the water in each cup in the “After” column.
5. Remove the washers from the water. Then take and record the temperature of the washers in the “After” column.
6. Empty the cup in a waste container or sink.

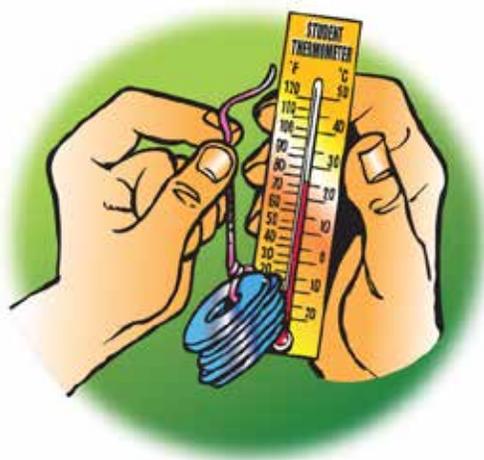


Room-temperature washers placed in hot water		
Temperature of...	Before	After
Water in your cup		
Water in the control cup		
Metal washers		

1. Why do you think the temperature of the water in your cup changes more than the water in the control cup?

Hot washers placed in room-temperature water

1. Pour about 30 milliliters of room-temperature water into your Styrofoam cup.
2. Place a thermometer into the water and record the temperature of the water in each cup in the “Before” column in the chart below.
3. Get a set of hot washers from your teacher and quickly use a thermometer to measure the temperature of the washers. Record this in the “Before” column.
4. With the thermometer still in the water, hold the string and lower the hot metal washers all the way into the water.
5. Observe any change in the temperature of the water. Leave the washers in the water until the temperature stops changing. Record the temperature of the water in the “After” column in the chart. Also record the temperature of the water in the control cup.
6. Remove the washers from the water. Then take and record the temperature of the washers.



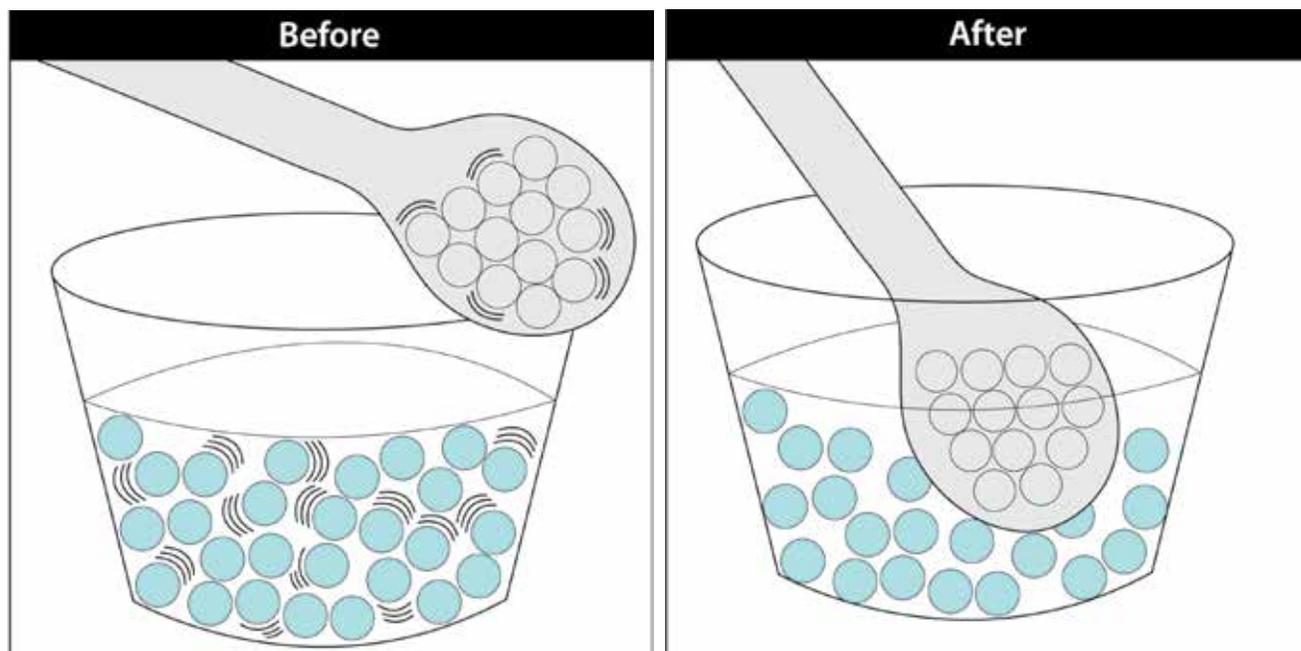
Hot washers placed in room-temperature water		
Temperature of...	Before	After
Water in your cup		
Water in the control cup		
Metal Washers		

EXPLAIN IT WITH ATOMS & MOLECULES

Room-temperature spoon placed in hot water

In the first part of the animation, you saw what happens when a spoon is placed in hot water.

2. Explain, on the molecular level, how energy was transferred from the hot water to the room-temperature spoon.
3. Draw motion lines near the atoms and molecules in the “After” illustration to show how the speed of the molecules in the spoon and water changed.



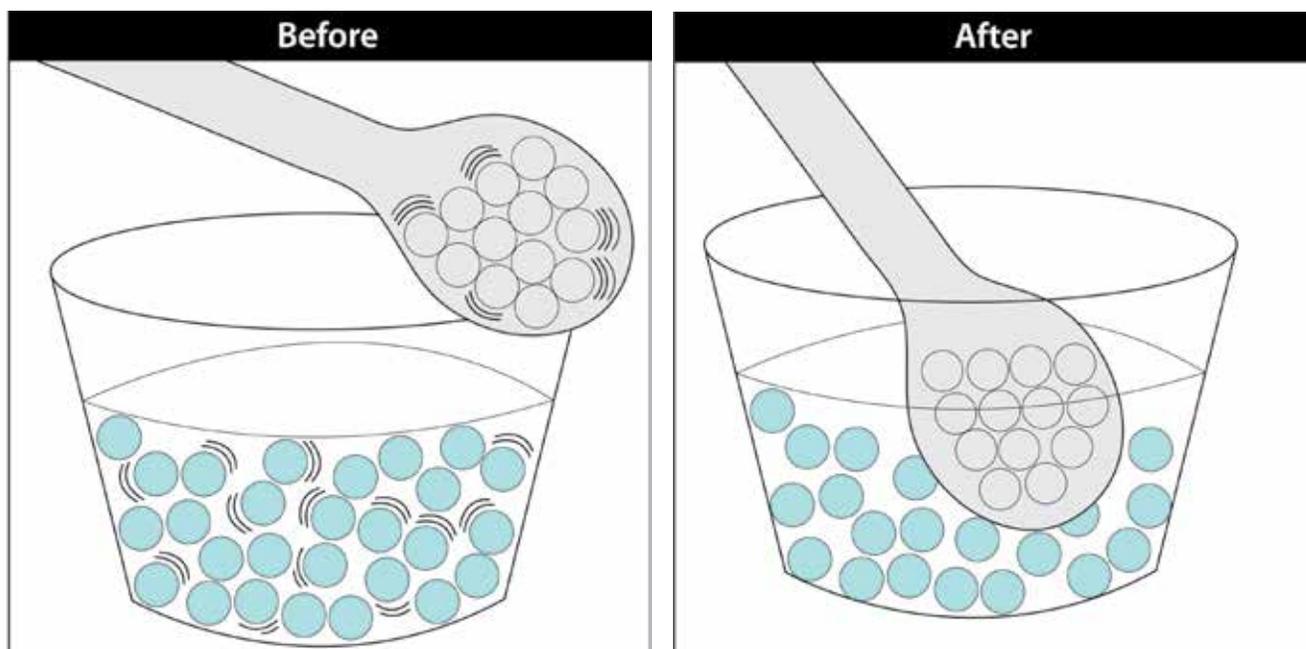
4. Now that you know what happens when a spoon is placed in hot water, explain what happened in the activity:

- Why did the metal washers get warmer?
- Why did the water get cooler?

Hot spoon placed in room-temperature water

In the next part of the animation, you saw what happens when a hot spoon is placed in room-temperature water.

5. Explain, on the molecular level, how the heat was conducted from the hot spoon to the room-temperature water.
6. Draw motion lines near the atoms and molecules in the “After” illustration to show how the speed of the atoms in the spoon and molecules in the water changed.



7. Now that you know what happens when a hot spoon is placed in room-temperature water, explain what happened in the activity:
- Why did the hot metal washers get cooler?
 - Why did the water get warmer?
8. You saw an animation that showed that temperature is a measure of the average kinetic energy of the atoms or molecules of a substance. Does this mean that all of the molecules in a cup of water are moving at the same speed or at a variety of speeds? Explain.

TAKE IT FURTHER

9. Touch your metal chair or desk leg and then touch your wooden or plastic desk top or some other wood or plastic.
- Which feels colder, the metal or the wood/plastic?
 - Explain why the metal feels colder even though it is the same temperature as the wood or plastic.
Hint: Certain materials are better at conducting heat than others.

10. Even though room-temperature water and room-temperature air are about the same temperature, the water feels colder when you put your finger in it. Use what you know about conduction to explain why the water feels colder than the air.

Hint: Certain materials are better at conducting heat than others.

11. Let's say that you put a cup of cold water in one room and a cup of hot water in another room. Both rooms are room-temperature. Why does the cold water get warmer and the hot water get cooler?

Additional Teacher Background

Chapter 2 Lesson 1, p.65

Specific Heat and Heat Capacity

When room temperature metal washers are placed in hot water, the temperature of the washers goes up and the temperature of the water goes down. This makes sense because energy was transferred from the hot water to the cooler washers. But the amount of temperature decrease of the water may not match the amount of temperature increase by the washers. Even though the same amount of energy left the water as went into the washers, the change in temperature of the two substances is different. This is because the water and the washers have a different specific heat.

Specific heat is the amount of energy required to raise the temperature of 1 gram of a substance by 1 °C. It makes sense that different substances have different specific heats because the size, mass, attractions, and arrangements of their atoms or molecules are different. Based on these differences, the amount of energy required to increase the motion of these atoms or molecules by a certain amount is different.

Chapter 2, Lesson 2: Changing State—Evaporation

Key Concepts

- Evaporation occurs when molecules in a liquid gain enough energy that they overcome attractions from other molecules and break away to become a gas.
- Adding energy increases the rate of evaporation.
- To conduct a valid experiment, variables need to be identified and controlled.

Summary

Students will help design an experiment to see if adding energy (heating) affects the rate of evaporation. Students will look at molecular animations to help explain why the heating water increases the rate of evaporation. Students will be introduced to a more detailed model of the water molecule. Students will create 3-D Styrofoam models of water molecules.

Objective

Students will be able to identify and control variables to design a test to see if heating water affects the rate of evaporation. Students will be able to explain, on the molecular level, why adding energy increases the rate of evaporation.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles. Use caution when handling hot water.

Materials for Each Group

- 2 quart-size zip-closing plastic storage bags
- Hot water
- Room-temperature water
- 2 squares of brown paper towel
- 2 droppers

Materials for each student

- 2 Styrofoam balls (1½-inch)
- 4 Styrofoam balls (1-inch)
- 2 flat toothpicks
- School glue
- Permanent marker

ENGAGE

1. Predict what might happen to a wet paper towel by the end of the class.

Show students two pieces of brown paper towel. Dampen one with water so that the color appears darker than the dry piece of paper towel. Select a student to feel the difference between the two paper towels now, and again at the end of the class period. Place both paper towels up in a prominent location.

Ask students:

- **At the end of class, do you think the paper towel will still be wet or will it be dry?**
Students should agree that the wet paper towel will likely become dry. They may say that the water will evaporate. Explain to students that when water evaporates, it changes from a liquid to a gas. Point out that the word “evaporate” has the word “vapor” in it—water changes to water vapor but it is still water.
- **What are some other examples of evaporation?**
Students may think of common examples of evaporation such as clothes in a dryer, wet hair drying on its own, or a puddle drying up in the sun.
- **When water evaporates, where do the water molecules go?**
Make it clear that, although you can't see the water anymore after it has dried up or evaporated, it still exists. The water molecules separate and are in the air as a gas called water vapor.

Tell students that they are going to find out what happens to water molecules as they evaporate by exploring how to make water evaporate faster.

2. Help students design an experiment to find out whether adding energy increases the rate of evaporation.

Tell students that they will test the evaporation of just 1 drop of water on a brown paper towel so that they can see results quickly.

Ask students:

- **What could you do to make a small amount of water evaporate faster from a paper towel?**
Students will know that they should somehow heat the water on the paper towel.
- **Will you need to put a drop of water on just one paper towel or on two?**
As you listen to students, help them realize that they will need to wet two paper towel samples but that only one will be heated. The unheated paper towel is the “control.” If they wet two paper towels and heat one of them, they will be able to see whether adding energy affects the rate of evaporation.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

EXPLORE

3. Have students conduct an experiment to see if adding energy increases the rate of evaporation.



Question to investigate

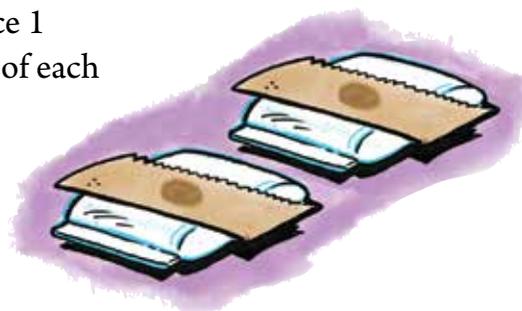
Does adding energy increase the rate of evaporation?

Materials for each group

- 2 quart-size zip-closing plastic storage bags
- Hot water (about 50 °C)
- Room-temperature water
- 2 squares of brown paper towel
- 2 droppers

Procedure

1. Add room-temperature water to a zip-closing plastic bag until it is about $\frac{1}{4}$ -filled. Get as much air out as possible, and seal the bag securely. Lay the bag down flat.
2. Add hot tap water to a different zip-closing plastic bag until it is about $\frac{1}{4}$ -filled. Get as much air out as possible, and seal the bag securely. Lay the bag down flat. This bag will serve as an energy source. The bag with the room-temperature water will serve as the control.
3. Place 2 pieces of paper towel on your table. You and your partner should each use a dropper to place 1 drop of room-temperature water in the center of each piece of paper towel at the same time.
4. Allow the drops to spread for about 5–10 seconds until they don't seem to be spreading any more.
5. At the same time, place one paper towel on each bag.
6. Observe every few minutes. Compare the amount of water on each paper towel.



Expected results

The water mark on the brown paper lying on the hot water bag should disappear faster than the mark on the paper lying on the room-temperature water bag. This will take about 3–5 minutes.

4. While waiting for evaporation, discuss the design of this experiment.

While students are waiting to see which drop of water evaporates faster, ask students about the design of the experiment.

Ask students:

- **How did we control variables?**
- **Why did we use the same type of paper towel for each sample?**
- **Why did we put the same amount of water on each piece of paper towel?**
- **Both drops of water on the paper towels were originally the same temperature. Was this a good idea?**
- **Why did we put the drops on the paper towel at the same time and in the same area?**

The type of paper towel material, amount of water, initial temperature of the water, and where the water is placed on the paper towel may all have an effect on the rate of evaporation. All these different factors are variables in the experiment. All these variables need to be kept the same so that the experiment is as fair as possible.

- **Why did we place one paper towel on a bag filled with room-temperature water?**
Even the surface each paper towel is placed on should be the same. This is why one paper towel is placed on a room-temperature bag instead of on a room-temperature table or desk. The only difference should be the amount of energy the paper towels are exposed to.

Be sure students understand the purpose of the control. The control is necessary because if there was only one sample that was heated, there would be nothing to compare it with. There would be no way of knowing whether adding energy made any difference in the rate of evaporation if there wasn't another sample to compare it to that was not heated.

5. Discuss student observations.

Ask students:

- **Does adding energy increase the rate of evaporation? How do you know?**
Yes. We can say that heating water increases the rate of evaporation because the drop of water that was heated evaporated first. Since the experiment controlled variables, heating water must increase the rate of evaporation.

- **Knowing what you do about energy and molecular motion, why do you think the water that was heated evaporated faster?**

Students should remember that adding energy increases the motion of molecules. They should realize that the water molecules on the paper towel on the warm bag are moving faster than the ones on the room-temperature bag. Students should conclude that more of these faster-moving molecules break away from the other molecules and go into the air.

EXPLAIN

6. **Show an animation to explain why adding energy increases the rate of evaporation.**



Show the animation *Evaporation*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson2#evaporation

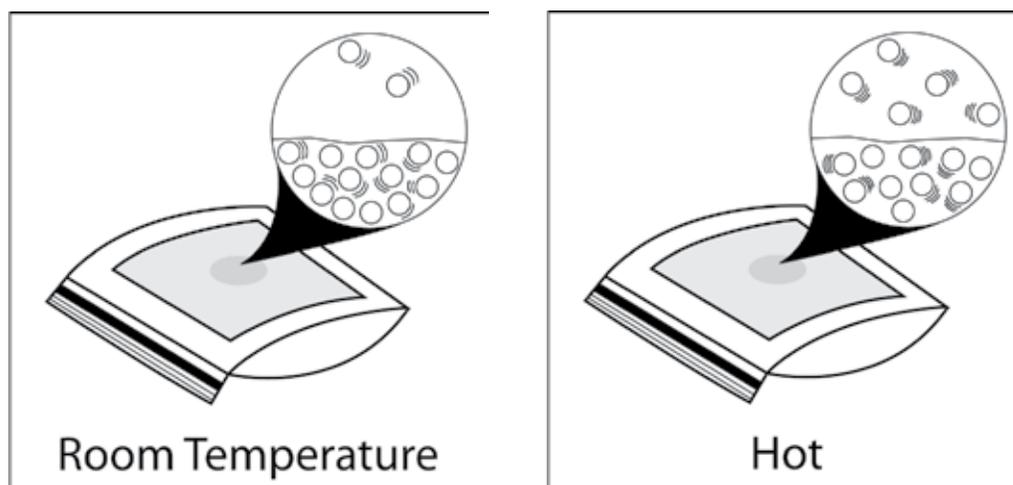
Tell students that adding energy to the water on the paper towel increases the motion of the water molecules. When the molecules have enough energy, they can move fast enough to break away from the attractions holding them to other molecules.

7. **Have students describe their observations on the molecular level.**

Project the image *Heating and Evaporation* from the activity sheet.

www.middleschoolchemistry.com/multimedia/chapter2/lesson2#heating_and_evaporation

Point out the difference in the number of motion lines in the water on each paper towel. Explain that the heated water molecules have more energy and move faster than the room-temperature water. These faster moving molecules are able to overcome the attractions they have for other water molecules and evaporate.



Have students include words or phrases with these pictures to indicate why heating the water on the paper towel increases the rate of evaporation.

8. Look at the paper towels from the start of the lesson.

Have the student who felt the two pieces of brown paper towel at the beginning feel them again. This student should report that the moist paper towel is drier or is completely dry.

Ask students:

- **The wet paper towel was not heated. Why did the water evaporate?**

Remind students of the model of average kinetic energy they saw in the last lesson. Explain to students that at room temperature, water molecules are moving at a variety of different speeds but most are moving fast enough to evaporate. As the molecules transfer energy between each other, even slower molecules will gain enough energy to evaporate.

9. Show a different model of a water molecule and review changes in state using this model.

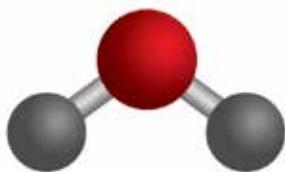
Tell students that they have been using a very simple model of water as just a circle or sphere but there are other models of water that show more detail about the structure of the molecule.

Show the animation of *Models of Water Molecules*.

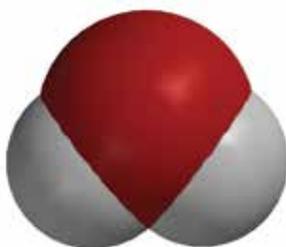
www.middleschoolchemistry.com/chapter2/lesson2#models_of_water_molecules

Show students that water is made up of 1 oxygen atom (red) and 2 hydrogen atoms (gray). Point out the ball and stick model and the space-filling model.

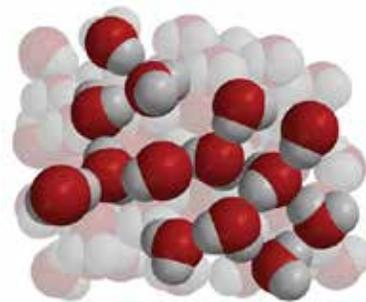
The ball-and-stick model is used to highlight the angles at which the atoms are bonded together within a molecule. The space-filling model is used to highlight the space taken up by the electron cloud around the atoms within a molecule.



Ball-and-stick model



Space-filling model



Liquid water

The shape of the water molecule and its attraction to other water molecules give water its characteristic properties.

Project animation Liquid water

Explain that water molecules, as a liquid, are very close together because of their attractions for one another but are able to slide past each other.

Note: You can mention to students that when water molecules attract each other, the oxygen part of one water molecule attracts the hydrogen part of another. The reason for this will be explored in detail in Chapter 5.

Project animation Water vapor

Explain that water molecules, as a gas, are much further apart and usually just bounce off each other when they collide. Be sure to point out that when the water evaporated, the molecules themselves did not break apart into atoms. The molecules separated from other molecules but stayed intact as a molecule.

EXTEND

10. Have students make their own space-filling models of water molecules using Styrofoam balls.

Have each student make 2 water molecules.

Question to investigate

How do water molecules move as water freezes, melts, evaporates, and condenses?

Materials note:

Styrofoam balls are available from craft stores and many science suppliers. You will need 1-inch and 1½-inch balls. These are available from Flinn Scientific, Product #AP2279 and AP2280. Each student will need 2 large and 4 small Styrofoam balls to make 2 water molecules each.

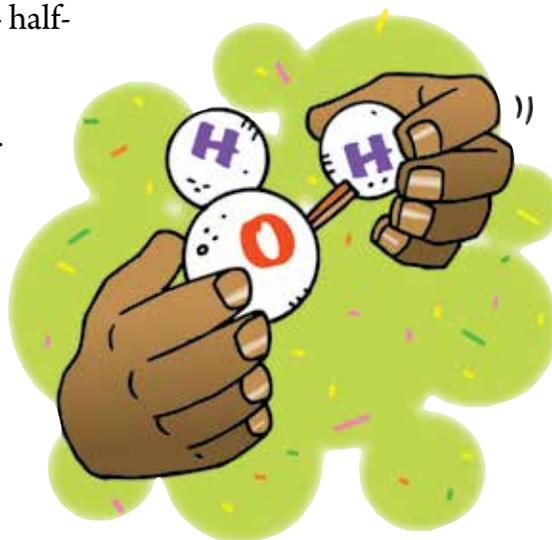
Point out that the large Styrofoam ball represents the oxygen atom and that the smaller Styrofoam balls represent the hydrogen atoms. Explain that the vast majority of each ball represents the electron cloud around the atom. Although it cannot be seen in the Styrofoam ball model, the center of each ball represents the extremely tiny nucleus where the protons and neutrons are. Almost the entire ball, except for the extremely tiny center, represents the area where the electrons are.

Materials for each student

- 2 Styrofoam balls (1½-inch)
- 4 Styrofoam balls (1-inch)
- 2 flat toothpicks
- School glue
- Permanent marker

Procedure

1. Break toothpicks in half so that there are 4 half-toothpicks.
2. Use a permanent marker to write an O on each of the large balls and an H on each of the small balls.
3. Push a half-toothpick about halfway into each small ball.
4. Push two small balls onto each larger ball at the angle shown.
5. Add 1 or 2 drops of glue where the hydrogen atoms meet the oxygen atoms. Allow the glue to dry over night.
6. Have students contribute their two water molecules to the group.



Activity Sheet
Chapter 2, Lesson 2
Changing State—Evaporation

Name _____

Date _____

ACTIVITY

Question to investigate

Does adding energy increase the rate of evaporation?

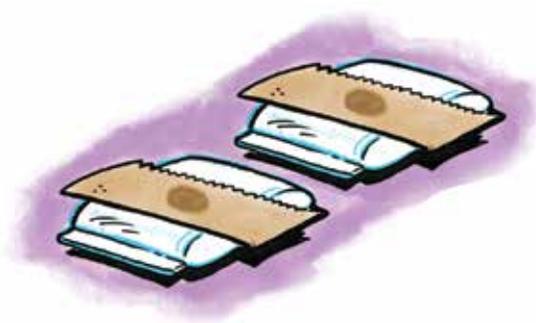


Materials for each group

- 2 quart-size zip-closing plastic storage bags
- Hot water (about 50 °C)
- Room-temperature water
- 2 squares of brown paper towel
- 2 droppers

Procedure

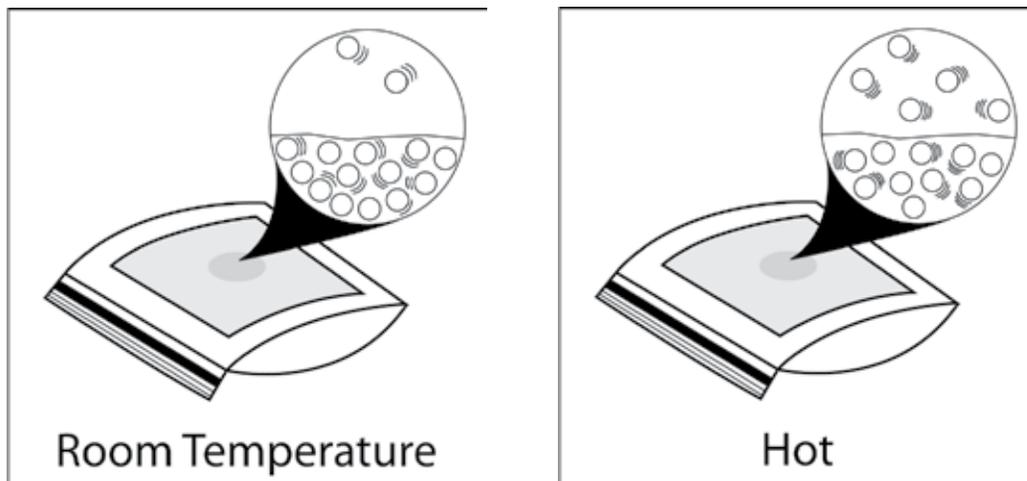
1. Add room-temperature water to a zip-closing plastic bag until it is about $\frac{1}{4}$ -filled. Get as much air out as possible, and seal the bag securely. Lay the bag down flat.
2. Add hot tap water to a different zip-closing plastic bag until it is about $\frac{1}{4}$ -filled. Get as much air out as possible, and seal the bag securely. Lay the bag down flat. This bag will serve as a heat source. The bag with the room-temperature water will serve as the control.
3. Place 2 pieces of paper towel on your table. You and your partner should each use a dropper to place 1 drop of room-temperature water in the center of each piece of paper towel at the same time.
4. Allow the drops to spread for about 10 seconds until they don't seem to be spreading any more.
5. At the same time, place one paper towel on each bag.
6. Observe every few minutes. Compare the amount of water on each paper towel.



1. One of the variables in the experiment was the amount of water placed on the brown paper towels. Why was it important to use the same amount of water on both pieces of paper towel?
2. Another variable was when the paper towels were placed on the plastic bags. Why was it important to put each paper towel on the plastic bag at the same time?
3. Does adding energy increase the rate of evaporation? What evidence do you have from the experiment to support your answer?

EXPLAIN IT WITH ATOMS & MOLECULES

You saw an animated model of your experiment showing water molecules evaporating from the paper towels.



ACTIVITY

Question to investigate

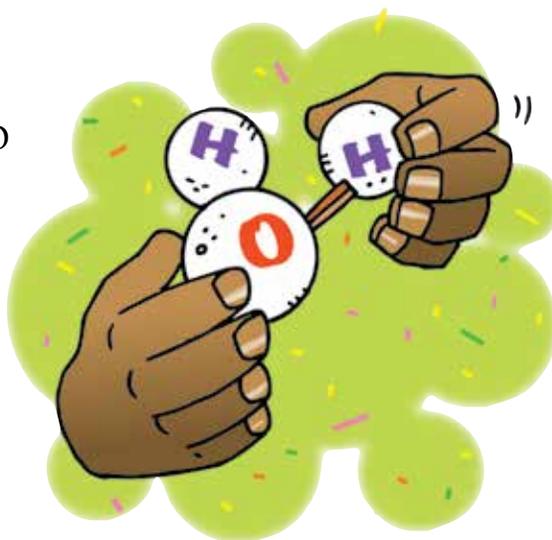
How do water molecules move as water freezes, melts, evaporates, and condenses?

Materials for each student

- 2 Styrofoam balls (1½-inch)
- 4 Styrofoam balls (1-inch)
- 2 flat toothpicks
- School glue
- Permanent marker

Procedure

1. Break toothpicks in half so that there are 4 half-toothpicks.
2. Use a permanent marker to write an O on each of the large balls and an H on each of the small balls.
3. Push a half-toothpick about half-way into each small ball.
4. Push two small balls onto each larger ball at the angle shown.
5. Add 1 or 2 drops of glue where the hydrogen atoms meet the Oxygen atom. Allow the glue to dry over night.



Have students contribute their two water molecules to the group.

Chapter 2, Lesson 3: Changing State—Condensation

Key Concepts

- Condensation is the process in which molecules of a gas slow down, come together, and form a liquid.
- When gas molecules transfer their energy to something cooler, they slow down and their attractions cause them to bond to become a liquid.
- Making water vapor colder increases the rate of condensation.
- Increasing the concentration of water vapor in the air increases the rate of condensation.

Summary

Students investigate the condensation of water vapor on the inside of a plastic cup. Then they design an experiment to see if cooling water vapor even more affects the rate of condensation. Students also relate evaporation and condensation to the water cycle.

Objective

Students will be able to describe on the molecular level how cooling water vapor causes condensation. Students will also describe the roles evaporation and condensation play in the water cycle.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- 1 short wide-rimmed clear plastic cup
- 1 tall smaller-rimmed clear plastic cup
- Hot water (about 50 °C)
- Magnifier

Materials for the Demonstration

- 2 clear plastic cups
- Room-temperature water
- Ice cubes
- Gallon-size zip-closing plastic bag

About this Lesson

Try the demonstration before presenting it to your students because it will not work if the humidity is too low. You could instead show students the video *Condensation on a Cold Cup* at www.middleschoolchemistry.com/chapter2/lesson3#condensation_cup.

The activity for the students will work no matter how dry or humid the air.

ENGAGE

1. Prepare for the demonstration about 5–10 minutes before class.

Materials for the demonstration

- 2 clear plastic cups
- Room-temperature water
- Ice cubes
- Gallon-size zip-closing plastic bag

Procedure

1. Place water and ice cubes into two identical plastic cups.
2. Immediately place one of the cups in a zip-closing plastic bag and get as much air out of the bag as possible. Close the bag securely.
3. Allow the cups to sit undisturbed for about 5–10 minutes.



Expected results

The cup inside the bag should have very little moisture on it because not much water vapor from the air was able to contact it. The cup exposed to air should have more moisture on the outside because it was exposed to the water vapor in the air, which condensed on the outside of the cup.

2. Show students the two cold cups of water and ask why water appears on the outside of only one of them.

Show students the two cups you prepared and ask:

- **Which cup has the most moisture on the outside of it?**
Students should realize that the cup exposed to more air has the most moisture on the outside of it.
- **Why do you think the cup that is exposed to more air has more water on the outside of it?**
Make sure students understand that this moisture came from water vapor in the air that condensed on the outside of the cup. Remind students that water vapor is one of the gases that makes up air. The cup in the bag has very little to no moisture on it because it is exposed to much less air. Less air means less water vapor.
- **Some people think that the moisture that appears on the outside of a cold cup is water that has leaked through the cup. How does this demonstration prove that this idea is not true?**

Because there is little to no moisture on the outside of the cup in the bag, students should conclude that water could not have leaked through the cup. If the moisture came from leaking, there would be water on the outside of both cups.

3. Introduce the process of condensation.

If students do not know what the process of condensation is, you can tell them it is the opposite of evaporation. In evaporation, a liquid (like water) changes state to become a gas (water vapor). In condensation, a gas (like water vapor) changes state to become a liquid (water).

Explain that as water molecules in the air cool and slow down, their attractions overcome their speed and they join together, forming liquid water. This is the process of *condensation*.

Ask students:

- **What are some examples of condensation?**

Coming up with examples of condensation is a bit harder than examples of evaporation. One common example is water that forms on the outside of a cold cup or the moisture that forms on car windows during a cool night. Other examples of condensation are dew, fog, clouds, and the fog you see when you breathe out on a cold day.

- **You may have made a cold window “cloudy” by breathing on it and then drawn on the window with your finger. Where do you think that cloudiness comes from?**

Help students realize that the moisture on the window, and all of the examples of condensation they gave, comes from water vapor in the air.

- **A real cloud is made up of tiny droplets of water. Where do you think they come from?**

The water in a cloud comes from water vapor in the air that has condensed.

Give each student an activity sheet.

Have students answer questions about the demonstration on the activity sheet. They will also record their observations and answer questions about the activity. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

4. Have students collect a sample of water vapor and observe the process of condensation.



Question to investigate

What happens when water vapor condenses?

Materials for each group

- 1 short wide-rimmed clear plastic cup
- 1 tall smaller-rimmed clear plastic cup
- Hot water (about 50°C)
- Magnifier

Procedure

1. Fill a wide clear plastic cup about $\frac{2}{3}$ full of hot tap water. Place the tall cup upside down inside the rim of the bottom cup as shown.
2. Watch the cups for 1–2 minutes.
3. Use a magnifier to look at the sides and top of the top cup.
4. Take the top cup off and feel the inside surface.



Expected results

The top cup will become cloudy-looking as tiny drops of liquid water collect on the inside surface of the cup.

5. Discuss with students what they think is happening inside the cups.

Ask students:

- **What do you think is on the inside of the top cup?**
Students should agree that the inside of the top cup is coated with tiny drops of liquid water.
- **How do you think the drops of water on the inside of the top cup got there?**
Students should realize that some of the water in the cup evaporated, filling the inside of the top cup with invisible water vapor. Some of this water vapor condensed into tiny drops of liquid water when it condensed on the inside of the top cup.

Explain that water vapor leaves the hot water and fills the space above, contacting the inside surface of the top cup. Energy is transferred from the water vapor to the cup, which cools the water vapor. When the water vapor cools enough, the attractions between the molecules bring them together. This causes the water vapor to change state and become tiny drops of liquid water. The process of changing from a gas to a liquid is called *condensation*.

EXPLAIN

6. Show an animation to help students understand what happens when gases condense to their liquid state.

Show the animation *Condensation*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson3#condensation

Explain that the fast-moving molecules of water vapor transfer their energy to the side of the cup, which is cooler. This causes the water vapor molecules to slow down. When they slow down enough, their attractions overcome their speed and they stay together as liquid water on the inside surface of the cup.



7. Discuss how to design an experiment to find out whether increased cooling of the water vapor affects the rate of condensation.

The goal of this discussion is to help students better understand the experimental design outlined in the procedure.

Ask students:

How could we set up an experiment to see if making water vapor even colder affects the rate of condensation?

- **How can we get the water vapor we need for this experiment?**
Students may suggest collecting water vapor as in the previous activity or collecting it over a pot of boiling water or some other way.
- **Will we need more than one sample of water vapor? Should we cool one sample of water vapor, but not the other?**
Help students understand that they will need 2 samples of water vapor, only one of which is cooled.
- **How will we cool the water vapor?**
Students may have many ideas for cooling water vapor, like placing a sample in a refrigerator or cooler filled with ice, or placing a sample of water vapor outside if the weather is cool enough.
- **How will you know which sample of water vapor condensed faster?**
By comparing the size of the drops of water formed in both samples, students can determine whether cooling water vapor increases the rate of condensation.

8. Have students do an activity to find out whether cooling water vapor increases the rate of condensation.



Question to investigate

Does making water vapor colder increase the rate of condensation?

Materials for each group

- 2 short wide-rimmed clear plastic cups
- 2 tall smaller-rimmed clear plastic cups
- Hot water (about 50 °C)
- Magnifier
- Ice

Read more about evaporation and condensation in the additional teacher background section at the end of the lesson.

Procedure

1. Fill two wide clear plastic cups about $\frac{2}{3}$ full of hot tap water.
2. Quickly place the taller cups upside down inside the rim of each cup of water, as shown.
3. Place a piece of ice on top of one of the cups.
4. Wait 2–3 minutes.
5. Remove the ice and use a paper towel to dry the top of the cup where the ice may have melted a bit.
6. Use a magnifier to examine the tops of the two upper cups.



Expected results

There will be bigger drops of water on the inside of the top cup below the ice.

9. While waiting for results, have students predict whether increased cooling will increase the rate of condensation.

Ask students to make a prediction:

- **What effect do you think adding the ice cube will have on the rate of condensation?**
- **Explain on the molecular level, why you think extra cooling might affect the rate of condensation.**

10. Discuss students' observations and draw conclusions.

Ask students:

- **Which top cup appears to have more water on it?**
The cup with the ice.

- **Why do you think the cup with the ice has bigger drops of water on the inside than the cup without ice?**

When the water vapor is cooled by the ice, the water molecules slow down more than in the cup without the ice. This allows their attractions to bring more molecules together to become liquid water.

- **Does cooling water vapor increase the rate of condensation? Yes.**

What evidence do you have from the activity to support your answer?

Students should realize that the bigger drops of water on the top cup with the ice indicate a greater amount of condensation. Because the water vapor in both sets of cups was condensing for the same length of time, the water vapor in the cup with the bigger drops must have condensed at a faster rate.

11. Explain examples of condensation on the molecular level.

Ask students:

- *Fogging up a cold window*

When you breathe out, there is water vapor in your breath. When you breathe on a cold window in the winter, the window gets tiny droplets of moisture on it or “fogs up.” What happens to the molecules of water vapor as they get near the cold window?

The water molecules in your breath are the gas water vapor. They slow down as they transfer some of their energy to the cold window. The attractions between the slower-moving water vapor molecules bring them together to form tiny droplets of liquid water.

- *Warm breath in cold air*

When you breathe out in the winter, you see “smoke,” which is really a fog of tiny droplets of liquid water. What happens to the molecules of water vapor from your breath when they hit the cold air?

The water vapor in your breath is warmer than the outside air. The water vapor molecules transfer energy to the colder air. This makes the water vapor molecules move more slowly. Their attractions overcome their motion and they join together or condense to form liquid water.

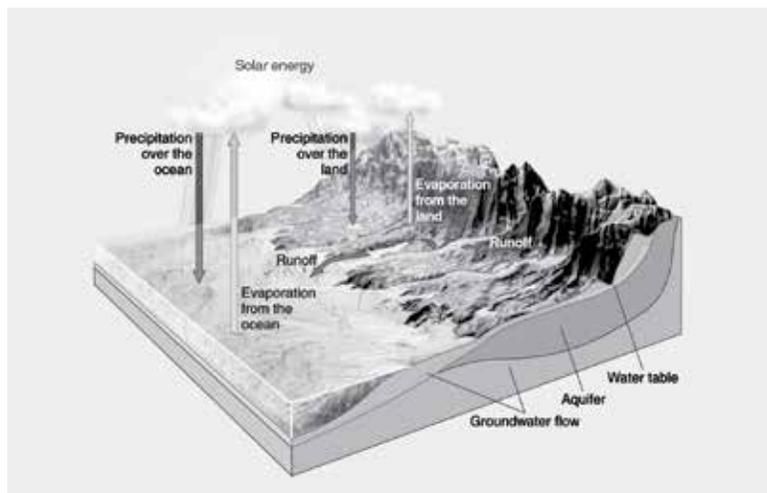
12. Explain to students that the evaporation and condensation occur naturally in the water cycle.

Project the image *Water Cycle* from the activity sheet.

www.middleschoolchemistry.com/multimedia/chapter2/lesson3#water_cycle

One common place you see the results of evaporation and condensation is in the weather. Water vapor in the air (humidity), clouds, and rain are all the result of evaporation and condensation. What happens to the water molecules during the evaporation and condensation stages of the water cycle?

Energy from the sun causes water to evaporate from the land and from bodies of water. As this water vapor moves high into the air, the surrounding air cools it, causing it to condense and form clouds. The tiny droplets of water in clouds collect on bits of dust in the air. When these drops of water become heavy enough, they fall to the ground as rain (or hail or snow). The rain flows over the land towards bodies of water, where it can evaporate again and continue the cycle.



EXTEND

13. Introduce the idea that the amount of water vapor in the air affects the rate of condensation.

Ask students if they know what a terrarium is. Tell students that a terrarium is a closed container with moss or other plants in which water continually evaporates and condenses. At first, the evaporation rate is higher than the rate of condensation. But as the concentration of water molecules increases in the container, the rate of condensation increases. Eventually, the rate of condensation equals the rate of evaporation and the water molecules go back and forth between the liquid and the gas.

Read more about evaporation and condensation equilibrium in the additional teacher background section at the end of the lesson.

Show the animation *Evaporation & Condensation*

www.middleschoolchemistry.com/multimedia/chapter2/lesson3#evaporation_condensation

Explain that the animation moves up through a sample of water to the surface. Water molecules evaporate (leave the liquid) and condense (reenter the liquid) at the same time.



The animation shows the beginning of the process where water molecules evaporate at a faster rate than they condense. Explain to students that if the process were to continue, the rate of evaporation and condensation would become equal.

So temperature isn't the only factor that affects condensation. The concentration of water molecules in the air is also an important factor. The higher the concentration of water molecules in the air (humidity), the higher the rate of condensation.

This is why clothes dry more slowly on a humid day. The high concentration of water vapor in the air causes water to condense on the clothes. So even though water is evaporating from the clothes, it is also condensing on them and slowing down the drying.

14. Have students design an activity to see why wind helps things dry more quickly.

Explain to students that when water evaporates from something like a paper towel, the area in the air immediately above the paper towel has a little extra water vapor in it from the evaporating water. Some of this water vapor condenses back onto the paper so the paper doesn't dry as quickly. If that water vapor is blown away by moving air like wind, there will be less condensation and the paper will dry more quickly.

Ask students:

- **How would you design an experiment that can test whether a paper towel dries more quickly if the air around the paper towel is moving?**

As you listen to suggestions from students, be sure that they identify and control variables. The paper should be in the same situation except for air moving over one piece but not the other. It is not a good idea to blow on one because the breath could be a different temperature than the surrounding air and also contains water vapor. These are both variables that would affect the experiment. It is better to wave one of the paper towels back and forth for a few minutes and have someone else hold the other or tape it so it hangs freely.

Materials

- 2 pieces of brown paper towel
- Water
- Dropper

Procedure

1. Place one drop of water on two pieces of brown paper towel.
2. Have your partner hold one while paper while you swing the other one through the air.
3. After about 30 seconds compare the paper towels to see if you can see any difference in how wet or dry the papers are.
4. Repeat step 3 until you notice a difference between the wet spots on the paper towel.

Expected results

The water on the paper towel with more air moving over it should dry faster than the other paper towel on the table. The paper towel on the table had air with a little more humidity over it condensing back onto the paper. This slowed down the drying process. The paper waved in the air didn't have humid air around it and condensing back on it as much so it dried more quickly.

EXTRA EXTEND

15. Use the processes of evaporation and condensation to purify water.

Evaporation and condensation can be used to purify water. Imagine what might happen if colored water evaporates and then condenses.

Question to investigate

If colored water evaporates and condenses, will there be any color in the water that is produced?

Materials for each group

- 1 short wide-rimmed clear plastic cup
- 1 tall smaller-rimmed clear plastic cup
- Hot water
- Food coloring
- Ice cube
- White napkin or paper towel



Procedure

1. Add hot tap water to a wide clear plastic cup until it is about $\frac{2}{3}$ full.
2. Add 1 drop of food coloring and stir until the water is completely colored.
3. Turn another clear plastic cup upside down on the cup of hot water as shown. Place an ice cube on the top cup to make condensation happen faster.
4. Wait 1–3 minutes for water vapor to condense to liquid water on the inside surface of the top cup.
5. Use a white paper towel to wipe the inside of the cup to check for any color.

Expected results

The water that collects on the inside of the top cup will be colorless. The color will remain in the bottom cup.

Explain that the process described in the procedure is called *distillation*. During distillation, water that has substances dissolved in it can be purified (as long as these substances don't easily evaporate). When the water evaporates and condenses, the food coloring is left behind and the pure water can be collected and used.

DEMONSTRATION

1. Your teacher showed you two cups of water with ice in them. One cup was in a bag with as much air taken out as possible. The other cup was left out in the air. After a few minutes, water was on the outside of the cup left in the air. Much less water was on the outside of the cup in the bag.

Why do you think the cup that is exposed to more air has water on the outside of it?



2. Condensation happens when water molecules in the air slow down so much that their attractions overcome their speed. This makes them join together, forming liquid water.

List two common examples of condensation.

ACTIVITY

Question to investigate

What happens when water vapor condenses?



Materials for each group

- 1 short wide-rimmed clear plastic cup
- 1 tall smaller-rimmed clear plastic cup
- Hot water (about 50 °C)
- Magnifier

Procedure

1. Fill a wide clear plastic cup about $\frac{2}{3}$ full of hot tap water. Place the tall cup upside down inside the rim of the bottom cup as shown.
2. Watch the cups for 1–2 minutes.
3. Use a magnifier to look at the sides and top of the top cup.
4. Take the top cup off and feel the inside surface.



WHAT DID YOU OBSERVE?

3. After a couple of minutes, what did you observe on the inside of the top cup?

4. How could the tiny drops of water get to the inside of the top cup? Use ideas about evaporation and condensation in your explanation.

ACTIVITY

Question to investigate

Does making water vapor colder increase the rate of condensation?



Materials for each group

- 2 short wide-rimmed clear plastic cups
- 2 tall smaller-rimmed clear plastic cups
- Hot water
- Magnifier
- Ice

Procedure

1. Fill two wide clear plastic cups about $\frac{2}{3}$ full of hot tap water.
2. Quickly place the taller cups upside down inside the rim of each cup of water, as shown.
3. Place a piece of ice on top of one of the cups.
4. Wait 2–3 minutes.
5. Remove the ice and use a paper towel to dry the top of the cup where the ice may have melted a bit.
6. Use a magnifier to examine the tops of the two upper cups.



WHAT DID YOU OBSERVE?

5. Does cooling water vapor increase the rate of condensation?

What evidence do you have from the activity to support your answer?

EXPLAIN IT WITH ATOMS & MOLECULES

6. The animation showed water molecules as a gas condensing to form liquid water on the inside of the top cup. Since the water molecules were all separated as a gas, why did they come together to form a liquid?

7. Why do you think the cup with the ice has bigger drops of water on the inside than the cup without ice?

TAKE IT FURTHER

Fogging up a cold window

8. When you breathe on a cold window in the winter, the window gets tiny droplets of moisture on it or “fogs up.” Using what you know about condensation, explain why you think the cold window gets foggy.

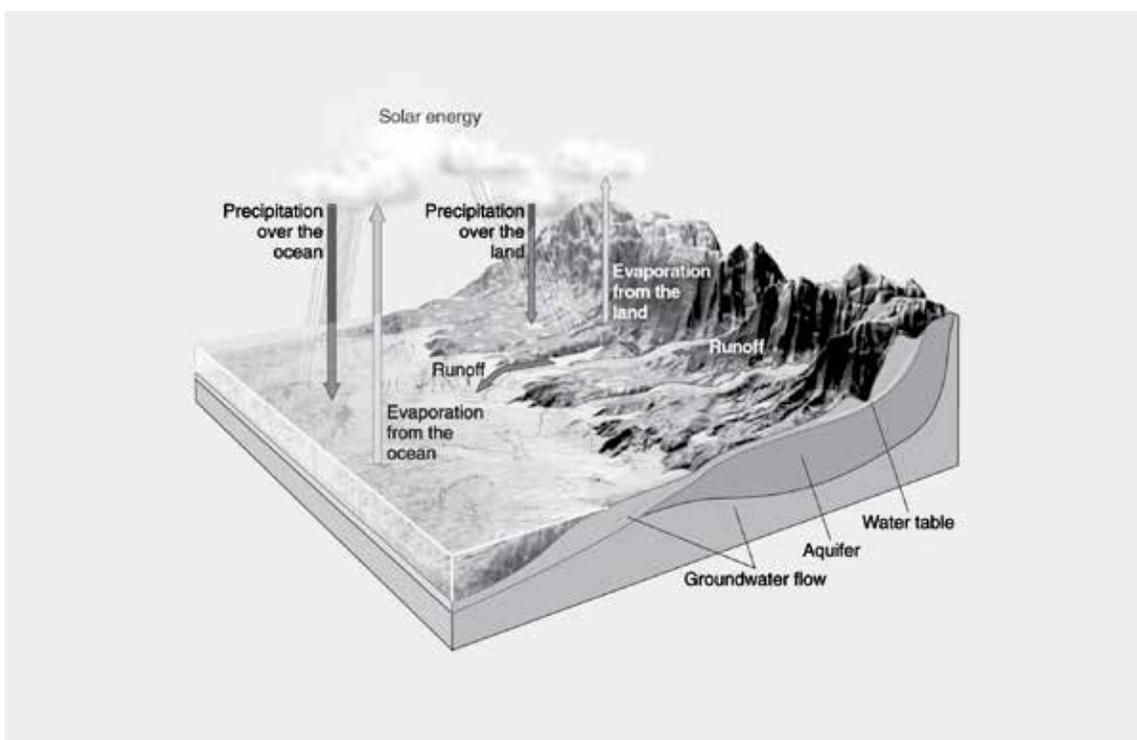
Hint: There is water vapor in your breath.

Warm breath in cold air

9. When you breathe out in the winter, you see “smoke,” which is really tiny droplets of liquid water. Using what you know about condensation, explain why you think this happens.

Evaporation and condensation in the water cycle

10. One common place you see the results of condensation is in the weather. Water vapor in the air (humidity), clouds, and rain are all the result of evaporation and condensation.



Using what you know about evaporation and condensation, explain what causes rain.

TAKE IT FURTHER

Question to investigate

Why do damp things dry more quickly on a windy day?



Materials

- 2 pieces of brown paper towel
- Water
- Dropper

Procedure

1. Place one drop of water on two pieces of brown paper towel.
2. Have your partner hold one while paper while you swing the other one through the air.
3. After about 30 seconds compare the paper towels to see if you can see any difference in how wet or dry the papers are.
4. Repeat step 3 until you notice a difference between the wet spots on the paper towel.

11. Why does moving air over a wet surface make it dry more quickly?

Hint: your answer should mention both evaporation and condensation.

TAKE IT FURTHER II

Question to investigate

If colored water evaporates and condenses, will there be any color in the water that is produced?

Materials for each group

- 1 short wide-rimmed clear plastic cup
- 1 tall smaller-rimmed clear plastic cup
- Hot water
- Food coloring
- Ice cube
- White napkin or paper towel



Procedure

1. Add hot tap water to a wide clear plastic cup until it is about $\frac{2}{3}$ full.
 2. Add 1 drop of food coloring and stir until the water is completely colored.
 3. Turn another clear plastic cup upside down on the cup of hot water as shown. Place an ice cube on the top cup to make condensation happen faster.
 4. Wait 1–3 minutes for water vapor to condense to liquid water on the inside surface of the top cup.
 5. Use a white paper towel to wipe the inside of the cup to check for any color.
12. Is there any color in the water that forms on the inside of the top cup?
13. If you were stranded on an island and only had saltwater, how could you make water to drink?

Additional Teacher Background

Chapter 2, Lesson 3, p. 99

Exploring evaporation and condensation gives middle school students an opportunity to understand some common phenomena on the molecular level. It can also give you an opportunity to review and better understand some big ideas in chemistry and physical science that are relevant to evaporation and condensation as well as other contexts in chemistry.

Big idea 1:

If two atoms or molecules, like water molecules, are attracted to each other and are “bonded”, it takes energy to pull them apart. If two atoms or molecules are attracted to each other and are not bonded, energy is released when they come together and bond. In chemistry, this concept is often stated as:

It takes energy to break bonds.

Energy is released when bonds are formed.

Note: In the context of evaporation and condensation, the use of the term “bond” refers to the interaction and close association between water molecules. It does not refer to the covalent bond which holds the oxygen atom and the hydrogen atoms together within the water molecule. The bond breaking and bond making involved in evaporation and condensation deals with the attractions and interactions between water molecules.

Big idea 2:

Another big idea is that the energy atoms and molecules have based on their motion is called kinetic energy. The energy they have based on their attraction to each other is called potential energy.

Big idea 3:

This is like a combination of Big ideas 1 and 2. When we say it takes energy to break bonds and energy is released when bonds are formed, it really means energy is converted between kinetic and potential energy. For example, it takes a certain amount of kinetic energy to separate two water molecules. When they are separated, the kinetic energy used to separate them is converted to the potential energy of attraction between them. If they come together again, this potential energy is converted back to kinetic energy. Energy is not created or destroyed; it is converted from one form to another.

These ideas can help explain why evaporation has a cooling effect and condensation has a warming effect.

Think about a single water molecule in a sample of water. Assume that the molecule has average kinetic energy. This molecule gets hit by some fast-moving water molecules and gains some extra kinetic energy. The water molecule now has above-average kinetic energy. This extra kinetic energy is “used” to break its “bonds” to other water molecules causing it to evaporate. This extra kinetic energy is converted to potential energy of attraction between the water molecule as a gas and the other water molecules in the liquid. The extra kinetic energy is not in the water anymore so the temperature of the water decreases.

Now imagine that same water molecule as a molecule of water vapor with average kinetic energy. As this molecule is attracted to other water molecules its potential energy decreases while its kinetic energy increases. The water molecule now has above-average kinetic energy. It hits other water molecules and transfers this extra kinetic energy to them which enables it to bond to other water molecules. The extra kinetic energy is now in the water so the temperature of the water increases.

Additional Teacher Background

Chapter 2, Lesson 3, p. 101

Dynamic equilibrium

It is common to see a puddle of water or wet clothes dry up from evaporation. It is also common to see water vapor condense on a cold surface to form liquid water. In both these cases, there is a net change in one direction—either from a liquid to a gas (evaporation) or from a gas to a liquid (condensation).

But under certain conditions, evaporation and condensation balance each other so there is no *net* change in either direction. The classic example is water placed in a closed container at room temperature. Even at room temperature, some fraction of the water molecules at the surface will gain enough energy from other molecules to evaporate and will enter the air inside the container. And some fraction of these water vapor molecules will lose enough energy to molecules at the surface to condense and become part of the liquid water.

At first, there are not many molecules of water vapor so the rate of condensation is slower than the rate of evaporation. But as more molecules evaporate, the concentration of molecules in the vapor increases and more molecules are available to condense to liquid water. Eventually, the air inside the container has enough water vapor molecules that the number losing energy and condensing equals the number gaining energy and evaporating. At this point, the air in the container is saturated and has a 100% relative humidity and evaporation and condensation are in equilibrium. Even though there is no *net* change, evaporation and condensation are still occurring. For this reason, the equilibrium is referred to as a *dynamic* equilibrium.

Dynamic equilibrium at different temperatures

Evaporation and condensation achieve a dynamic equilibrium at any temperature. For example, if the room temperature container in the above example is cooled, the rate of evaporation decreases. This means that the rate of condensation is greater than the rate of evaporation. But as more water vapor molecules condense, there are fewer water vapor molecules in the air and the rate of condensation slows down. Eventually, the rate of condensation and evaporation become equal at the lower temperature with fewer water molecules evaporating and condensing than at room temperature.

If the temperature of the container was now increased above room temperature, the rate of evaporation would again be greater than the rate of condensation. But as more water vapor molecules entered the air in the container, the rate of condensation would increase. Eventually, the rate of evaporation and condensation would become equal at the higher temperature with more water molecules evaporating and condensing than at room temperature.

Chapter 2, Lesson 4: Changing State—Freezing

Key Concepts

- Freezing is the process that causes a substance to change from a liquid to a solid.
- Freezing occurs when the molecules of a liquid slow down enough that their attractions cause them to arrange themselves into fixed positions as a solid.

Summary

Students will mix ice and salt in a metal can to make it very cold. They will then see liquid water and ice form on the outside of the can. Students will watch an animation of water molecules arranged as ice.

Objective

Students will be able to explain on the molecular level why a low enough temperature can cause the water vapor in air to condense to liquid water and then freeze to form ice.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- Empty clean metal soup can
- Salt
- Ice
- Metal spoon or sturdy stick
- Teaspoon
- Paper towel

Materials for the Teacher

- Pliers
- Duct tape

About this Lesson

If the level of humidity in your classroom is too low, you cannot do the activities in the Explore section of this lesson. However, you can still teach the lesson by showing students the video *Ice on a Can*. It may be helpful to show students the difference in your results:

www.middleschoolchemistry.com/chapter2/lesson4#ice_can

ENGAGE

1. Show students that liquid water expands when it freezes to become solid ice.

Teacher Preparation

- Place 50 milliliters of water into a plastic 100 ml graduated cylinder and place it in the freezer over night.
- The next day, bring it into class and show students that the level of ice is higher than the level of water you started with. Explain to students that as water freezes, it expands and takes up more space than it did as liquid water.

Show the movie Ice Bomb

www.middleschoolchemistry.com/multimedia/chapter2/lesson4#ice_bomb

This video is from the Chemistry Comes Alive! series and is used with permission from the Division of Chemical Education of the American Chemical Society.

Ask students:

- **Why do you think freezing water in the metal container caused it to burst?**
Water molecules move further apart when water freezes. This movement caused the metal container to burst.
- **Why are roads likely to develop potholes during cold winters?**
Hint: Think about what happened to the metal container.
When water gets in small cracks in the road and freezes it expands and breaks the asphalt. When this continues to happen below the surface, it eventually forms a pothole.

Ask students:

- **What do you think happens to water molecules when liquid water changes to solid ice?**
Students learned that when water vapor is cooled, attractions between water molecules cause them to condense and become liquid water. Students may say that the water molecules slow down enough that their attractions hold them together as ice.

Note: Students may say that water molecules get closer together to form ice. Water is unusual because its molecules move further apart when it freezes. The molecules of just about every other substance move closer together when they freeze. This will be covered in more detail in Chapter 3, Density.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on instructions. Look at the teacher guide to find the questions and answers.



EXPLORE

2. Have students chill a metal can so that ice forms on it.



Question to investigate

How can you make the water vapor in air condense and then freeze?

Materials for each group

- Empty clean metal soup can
- Salt
- Ice
- Metal spoon or sturdy stick
- Teaspoon
- Paper towel

Materials for the teacher

- Pliers
- Duct tape

Teacher preparation

Use pliers to bend sharp edges on the can down. Then cover the rim with 2–3 layers of duct tape to prevent possible injuries.

Procedure

1. Dry the outside of a can with a paper towel.
2. Place 3 heaping teaspoons of salt in the bottom of the can. Fill the can about halfway with ice.
3. Add another 3 heaping teaspoons of salt.
4. Add more ice until the can is almost filled and add another 3 teaspoons of salt.
5. Hold the can securely and mix the ice-salt mixture with a metal spoon or sturdy stick for about 1 minute. Remove the spoon, and observe the outside of the can. Do not touch it yet.
6. Wait 3–5 minutes. While you wait, watch the animations.



Note: After completing Step 5, you may choose to have students place a thermometer inside the can. The temperature of the salt and ice mixture will be below the normal freezing point of water, which is 0 °C.

Read more about why salt lowers the temperature of an ice water mixture.

Expected results

A thin layer of ice will appear on the outside of the can. Students may also see liquid water on the upper part of the can where it isn't as cold.

3. Discuss student observations and ask how the attractions and motion of molecules can explain the changes in state.

Ask students:

- **Look at and touch the outside of the can. What do you observe?**

A thin layer of ice covers the coldest part of the can. Some small drops of water may appear higher on the can where it is not as cold.

- **Describe what happens to water molecules as they move from being water vapor near the can to ice on the can.**

Water vapor molecules in the air near the can cooled when energy from the air transferred to the cold can. These water molecules slowed down, condensed to liquid water, and then froze to become ice.

- **Your can might have some water and some ice on the outside of it. Explain how this is possible.**

Tiny drops of water appear on the part of the can above the ice because the molecules slow down and condense to liquid water. Ice appears on the colder part of the can because the water vapor that came in contact with this part of the can was cooled so much that it froze.

Give students time to answer questions about the activity and the animations.



EXPLAIN

4. Show a molecular model animation to help students visualize what happens when water freezes.



Project the animation *Ice structure*

www.middleschoolchemistry.com/multimedia/chapter2/lesson4#ice_structure

Point out that when water freezes, the water molecules have slowed down enough that their attractions arrange them into fixed positions. Water molecules freeze in a hexagonal pattern and the molecules are further apart than they were in liquid water.

Note: The molecules in ice would be vibrating. The vibrations are not shown here but are shown on the next animation.

Project the animation *Ice different angles*

www.middleschoolchemistry.com/multimedia/chapter2/lesson4#ice_at_different_angles

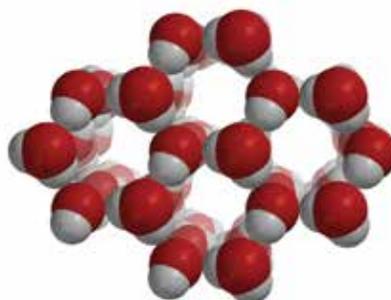
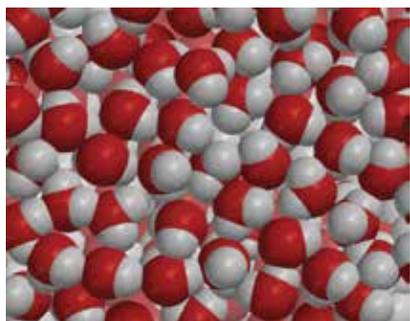
Explain that this animation shows different views of the ice crystal. Point out that even though the ice is cold the molecules still have motion. They vibrate but cannot move past one another.

5. Have students compare molecular models of liquid water and ice.

Project the image *Water and Ice*

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#ice_and_water

ZX



Ask students:

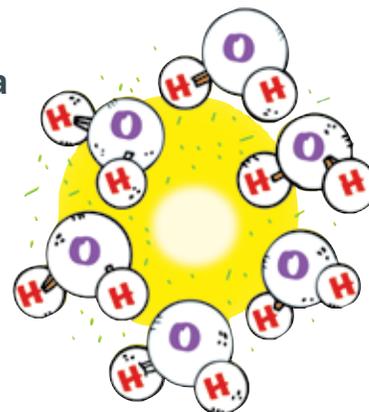
- What are some of the differences between liquid water and solid ice?

The molecules in liquid water are closer together than they are in ice. Compared to other substances, water is unusual in this way. The molecules in the liquid are moving past one another. The hydrogen end of one water molecule is attracted to the oxygen end of another but only for a short time because they are moving.

The molecules in ice are further apart than in liquid water. This is why ice floats in water. The molecules in ice are in fixed positions but still vibrate.

6. Have each group arrange their water molecules into a six-sided ring of ice.

Students do not need to orient the molecules exactly as they are in the space-filling model but they should try to have a hydrogen atom of one molecule near an oxygen atom of another. Ask students to handle their models gently because they will need them for other lessons.



7. Discuss why different liquids have different freezing points.

Tell students that the temperature at which a substance freezes is called the *freezing point*. The freezing point of water is 0 °C (32 °F). Corn oil and isopropyl alcohol have lower freezing points than water. This means that they need to be cooled to lower temperatures to make them freeze.

Water	0 °C
Corn oil	about -20 °C
Isopropyl alcohol	-88.5 °C

Ask students:

- **Why do you think different liquids have different freezing points?**

Help students realize that each liquid is made up of different molecules. The molecules of a liquid are attracted to each other by different amounts. The molecules have to slow down to different levels before their attractions can take hold and organize them into fixed positions as a solid.

8. Have students consider the freezing point of a gas.

Tell students that the air around them is made of different kinds of gases. The attractions between the molecules of gases in air (except water vapor) are so weak that they need to be cooled to very low temperatures in order to condense to a liquid or freeze to a solid.

Nitrogen gas makes up about 80% of the air. If nitrogen is made cold enough, the weak attractions between its molecules can cause it to condense to a liquid. Nitrogen condenses to a liquid at -196 °C and it freezes at -210 °C.

Show the video *Liquid Nitrogen*.

Have students observe how cold liquid nitrogen is by watching the following video.

www.middleschoolchemistry.com/multimedia/chapter2/lesson4#liquid_nitrogen

EXTEND

9. Show some pictures of frost and introduce how substances can sometimes change directly from a gas to a solid.

Tell students that under some conditions a gas can turn directly to a solid without going through the liquid phase. Explain that this process is called *deposition*. Some of the ice that formed on the outside of the can may have been a result of deposition.

Project the image *Frost*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson4#frost

Tell students that the frost that forms on the ground, windows, or grass in winter is formed by deposition.



Give students time to answer questions about freezing points, nitrogen, and deposition to complete their activity sheets for this lesson.



You could also show students images of snowflakes from www.its.caltech/~atomic/snowcrystals and the video of a snowflake forming at <http://www.its.caltech.edu/~atomic/snowcrystals/movies/movies.htm>.

Read more about how changes of state relate to the weather, in the additional teacher background section at the end of this lesson.



DEMONSTRATION

1. In the video, you saw a round metal container filled with water and placed in a very cold liquid mixed with dry ice. What happened when the water inside the container froze?

What caused this to happen?

2. Use the example of what happens to the metal container to explain why roads are likely to develop potholes during cold winters?

ACTIVITY

Question to investigate

How can you make the water vapor in air condense and then freeze?



Materials for each group

- Empty clean metal can
- Salt
- Ice
- Metal spoon or sturdy stick
- Teaspoon
- Paper towel

Procedure

1. Dry the outside of a can with a paper towel.
2. Place 3 heaping teaspoons of salt in the bottom of the can. Fill the can about halfway with ice.
3. Add another 3 heaping teaspoons of salt.
4. Add more ice until the can is almost filled and add another 3 teaspoons of salt.



5. Hold the can securely and mix the ice-salt mixture with a metal spoon or sturdy stick for about 1 minute. Remove the spoon, and observe the outside of the can. Do not touch it yet.
6. Wait 3–5 minutes. Watch the animations while you wait.

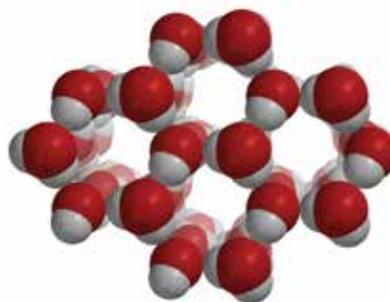
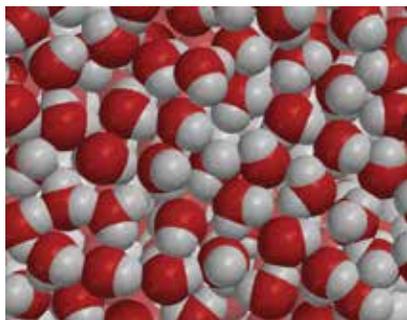


EXPLAIN IT WITH ATOMS & MOLECULES

3. Look at and touch the outside of the can. What do you observe?
4. Describe what happened to the water vapor in the air when it came in contact with the cold surface of the can. Be sure to mention how the molecules change speed and how they are attracted to each other.
5. Your can might have some water and some ice on the outside of it. Explain how this can be possible.
6. You have seen molecular model animations of water and ice. Fill out the chart to compare how the molecules move in water and ice. Select one of the options in each row and write it under “water” or “ice” in the chart.

Compare molecules in water and ice		
	Water	Ice
Speed of molecules faster slower		
Amount of movement remain in fixed positions move past each other		
Arrangement of molecules very organized random and unorganized		
Distance between molecules closer together slightly further apart		

7. Write captions under the pictures to explain how the movement and position of molecules changes as the water freezes to become ice.



8. The temperature at which a substance freezes is called the freezing point. Different liquids have different freezing points. Here are a few examples.

Water	0 °C
Corn oil	about -20 °C
Isopropyl alcohol	-88.5 °C

Why do you think different liquids have different freezing points?

9. Nitrogen is a gas at room temperature. It needs to be cooled to -196 °C to condense to a liquid and freezes at -210 °C. Do you think the attractions between nitrogen molecules are strong or weak?

Why?

TAKE IT FURTHER

10. Freezing is the process that occurs when a liquid changes to a solid. Frost forms through a process called *deposition*. What happens during the process of deposition?

Additional Teacher Background

Chapter 2 Lesson 4, p. 116

Why does salt make ice colder?

There is a principle in chemistry that comes up in several different contexts.

The simple version is that:

- It takes energy to break bonds
- Energy is released when bonds are formed.

These ideas can be used to help explain why the temperature decreases when salt is added to ice.

If an ice-and-water mixture is placed in a well-insulated container, some ice melts but some water also freezes. The breaking of “bonds” between water molecules to melt the ice uses some energy so the process of melting makes the ice/water mixture colder. But the making of “bonds” between water molecules to form ice is energy-releasing so the process of freezing makes the ice/water mixture warmer. When these two processes happen at the same rate, the ice/water mixture stays at the same temperature. But when salt is added, it dissolves into the water and forms a salt water solution. The salt water does not refreeze as fast as the rate at which the ice melts. The energy used to melt the ice is not balanced by an equal amount of energy released by freezing so the ice/saltwater solution gets colder.

This would actually work with any substance that dissolves well in cold water. Salt dissolves pretty well in cold water and is pretty cheap, so it is a popular choice.

Additional Teacher Background

Chapter 2, Lesson 4, p. 120

Relative humidity

One condition that you often hear in the weather report is relative humidity. Relative humidity is reported as a percentage, but a percentage of what? As you know, humidity refers to the amount of water vapor in the air. Relative humidity is the amount of water vapor in the air compared to, or relative to, the maximum amount the air could “hold” at that temperature.

For example, let’s say the relative humidity is 50% at a temperature of 60 °F. This means that the concentration of water vapor in the air is 50% of the maximum it could hold at that temperature. Since water vapor condenses more readily at lower temperatures, it can hold more water at higher temperatures. This means that air with a relative humidity of 50% at 80 °F would have more water vapor in it than air with a relative humidity of 50% at 60 °F.

Dew point

Another condition in the weather report is dew point. Dew point is like the opposite of relative humidity. It is the temperature that it would need to be for the amount of water vapor in the air to condense.

For example, if the air had a certain concentration of water vapor, it might condense at 40 °F. Then the dew point would be 40 degrees. But if the air contained more water vapor, it might condense at 45 degrees so this temperature would be the dew point.

Conditions for frost

When the relative humidity is low, the temperature required to make the water vapor in the air condense (dew point) is low. When a surface is at or below the dew point and the dew point is at or below the freezing point for water, frost can form on that surface.

Chapter 2, Lesson 5: Changing State—Melting

Key Concepts

- Melting is a process that causes a substance to change from a solid to a liquid.
- Melting occurs when the molecules of a solid speed up enough that the motion overcomes the attractions so that the molecules can move past each other as a liquid.

Summary

Students will see a small piece of ice melt on an aluminum surface. Based on what they have covered in Chapters 1 and 2, students will explain the energy transfer and molecular motion which cause the change in state from a solid to a liquid. Students will see and discuss an animation of ice melting and compare the state changes of water to the state changes of other substances. They will also investigate sublimation of dry ice through a teacher demonstration, or video if dry ice is not readily available.

Objective

Students will be able to explain on the molecular level the process of heat transfer and molecular motion that causes a solid to melt to form a liquid. Students will also be able to explain how the arrangement of water molecules is different from most other substances when it changes state from a solid to a liquid.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- 2 small pieces of ice
- 2 small clear plastic cups
- Water

Materials for the Demonstration

- Ice
- Dry ice
- Brown paper towel
- Cold water
- Hot water

ENGAGE

1. Have students watch a small piece of ice melting.



Show students the video *Ice Melting on Different Surfaces*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#ice_melting_on_different_surfaces

In this video, ice is placed on two similar-looking black surfaces—one aluminum and the other plastic. The ice melts faster on the aluminum because it is a better thermal conductor than the plastic.

2. Discuss student observations.

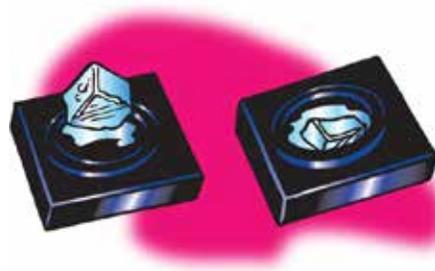
Ask students:

- **Where do you think the energy came from to melt the ice?**

The energy comes from the air and from the surface that the ice is placed on, both of which are at room temperature. Since room temperature is warmer than the temperature of the ice, energy is transferred from the surface and the air to the ice.

- **What do you think happened to the speed of the molecules in the ice when it was heated?**

The water molecules moved faster.



Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



Give students time to answer the first two questions on the activity sheet.

EXPLORE

3. Have students explore how to make ice melt faster.



Introduce the question to investigate:

- **How can you make the ice melt faster?**

Help students plan and conduct their experiment by asking:

- **How could you set up an experiment to test your method?**

Students might suggest breathing on the ice, holding it in their hand, or placing the ice in room-temperature or warm water. Any of these methods are fine, but try to have students think about including a control as part of the experiment. In each case, they would need two similar size pieces of ice—one that they warm in some way and one that they don't.

Here is one method students could try.

Question to investigate

Will placing ice in water make ice melt faster?

Materials

- 2 small pieces of ice
- 2 small clear plastic cups
- Water

Procedure

1. Add room-temperature water to a cup until it is about $\frac{1}{2}$ -full.
2. Place a small piece of ice in the water and another small piece of ice in a cup without water.



Expected results

The ice placed in the water will melt faster than the ice in air. Since the water and the air are both at room temperature, it may not be obvious why the ice melts faster in the water. There are so many more molecules in the water that can contact the ice that the transfer of heat to the ice is much more efficient and faster in the water than in the air.

Give students time to write their procedure and answer the question on the activity sheet.



EXPLAIN

4. Show an animation of ice melting.

Show the animation *Melting Ice*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#melting_ice



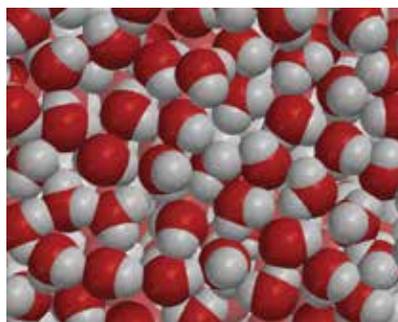
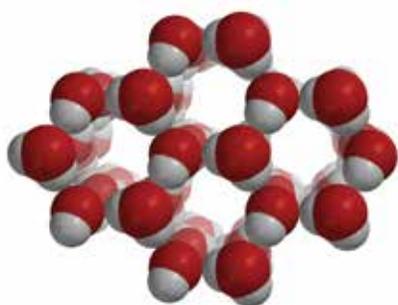
Point out that the water molecules in ice vibrate but don't move past each other. As the temperature increases they begin to vibrate more. Eventually their movement overcomes their attractions and they can no longer stay in their orderly crystal structure. As the ice melts, the orderly arrangement collapses and the water molecules move past each other and actually get closer together as liquid water.

Project the image *Ice and Water*

Ask students

- **How did the motion and arrangement of the water molecules change as the ice melted?**

As energy is transferred to the water molecules in the ice, the motion of the molecules increases. The motion of the molecules increases enough that it overcomes the attractions the water molecules have for each other causing the ice to melt.

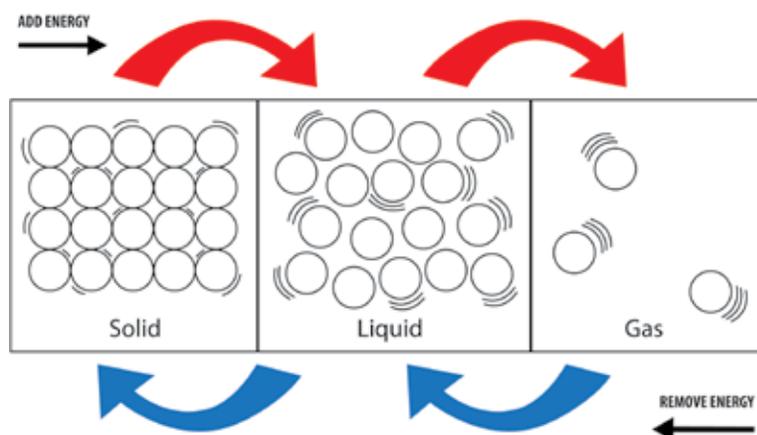


5. Compare the motion and arrangement of the molecules of a substance (not water) for each state of matter.

Project the image *States of Matter*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#states_of_matter

Explain that the diagram illustrates the motion and arrangement of atoms or molecules in a single substance (not water) when it changes between a solid, liquid, and gas.

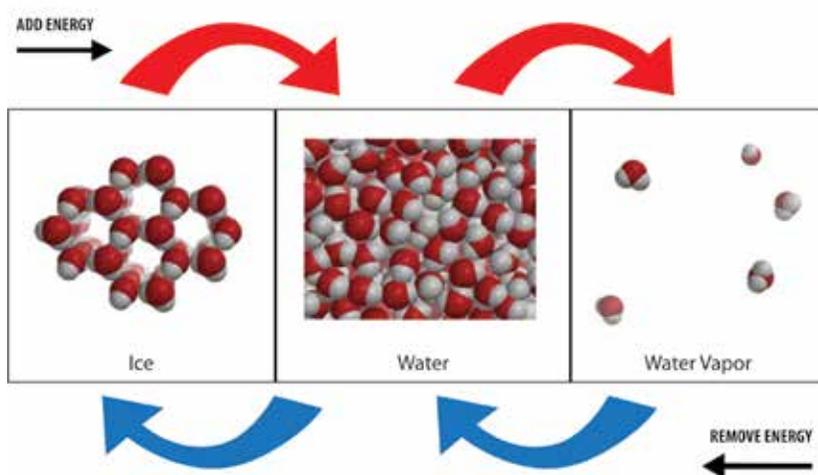


6. Have students compare the state changes of most substances to the state changes of water.

Project the image *States of Water*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#states_of_water

Tell students that the motion of water molecules in each state of matter is similar to what happens for most substances. Adding energy increases the motion of the molecules and causes them to move further apart. Removing energy decreases the motion of the molecules and causes them to move closer together. But, water does something very unusual when it freezes to become ice. The molecules, which were moving closer and closer together, move further apart as they organize themselves into the open ring pattern shown below for ice. This is why ice expands when it freezes.



Ask students:

How are the state changes of water similar to and different from the state changes in most other substances?

For water or any other substance, molecular motion increases when energy is added and decreases when energy is removed. The main difference between water and other substances is the arrangement between the molecules of the solid and the liquid. In water, the molecules in ice are further apart than they are in liquid water. This is unusual because the molecules of solids in most other substances are closer together than they are as a liquid.

Read more about energy and state changes in the additional teacher background section at the end of this lesson.

7. Have groups use their water molecules to model freezing, melting, evaporation, and condensation.

Procedure

- **Project the image *Ice*.**

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#ice

Have each group arrange their water molecules into a six-sided ring of ice. Ask students to handle their models gently because they will need them for other lessons.

- ***Ice Melts.***

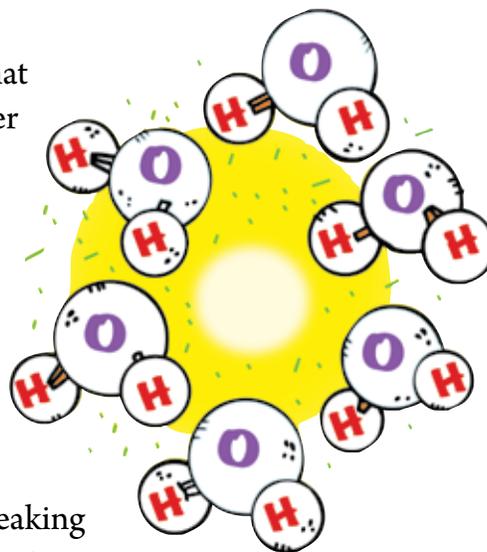
Have students use their models to represent what happens when ice melts. Point out that the water molecules are closer together than they were as ice. Students could show the water molecules moving past each other.

- **Water Evaporates**

Have students use their molecules to model what would happen if the water was heated and the molecules evaporated. Students should show the water molecules moving faster and breaking away from the other molecules and entering the air.

- **Water Vapor Condenses**

Have students use their molecules to model what would happen if water vapor was cooled enough to cause it to condense. Students should show the water molecules in the air slowing down and joining together but still moving past one another as liquid water.



Collect the water molecules. These models will be used again in Chapter 5.

EXTEND

8. Do a demonstration to compare the melting of regular ice and dry ice.

Let students know that dry ice is frozen carbon dioxide gas. Carbon dioxide gas must be very cold in order to become a solid (about $-78\text{ }^{\circ}\text{C}$ or $-109\text{ }^{\circ}\text{F}$).

Preparation

You will need some dry ice for this demonstration. If you cannot get any dry ice, show the video *Dry Ice*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#dry_ice

Question to investigate

Does dry ice melt the way regular ice does?

Materials

- Ice
- Dry ice
- Brown paper towel
- Cold water
- Hot water (about 50 °C)



Procedure

1. Place a piece of dry ice and a piece of regular ice on a brown paper towel.

Expected results

In a short amount of time, the ice will begin to melt and the paper towel around the ice will become wet and darker. The paper towel around the dry ice will stay dry and will not get darker. If you notice a small dark spot on the paper towel near the dry ice, it is possible that water vapor from the air condensed on the dry ice and melted onto the paper towel.

If students see misty white fog coming from the dry ice, let them know that it is not the carbon dioxide gas itself. Carbon dioxide is colorless, odorless, and invisible. The misty smoke or fog is actually water vapor in the air that gets cold enough to condense. The water vapor is cooled by the dry ice and the cold carbon dioxide gas. The fog tends to drift downward because it is carried by the carbon dioxide gas, which is more dense than the air around it.

9. Discuss student observations and introduce the idea that some substances can change directly from a solid to a gas.

Ask students:

- **Do regular ice and dry ice melt in the same way?**

No. The regular ice changes to a liquid, which you see on the brown paper towel. The dry ice does not seem to change to a liquid.

Explain to students that the reason that the dry ice does not make the paper towel wet is because it does not melt. When energy is transferred to dry ice, the solid carbon dioxide does not melt to liquid carbon dioxide. Instead, the solid changes directly to a gas. This process is called *sublimation*. Sublimation occurs when molecules of a solid move fast enough to overcome the attractions from other molecules and become a gas. Since frozen carbon dioxide never becomes a liquid under normal pressure, it is called dry ice.

10. Show students what happens when dry ice is placed in water.

Place a piece of dry ice in water or show the video *Dry Ice in Water*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#dry_ice_in_water

Expected results

Bubbles will form and a misty white fog will be produced. Since the water is much warmer than the dry ice, energy is transferred from the water to the dry ice, causing it to change from a solid to a gas and bubble through the water. After detergent is added, a mound of bubbles will form.

Students will be curious about all of the fog coming out of the cup. Tell them that some water changes to water vapor within the bubbles of carbon dioxide gas and then condenses. This causes fog within the bubbles which escapes when the bubble pops.

Ask students:

- **You saw that the dry ice sublimates very quickly in water.**

What could you do to make dry ice sublimate even faster?

There are several ways to make dry ice sublimate faster. One option is to put the dry ice in hot water.



Place a piece of dry ice in $\frac{1}{4}$ cup of cold water and another piece in $\frac{1}{4}$ cup of hot water. Or show the video *Dry Ice in Hot and Cold Water*.

www.middleschoolchemistry.com/multimedia/chapter2/lesson5#dry_ice_hot_cold_water

Expected results:

Much more fog will be produced from the cup with hot water.

Tell students that more fog is produced when dry ice is placed in hot water because the transfer of energy and sublimation happens faster. This causes the fog to be produced at a faster rate.

Activity Sheet
Chapter 2, Lesson 5
Changing State—Melting

Name _____

Date _____

DEMONSTRATION

1. You watched a piece of ice melt. Where do you think the energy came from to melt the ice?



2. What do you think happened to the speed of the molecules in the ice when it was heated?

ACTIVITY

Work with your group to design a way to make ice melt faster. You will need to show that your method really does make ice melt faster, so be sure to use a control. Check with your teacher before conducting your experiment.



Question to investigate

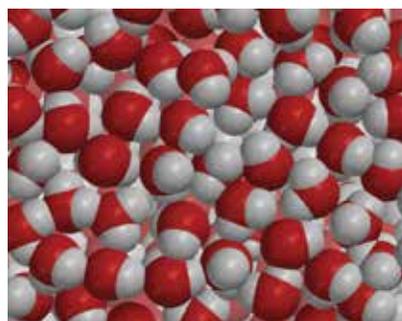
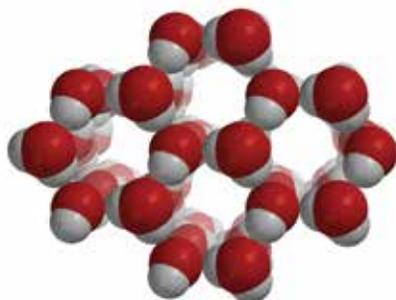
Will _____ make ice melt faster?

3. Does your method make ice melt faster?

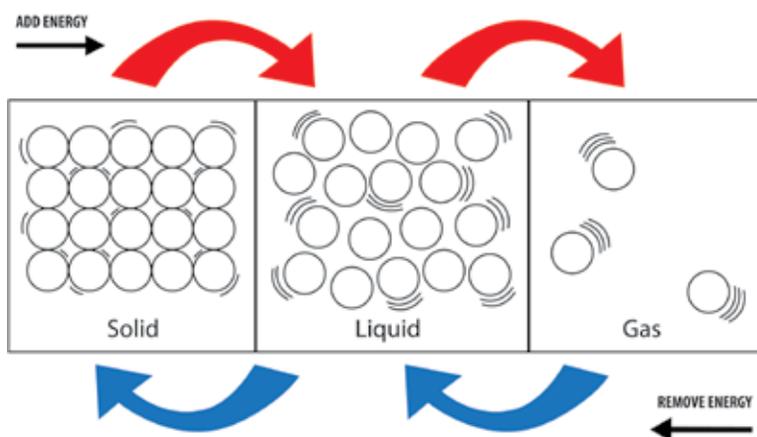
How do you know?

EXPLAIN IT WITH MOLECULES

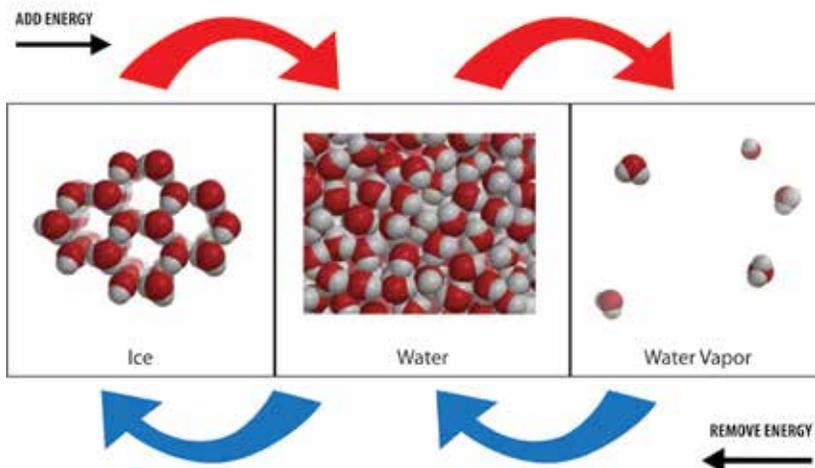
4. Write a caption underneath each picture to explain how the motion and arrangement of the water molecules changes as ice melts.



5. Look at the diagram below representing the motion and arrangement of the molecules of a substance (not water) when it is a solid, a liquid, and a gas. Write the name of the state change that takes place on each curved arrow.



6. The following diagram uses the space-filling model of water to represent the arrangement of water molecules when it is a solid, liquid, and a gas.



How are the state changes of water similar to the state changes in most other substances?

How are state changes of water different from the state changes in most other substances?

TAKE IT FURTHER

7. Do regular ice and dry ice melt in the same way?

How do you know?

8. You saw that the dry ice sublimates very quickly in water. Why does it sublimate even faster in hot water?

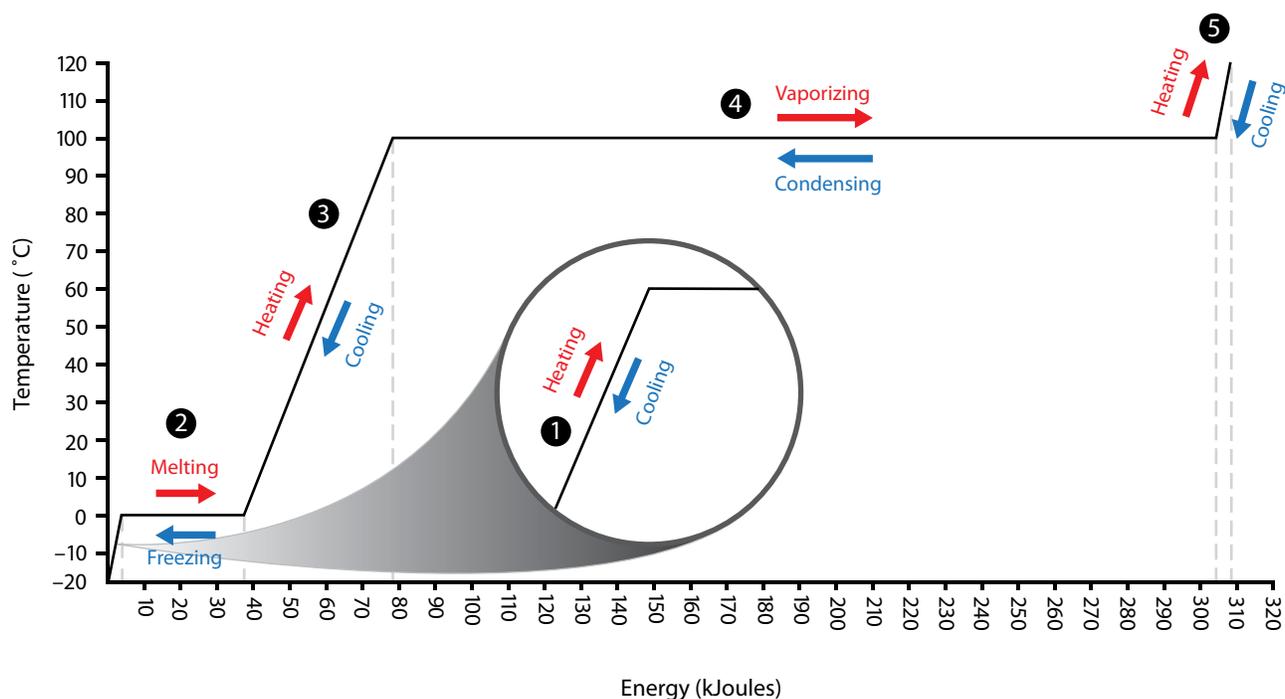
Additional Teacher Background

Chapter 2, Lesson 5, p. 130

Heating/Cooling curve

Throughout chapter 2, students have seen that adding energy (heating) increases the rate of melting and evaporation. They have also seen that removing energy (cooling) increases the rate of condensation and freezing. There are other interesting phenomena related to state changes that are not as easily explored in a classroom setting. One has to do with the relative amount of energy it takes to cause a substance to change from one state to another. Another is that the temperature of a substance remains constant during state changes.

Both these factors can be displayed on a graph showing the change in temperature while water is heated or cooled. This type of graph is called a heating curve. (You may sometimes see it called a phase diagram but a phase diagram is technically different.) The graph shows the energy added or removed on the x-axis and the corresponding temperature change on the y-axis. The units of energy are in kilojoules. The actual number of kilojoules required for the different state changes is not as important as seeing that some changes require much more or less energy than others. The graph is easiest to interpret if you look at the different steps separately and then all together.



Step 1: Adding energy increases the temperature of ice from $-20\text{ }^{\circ}\text{C}$ to $0\text{ }^{\circ}\text{C}$.

This takes a relatively small amount of energy because the energy goes into vibrating the molecules and not to break the bonds holding the molecules together. The molecules which are vibrating in fixed positions in the ice at $-20\text{ }^{\circ}\text{C}$ are made to vibrate somewhat faster at $0\text{ }^{\circ}\text{C}$.

Step 2: Adding energy causes the ice to change state by melting to liquid water with no change in temperature.

This process takes about 8 times as much energy as step 1. It makes sense that more energy is required because energy is being used to overcome the attractive forces holding the molecules in the crystal structure. The temperature does not change during this process because right at the melting point, the energy used in bond breaking does not increase the speed of the molecules, it just breaks the bond. The kinetic energy added is converted to potential energy which does not change the temperature.

Step 3: Adding energy increases the temperature of the water from $0\text{ }^{\circ}\text{C}$ to $100\text{ }^{\circ}\text{C}$.

This process takes about 25% more energy than step 2. So it takes more energy to raise the temperature of water from $0\text{ }^{\circ}\text{C}$ to $100\text{ }^{\circ}\text{C}$ than it does to melt the same amount of ice to liquid water. To melt ice, energy is added which causes the molecules to vibrate until the orderly arrangement of water molecules in the crystal collapses. But as liquid water, the water molecules are still attracted to each other and are still close together (in fact they are closer together in liquid water than in ice) but are able to slide past one another.

Once the ice has turned to liquid water, the energy added must still work against the attractions of the water molecules to raise the temperature of the water. This is why water has such a high specific heat which is the amount of energy required to raise 1 gram of a substance by $1\text{ }^{\circ}\text{C}$. Therefore raising liquid water from $0\text{ }^{\circ}\text{C}$ to $100\text{ }^{\circ}\text{C}$ takes a lot of energy.

Step 4: Adding energy causes the water to change state by vaporizing to become water vapor at the boiling point with no change in temperature.

This process takes nearly 5 times the amount of energy of step 3. Boiling takes more energy than the other processes because it is the only process in which the attractions between water molecules are completely overcome and water molecules are separated by relatively large distances. The temperature does not change during this process because right at the boiling point, the energy used in bond breaking does not increase the speed of the molecules, it just breaks the bond. The kinetic energy added is converted to potential energy which does not change the temperature.

Step 5: Adding energy causes the temperature of the water vapor to increase from 100 °C to 120 °C.

This process takes less energy than any of the other steps. Water molecules in the vapor phase are already far apart and do not feel significant attractions from one another. When energy is added to them, their motion readily increases.

Looking at the graph as a cooling curve

The entire graph can also be looked at as a cooling curve if you look at it in reverse.

Step 5: Removing energy (cooling) causes the temperature of water vapor to decrease from 120 °C to 100 °C.

It is exactly the same amount of energy removed that was added to cause the temperature to increase from 100 °C to 120 °C.

Step 4: Removing energy causes the water to change state by condensing to become liquid water at the boiling point with no change in temperature.

The temperature does not change during this phase change. This is because right at the condensation point, the energy removed allows the molecules to “bond” but does not change their speed.
Step 3: Removing energy from the liquid water causes its temperature to decrease from 100 °C to 0 °C.

It is exactly the same amount of energy removed that was added to cause the temperature to increase from 0 °C to 100 °C.

Step 2: Removing energy from water at 0 °C causes liquid water to change state by freezing to solid ice.

The temperature does not change during this process because right at the freezing point, the energy removed allows to molecules to form bonds but does not change their speed.

Step 1: Removing energy from ice causes the temperature of the ice to decrease from 0 °C to -20 °C.

It is exactly the same amount of energy removed as was added to cause the temperature to increase from -20 °C to 0 °C.

Chapter 2 — Student Reading

Atoms and molecules are in motion

We warm things up and cool things down all the time, but we usually don't think much about what's really happening. If you put a room-temperature metal spoon into a hot liquid like soup or hot chocolate, the metal gets hotter. But what actually has to happen for the hot liquid to make the metal hotter?

By now you know that substances are made of atoms and molecules. These atoms and molecules are always in motion. You also know that when atoms and molecules are heated, they move faster and when they are cooled, they move slower. But how do the atoms and molecules actually get heated and cooled? In our example of heating a metal spoon in a hot liquid, what is the process that transfers energy from the water to the spoon?

Moving atoms and molecules have energy

To answer this question, you really have to think about the moving atoms and molecules as having *energy*. Anything that has mass and is moving, like a train, a moving ball, or an atom has a certain amount of energy. The energy of a moving object is called kinetic energy. If the speed of the object increases, its kinetic energy increases. If the speed of the object decreases, its kinetic energy decreases. So if the atoms or molecules of a substance are moving fast, they have more kinetic energy than when they are moving more slowly.



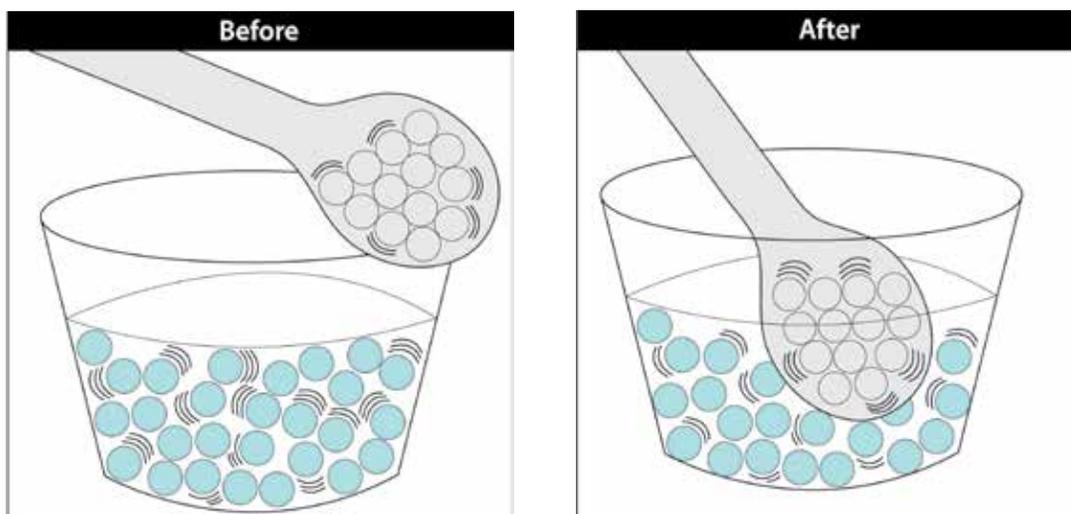
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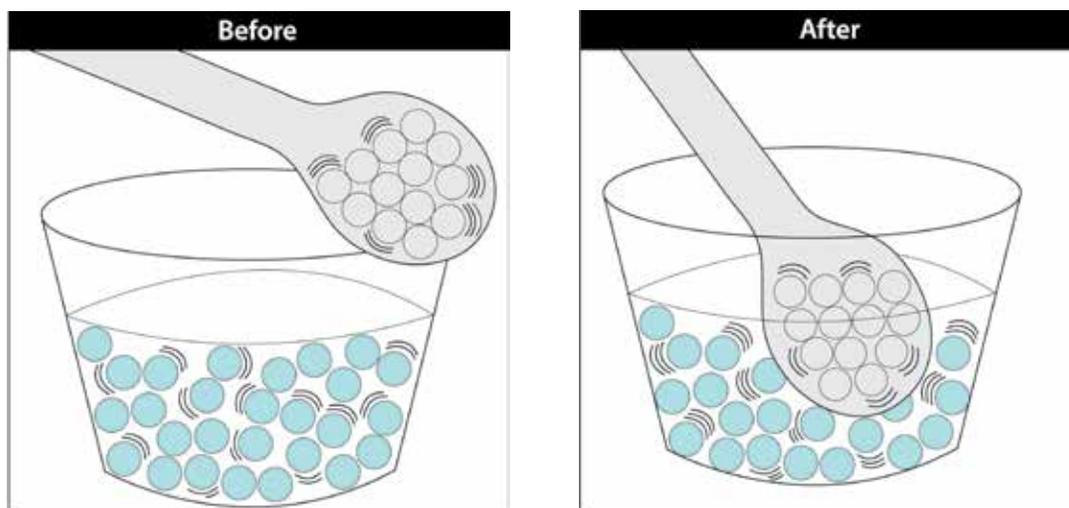
Energy can be transferred to make things warmer

In our example of a spoon in hot liquid, the molecules of hot liquid are moving quickly so they have a lot of kinetic energy. When the room-temperature spoon is placed in the liquid, the fast-moving molecules in the liquid contact the slower-moving atoms in the spoon. The fast-moving molecules hit the slower-moving atoms and speed them up. In this way, the fast-moving molecules transfer some of their kinetic energy to the slower atoms so that these slower atoms now have more kinetic energy. This process of transferring energy by direct contact is called **conduction**.



Energy can be transferred to make things cooler

Cooling things by conduction works the same way as warming but you just look at the substance *losing* energy instead of the substance *gaining* energy. This time, let's say that you take a hot metal spoon and put it in room-temperature water. The faster-moving atoms in the spoon contact the slower-moving molecules in the water. The atoms in the spoon transfer some of their energy to the molecules in the water. The spoon will get cooler and the water will get a little warmer.



Another example is cans of room-temperature soda pop placed in a cooler filled with ice. Kinetic energy is transferred from the warmer metal can to the cooler ice. This makes the can colder. Energy is then transferred from the warmer soda to the colder can. This transfer of energy from the soda results in slower motion of the molecules of the soda, which can be measured as a lower temperature and colder soda.

So the way to cool something is for its energy to be transferred to something colder. This is a rule about conduction: **Energy can only be transferred from something at a higher temperature to something at a lower temperature.** You can't cool something by adding "coldness" to it. You can only make something colder by allowing its energy to be transferred to something colder.

Temperature is a measure of the average kinetic energy of the atoms or molecules of a substance. This brings up the question of what exactly is temperature. Temperature is related to the kinetic energy of the moving atoms or molecules of a substance. By taking the temperature of something, you are actually getting information about the kinetic energy of its atoms and molecules, but not the kinetic energy of any particular one. There are more than a billion trillion atoms or molecules in even a small sample of a substance. They are constantly moving and bumping into each other and transferring little amounts of energy between each other. So at any time, the atoms and molecules don't all have the same kinetic energy. Some are moving faster and some are moving slower than others but most are about the same. So when you measure the temperature of something, you are actually measuring the average kinetic energy of its atoms or molecules.

Heat is the energy that is transferred from a substance at a higher temperature to a substance at a lower temperature.

If temperature is the average kinetic energy of atoms and molecules, then what is heat?

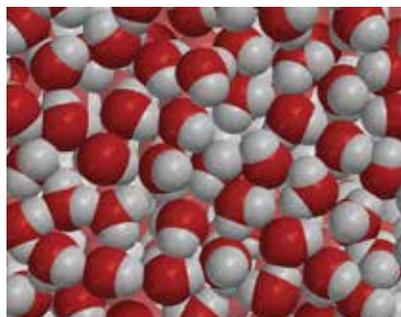
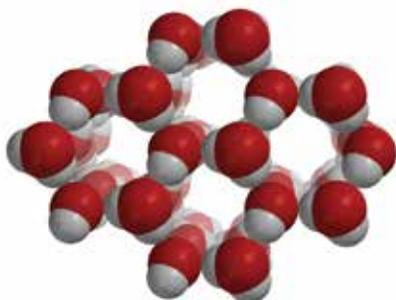
The word "heat" has a specific meaning in science even though we use the word all the time to mean different things in our daily life. The scientific meaning of heat has to do with energy that is being transferred. During conduction, the energy transferred from faster-moving atoms to slower-moving atoms is called heat.

Changing State

Changing from a solid to a liquid—Melting

In solids, the atoms or molecules that make up the substance have strong attractions to each other and stay in fixed positions. These properties give solids their definite shape and volume.

When a solid is heated, the motion of the particles (atoms or molecules) increases. The atoms or molecules are still attracted to each other but their extra movement begins to compete with their attractions. If enough energy is added, the motion of the particles begins to overcome the attraction and the particles move more freely. They begin to slide past each other as the substance begins to change state from a solid to a liquid. This process is called *melting*.



The particles of the liquid are only slightly further apart than in the solid. (Water is the exception because the molecules in liquid water are actually *closer together* than they are in ice) The particles of the liquid have more kinetic energy than they did as a solid but their attractions are still able to hold them together enough so that they retain their liquid state and do not become a gas.

Different solids melt at different temperatures

The temperature at which a substance begins to melt is called the *melting point*. It makes sense that different substances have different melting points. Since the atoms or molecules that substances are made of have a different amount of attraction for each other, a different amount of energy is required to make them change from a solid to a liquid. A good example is the melting point of salt and sugar. The melting point of sugar is 186 °C. The melting point for regular table salt is 801 °C. Metals like iron and lead also have different melting points. Lead melts at 327 °C and iron melts at 1,538 °C.

Some solids, like glass do not have a precise melting point but begin to melt over a range of temperatures. This is because the molecules that make up glass are not arranged in as orderly a way as those in crystals like salt or sugar or metals like iron. Depending on the type of glass, the melting point is usually between 1,200–1,600 °C.



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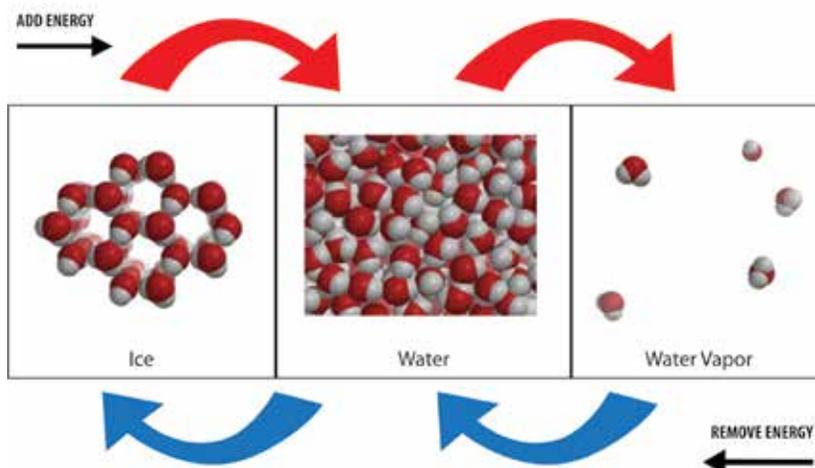
Changing from a solid to a gas—Sublimation

Some substances can also change directly from a solid to a gas. This process is called *sublimation*. One of the more popular examples of sublimation is dry ice which is frozen carbon dioxide (CO₂). To make dry ice, carbon dioxide gas is placed under high pressure and made very cold (about –78.5 °C). When a piece of dry ice is at room temperature and normal pressure, the molecules of CO₂ move faster and break away from each other and go directly into the air as a gas. Regular ice cubes in the freezer will also sublime but much more slowly than dry ice.

Changing from a liquid to a gas—Evaporation

You see evidence of evaporation all the time. Evaporation causes wet clothes to dry and the water in puddles to “disappear”. But the water doesn’t actually disappear. It changes state from a liquid to a gas.

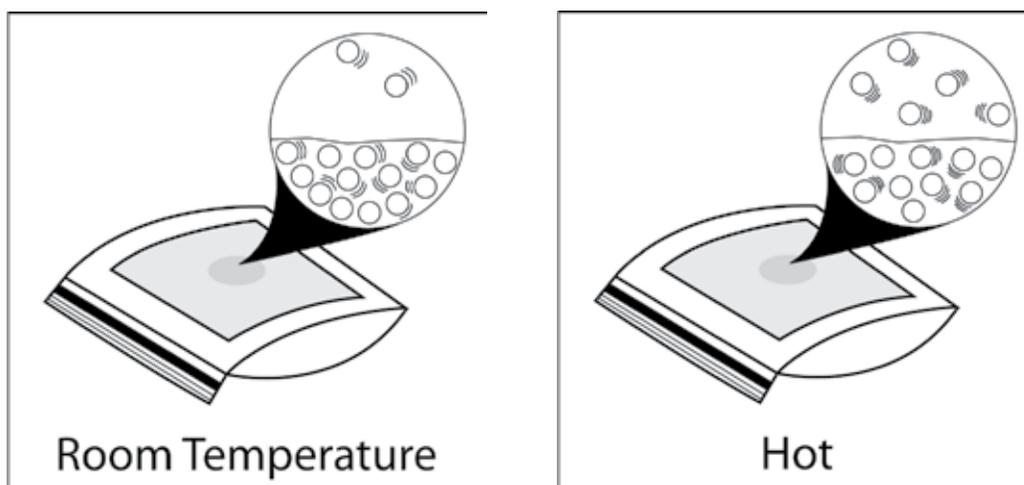
The molecules in a liquid evaporate when they have enough energy to overcome the attractions of the molecules around them. The molecules of a liquid are moving and bumping into each other all the time, transferring energy between one another. Some molecules will have more energy than others. If their motion is energetic enough, these molecules can completely overcome the attractions of the molecules around them. When this happens, the molecules go into the air as a gas. This process is called *evaporation*.



Heating increases the rate of evaporation

You've probably noticed that higher temperatures seem to make evaporation happen faster. Wet clothes and puddles seem to dry more quickly when they are heated by the sun or in some other way.

You can test whether heat affects the rate of evaporation by placing a drop of water on two paper towels. If one paper is heated and the other remains at room temperature, the water that is heated will evaporate faster.



When a liquid is heated, the motion of its molecules increases. The number of molecules that are moving fast enough to overcome the attractions of other molecules increases. Therefore, when water is heated, more molecules are able to break away from the liquid and the rate of evaporation increases.

Different liquids have different rates of evaporation

It makes sense that different liquids have different rates of evaporation. Different liquids are made of different molecules. These molecules have their own characteristic strength of attraction to one another. These molecules require a different amount of energy to increase their motion enough to overcome the attractions to change from a liquid to a gas.

Liquids will evaporate over a wide range of temperatures

Even at room temperature, or lower, liquids will evaporate. You can test this by wetting a paper towel and hanging it up indoors at room temperature. Evaporation at room temperature might seem strange since the molecules of a liquid need to have a certain amount of energy to evaporate. Where do they get the energy if the liquid is not warmed? But remember that the temperature of a substance is the average kinetic energy of its atoms or molecules. Even cold water, for instance, has a small percentage of molecules with much more kinetic energy than the others. With all the random bumping of a billion trillion molecules, there are always a few molecules which gain enough energy to evaporate. The *rate* of evaporation will be slow but it will happen.

Changing from a gas to a liquid—Condensation

If you have seen water form on the outside of a cold cup, you have seen an example of condensation. Water molecules from the air contact the cold cup and transfer some of their energy to the cup. These molecules slow down enough that their attractions can overcome their motion and hold them together as a liquid. This process is called *condensation*.

Cooling increases the rate of condensation

You can test to see if cooling water vapor increases the rate of condensation by making two similar samples of water vapor and cooling one more than the other. In the illustration, two samples of water vapor are trapped inside the cups. Ice is placed on one top cup but not the other.

In a few minutes, there are water drops on the inside top of both cups but more water can be seen on the inside of the top cup with the ice. This shows that cooling water vapor increases the rate of condensation. When a gas is cooled, the motion of its molecules decreases. The number of molecules moving slow enough for their attractions to hold them together increases. More molecules join together to form a liquid and the rate of condensation increases.



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The amount of water vapor in the air affects the rate of condensation

Temperature isn't the only factor that affects the rate of condensation. At a given temperature, the more water molecules in the air, the greater the rate of condensation. If there are more molecules, a greater number of molecules will be moving at these different speeds. More molecules will be moving slowly enough to condense.

Different gases condense at different temperatures

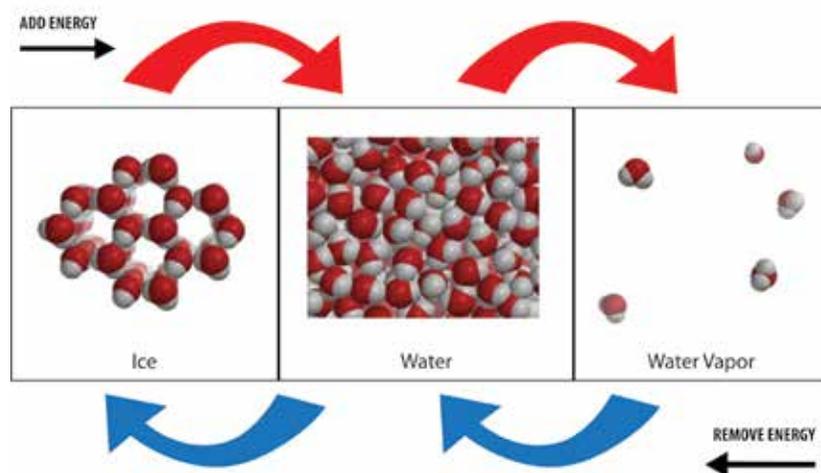
Each gas is made up of its own molecules which are attracted to each other a certain amount. Each gas needs to be cooled to a certain temperature in order for the molecules to slow down enough so that the attractions can hold them together as a liquid.

Changing from liquid to solid—Freezing

If a liquid is cooled enough, the molecules slow down to such an extent that their attractions begin to overcome their motion. The attractions between the molecules cause them to arrange themselves in more fixed and orderly positions to become a solid. This process is called *freezing*.

Water molecules move further apart as water freezes

The freezing of water is very unusual because water molecules move farther apart as they arrange themselves into the structure of ice as water freezes. The molecules of just about every other liquid move closer together when they freeze.



Different liquids have different freezing points

It makes sense that different liquids have different freezing points. Each liquid is made up of different molecules. The molecules of different liquids are attracted to each other by different amounts. These molecules have to slow down to different extents before their attractions can take hold and organize themselves into fixed positions as a solid.

Changing from a gas to a solid—Deposition

With the right concentration of gas molecules and temperature, a gas can change directly to a solid without going through the liquid phase. This process is called *deposition*. It is the opposite of sublimation. One of the most common examples of deposition is the formation of frost. When there is the right combination of water vapor in the air and temperature, the water can change to frost without first condensing to liquid water.



Evaporation, condensation, and the weather

Clouds

Clouds form when liquid water evaporates to become water vapor and moves up into the sky in upward-moving air. Air at higher altitudes is usually cooler than air near the ground. So as the water vapor rises, it cools and condenses, forming tiny drops of water. These droplets are suspended in the air as clouds. Clouds at higher altitudes where the air is even colder also contain ice crystals. Clouds at very high levels are composed mostly of ice crystals.



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Rain

Rain begins as tiny droplets of water suspended in the air as clouds. These droplets are so small that they don't fall yet to the ground as rain. They are similar to the tiny droplets in fog or mist. But when these droplets collect and join together, they become bigger and heavier drops. Eventually these drops become so heavy that they fall to the ground as rain.



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Snow

Like rain, snow begins with condensing water vapor that forms clouds. However, when it's cold enough, water vapor not only condenses but also freezes, forming tiny ice crystals. More and more water vapor condenses and freezes on these seed crystals, forming the beautiful ice crystals with six-sided symmetry that we know as snowflakes.



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Hail

Hail forms when a small drop of water freezes, falls, and then gets pushed back up into a cloud. More water droplets collect and freeze on this ice crystal, which makes it heavier, and it begins to fall again. The violent air in a thundercloud (cumulonimbus cloud) repeatedly bounces the ice crystal upward. Each time it gets another coating of freezing water. Finally, the ice crystal is so heavy that it falls all the way down to the ground as hail.

Dew

Dew is produced when moist air close to the ground cools enough to condense and forms liquid water. Dew is different than rain because dew doesn't fall onto the ground in drops. It slowly accumulates to form drops on objects near the ground. Dew often forms on blades of grass and leaves and can make beautiful designs on spider webs.



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Frost

If the temperature of surfaces on the ground is low enough, water vapor in the air can change directly to the solid frost without first condensing to a liquid.



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Fog

Usually the air near the ground is warmer than the air above it, but the conditions that cause fog are just the reverse. Fog forms when warm, moist air passes over cold ground or snow. As the water vapor in the air cools, it condenses, forming very tiny drops of water suspended in the air. Fog is very much like a cloud, but closer to the ground.



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Mist on a Pond

Water evaporates even when the air is cold. To form mist, the water in a pond, pool, or hot tub must be warmer than the air above it and the air must be cold enough to cause the water vapor to condense as it rises. The mist seems to disappear as the water droplets evaporate to become water vapor again.



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Chapter 3, Lesson 1: What is Density?

Key Concepts

- Density is a characteristic property of a substance.
- The density of a substance is the relationship between the mass of the substance and how much space it takes up (volume).
- The mass of atoms, their size, and how they are arranged determine the density of a substance.
- Density equals the mass of the substance divided by its volume; $D = m/v$.
- Objects with the same volume but different mass have different densities.

Summary

Students will observe a copper and an aluminum cube of the same volume placed on a balance. They will see that the copper has a greater mass. Students will try to develop an explanation, on the molecular level, for how this can be. Students are then given cubes of different materials that all have the same volume. Students determine the density of each cube and identify the substance the cube is made from.

Objective

Students will be able to calculate the density of different cubes and use these values to identify the substance each cube is made of. Students will be able to explain that the size, mass, and arrangement of the atoms or molecules of a substance determines its density.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- Cubes marked A–H that you will share with other groups
- Balance that can measure in grams
- Calculator

Materials for the Demonstration

- Copper cube and aluminum cube of the same volume
- Balance

Notes about the materials

Cubes

For this lesson, you will need a set of cubes of different materials that are all the same volume. These sets of cubes are available from a variety of suppliers. Flinn Scientific sells a Density Cube Set, Product #AP6058. This set comes with 10 cubes—4 metal, 3 plastic, and 3 wood. It is easier for students if you reduce the number to 8 by using all the samples of metal but only two wood and two plastic cubes. We suggest using the nylon (off-white, least dense) plastic cube and the PVC (gray, most dense) plastic cube. For the wood, we suggest using the oak (darker and most dense) and either the pine or poplar (paler, less dense). In the activity, each group will need to measure the mass of each of the eight cubes. Groups will need to measure and record their data for a cube and pass it along to another group until each group has used each of the cubes.

Balances

Use a simple, plastic, two-sided balance that looks like a see-saw for the demonstration. One of the least expensive is Delta Education, Stackable Balance (21-inch) Product #WW020-0452. Have students use any balance that can measure in grams.

Metric ruler

Students will use a metric ruler in the engage portion of the activity when they measure the length, width, and height of a cube along with you.

About this Lesson

This is the first lesson in which students see models of molecules that are more complex than a water molecule. Some of these molecules may look a little intimidating. Let students know that they do not need to memorize or draw these molecules. For the purpose of this chapter, students only need to think about the size and mass of the atoms that make up the molecule and how they are arranged in the substance.

ENGAGE

1. Do a demonstration to show that cubes of the same volume but made of different metals have different masses.

Question to investigate

Do cubes of exactly the same size and shape, have the same mass?



Materials for the demonstration

- Copper cube and aluminum cube of the same volume
- Balance

Procedure

Place the copper and aluminum cube on opposite sides of a simple balance.

Expected results

The copper cube will have a greater mass than the aluminum cube.



2. Lead a discussion about why the copper cube has a greater mass than the aluminum cube.

Tell students that both cubes are exactly the same size and both are solid with no hollow spots. Explain that the aluminum cube is made of only aluminum atoms and the copper cube is made of only copper atoms.

Ask students:

- **How can two objects, which are exactly the same size and shape, have a different mass?** Help students understand that the difference in mass must have something to do with the atoms in each cube. There are three possible explanations about the copper and aluminum atoms in the cubes that could explain the difference in mass.
 - Copper atoms might have more mass than aluminum atoms.
 - Copper atoms might be smaller so more can fit in the same volume.
 - Copper and aluminum atoms might be arranged differently so more copper atoms fit in the same size cube.

Explain that any one of these explanations alone, or two or three together, could be the reason why the copper cube has more mass.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



3. Project an illustration and use the pictures of the copper and aluminum atoms to introduce the concept of density.



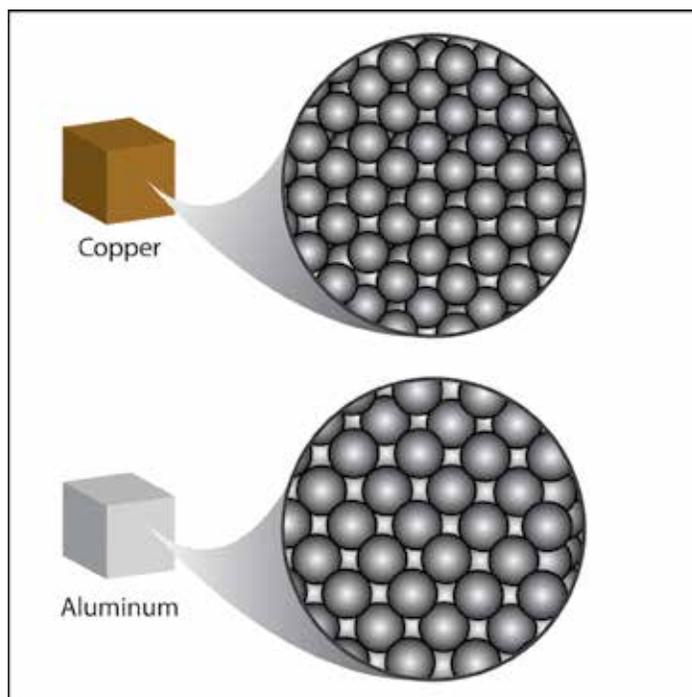
Have students turn to the illustration of copper and aluminum cubes and their atoms on their activity sheet.

Show students the image *Aluminum and Copper Atoms*

www.middleschoolchemistry.com/chapter3/lesson1#aluminum_and_copper

Point out that the copper atoms are slightly smaller than aluminum atoms. This smaller size means that more copper atoms can fit in the same amount of space. So, the copper cube contains more atoms than the aluminum cube. Although they are smaller, individual copper atoms actually have more mass than individual aluminum atoms. The combination of more atoms, each with a greater mass, makes a copper cube weigh more than an aluminum cube of the same size and shape.

Explain to students that this idea of how heavy something is compared to the amount of space it takes up is called *density*. The density of an object is the mass of the object compared to its volume. The equation for density is: $Density = mass/volume$ or $D = m/v$. Each substance has its own characteristic density because of the size, mass, and arrangement of its atoms or molecules.



4. Show animations and demonstrate how to measure volume and mass of a cube.

Explain to students that volume is a measure of the amount of space an object takes up. It is always in three dimensions. To find the volume of an object like a cube or a box, you measure the length, width, and height and then multiply them ($V = l \times w \times h$). If measured in centimeters, the answer will be in cubic centimeters (cm^3).



Note: Students often confuse volume and area. Check their understanding to make sure they know the difference. Make sure they understand that area is measured in two dimensions (length \times width) with an answer in cm^2 . Area is a measure of the amount of surface. But volume is measured in three dimensions (length \times width \times height) with an answer in cm^3 . Volume is a measure of the entire object, including the surface and all the space the object takes up.

Show the animation *Cube*.

www.middleschoolchemistry.com/chapter3/lesson1#cube

While the animation is playing, you can demonstrate the measuring process with a cube and ruler. Have students measure along with you to confirm the volume of the cubes.

Volume

The cubes are 2.5 centimeters on each side. Show students that in order to calculate the volume, you multiply the length (2.5 cm) \times width (2.5 cm) \times height (2.5 cm) to get 15.625 cm^3 . Rounding this number to 15.6 cm^3 is accurate enough and will make the density calculations easier. Record the volume of the cube in cubic centimeters (cm^3).



Mass

Demonstrate how to use the balance that students will be using to measure the mass of the cube. Record the mass of the cube in grams (g).

Density

Show students how to calculate density by dividing the mass by the volume. Point out that the answer will be in grams per cubic centimeter (g/cm^3).

EXPLORE

5. Have students calculate the density of eight different cubes and use the characteristic property of density to correctly identify them.

Student groups will not need to measure the volume of the cubes. The volume of each cube is the same, 15.6 cm^3 , and is given in their chart on the activity sheet. They will need to measure the mass of each of the eight different cubes and calculate their densities. Students will use their values for density to identify each cube.

Note: The densities students calculate may not be exactly the same as the given densities in this chart. However, their calculations will be close enough that they should be able to identify most of the cubes.

Question to investigate

Can you use density to identify eight cubes made of different materials?

Materials for the class

- Set of eight cubes of equal volume
- Calculator

Teacher preparation

Use a piece of masking tape and a permanent marker to mark the eight cubes with the letters A–H.

Materials for each group

- Cubes marked A–H that you will share with other groups
- Balance that can measure in grams
- Calculator

Procedure

1. The volume of each cube is given in the chart. It is 15.6 cm^3 .
2. Find the mass in grams of each cube using a scale or balance. Record this mass in the chart.
3. Trade cubes with other groups until you have measured the mass of all eight cubes.
4. Calculate the density using the formula $D = m/v$ and record it in the chart.

Sample	Volume (cm ³)	Mass (g)	Density (g/cm ³)	Material
A	15.6			
B	15.6			
C	15.6			
D	15.6			
E	15.6			
F	15.6			
G	15.6			
H	15.6			

Material	Approximate density (g/cm ³)
Aluminum	2.9
Brass	8.8
Copper	9.3
Steel	8.2
PVC	1.3
Nylon	1.2
Oak	0.7–0.9
Pine or poplar	0.4–0.6

5. Compare the value you found for density with the given value in the chart below to identify which cube is made out of which material. Write the name of the material in your chart for cubes A–H.

Expected results: Student values for density for each cube will not be exact, but will be close enough that they should be able to identify each of the cubes. You may notice that the approximate densities given for each cube in this lesson are slightly different than those given in the cube set. Most of this difference is probably due to the value for the volume of each cube. Since it is likely that these are 1-inch cubes, each side should be 2.54 cm. We rounded to 2.5 cm because students can make this measurement more easily.

EXPLAIN

6. Discuss how the mass, size, and arrangement of atoms and molecules affect the densities of metal, plastic, and wood

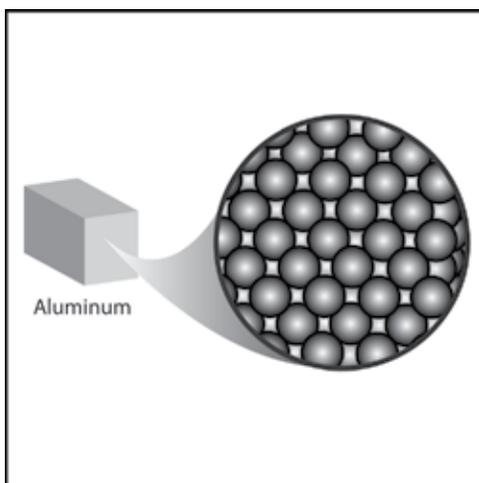


Explain to students that each substance has its own density because of the atoms and molecules it is made from. The metal, plastic, and wood cubes that students measured each have their own unique density. In general, the density of metal, plastic, and wood can be explained by looking at the size and mass of the atoms and how they are arranged.

Project the image *Metal*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson1#metal

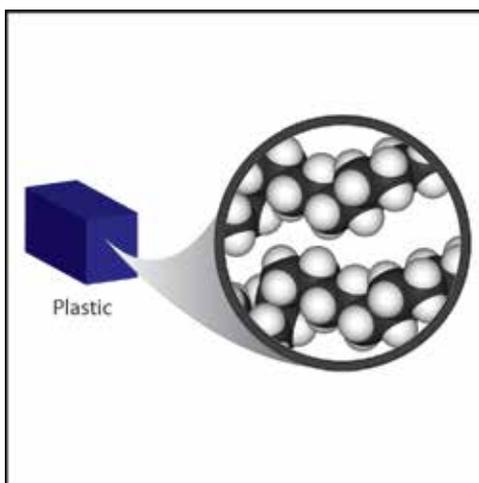
Most common metals like aluminum, copper, and iron are more dense than plastic or wood. The atoms that make up metals are generally heavier than the atoms in plastic and wood and they are packed closer together. The difference in density between different metals is usually based on the size and the mass of the atoms but the arrangement of the atoms in most metals is mostly the same.



Project the image *Plastic*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson1#plastic

Most plastics are less dense than metal but can have similar density to wood. Plastics are made from individual molecules bonded together into long chains called *polymers*. These polymer chains are arranged and packed together to make the plastic. One common plastic, *polyethylene*, is made up of many individual molecules called *ethylene* which bonded together to make the long polymer chains. Like most plastics, the polymers in polyethylene are made of carbon and hydrogen atoms.



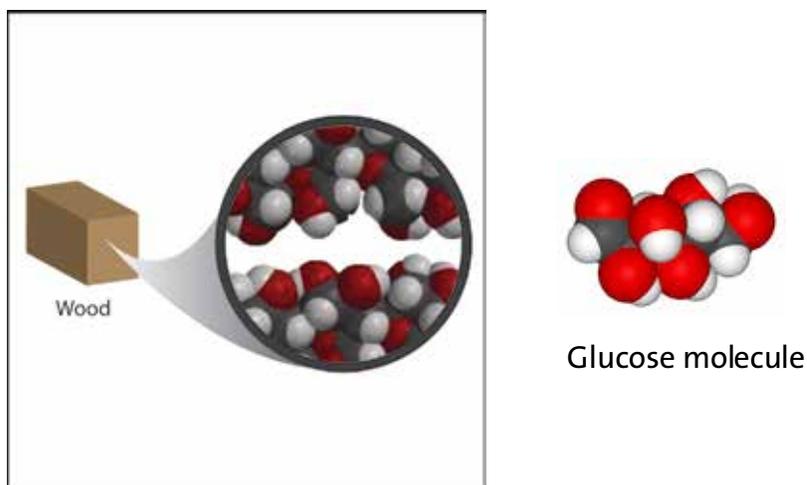
Ethylene molecule

The carbon and hydrogen atoms are very light, which helps give plastics their relatively low density. Plastics can have different densities because different atoms can be attached to the carbon-hydrogen chains. The density of different plastics also depends on the closeness of packing of these polymer chains.

Project the image Wood.

www.middleschoolchemistry.com/multimedia/chapter3/lesson1#wood

Wood is made mostly from carbon, hydrogen, and oxygen atoms bonded together into a molecule called *glucose*. These glucose molecules are bonded together to form long chains called *cellulose*. Many cellulose molecules stacked together give wood its structure and density.



In general, the density of wood and plastic are similar because they are made of similar atoms arranged in long chains. The difference in density is mostly based on the arrangement and packing of the polymer chains. Also, since wood is from a living thing, its density is affected by the structure of plant cells and other substances that make up wood.

Ask students:

The size, mass, and arrangement of atoms affect the density of a substance.

- **How might these factors work together to cause a substance to have a *high density*?**

A substance with smaller more massive atoms that are close together is going to have a higher density.

- **How might these factors work together to cause a substance to have a *low density*?**

A substance with larger, lighter atoms that are farther apart is going to have a lower density.

EXTEND

7. Have students explain on the molecular level why two blocks of different materials that have the same mass can have different densities.

Remind students that they looked at cubes that had the *same volume* but *different masses*. Point out that their activity sheet has drawings of two blocks (Sample A and Sample B) made of different substances that both have the *same mass*, but *different volumes*.

Ask students:

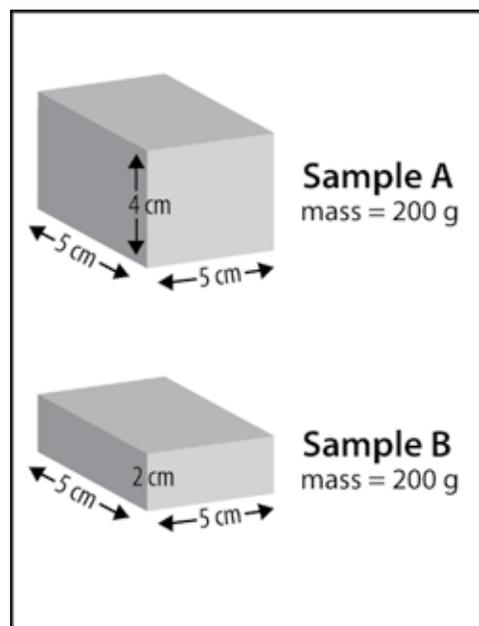
- **What is the density of Sample A?**
Volume = $5 \times 5 \times 4 = 100 \text{ cm}^3$
Mass = 200 g
Density = $200 \text{ g}/100 \text{ cm}^3 = 2 \text{ g}/\text{cm}^3$

- **What is the density of Sample B?**
Volume = $5 \times 5 \times 2 = 50 \text{ cm}^3$
Mass = 200 g
Density = $200 \text{ g}/50 \text{ cm}^3 = 4 \text{ g}/\text{cm}^3$
Give two possible explanations for why one sample is more dense than the other.

Hint: The size, mass, and arrangement of molecules affect the density of a substance. There are several possible answers for why sample B is more dense than sample A.

- Sample B atoms might have more mass than Sample A atoms.
- Sample B atoms might be smaller than Sample A atoms so more can fit in the same volume.
- Sample B atoms might be arranged differently so more Sample B atoms than Sample A atoms fit in the same size cube.

Any one of these explanations alone, or any combination, could be the reason why Sample B is more dense than Sample A.



Activity Sheet
Chapter 3, Lesson 1
What is Density?

Name _____

Date _____

DEMONSTRATION

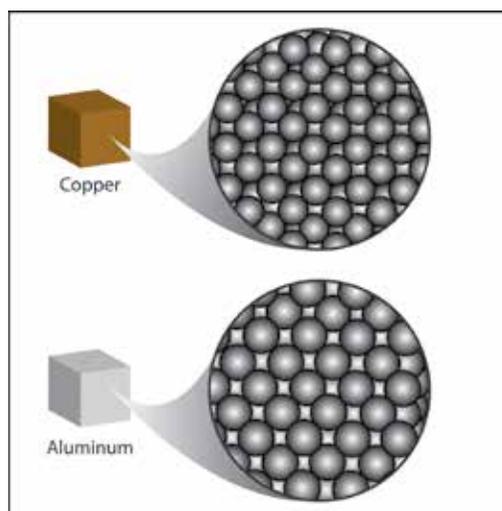
Your teacher placed a copper and an aluminum cube on a balance. Even though the cubes are the same size and shape, the copper has a greater mass than the aluminum. Both cubes are solid and are not hollowed out anywhere inside. The copper cube is made up of only copper atoms and the aluminum cube is made up of only aluminum atoms.

1. Look at the drawing of the copper and aluminum cubes and their atoms. What are two possible explanations for why the copper cube has a greater mass than the aluminum cube?

Hint: Just because the aluminum atoms are larger, they are not necessarily heavier.

Explanation 1:

Explanation 2:



2. The density of a substance like copper or aluminum is its mass divided by its volume (how much space it takes up).
Density = mass/volume or $D = m/v$.

Which is more dense, copper or aluminum?

How do you know?

You saw an animation about how to find the volume, mass, and density of a cube.

3. How do you find the volume of a cube?

4. How do you find the mass of a cube?

5. Once you know the volume and mass of a cube, how do you find the density of the cube?

6. Calculate the density of a cube using the following information:

- Each side is 4 cm long.
- The mass of the cube is 128 g.

ACTIVITY

Your group will work with eight cubes each with the same volume, but made of a different material. Carefully measure the mass of each cube and calculate the density. Then use density to correctly identify each of the 8 cubes.

Question to investigate

Can you use density to identify eight cubes made of different materials?

Materials for the class

- Set of eight cubes of equal volume
- Calculator

Teacher preparation

Use a piece of masking tape and a permanent marker to mark the eight cubes with the letters A–H.

Materials for each group

- Cubes marked A–H that you will share with other groups
- Balance that can measure in grams
- Calculator

Procedure

1. The volume of each cube is given in the chart. It is 15.6 cm^3 .
2. Find the mass in grams of each cube using a scale or balance. Record this mass in the chart.
3. Trade cubes with other groups until you have measured the mass of all eight cubes
4. Calculate the density using the formula $D = m/v$ and record it in the chart.

Sample	Volume (cm^3)	Mass (g)	Density (g/cm^3)	Material
A	15.6			
B	15.6			
C	15.6			
D	15.6			
E	15.6			
F	15.6			
G	15.6			
H	15.6			

5. Compare the value you found for density with the given value in the chart below to identify which cube is made out of which material. Write the name of the material in your chart for cubes A–H.

Material	Approximate Density (g/cm ³)
Aluminum	2.9
Brass	8.8
Copper	9.3
Steel	8.2
PVC	1.3
Nylon	1.2
Oak	0.7–0.9
Pine or poplar	0.4–0.6

EXPLAIN IT WITH ATOMS & MOLECULES

7. The size, mass, and arrangement of atoms affect the density of a substance.

If a substance has a *high density*, what can you guess about the size, mass, and arrangement of the atoms that make up the substance?

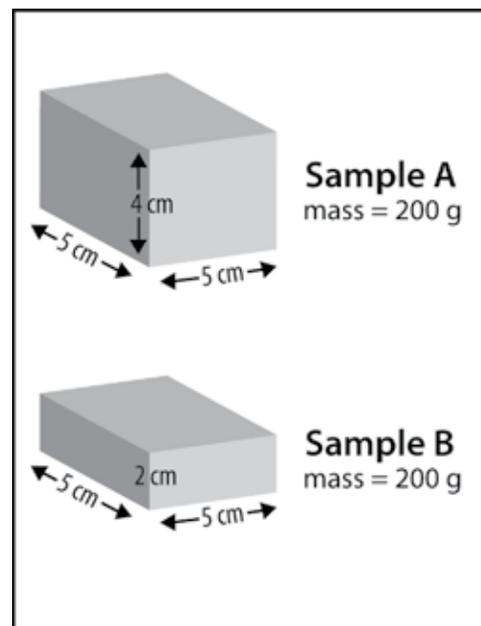
If a substance has a *low density*, what can you guess about the size, mass, and arrangement of the atoms that make up the substance?

TAKE IT FURTHER

In this activity, you investigated cubes made out of different substances. The cubes had the same volume, but different masses. When you calculated the density of each cube, you found that this was different, too.

8. Now imagine two blocks (Sample A and Sample B) made of different substances that both have the *same mass*, but *different volumes*.

a. What is the density of Sample A?



b. What is the density of Sample B?

c. Give two possible explanations for why one sample is more dense than the other.

Hint: The size, mass, and arrangement of molecules affect the density of a substance.

Chapter 3, Lesson 2: Finding Volume—The Water Displacement Method

Key Concepts

- A submerged object displaces a volume of liquid equal to the volume of the object.
- One milliliter (1 mL) of water has a volume of 1 cubic centimeter (1 cm³).
- Different atoms have different sizes and masses.
- Atoms on the periodic table are arranged in order according to the number of protons in the nucleus.
- Even though an atom may be smaller than another atom, it might have more mass.
- The mass of atoms, their size, and how they are arranged determine the density of a substance.
- Density equals the mass of the object divided by its volume; $D = m/v$.
- Objects with the same mass but different volume have different densities.

Summary

Students use the water displacement method to find the volume of different rods that all have the same mass. They calculate the density of each rod, and use the characteristic density of each material to identify all five rods. Then students consider the relationship between the mass, size, and arrangement of atoms to explain why different rods have different densities. Students will be briefly introduced to the periodic table.

Objective

Students will be able to explain that materials have characteristic densities because of the different mass, size, and arrangement of their atoms. Students will be able to use the volume displacement method to find the volume of an object.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- Set of 5 different rods that all have the same mass
- Graduated cylinder, 100 mL
- Water in a cup
- Calculator

Notes about the materials:

For this lesson you will need a set of five solid rods, each with the same mass, same diameter, but a different volume. Each rod is made of a different material. There are several versions of these rods available from different suppliers. This activity uses a set from Flinn Scientific, Equal Mass Kit, Product # AP 4636 but can be adapted to any set of equal mass rods. Since there are only five samples in the Equal Mass kit, you may need two kits so that each group can work with a sample.

This chart will help you identify each rod. Do not reveal this information to the students. They will discover the identity of each rod and the inverse relationship between the density and the length of each rod later in this lesson.

Sample	Material	Approximate density (g/cm ³)	Relative length
Smallest metal	Brass	7.5	shortest longest
Shiny gray metal	Aluminum	3.0	
Dark gray	PVC	1.4	
Tall off-white	Nylon	1.1	
Tallest white	Polyethylene	0.94	

ENGAGE

1. Show students five rods that have the same mass but different volumes.

Show students the five rods and explain that they all have the same mass. Then hold up the longest, middle-sized, and shortest rods and remind students that they have the same mass.

Ask students to make a prediction:

- Which rod is the most dense? Least dense? In between?

Students may reason that since the mass of each rod is the same, the volume of each rod must have something to do with its density. Some may go so far as to say that the rod with the smallest volume must have the highest density, because the same mass is packed into the smallest volume. Or that the rod with the largest volume must have the lowest density, because the same mass is spread out over the largest volume.

Tell students that like the cubes in the previous activity, they will need to know the volume and mass of each of the samples. They will also calculate the density of each sample and use this value to figure out which material each rod is made of.

2. Show an animation and demonstrate how to measure volume using the water displacement method.

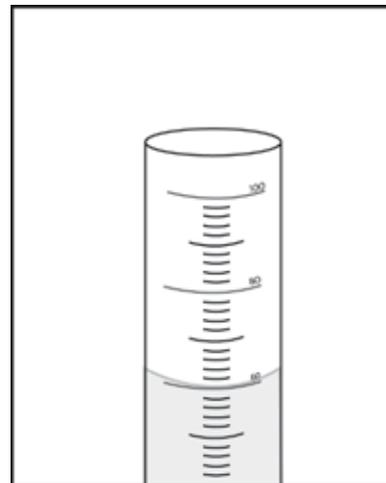
Project the animation *Water Displacement*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson2#water_displacement

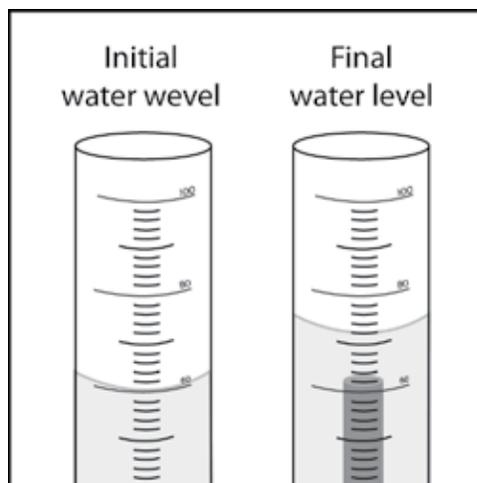
Play the animation as you demonstrate the water displacement method using a cup of water, a graduated cylinder, and a rod, the way students will do in the activity. Use the dark gray plastic sample so that students can see it better.

Volume

1. Demonstrate what students will do by pouring water from a cup into a 100-mL graduated cylinder until it reaches a height that will cover the sample. This is the “initial water level.”
2. Tell students that the surface of water in a tube may not be completely flat. Instead, the surface may curve in a shallow U-shape called the *meniscus*. When measuring, read the line just at the bottom of the meniscus.
3. Tilt the graduated cylinder and slowly slide the sample into the water. Hold the graduated cylinder upright. Record the level of the water. Point out that this is the “final water level.”



4. Tell students that you want to find out how much the water level changed. Subtract the initial water level from the final water level to find the volume of the rod.
 - Volume of sample = final water level – initial water level



Students may be confused that the unit for volume in the graduated cylinder is milliliters (mL), when in the previous lesson students calculated volume in cubic centimeters (cm^3). Explain to students that 1 mL is the same as 1 cm^3 . Click on the oval-shaped button on the first screen of the animation marked “1 mL = 1 cm^3 .”

Ask students:

- **When you place a sample in the water, why does the water level go up?**
The volume that the rod takes up pushes or displaces the water. The only place for the water to go is up. The amount or volume of water displaced is equal to the volume of the sample.
- **Is the volume of the sample equal to the final water level?**
No. Students should realize that the volume of the rod is not equal to the level of the water in the graduated cylinder. Instead, the volume of the rod equals the amount that the water went up in the graduated cylinder (the amount displaced). To find the amount of water displaced, students should subtract the initial level of the water (60 mL) from the final level of the water.
- **What units should you use when you record the volume of the sample?**
Because they will be using the volume to calculate density, students should record the volume of the sample in cm^3 .

Mass

Student groups will not need to measure the mass of the rods. The mass of each rod is the same, 15 grams, and is given in their chart on the activity sheet. They will need to measure the volume of each of the five different rods and calculate their densities. Students will use their values for density to identify each rod.

Density

Demonstrate how to calculate density ($D = m/v$) by dividing the mass by the volume. Point out that the answer will be in grams per cubic centimeter (g/cm^3).



Give one activity sheet to each student.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

Give students time to answer questions 1–5 on the activity sheet before starting the activity.

EXPLORE

3. Have students calculate the density of five different rods and use the characteristic property of density to correctly identify them.



Note: The densities for the three plastics are similar, so students need to be very careful when measuring their volume using the water displacement method. Also, it is difficult to measure the volume of the smallest rod. Give students a hint that it is between 1.5 and 2.0 mL.

Question to investigate

Can you use density to identify all five rods?

Materials for each group

- Set of five different rods that all have the same mass
- Graduated cylinder, 100 mL
- Water in a cup
- Calculator

Teacher preparation

- Use a permanent marker to mark the five rods with letters A, B, C, D, and E. Keep track of which letter corresponds to which sample without letting students know. If you are using two or more sets of rods, be sure to mark each sample of the same material with the same letter.
- After a group finds the volume of a sample, they should then pass that sample to another group until all groups have found the volume of all five rods.
- For the longest sample, which floats, students can use a pencil to gently push the sample just beneath the surface of the water to measure its full volume.

Procedure

Volume

1. Pour enough water from your cup into the graduated cylinder to reach a height that will cover the sample. Read and record the volume.
2. Slightly tilt the graduated cylinder and carefully place the sample into the water.
3. Place the graduated cylinder upright on the table and look at the level of the water. If the sample floats, use a pencil to gently push the top of the sample just under the surface of the water. Record the number of milliliters for this final water level.
4. Find the amount of water displaced by subtracting the initial level of the water from the final level. This volume equals the volume of the cylinder in cm^3 .
5. Record this volume in the chart on the activity sheet.
6. Remove the sample by pouring the water back into your cup and taking the sample out of your graduated cylinder.



Density

7. Calculate the density using the formula $D = m/v$. Record the density in (g/cm^3) .
8. Trade samples with other groups until you have measured the volume and calculated the density of all five samples.

Sample	Initial water level (mL)	Final water level (mL)	Volume of the rods (cm^3)	Mass (g)	Density (g/cm^3)
A				15.0	
B				15.0	
C				15.0	
D				15.0	
E				15.0	

Identify the samples

9. Compare the values for density you calculated to the values in the chart. Then write the letter name for each sample in the chart.

Note: The densities students calculate may not be exactly the same as the given densities in the chart. As students are working, check their values for volume to be sure that they are using the difference between the final and initial water levels, not just the final level.

Material	Approximate density (g/cm ³)	Sample (Letters A–E)
Brass	8.8	
Aluminum	2.7	
PVC	1.4	
Nylon	1.2	
Polyethylene	0.94	

4. Discuss whether students' values for density support their predictions from the beginning of the lesson.

Discuss student values for density for each of the samples. Point out that different groups may have different values for density, but that most of the values are close to the values in the chart.

Ask students:

- **Each group measured the volume of the same samples. What are some reasons that groups might have different values for density?**

Students should realize that small inaccuracies in measuring volume can account for differences in density values. Another reason is that the graduated cylinder, itself, is not perfect. So there is always some uncertainty in measuring.

Remind students that in the beginning of the lesson they made a prediction about the density of the small, medium, and long sample. Students should have predicted that the longest cylinder has the lowest density, the shortest cylinder has the highest density, and the middle is somewhere in between.

Ask students:

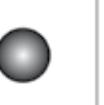
- **Was your prediction about the density of these three samples correct?**
Have students look at their chart with the values for mass, volume, and density for each cylinder. Have them look for a relationship between the volume and the density. Students should realize that the shortest cylinder has the greatest density and the longest cylinder has the lowest density.
- **Is it fair to say that if two samples have the same mass that the one with the larger volume will have a lower density? Yes. Why?** Because the samples have the same mass, their volumes will give you an idea about their densities according to the equation $D = m/v$. If a larger number for volume is in the denominator, the density will be lower.
- **Is it fair to say that the one with the smaller volume will have a higher density? Yes. Why?** If a smaller number for volume is in the denominator, the density will be higher.

EXPLAIN

5. Have students look at the size and mass of atoms to help explain why each sample has a different density.

Project the image *Atomic Size and Mass*.

www.middleschoolchemistry.com/multimedia/chapter3/chapter2#atomic_size_and_mass

ATOMIC SIZE & MASS ELEMENTS 1-20							
HYDROGEN 1  1.01							HELIUM 2  4.00
LITHIUM 3  6.94	BERYLLIUM 4  9.01	BORON 5  10.81	CARBON 6  12.01	NITROGEN 7  14.01	OXYGEN 8  16.00	FLUORINE 9  19.00	NEON 10  20.18
SODIUM 11  22.99	MAGNESIUM 12  24.31	ALUMINUM 13  26.98	SILICON 14  28.09	PHOSPHORUS 15  30.97	SULFUR 16  32.07	CHLORINE 17  35.45	ARGON 18  39.95
POTASSIUM 19  39.10	CALCIUM 20  40.08						

Tell students that this chart is based on the periodic table of the elements but that it only includes the first 20 elements out of about 100. A representation of an atom for each element is shown. For each element, the atomic number is above the atom and the atomic mass is below. This chart is special because it shows both the size and mass of atoms compared to other atoms.

Note: Students may want to know more about why atoms have different atomic numbers and different sizes. These questions will be covered in later chapters but you can tell them that the atomic number is the number of protons in the center or nucleus of the atom. Each element has a certain number of protons in its atoms, so each element has a different atomic number. The dif-

ference in size is a little harder to explain. Atoms have positively charged protons in the nucleus and negatively charged electrons moving around the nucleus. It's really the space the electrons occupy that makes up most of the size of the atom. As the number of protons in the atom increases, both its mass and the strength of its positive charge increases. This extra positive charge pulls electrons closer to the nucleus, making the atom smaller. The atoms get bigger again in the next row because more electrons are added in a space (energy level) further from the nucleus.

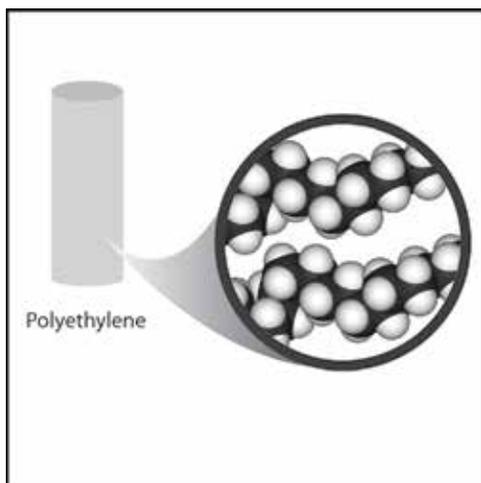
Let students know that they will learn more about the periodic table and atoms in Chapter 4. For now, all students need to focus on is the size and mass of the atoms.

Tell students that the difference in density between the small, medium, and large samples that they measured can be explained based on the atoms and molecules they are made from.

Project the image *Polyethylene (longest rod)*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson2#polyethylene

Polyethylene is made of long molecules of only *carbon* and *hydrogen* atoms. In the *Atomic Size and Mass* chart, the mass of carbon is pretty low, and the mass of hydrogen is the lowest of all the atoms. These low masses help explain why polyethylene has a low density. Another reason is that these long, skinny molecules are loosely packed together.

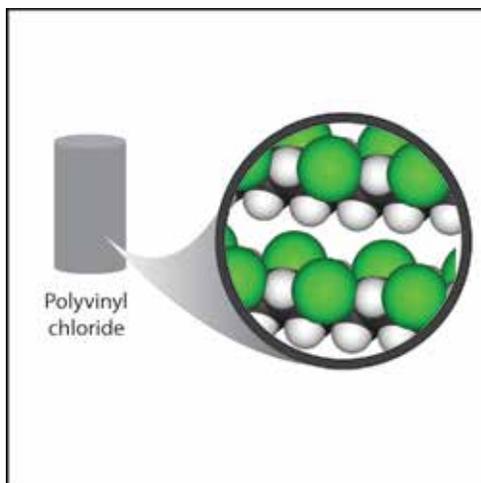


Project the image *Polyvinyl Chloride (medium-length rod)*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson2#polyvinyl_chloride

Polyvinyl chloride is made up of *carbon*, *hydrogen*, and *chlorine* atoms. If you compare polyvinyl chloride to polyethylene, you will notice that there are chlorine atoms in some places where there are hydrogen atoms in the polyethylene. In the chart, chlorine has a

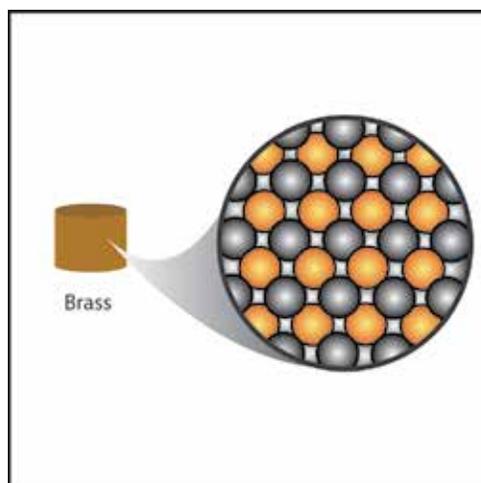
large mass for its size. This helps make polyvinyl chloride more dense than polyethylene. The density of different plastics is usually caused by the different atoms that can be connected to the carbon–hydrogen chains. If they are heavy atoms for their size, the plastic tends to be more dense; if they are light for their size, the plastic tends to be less dense.



Project the image *Brass* (shortest rod).

www.middleschoolchemistry.com/multimedia/chapter3/lesson2#brass

Brass is a combination of *copper* and *zinc* atoms. Copper and zinc come up later in the periodic table, so they are not shown in the chart, but they are both heavy for their size. The atoms are also packed very closely together. For these reasons, brass is more dense than either polyethylene or polyvinyl chloride.

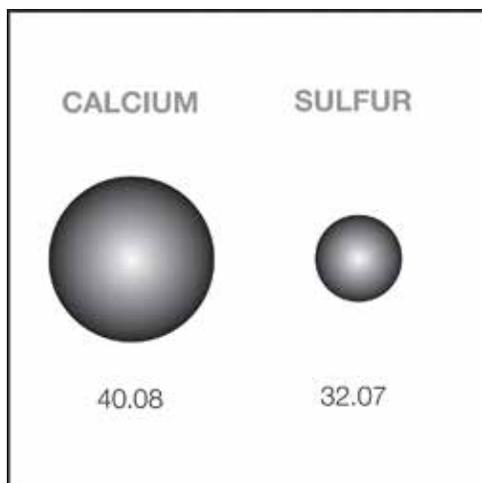


EXTEND

6. Discuss the density of calcium compared to the density of sulfur.

Have students refer to the illustration of Calcium and Sulfur on their activity sheets.

Explain that a calcium atom is both bigger and heavier than a sulfur atom. But a piece of solid sulfur is more dense than a solid piece of calcium. The density of sulfur is about 2 g/cm^3 and the density of calcium is about 1.5 g/cm^3 .



Ask students:

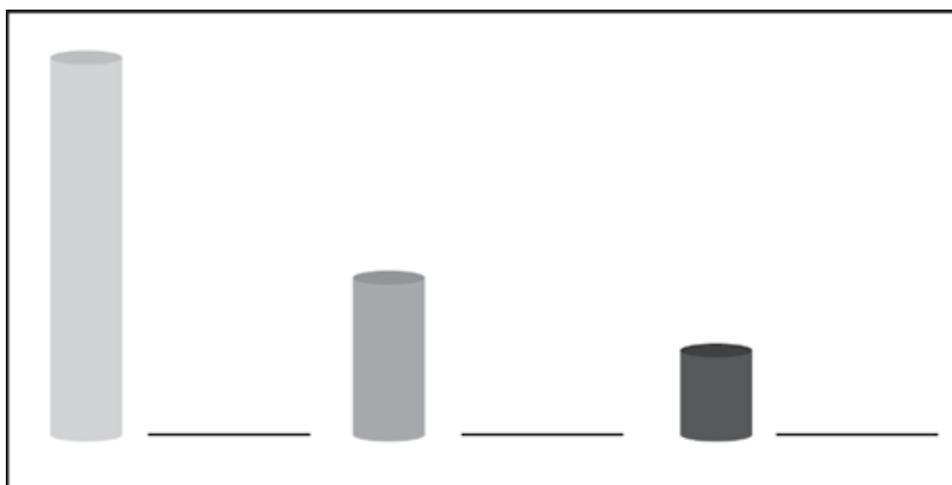
- **Based on what you know about the size, mass, and arrangement of atoms, explain why a sample of sulfur is more dense than a sample of calcium.**

Even though a sulfur atom has less mass than a calcium atom, many more sulfur atoms can pack together in a certain amount of space. This gives sulfur more mass per volume than calcium, making it more dense.

DEMONSTRATION

Think about the longest, middle-sized, and shortest rods your teacher showed you. All of these samples have the same mass, but their volumes are different.

1. Predict the densities of each sample by writing a phrase from the box on the line next to each sample.

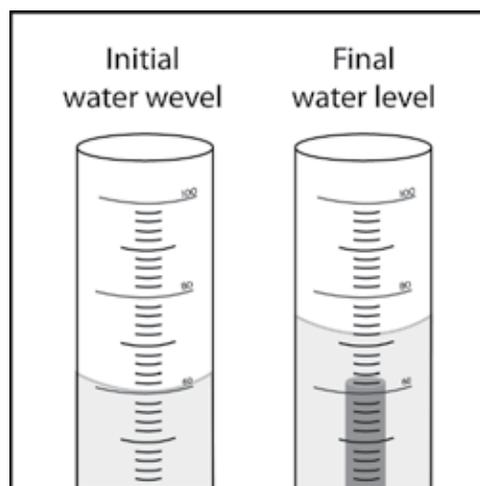


Least dense Medium dense Most dense

2. Explain why you think each rod is either the most, medium, or least dense.

3. The animation showed you how to find the volume of a sample using the water displacement method.

Look at the illustrations showing the water level in a graduated cylinder before and after a sample is submerged in water. What does this difference in water level tell you about the sample?



How much would the water level rise if you submerged a cube with a volume of 1 cm^3 in a graduated cylinder filled with 40 mL of water?

4. What is the density of the sample described below?
Be sure to write the units in g/cm^3 .
- Water level rose from 60 mL to 85 mL
 - Mass = 50 g

ACTIVITY

Your group will work with five rods each with the same mass, but made of a different material. Carefully measure the volume of each sample and calculate the density. Then use density to correctly identify each of the five samples.

Question to investigate

Can you use density to identify all five rods?



Materials for each group

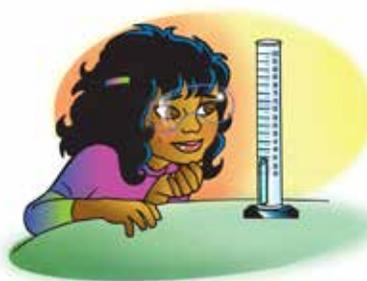
- Set of five different rods that all have the same mass
- Graduated cylinder, 100 mL
- Water in a cup
- Calculator

Hint: the volume of the smallest rod is between 1.5–2.0 cm³.

Procedure

Volume

1. Pour enough water from your cup into the graduated cylinder to reach a height that will cover the sample. Read and record the volume.
2. Slightly tilt the graduated cylinder and carefully place the sample into the water.
3. Place the graduated cylinder upright on the table and look at the level of the water. If the sample floats, use a pencil to gently push the top of the sample just under the surface of the water. Record the number of milliliters for this final water level.
4. Find the amount of water displaced by subtracting the initial level of the water from final level. This volume equals the volume of cylinder in cm³.
5. Record this volume in the chart on the activity sheet.
6. Remove the sample by pouring the water back into your cup and taking the sample out your graduated cylinder.



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the

of

Density

- Calculate the density using the formula $D = m/v$. Record the density in (g/cm^3).
- Trade samples with other groups until you have measured the volume and calculated the density of all five samples.

Sample	Initial water level (mL)	Final water level (mL)	Volume of the rods (cm^3)	Mass (g)	Density (g/cm^3)
A				15.0	
B				15.0	
C				15.0	
D				15.0	
E				15.0	

Identify the samples

- Compare the values for density you calculated to the values in the chart. Then write the letter name for each sample in the chart.

Material	Approximate density (g/cm^3)	Sample (Letters A–E)
Brass	8.8	
Aluminum	2.7	
PVC	1.4	
Nylon	1.2	
Polyethylene	0.94	

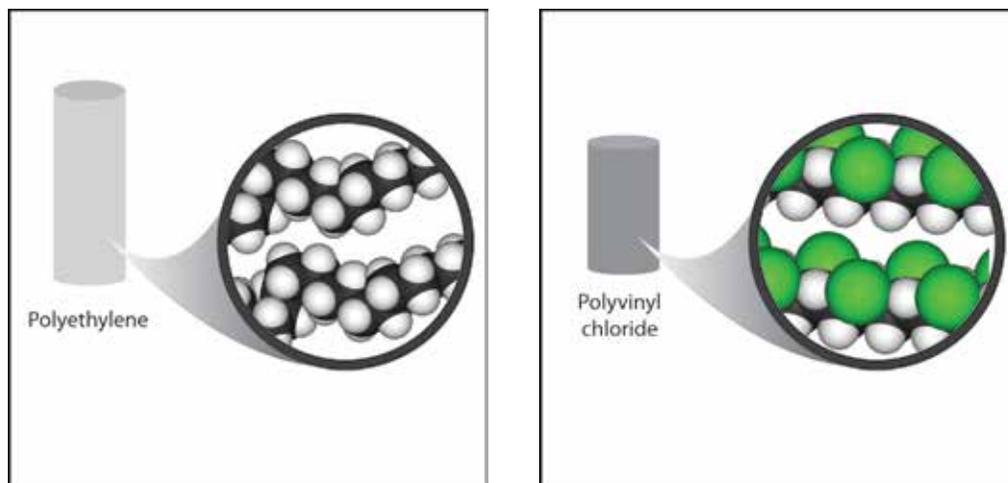
- On the first page of this activity sheet, you made a prediction about the density of a small, medium, and long rod. Based on your calculations for density in your chart, were your predictions correct? If a short rod and a long rod have the same mass, explain why the short one will be more dense than the long one.

EXPLAIN IT WITH ATOMS & MOLECULES

The difference in density between the small, medium, and large rods can be explained based on the atoms and molecules they are made from. Refer to the chart of atomic size and mass to answer the following question about each substance.

6. Polyethylene is made of carbon and hydrogen atoms. Polyvinyl chloride is also made of carbon and hydrogen atoms, but also has chlorine atoms.

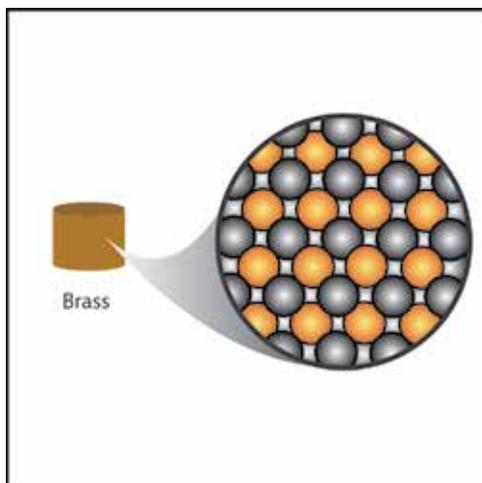
Look at the size and mass of these atoms in the chart to explain why polyvinyl chloride is more dense than polyethylene.



ATOMIC SIZE & MASS ELEMENTS 1-20

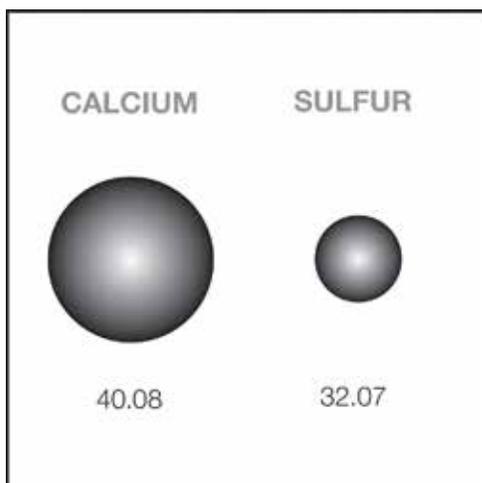
<p>HYDROGEN 1</p>  <p>1.01</p>							<p>HELIUM 2</p>  <p>4.00</p>
<p>LITHIUM 3</p>  <p>6.94</p>	<p>BERYLLIUM 4</p>  <p>9.01</p>	<p>BORON 5</p>  <p>10.81</p>	<p>CARBON 6</p>  <p>12.01</p>	<p>NITROGEN 7</p>  <p>14.01</p>	<p>OXYGEN 8</p>  <p>16.00</p>	<p>FLUORINE 9</p>  <p>19.00</p>	<p>NEON 10</p>  <p>20.18</p>
<p>SODIUM 11</p>  <p>22.99</p>	<p>MAGNESIUM 12</p>  <p>24.31</p>	<p>ALUMINUM 13</p>  <p>26.98</p>	<p>SILICON 14</p>  <p>28.09</p>	<p>PHOSPHORUS 15</p>  <p>30.97</p>	<p>SULFUR 16</p>  <p>32.07</p>	<p>CHLORINE 17</p>  <p>35.45</p>	<p>ARGON 18</p>  <p>39.95</p>
<p>POTASSIUM 19</p>  <p>39.10</p>	<p>CALCIUM 20</p>  <p>40.08</p>						

7. Brass is made of copper and zinc atoms. These atoms are pretty heavy for their size, but they are also packed together differently than the molecules of the plastics. How does the way these atoms pack together help make the brass more dense than the plastics?



TAKE IT FURTHER

8. Based on the *Atomic Size and Mass* chart, a calcium atom is both bigger and heavier than a sulfur atom. But a piece of solid sulfur is more dense than a solid piece of calcium. In fact, sulfur is about 2 g/cm^3 , and calcium is about 1.5 g/cm^3 .



Based on what you know about the size, mass, and arrangement of atoms, explain why a sample of sulfur is more dense than a sample of calcium.

Chapter 3, Lesson 3: Density of Water

Key Concepts

- Just like solids, liquids also have their own characteristic density.
- The volume of a liquid can be measured directly with a graduated cylinder.
- The molecules of different liquids have different size and mass.
- The mass and size of the molecules in a liquid and how closely they are packed together determine the density of the liquid.
- Just like a solid, the density of a liquid equals the mass of the liquid divided by its volume; $D = m/v$.
- The density of water is 1 gram per cubic centimeter.
- The density of a substance is the same regardless of the size of the sample.

Summary

Students measure the volume and mass of water to determine its density. Then they measure the mass of different volumes of water and discover that the density is always the same. Students make a graph of the relationship between the volume and the mass of water.

Objective

Students will be able to measure the volume and mass of water and calculate its density. Students will be able to explain that since any volume of water always has the same density, at a given temperature, that density is a characteristic property of water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- Graduated cylinder, 100 ml
- Water
- Balance that measures in grams (able to measure over 100 g)
- Dropper

Materials for the Demonstration

- Water
- Two identical buckets or large containers

ENGAGE

1. Do a demonstration to introduce the idea that water has density.

Materials

- Water
- Two identical buckets or large containers



Teacher preparation

- Half-fill one bucket and add only about 1 cup of water to the other.

Procedure

1. Select a student to lift both buckets of water.
2. Ask the student volunteer which bucket appears to have more mass.

Expected results

The bucket containing more water has more mass.

Ask students:

- **In Lessons 1 and 2, you found the density of solids, by measuring their mass and volume. Do you think a liquid, like water can have a density?**

Students should realize that water has volume and mass. Because $D=m/v$, water must also have density.

- **How do you think you can find the density of a liquid like water?**

Students are not expected to be able to fully answer this question at this point. It is meant as a lead-in to the investigation. But students may realize that they should somehow find the mass and volume of the water first.

- **Could both the small and large amounts of water your classmate lifted have the same density?**

Students may point out that the bucket with more water has more mass but a greater volume. The bucket with less mass has less volume. So it is possible that different amounts of water could have the same density.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Discuss with students how to find the volume and mass of water.

Tell students that they are going to try to find the density of water.

Ask students:

- **What two things do you need to know in order to find the density of water?**

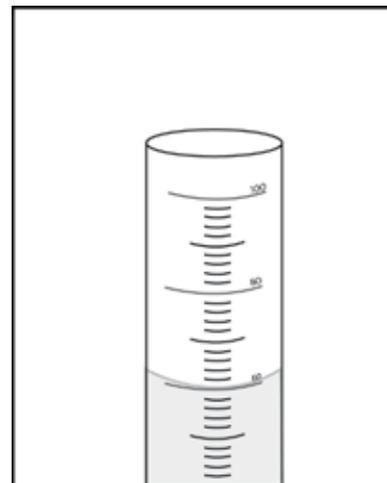
Students should realize that they need both the volume and mass of a sample of water to find its density.

- **How can you measure a *volume* of water?**

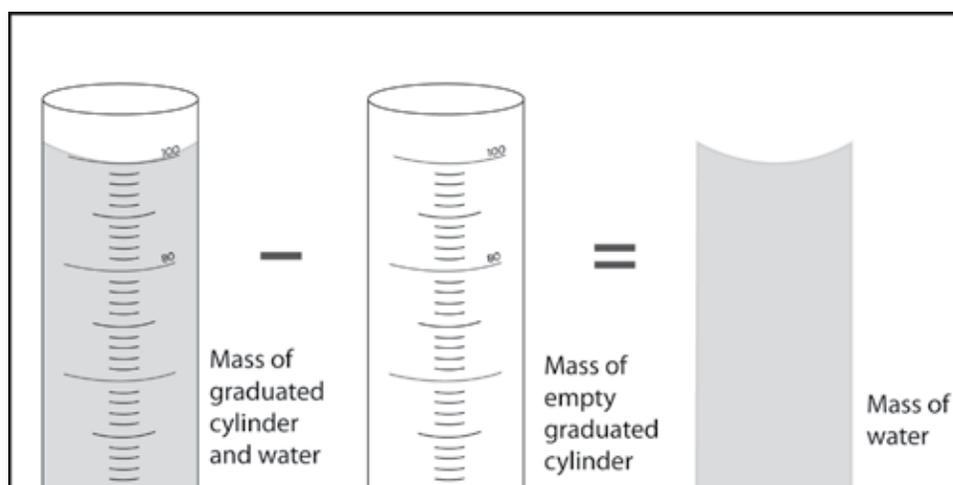
Suggest that students use a graduated cylinder to measure volume in milliliters. Remind students that each milliliter equals 1 cm^3 .

- **How can you measure the *mass* of water?**

Suggest that students use a balance to measure the mass in grams. Tell students that they can find mass by weighing the water. However, since water is a liquid, it needs to be in some sort of container. So in order to weigh the water, they have to weigh the container, too. Explain to students that they will have to subtract the mass of an empty graduated cylinder from the mass of the cylinder and water to get the mass of just the water.



Measuring the mass of water



3. Have students find the mass of different volumes of water to show that the density of water does not depend on the size of the sample.

Question to investigate

Do different amounts of water have the same density?

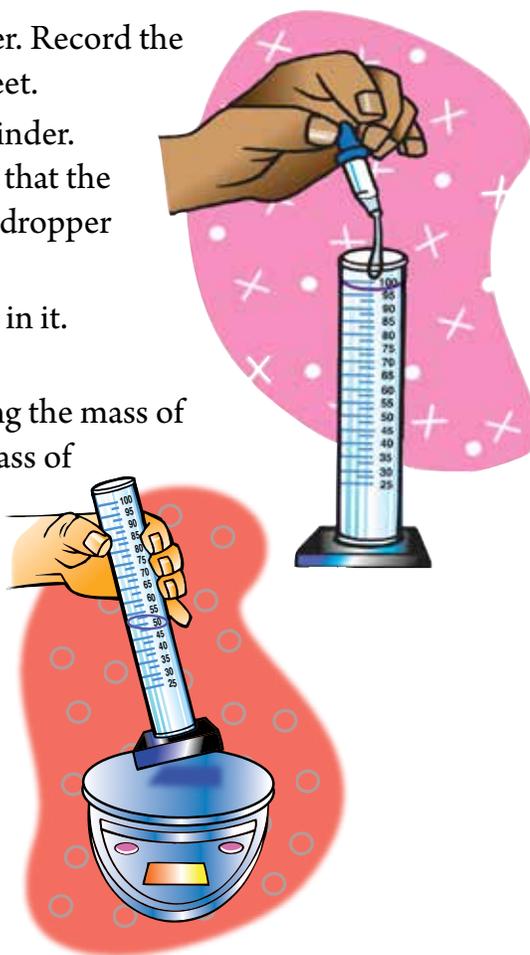


Materials for each group

- Graduated cylinder, 100 ml
- Water
- Balance that measures in grams (able to measure over 100 g)
- Dropper

Procedure

1. Find the mass of an empty graduated cylinder. Record the mass in grams in the chart on the activity sheet.
2. Pour 100 mL of water into the graduated cylinder. Try to be as accurate as possible by checking that the meniscus is right at the 100-mL mark. Use a dropper to add or remove small amounts of water.
3. Weigh the graduated cylinder with the water in it. Record the mass in grams.
4. Find the mass of only the water by subtracting the mass of the empty graduated cylinder. Record the mass of 100 mL of water in the chart.
5. Use the mass and volume of the water to calculate density. Record the density in g/cm^3 in the chart.
6. Pour off water until you have 50 mL of water in the graduated cylinder. If you accidentally pour out a little too much, add water until you get as close as you can to 50 mL.
7. Find the mass of 50 mL of water. Record the mass in the activity sheet. Calculate and record the density.
8. Next, pour off water until you have 25 mL of water in the graduated cylinder. Find the mass of 25 mL of water and record it in the chart. Calculate and record the density.



Finding the density of different volumes of water			
Volume of water	100 milliliters	50 milliliters	25 milliliters
Mass of graduated cylinder + water (g)			
Mass of empty graduated cylinder (g)			
Mass of water (g)			
Density of water (g/cm ³)			

Expected results

The density of water should be close to 1 g/cm³. This is true for 100, 50, or 25 mL.

Ask students:

- **Look at your values for density in your chart. Does the density of the different volumes of water seem to be about the same?**

Help students see that most of the different values for density are near 1 g/cm³. They may wonder why their values are not all exactly 1 g/cm³. One reason could be inaccuracies in measuring. Another reason is that the density of water changes with temperature. Water is most dense at 4 °C and at that temperature has a density of 1 g/cm³. At room temperature, around 20–25 °C, the density is about 0.99 g/cm³.

- **What is the density of water in g/cm³?**

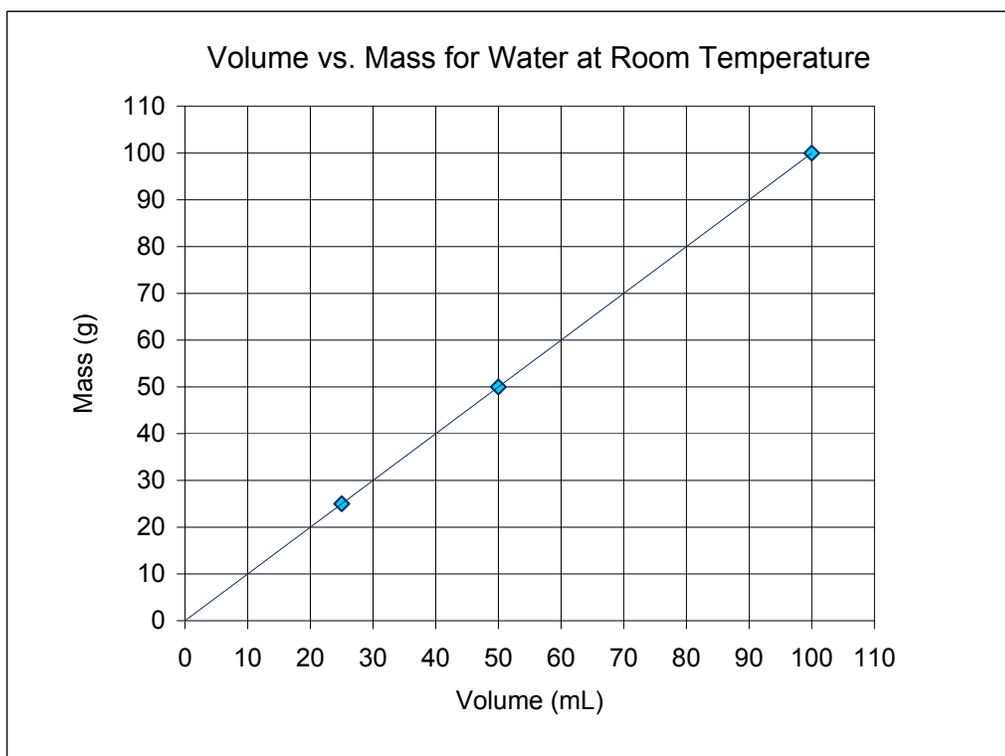
Students answers will vary, but their values should mostly be around 1 g/cm³.



4. Have students graph their results.

Help students make a graph of the data on their activity sheet. The x-axis should be volume and the y-axis should be mass.

When students plot their data, there should be a straight line showing that as volume increases, mass increases by the same amount.



5. Discuss student observations, data, and graphs.

Ask students:

- **Use your graph to find the mass of 40 mL of water. What is the density of this volume of water?**
The mass of 40 mL of water is 40 grams. Since $D = m/v$ and $\text{mL} = \text{cm}^3$, the density of water is 1 g/cm^3 .
- **Choose a volume between 1 and 100 mL. Use your graph to find the mass. What is the density of this volume of water?**
Whether students weigh 100, 50, 25 mL or any other amount, the density of water will always be 1 g/cm^3 .

Tell students that density is a characteristic property of a substance. This means that the density of a substance is the same regardless of the size of the sample.

Ask students:

- **Is density a characteristic property of water? How do you know?**
Density is a characteristic property of water because the density of any sample of water (at the same temperature) is always the same. The density is 1 g/cm^3 .

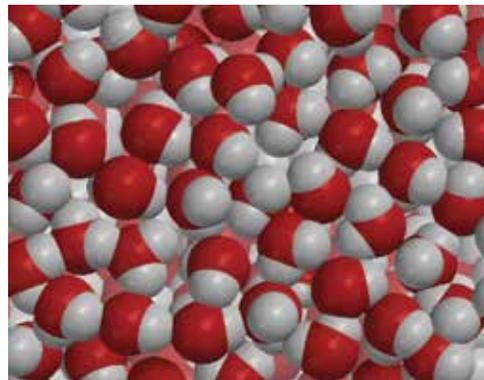
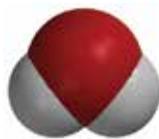
EXPLAIN

6. Explain why the density of any size sample of water is always the same.

Project the image *Density of Water*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson3#density_of_water

Water molecules all have the same mass and size. Water molecules are also packed pretty close together. They are packed the same way throughout an entire sample of water. So, if a volume of water has a certain mass, twice the volume will have twice the mass, three times the volume has three times the mass, etc. No matter what size sample of water you measure, the relationship between the mass and volume will always be the same. Because $D=m/v$, the density is the same for any amount of water.



Project the animation *Liquid Water*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson3#liquid_water

Water molecules are always moving. But on the average they are packed the same throughout. Therefore, the ratio between the mass and volume is the same, making the density the same. This is true no matter the size of the sample or where you select your sample from.

EXTEND

7. Have students consider whether the density of a large piece of a solid substance is the same as the density of a smaller piece.



Give students time to calculate the density of each of the three samples drawn on their activity sheet and answer the related questions.

Ask students:

- **The density of a liquid is the same no matter what the size of the sample. Could this be true for solids, too? Calculate the density of each of the three samples to find out.**

Yes. The density of a solid substance is the same no matter how big or small the sample.

- **Sample A has a mass of 200 g. What is the density of Sample A?**

$$D = m/v$$

$$D = 200\text{g}/100\text{cm}^3$$

$$D = 2 \text{ g/cm}^3$$

- **If you cut Sample A in half and looked at only one half, you would have Sample B. What is the density of Sample B?**

If students do not know what the mass is, tell them that it is half the mass of Sample A. Because Sample A was 200 g, Sample B is one half the mass of Sample A. Because Sample A was 200 g, Sample B is one half the mass (100 g).

$$D = m/v$$

$$D = 100\text{g}/50 \text{ cm}^3$$

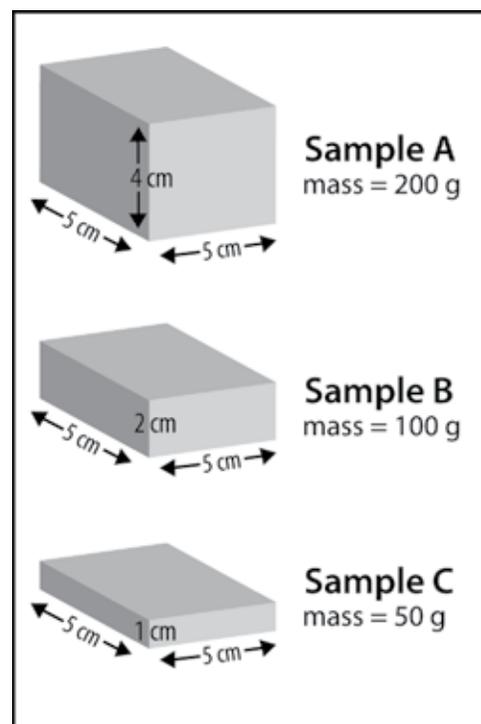
$$D = 2 \text{ g/cm}^3$$

- **If you cut Sample B in half you would have Sample C. What is the density of Sample C?**

$$D = m/v$$

$$D = 50\text{g}/25 \text{ cm}^3$$

$$D = 2 \text{ g/cm}^3$$



DEMONSTRATION

1. One of your classmates lifted different amounts of water. The largest amount of water also had the most mass.

You know how to find the density of solids. Do you think a liquid, like water, can have a density?

How do you think you could find the density of a liquid like water?

Could both the small and large amounts of water your classmate lifted have the same density?

Explain.

ACTIVITY

Question to investigate

Do different amounts of water have the same density?



Materials for each group

- Graduated cylinder, 100 mL
- Water
- Balance that measures in grams (able to measure over 100 g)
- Dropper

Procedure

1. Find the mass of an empty graduated cylinder. Record the mass in grams in the chart on the activity sheet.
2. Pour 100 mL of water into the graduated cylinder. Try to be as accurate as possible by checking that the meniscus is right at the 100 mL mark. Use a dropper to add or remove small amounts of water.

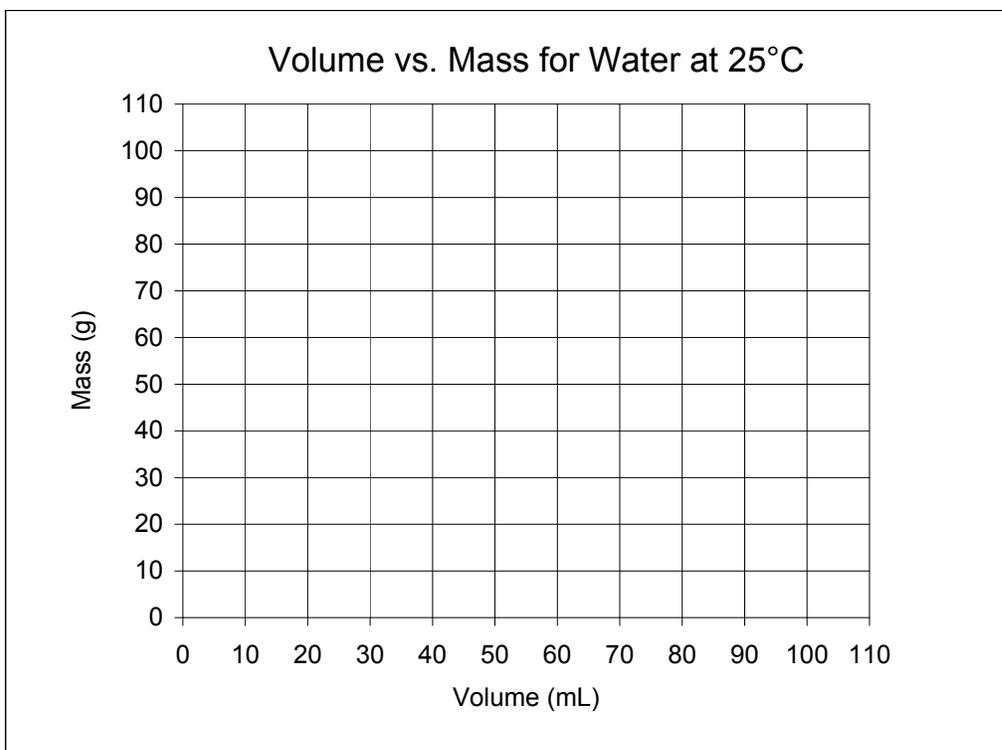
3. Weigh the graduated cylinder with the water in it. Record the mass in grams.
4. Find the mass of only the water by subtracting the mass of the empty graduated cylinder. Record the mass of 100 mL of water in the chart.
5. Use the mass and volume of the water to calculate density. Record the density in g/cm^3 in the chart.
6. Pour off water until you have 50 mL of water in the graduated cylinder. If you accidentally pour out a little too much, add water until you get as close as you can to 50 mL.
7. Find the mass of 50 mL of water. Record the mass in the activity sheet. Calculate and record the density.
8. Next, pour off water until you have 25 mL of water in the graduated cylinder. Find the mass of 25 mL of water and record it in the chart. Calculate and record the density.

Finding the density of different volumes of water			
Volume of water	100 milliliters	50 milliliters	25 milliliters
Mass of graduated cylinder + water (g)			
Mass of empty graduated cylinder (g)			
Mass of water (g)			
Density of water (g/cm^3)			

2. Look at your values for density in your chart. Does the density of the different volumes of water seem to be about the same?
3. What do you think is the density of water in g/cm^3 ?

4. Using the data from your chart, graph the volume and mass for 100 mL, 50 mL, and 25 mL.

Volume vs. Mass for Water at Room Temperature



5. Look at the graph you made.
If you measured 40 milliliters of water, what do you think its mass would be?
What would its density be?

Volume: 40 mL

Mass _____

Density _____

6. Choose any volume of water between 1 and 100 milliliters. Based on the graph, what would its mass be? What would its density be?

Volume _____

Mass _____

Density _____

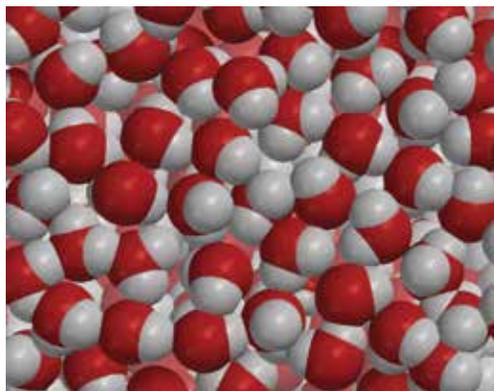
7. Density is a “characteristic property” of a substance. This means that the substance will have the same density no matter how big or small the sample is. Would you say that density is a characteristic property of water? Why or why not?

EXPLAIN IT WITH ATOMS & MOLECULES

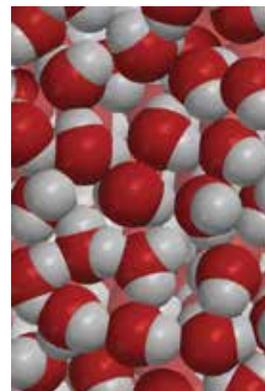
Each individual molecule has the same size and mass. The water molecules are packed very close together the same way throughout an entire sample of water.

8. Sample B is half the volume of Sample A.

Do the samples have the same mass?



Sample A



Sample B

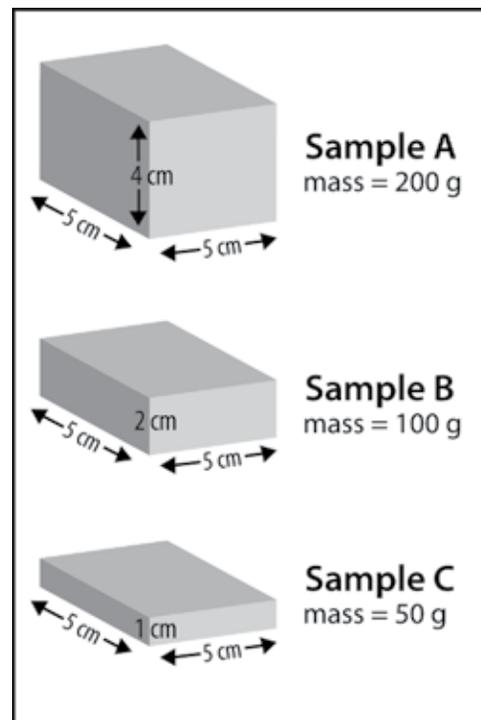
Do the samples have the same density?

TAKE IT FURTHER

9. The density of a liquid is the same no matter the size of the sample. Could this be true for solids, too? Calculate the density of each of the three samples to find out.

Sample A has a mass of 200 g. What is the density of Sample A?

If you cut Sample A in half and looked at only one half you would have Sample B. What is the density of this Sample B?



If you cut Sample B in half you would have Sample C. What is the density of Sample C?

Chapter 3, Lesson 4: Density: Sink and Float for Solids

Key Concepts

- The density of an object determines whether it will float or sink in another substance.
- An object will float if it is less dense than the liquid it is placed in.
- An object will sink if it is more dense than the liquid it is placed in.

Summary

Students will investigate a wax candle and a piece of clay to understand why the candle floats and the clay sinks even though the candle is heavier than the piece of clay. Students will discover that it is not the weight of the object, but its density compared to the density of water, that determines whether an object will sink or float in water.

Objective

Students will be able to determine whether an object will sink or float by comparing its density to the density of water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles.

Materials for Each Group

- 2 tea light candles in their metal containers
- Clay
- Water in cup
- Small balance
- Tape
- Dropper

Notes About the Materials

A simple balance is required for the demonstration. One of the least expensive is Delta Education, Stackable Balance (21-inch) Product # 020-0452-595. Students can use the smaller version of the same balance, Delta Education, Stackable Balance (12-inch), Product # 023-0724-595. You will need tea light candles for the demonstration and for each student group. Look for candles in which the wax completely fills the metal container.

ENGAGE

1. Do a demonstration to show that the wax is heavier than the clay but that the wax floats and the clay sinks.



Materials for the demonstration

- 1 tea light candle
- Clay
- Clear plastic container
- Water
- Large balance

Teacher preparation

- Use a small enough piece of clay so that you are sure that the candle weighs more than the clay.
- Pour water into a clear plastic container (or large cup) until it is about $\frac{1}{2}$ -full.



Procedure

1. Place a piece of clay that weighs less than a tea light candle on one end of a balance.
2. Remove the candle from its metal container and place the candle on the other end of the balance.
3. Ask students which is heavier, the clay or the candle. Ask them to predict which will sink and which will float. Then, place the clay and candle in a clear container of water.



Expected results

Even though the candle weighs more than the clay, the candle floats and the clay sinks.

2. Discuss students' ideas about why the heavier candle floats and the lighter clay sinks.

Ask students:

- **Why do you think the bigger, heavier candle floats and the smaller, lighter clay sinks?**

Students may say that the clay is more dense than wax. Tell students that the density of the objects is important, but that in sinking and floating, the density of the *water* matters, too. In fact, by comparing the density of an object to the density of water, you can predict whether an object will sink or float. One way to compare them is to weigh equal volumes of each.

Give an activity sheet to each student.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

3. Have students compare the density of water, wax, and clay.

Question to investigate

Why does a heavier candle float and a lighter piece of clay sink?

Materials for each group

- 2 tea light candles in their metal containers
- Clay
- Water in cup
- Small balance
- Tape
- Dropper

Procedure

Compare the density of wax and water

1. Roll two pieces of tape and stick them to the center of the pan at each end of the balance.
2. Attach each tea light candle to the tape so that each candle is in the center of the pan.
3. Use the wick to pull one candle out of its container.
4. Carefully pour water into the empty metal container until it fills the container to the same level as the candle in the other container. You may use a dropper to add the last bit of water and prevent spilling. The goal is to compare the mass of equal volumes of wax and water.



Expected results

The water has a greater mass than an equal volume of wax. So, the density of water must be greater than the density of wax.

Ask students:

- **Which weighs more, wax or an equal volume of water?**
Water weighs more than an equal volume of wax.
- **Which is more dense, wax or water?**
Water is more dense.

If students have trouble understanding this relationship between the mass and density of equal volumes, have them think about the demonstration from Chapter 3, Lesson 1 with the aluminum and copper cubes. Both had the same volume, but the copper cube weighed more. Because the copper had more mass, it also had a greater density.

Compare the density of clay and water

1. Make sure you have one piece of tape in the center of each pan on the balance.
2. Fill one container with clay and place it on the tape so that it is in the center of the pan.
3. Place an empty container on the tape at the opposite end of the balance.
4. Slowly and carefully add water to the empty container until it is full.



Expected results

The clay has a greater mass than an equal volume of water. So, the density of clay is greater than the density of water.

Ask students:

- **Which weighs more, the clay or an equal volume of water?**
The clay weighs more than an equal volume of water.
- **Which is more dense, clay or water?**
Clay is more dense.
- **Knowing the density of an object can help you predict if it will sink or float in water.**
 - **If an object is more dense than water, would you expect it to sink or float?**
Objects that are more dense than water sink.
 - **If an object is less dense than water, would you expect it to sink or float?**
Objects that are less dense than water float.

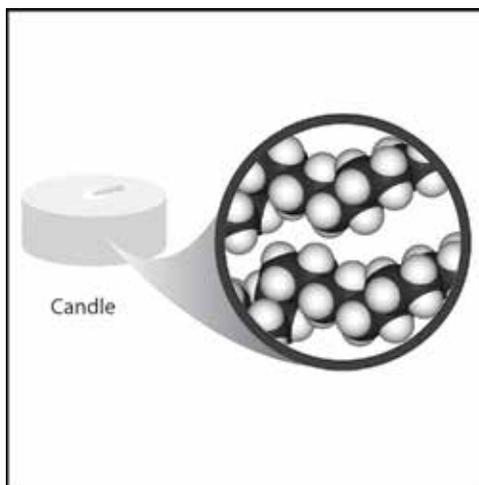
EXPLAIN

4. Compare the density of wax, water, and clay on the molecular level.

Project the image Wax.

www.middleschoolchemistry.com/multimedia/chapter3/lesson4#wax

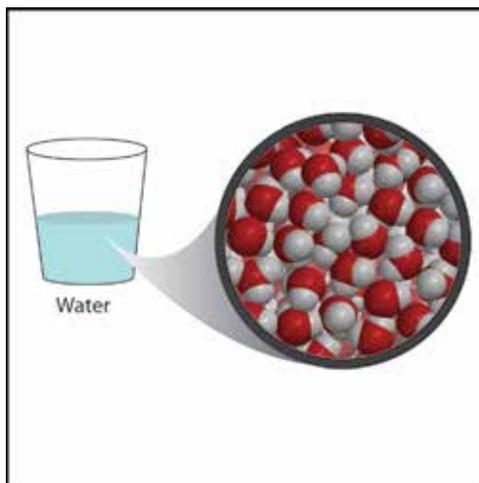
Wax is made of carbon and hydrogen atoms connected together in long chains. These long chains are tangled and intertwined and packed together to make the wax.



Project the image Water.

www.middleschoolchemistry.com/multimedia/chapter3/lesson4#water

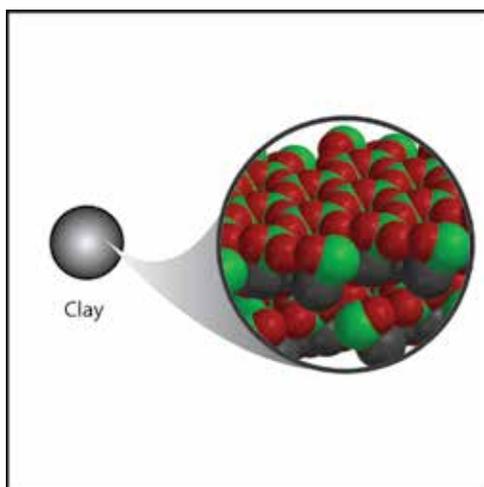
Even though they both have lots of hydrogen atoms, water is more dense than wax because the oxygen in water is heavier and smaller than the carbon in the wax. Also, the long chains of the wax do not pack as efficiently as the small water molecules.



Project the image *Clay*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson4#clay

Clay has oxygen atoms like water, but it also has heavier atoms like silicon and aluminum. The oxygen atoms are bonded to the silicon and aluminum to make molecules with a lot of mass. These are packed closely together, which makes the clay more dense than water.



EXTEND

5. Have students explain, in terms of density, why a very heavy object like a big log floats and why a very light object like a tiny grain of sand sinks.

Ask students:

- A giant log can float on a lake, while a tiny grain of sand sinks to the bottom. Explain why a heavy object like the log floats while a very light grain of sand sinks.

Students should recognize that a log will float because wood is less dense than water. If you could weigh a large amount of water that has the same volume as the log, the log will weigh less than the water. Therefore, the log floats. A grain of sand will sink because sand is more dense than water. If you could weigh a small amount of water that has the same volume as the grain of sand, the sand will weigh more than the water. Therefore, the sand sinks.

Students should realize that if an object weighs more than an equal volume of water, it is more dense and will sink, and if it weighs less than an equal volume of water, it is less dense and will float.

- Remember that the density of water is about 1 g/cm^3 . Predict whether the following objects will sink or float.

Will these objects sink or float?		
Object	Density (g/cm ³)	Sink or float
Cork	0.2–0.3	Float
Anchor	7.8	Sink
Spruce wood oar	0.4	Float
Apple	0.9	Float
Orange	0.84	Float
Orange without peel	1.16	Sink

Ask students:

- **If a peach has a volume of 130 cm³ and sinks in water, what can you say about its mass?**
Its mass must be more than 130 grams.
- **If a banana has a mass of 150 grams and floats in water, what can you say about its volume?**
Its volume must be more than 150 cm³.

Note: Students may wonder why boats made out of dense material like steel can be made to float. This is a good question and there are several ways of answering it. A key to understanding this phenomenon is that the density of the material and the density of an object made of that material are not necessarily the same. If a solid ball or cube of steel is placed in water, it sinks. But if that same steel is pounded and flattened thin and formed into a big bowl-like shape, the overall volume of the bowl is much greater than the volume of the steel cube. The mass of the steel is the same but the big increase in volume makes the density of the bowl less than the density of water so the bowl floats. This is the same reason why a steel ship is able to float. The material is shaped in such a way so that the density of the ship is less than the density of water.

Read more about sinking and floating in the additional teacher background section at the end of this lesson.

ACTIVITY

Question to investigate

Why does a heavier candle float and a lighter piece of clay sink?



Materials for each group

- 2 tea light candles in their metal containers
- Clay
- Water in cup
- Small balance
- Tape
- Dropper

Procedure

Compare the density of wax and water

1. Roll two pieces of tape and stick them to the center of the pan at each end of the balance.
2. Attach each tea light candle to the tape so that each candle is in the center of the pan.
3. Use the wick to pull one candle out of its container.
4. Carefully pour water into the empty metal container until it fills the container to the same level as the candle in the other container. You may use a dropper to add the last bit of water and prevent spilling. The goal is to compare the mass of equal volumes of wax and water.

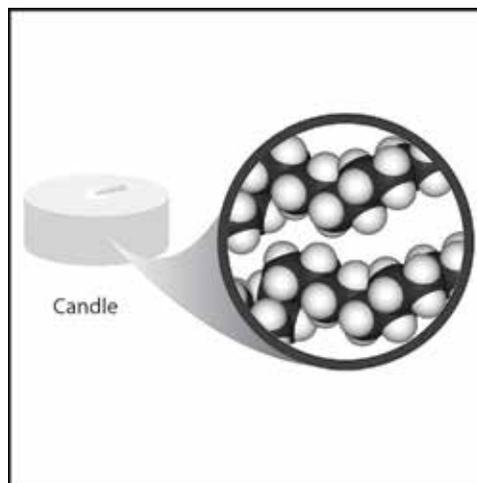
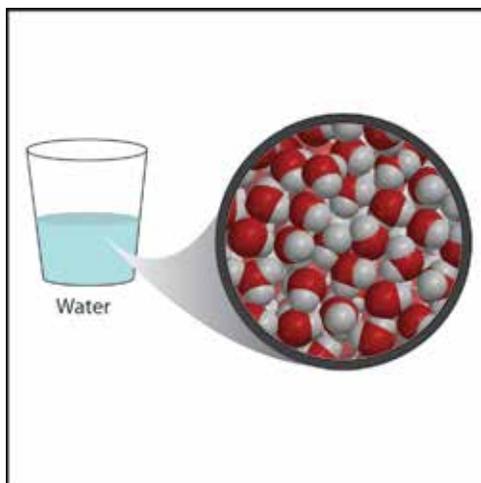


1. Which weighs more, wax or an equal volume of water?

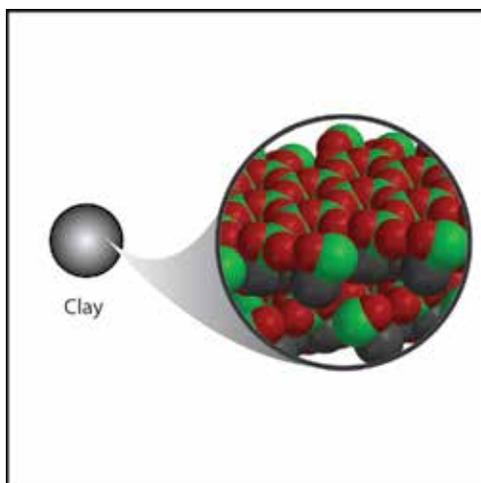
Which is more dense, wax or water?

EXPLAIN IT WITH ATOMS & MOLECULES

4. Water is made up of small molecules of oxygen and hydrogen. Water molecules are closely packed together. Wax is made of carbon and hydrogen atoms connected together in long chains. Explain on the molecular level why wax is less dense than water.



Clay is made of oxygen and heavier atoms such as silicon and aluminum. Explain on the molecular level why clay is more dense than water.



TAKE IT FURTHER

5. A giant log can float on a lake while a tiny grain of sand sinks to the bottom. Explain why a heavy object like the log floats while a very light grain of sand sinks.
6. Remember that the density of water is 1 g/cm^3 . Predict whether the following objects will sink or float.

Will these objects sink or float?		
Object	Density (g/cm^3)	Sink or float
Cork	0.2–0.3	
Anchor	7.8	
Wooden oar	0.4	
Apple	0.9	
Orange	0.84	
Orange without peel	1.16	

7. If a peach has a volume of 130 cm^3 and sinks in water, what can you say about its mass?

8. If a banana has a mass of 150 grams and floats in water, what can you say about its volume?

Additional Teacher Background

Chapter 3, Lesson 4, p. 205

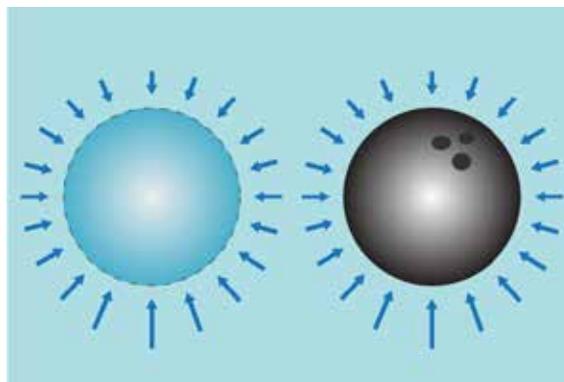
Another way to analyze the floating ship question is based on comparing the mass of the ship to the mass of water it displaces. When an object floats, the weight of the object equals the weight of the water that the object displaces. Even though a ship is very massive, the volume is so huge that the weight of water displaced by the ship is equal to the weight of the ship.

For example, a solid cube of steel displaces a volume of water equal to the volume of the cube. But since steel is more dense than water, the weight of the cube is greater than the weight of the displaced water, and the cube sinks. But if the cube is flattened and shaped into a large enough boat and placed in water, the boat will displace more water. Its weight hasn't changed, but the volume of water it displaces has increased. When the weight of the water displaced by the boat is equal to the weight of the boat, the boat floats.

But why does the weight of the water matter and what *really* is making the boat float? This question requires a more physics-like analysis. There is a force that makes objects float called the *buoyant force*. Some curricula attempt to teach aspects of density using the concept of buoyant force but we do not recommend it for middle school students. Here is a short explanation for you:

An object in a fluid is acted on by the pressure (force /unit area) exerted by the fluid surrounding it. Because the pressure (force) is greater the deeper you go, there is more force beneath the object pushing up than the force above the object pushing down. Therefore, the buoyant force on an object is a net upward force.

In the diagram, an imaginary ball of water (on the left) is suspended or floating in the water surrounding it. Since water has a particular density, this volume of water has a certain mass and weight. The net upward force due to the buoyant force of the surrounding water is strong enough to counteract the weight of the ball of water and keep it suspended or floating in place. This means that the buoyant force is equal to the weight of the ball of water.



But what happens if the ball of water is replaced by an object that is more dense than water, like a bowling ball?

Since the bowling ball is more dense, it weighs more than the ball of water of the same volume. The buoyant force of the surrounding water is not strong enough to support this extra weight and the bowling ball sinks.

If the ball of water is replaced by a ball that is less dense than water, that ball must weigh less than the ball of water. That ball will be pushed up by the buoyant force and will float to the surface.

Chapter 3, Lesson 5: Density: Sink and Float for Liquids

Key Concepts

- Since density is a characteristic property of a substance, each liquid has its own characteristic density.
- The density of a liquid determines whether it will float on or sink in another liquid.
- A liquid will float if it is less dense than the liquid it is placed in.
- A liquid will sink if it is more dense than the liquid it is placed in.

Summary

Students will observe three household liquids stacked on each other and conclude that their densities must be different. They will predict the relative densities of the liquids and then measure their volume and mass to see if their calculations match their observations and predictions.

Objective

Students will be able to determine whether a liquid will sink or float in water by comparing its density to the density of water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles. When using isopropyl alcohol, read and follow all warnings on the label. Isopropyl alcohol is flammable. Keep it away from any flames or spark sources. Have students wash hands after the activity.

Materials for Each Group

- Balance
- Isopropyl alcohol, 70% or higher
- Water
- Graduated cylinder
- 2 identical tall clear plastic cups
- 2 tea light candles

Materials for the Demonstration

- Balance
- Isopropyl alcohol, 70% or higher
- Water
- Graduated cylinder
- 2 identical tall clear plastic cups
- 2 tea light candles

Notes about the Materials

Isopropyl alcohol

The demonstrations and activity work best with 91% isopropyl alcohol solution, which is available in many grocery stores and pharmacies. If you can't find 91% solution, 70% will work, but your candle might not sink in it. If that happens, do not do that demonstration. Although the isopropyl alcohol solution is 91% alcohol and 9% water, you can disregard the small amount of water for the purpose of this lesson.

Balance

A simple balance is all that is required for the second demonstration. One of the least expensive is Delta Education, Stackable Balance (21-inch) Product # 020-0452-595. Students can use the smaller version of the same balance, Delta Education, Stackable Balance (12-inch), Product # 023-0724-595.

ENGAGE

1. Do two demonstrations to show that different liquids have different densities.



Materials

- Balance
- Isopropyl alcohol, 70% or higher
- Water
- Graduated cylinder
- 2 identical tall clear plastic cups
- 2 tea light candles

Teacher preparation

- Use a graduated cylinder to measure 50 mL of water and pour it into a clear plastic cup.
- Measure 50 mL of isopropyl alcohol and pour it into another identical clear plastic cup.

Procedure

Demonstrate the density of two liquids with sinking and floating

1. Place a tea light candle in a cup with water and another tea light candle in a cup with alcohol.
2. Hold up the two cups.



Expected results

The candle will float on water and sink in alcohol.

Ask students:

- **What might be causing one candle to float and the other to sink?**

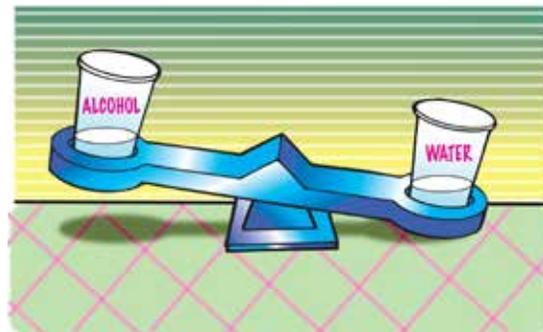
Explain that the two candles are the same. Students should reason that the liquids must be different and have different densities. Explain that the cup with the floating candle contains water and the cup with the sinking candle contains isopropyl alcohol.

- **Do you think these two liquids have the same or different densities?**
Students should conclude that the liquids must have different densities. They may even realize that water is more dense and alcohol is less dense than the wax candle.

Procedure

Demonstrate the density of two liquids by comparing the mass of equal volumes

3. Remove the candles from each liquid and tell students that each cup contains the same volume of liquid.
4. Carefully place the cups of water and alcohol on opposite ends of a balance.



Expected results

The water will weigh more than the alcohol.

Ask students:

- **Which liquid is more dense?**
Students should agree that the water is more dense than the alcohol.
- **How do you know?**
Since the water has more mass than an equal volume of alcohol, water must be more dense.

EXPLORE

2. **Demonstrate that liquids can float or sink in other liquids by making a density column with water, oil, and alcohol.**



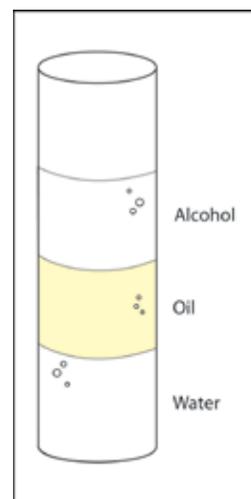
Materials for the demonstration

- Graduated cylinder
- Water
- Vegetable oil
- Isopropyl alcohol, 70% or higher

Note: If you would like the liquids to be more visible, add 1 drop of food coloring to the water and another drop of a different color to the alcohol.

Procedure

1. Pour about 15 mL of water into the graduated cylinder. Gradually add about 15 mL of oil. Then slowly pour about 15 mL of alcohol on top. The liquids should form layers in the graduated cylinder.
2. Show students the layered liquids in the graduated cylinder and point out that the alcohol floats on the oil while the water sinks.



Expected results

Alcohol floats on oil and water sinks in oil. Water, alcohol, and oil layer well because of their densities, but also because the oil layer does not dissolve in either liquid. The oil keeps the water and alcohol separated so that they do not dissolve in one another.

Ask students:

- **Why does the alcohol float on the oil?**
They should conclude that the alcohol floats because it is less dense than the oil.
- **Why does the water sink in the oil?**
Water sinks because it is more dense than oil. Explain that, just like solids, liquids are made from atoms and molecules, which have a certain mass and size. Depending on the mass of the molecules that make up a liquid and how closely they pack together, liquids have their own densities.
- **In the activity, you will compare the mass of equal volumes of each liquid. Which liquid do you think will have the most mass? The least mass? In between?**
Students should predict that the water will weigh the most, the alcohol will weigh the least, and the vegetable oil will weigh somewhere in between.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. *The Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

Give students time to answer the questions about the demonstration before conducting the activity.



3. Calculate the density of water, alcohol, and oil.

Question to investigate

Why does water sink in oil and alcohol float in oil?



Materials for each group

- Water
- Vegetable oil
- Isopropyl alcohol (70% or higher)
- Graduated cylinder
- Balance that measures in grams

This activity is written for students to make actual measurements of the mass and volume and to calculate the density of each liquid. Emphasize to students that they should be sure to accurately measure the volume and mass of each liquid.

Procedure

1. Find the mass of an empty graduated cylinder. Record the mass in grams in the chart on the activity sheet.
2. Pour 20 mL of water into the graduated cylinder. Try to be as accurate as possible by checking that the meniscus is right at the 20 mL mark.
3. Weigh the graduated cylinder with the water in it. Record the mass in grams.
4. Find the mass of only the water by subtracting the mass of the empty graduated cylinder. Record the mass of 20 mL of water in the chart.
5. Use the mass and volume of the water to calculate density. Record the density in g/cm^3 in the chart.
6. Follow steps 2–5 for alcohol and then oil. Be sure to measure the oil last because it does not rinse easily from the graduated cylinder.

	Water	Alcohol	Oil
Mass of graduated cylinder + liquid (g)			
Mass of empty graduated cylinder (g)			
Mass of liquid (g)			
Density of liquid (g/cm^3)			

4. Discuss whether the calculated densities support the order the liquids layer in the graduated cylinder.

Ask students:

- **Do the densities you calculated explain why liquids float and sink in one another? Explain.**

Yes, the water is the most dense and sinks in the oil. The alcohol is the least dense and floats on the oil.

EXPLAIN

5. Compare the density of water, alcohol, and oil on the molecular level.

Depending on the mass and size of the molecules that make up different liquids and how closely they pack together, liquids have their own characteristic densities.

Project the image *Oil*

www.middleschoolchemistry.com/multimedia/chapter3/lesson5#oil

Tell students that molecules of oil are mostly made of carbon and hydrogen atoms bonded together.

Project the image *Water*

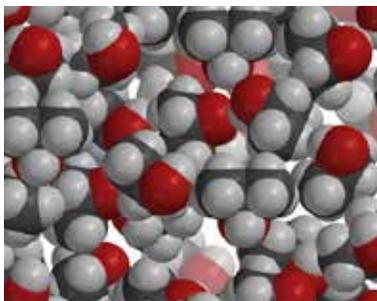
www.middleschoolchemistry.com/multimedia/chapter3/lesson5#water

Water molecules are made up of oxygen and hydrogen atoms bonded together. Oxygen is heavier and smaller than carbon, so a volume of water molecules is heavier than the same volume of oil molecules. This makes water more dense than oil. Also, water molecules are very attracted to each other and pack very close together. This is another reason why water is more dense than oil.

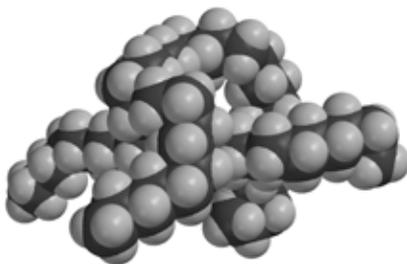
Project the image *Alcohol*

www.middleschoolchemistry.com/multimedia/chapter3/lesson5#alcohol

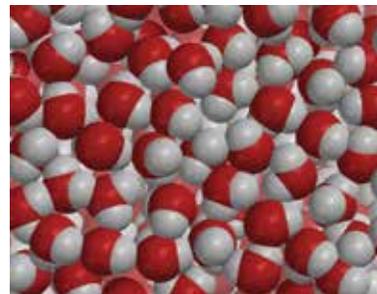
Alcohol is less dense than oil. Alcohol molecules are mostly carbon and hydrogen atoms so they are similar to oil. But they also contain an oxygen atom, which makes them a little heavy. For this reason, you might think that alcohol would be more dense than oil. But alcohol molecules do not pack very tightly together. Because of their shape and size, alcohol molecules do not pack as efficiently as oil molecules, making alcohol less dense than oil.



Alcohol



Oil



Water

EXTEND

6. As a demonstration, change the density of water so that a sinking carrot slice floats.

You may choose to do the following either as a demonstration or as an activity that students can do.

Materials

- Tall clear plastic cup
- Water
- Carrot slice about $\frac{1}{4}$ inch thick
- Salt
- Spoon

Procedure

1. Pour water into a clear tall plastic cup until it is about $\frac{1}{2}$ -filled.
2. Place a slice of carrot in the water.

Ask students:

Is the carrot more or less dense than the water?

Since the carrot sinks, students should conclude that the carrot is *more* dense than water.

3. Add about 1 teaspoon of salt to the water and stir. Continue to stir until the carrot floats to the surface of the salt water. If the carrot does not float to the surface, add more salt and stir.



Expected results

The slice of carrot should float in the saltwater.

Ask student:

- **Is it more or less dense than saltwater?**
Since the carrot floats in saltwater, students should conclude that the carrot is less dense than saltwater.
- **How does adding salt change the density of the water?**
Dissolving salt in water increases both the mass and volume of the water, but it increases the mass more. Because $D = m/v$, increasing the mass more than the volume results in an increase in density.
- **What would you expect if you placed equal volumes of water and saltwater on opposite ends of a balance?**
If equal volumes of water and saltwater were placed on a balance, the saltwater would be heavier.

DEMONSTRATION

1. Your teacher showed you one candle floating in water and another identical candle sinking in alcohol.

Do water and alcohol have the same or different densities?

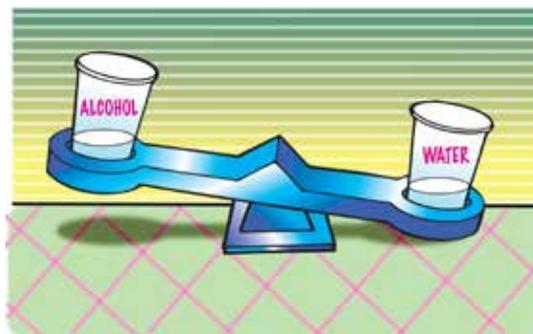
Which liquid is more dense?

How do you know?



2. Your teacher placed equal volumes of water and alcohol on a balance.

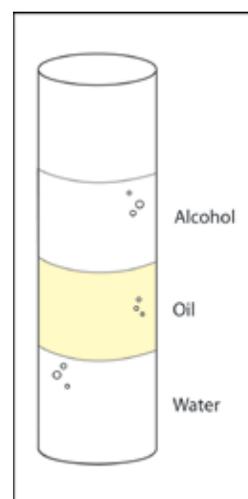
Explain how this demonstration proves that water is more dense than alcohol. Be sure to mention both volume and mass.



3. Your teacher showed you a graduated cylinder with alcohol, oil, and water.

Why does the alcohol float on the oil?

Why does the water sink in the oil?



ACTIVITY

Question to investigate

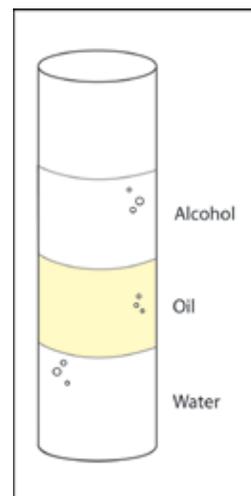
Why does water sink in oil, and alcohol float in oil?

Materials for each group

- Water
- Vegetable oil
- Isopropyl alcohol
- Graduated cylinder
- Balance that measures in grams

Procedure

1. Find the mass of an empty graduated cylinder. Record the mass in grams in the chart on the activity sheet.
2. Pour 20 mL of water into the graduated cylinder. Try to be as accurate as possible by checking that the meniscus is right at the 20-mL mark.
3. Weigh the graduated cylinder with the water in it. Record the mass in grams.
4. Find the mass of only the water by subtracting the mass of the empty graduated cylinder. Record the mass of 20 mL of water in the chart.
5. Use the mass and volume of the water to calculate density. Record the density in g/cm^3 in the chart.
6. Follow steps 2–5 for alcohol and then oil. Be sure to measure the oil last because it does not rinse easily from the graduated cylinder.

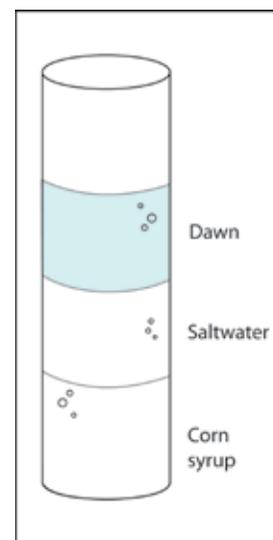


	Water	Alcohol	Oil
Mass of graduated cylinder + liquid (g)			
Mass of empty graduated cylinder (g)			
Mass of liquid (g)			
Density of liquid (g/cm^3)			

4. How do the densities you calculated explain why water sinks in oil and alcohol floats on oil?

5. Look at the layered liquids in the illustration. Write *most*, *least*, or *in-between* in the chart below to describe the density of each liquid.

Liquid	Density
Dawn	
Saltwater	
Corn syrup	



EXPLAIN IT WITH ATOMS & MOLECULES

6. Water molecules are smaller and have less mass than alcohol and oil molecules. Explain why water is more dense than alcohol and oil.

TAKE IT FURTHER

7. A carrot slice sinks in water and floats in saltwater.

Is the carrot more dense or less dense than *water*?



Is the carrot more dense or less dense than *saltwater*?

8. Does adding salt change the density of the water?

How do you know?

9. What would you expect if you placed equal volumes of water and saltwater on opposite ends of a balance?

10. Adding salt to water increases both its mass and volume; which do you think it increases more, the mass of the water or the volume?

Chapter 3, Lesson 6: Temperature Affects Density

Key Concepts

- Heating a substance causes molecules to speed up and spread slightly further apart, occupying a larger volume that results in a decrease in density.
- Cooling a substance causes molecules to slow down and get slightly closer together, occupying a smaller volume that results in an increase in density.
- Hot water is less dense and will float on room-temperature water.
- Cold water is more dense and will sink in room-temperature water.

Summary

Students place hot and cold colored water into room-temperature water. They observe that the hot water floats on the room-temperature water and the cold water sinks. Students will combine the concepts of temperature, molecular motion, and density to learn that hot water is less dense than room-temperature water and that cold water is more dense.

Objective

Students will be able to explain, on the molecular level, how heating and cooling affect the density of water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Make sure you and your students wear properly fitting goggles. Use caution when handling hot water.

About this Lesson

In this lesson, you can help students connect some of the concepts about density to ideas from Chapter 1. In Chapter 1, students saw that heat increases molecular motion. This increased motion competes with the attractions between molecules, causing the molecules to move a little further apart. They also saw that as a substance is cooled, molecules slow down and their attractions bring them closer together. These ideas can also be applied to the concept of density.

Materials for Each Group

- Cold water (colored blue) in foam cup
- Hot water (colored yellow) in foam cup
- Room temperature water in clear plastic cup (colorless)
- 2 droppers

Materials for the Demonstration

- Hot water (colored yellow)
- Cold water (colored blue)
- 2 identical clear baby food jars
- Water-resistant card (from a deck of cards or laminated index card)
- Paper towels

ENGAGE

1. Do a demonstration to show that hot water floats on cold water.

Tell students that in Chapter 3, they have seen that different substances have different densities. In this activity, they will see that the *same* substance can have *different* densities at different temperatures.

Tell students that you are going to try to place one jar filled with hot colored water upside down over another jar with cold colored water.

Ask students to make a prediction:

- Do you think the hot and cold water will mix or stay separate?

Either follow the procedure below or project the video for students. If you decide to do the demonstration, you may want to watch the video first in order to see how to set the jars up.

Project the video *Hot Water on Cold Water*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson6#hot_water_on_cold_water

Materials

- Hot water (about 50 °C, colored yellow)
- Cold water (about 5 °C, colored blue)
- 2 identical clear baby food jars
- Water-resistant card (from a deck of cards or laminated index card)
- Paper towels

Procedure

Hot water on top

1. Completely fill a baby food jar with hot tap water and add 2 drops of yellow food coloring.
2. Completely fill another baby food jar with very cold water and add 2 drops of blue food coloring. Stir the water in both jars so that the coloring is well-mixed in both. Place the cold water jar on a paper towel.
3. Hold a water-resistant card over the top of the hot water jar.
4. While holding the card against the jar opening, carefully turn the jar upside down.
5. With the card still in place, position the jar of hot water directly over the jar of cold water so that the tops line up exactly.
6. Slowly and carefully remove the card so that the hot water jar sits directly on top of the cold water jar.



Expected results

Although removing the card may result in a little mixing or spilling, the hot yellow water will remain in the top jar and the cold blue water will remain in the bottom jar.

Ask students:

- **Why do you think the hot water stayed on top of the cold water?**
Students should realize that there is a density difference between hot and cold water. Hot water is less dense so it floats on the denser cold water.

Ask students to make a prediction:

- **What might happen if you placed the cold blue water on top of the hot yellow water and then removed the card?**

Cold water on top

7. Use the same procedure as above, but place the jar of cold water, upside down over the jar of hot water.

Expected results

The cold blue water will immediately fall into the hot yellow water causing mixing. The water will quickly become green throughout.

Ask students:

- **Why do you think the hot and cold water mixed when the cold water was placed on top?**
When the cold water is placed on top, the colors mix because the cold water is more dense and sinks in the hot water.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Have students try adding cold and hot water to room-temperature water.

Question to investigate

Is there a density difference between hot and cold water?



Materials for each group

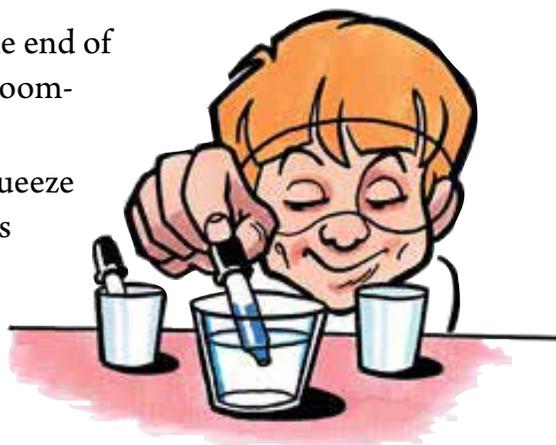
- Cold water (colored blue) in foam cup
- Hot water (colored yellow) in foam cup
- Room-temperature water in clear plastic cup (colorless)
- 2 droppers

Teacher preparation

- Add ice to water to make very cold water. Half-fill one foam cup with cold water (no ice cubes) and another with hot water for each group.
- Add 2 drops of yellow food coloring to the hot water and 2 drops of blue food coloring to the cold water.
- Fill a clear plastic cup about $\frac{2}{3}$ of the way with room-temperature water.
- Distribute the set of 3 cups to each group.

Procedure

1. Fill one dropper with blue cold water. Poke the end of the dropper about halfway into the colorless room-temperature water.
2. While observing from the side, very gently squeeze the dropper so that the cold water slowly flows into the room-temperature water.
3. Fill another dropper with yellow hot water. Poke the end of the dropper about halfway into the room-temperature water.
4. While observing from the side, very gently squeeze the dropper so that the hot water slowly flows into the room-temperature water.
5. Record your observations on the activity sheet.



Expected results

The cold blue water will flow down and collect at the bottom of the room-temperature water. The hot yellow water will rise and collect at the surface.

3. Discuss student observations.

Ask students:

About cold water

- **What did you notice when you placed the cold blue water in room-temperature water?**

The cold water sank in the room-temperature water.

- **Is cold water more, less, or the same density as room-temperature water?**
Cold water is more dense than room-temperature water.

About hot water

- **What did you notice when you placed the hot yellow water in room-temperature water?**
The hot water floated to the surface in the room-temperature water.
- **Is hot water more, less, or the same density as room-temperature water?**
Hot water is less dense than room-temperature water.

EXPLAIN

4. Explain the difference in density between hot and cold water on the molecular level.

Project the animation *Cold and Hot Water*.

www.middleschoolchemistry.com/multimedia/chapter3/lesson6#cold_and_hot_water

Cold water

Point out that the molecules of cold water move slower and are a little closer together than the hot or room-temperature water. Also point out that when the water is cooled, the water level falls slightly in the graduated cylinder.

Ask students:

- **In the animation, you saw that as water is cooled the water level goes down. Cold water has a smaller volume, but the mass stays the same. What does this tell you about the density of cold water?**
Students should understand that when the molecules come together as the water is cooled, the volume decreases. But the mass of the water does not change. Students should realize that decreasing the volume without increasing the mass is going to increase the density.
- **How does this help explain why cold water sinks in room-temperature water?**
The more-dense cold water sinks in the room-temperature water.

Hot water

Point out that the molecules in the hot water are moving faster and are a little farther apart than the molecules in room-temperature water. Make sure students notice that when the water is heated, the water level rises slightly in the graduated cylinder.

Ask students:

- **In the animation, you saw that as water is heated the water level rises. Hot water takes up more volume, but the mass stays the same. What does this tell you about the density of hot water?**

Based on the animation, students should understand that the spreading apart of the molecules increases the volume but does not affect the mass of the water. Students should realize that increasing the volume without increasing the mass is going to decrease the density.

- **How does this help explain why hot water floats on room temperature water?**
The less-dense hot water floats on the more dense room temperature water.

EXTEND

5. Do a demonstration to show students how dense cold water causes mixing.

Tell students that in winter, ice can form on the tops of ponds and lakes. In the spring when the ice melts, the cold water sinks. This causes mixing from the bottom which brings nutrients up to the surface. Tell students that you will model this process.

Materials

- Room-temperature water
- Ice cubes
- 2 identical tall clear plastic cups
- Small cup
- Food coloring, any color other than yellow
- Dropper
- Sheet of plain white paper

Procedure

1. Fill two tall clear plastic cups about $\frac{2}{3}$ of the way with room-temperature water.
2. Place about 15 drops of food coloring into a small empty cup.
3. Use a dropper to pick up the food coloring. Then carefully push the dropper into the water until the tip of the dropper is near the bottom of the cup.
4. Very gently squeeze the dropper so that all of the food coloring slowly flows to the bottom of the cup. Then carefully remove the dropper to prevent food coloring from mixing into the water. (It is ok if a little coloring gets mixed in the water.)
5. Repeat steps 2–4 for the other cup of water.
6. Gently place two ice cubes in the water in one of the cups. (Avoid agitating the water.)
7. Place a piece of white paper behind each cup and observe.



Expected results

The coloring in the cup with the ice will move up from the bottom and begin to mix throughout the water. The coloring in this cup will mix faster than the coloring in the cup without ice.

Ask students:

- **The food coloring mixed more quickly in the cup that had the ice. Use what you know about the density of water at different temperatures to explain why this happened.**

Ice is about 0 °C and the water in the cup is about 20 °C. As the ice melts, the water from the melted ice is colder than the water around it. This colder water is also more dense, so it sinks to the bottom. This sinking water pushes the food coloring out of the way, causing mixing.

DEMONSTRATION

You saw a jar of hot water placed upside down over a jar of cold water. The hot water stayed on top of the cold water without mixing.

1. Why did the hot water stay on top of the cold water?
2. Why do you think the hot and cold water mixed when the cold water was placed on top?



ACTIVITY

Question to investigate

Is there a density difference between hot and cold water?



Materials for each group

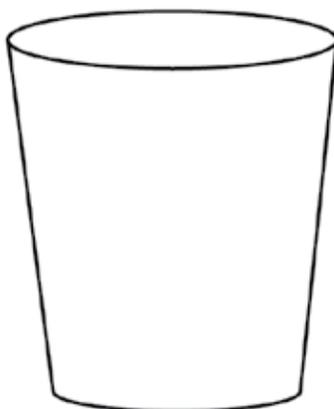
- Cold water (colored blue) in foam cup
- Hot water (colored yellow) in foam cup
- Room-temperature water in clear plastic cup (colorless)
- 2 droppers

Procedure

1. Fill one dropper with blue cold water. Poke the end of the dropper about halfway into the colorless room-temperature water.
2. While observing from the side, very gently squeeze the dropper so that the cold water slowly flows into the room-temperature water.
3. Fill another dropper with yellow hot water. Poke the end of the dropper about halfway into the room-temperature water.
4. While observing from the side, very gently squeeze the dropper so that the hot water slowly flows into the room-temperature water.
5. Record your observations on the activity sheet.



3. Draw what you observed in the cup of room-temperature water after adding blue cold water and yellow hot water.



Be sure to label the areas of cold and hot water.

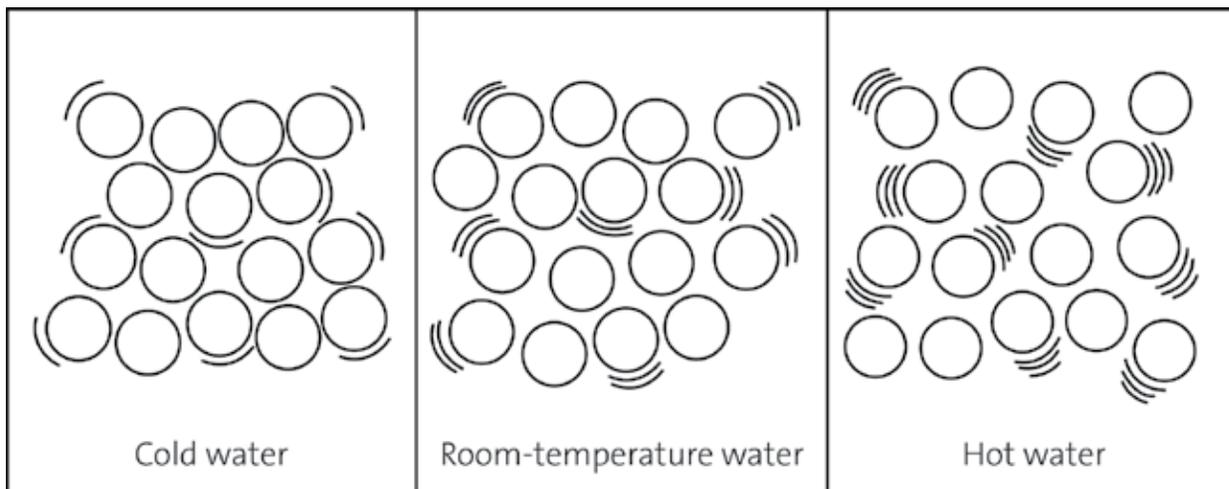
Is *cold* water more, less, or the same density as room-temperature water?

Is *hot* water more, less, or the same density as room-temperature water?

EXPLAIN IT WITH ATOMS & MOLECULES

In the animation, you saw water molecules being heated and cooled.

- Look at the model of water molecules in the diagram below to help you compare the volume, mass, and density of cold and hot water.



Write *more*, *less*, or *same* in the chart to describe the volume, mass, and density of cold and hot water compared to room temperature water.

Comparing cold and hot water to room-temperature water		
	Cold water	Hot water
Volume		
Mass		
Density		

5. Use what you know about density to answer the following questions.

Why does cold water sink in room-temperature water?

Why does hot water float on room-temperature water?

TAKE IT FURTHER

6. Your teacher did a demonstration with two cups of water that both had food coloring on the bottom. Ice was placed in one cup of water, but not the other. The food coloring mixed more quickly in the cup that had the ice. Use what you know about the density of water at different temperatures to explain why this happened.

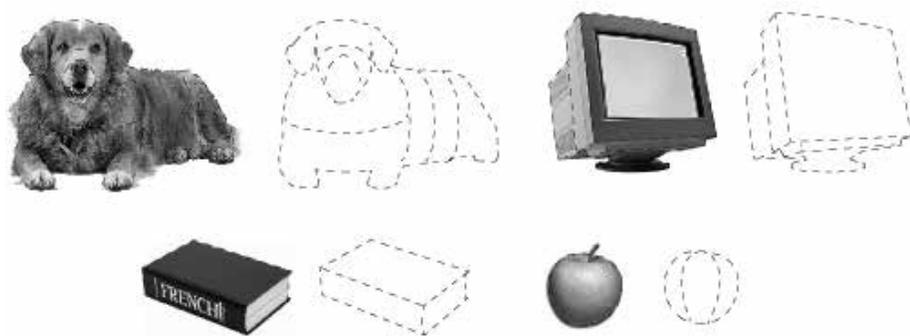
Note: Students may wonder why boats made out of dense material like steel can be made to float. This is a good question but not trivial to explain. To adequately answer this question, you will need to consider another aspect of sinking and floating. For an object to float, the mass of the water displaced by the object must equal the mass of the object. For example, a cube of steel will displace a certain volume of water equal to the volume of the cube. But the mass of the water displaced is less than the mass of the cube, so the cube sinks. But if you flatten out the steel cube and shape it into a large enough bowl and place the steel bowl in the water, the bowl will float. Its mass hasn't changed, but the volume of water it can displace has increased. The mass of the water displaced now equals the mass of the bowl and the bowl floats. You could stop your explanation here but there is another good question that can arise.

Why should the *mass* of water displaced matter? This question requires a more physics-like perspective to answer. The mass of the bowl is pushing down onto the water by a force equal to the mass of the bowl \times the acceleration of gravity ($\text{Force} = m \times a$). There is a law in physics that says that for every force on a stationary object, there is an equal and opposite force. In this case, the downward force of the bowl on the water is met with an equal force of the water pushing up on the bowl resulting in the bowl floating. If students have any question about an opposing upward force, ask them if they have ever tried to push an inflated ball like a beach ball under water.

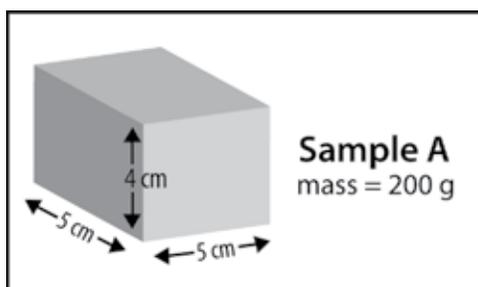
Chapter 3—Student Reading

If you hold a solid piece of lead or iron in your hand, it feels heavy for its size. If you hold the same size piece of balsa wood or plastic, it feels light for its size. The property of an object that causes this effect is called *density*. The density of an object depends on its mass and its volume. The mass is the amount of matter in the object. The volume is the amount of space that the object takes up in three dimensions.

All the objects around you take up a certain amount of space, no matter what shape they are. They all have length, width, and depth so they take up space in three dimensions.



These pictures show that every object has a volume that takes up 3-dimensional space. The volume of a wooden block, for instance is its length \times width \times height. For the block shown, the volume equals $5\text{cm} \times 5\text{cm} \times 4\text{cm} = 100$ cubic centimeters (cm^3).



A mathematical equation for density is: Density = mass/volume or $D = m/v$. If something has a large mass compared to its volume, it has a high density. This is like a set of weights which can be small but heavy.



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But if an object has a small mass compared to its volume it has a lower density. This is like an apple or a piece of wood which can seem light for its size.



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Different types of plastic, metal, wood, and other materials have different densities. The density of a material is based on the atoms or molecules the substance is made from. For example, a copper and an aluminum cube of the same volume feel very different when you hold them. The copper cube feels much heavier than the aluminum cube. If you put them on a balance, you see that the copper cube has more mass than the aluminum cube.

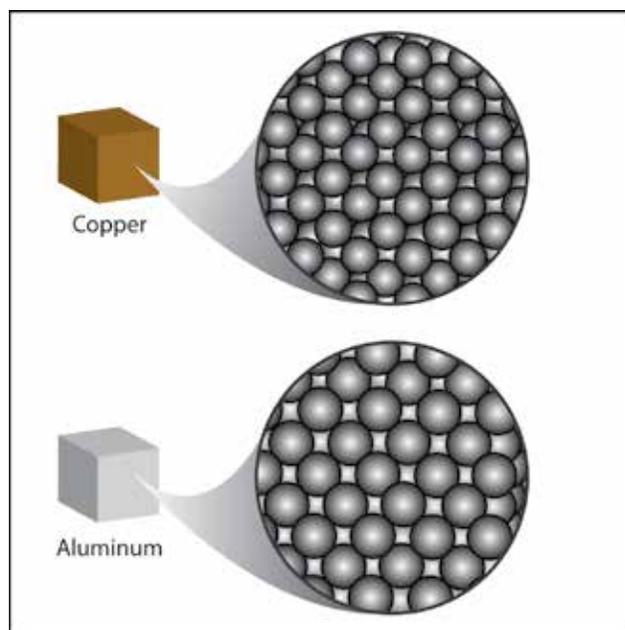
Since the cubes have the same volume and the copper has a greater mass, the copper cube is more dense than the aluminum cube. This is true since $D = m/v$. If the cubes have the same volume, the one with the greater mass must be more dense.



If you think about the atoms of the two metals, there can only be a couple of reasons why the copper is more dense than the aluminum:

- Copper atoms might have more mass than aluminum atoms.
- Copper atoms might be smaller than aluminum atoms so more can fit in the same volume.
- Copper atoms might be arranged differently than aluminum atoms so more can fit in the same volume.

Either one or any combination of these explanations could be the reason why the copper cube has more mass. It turns out that copper and aluminum atoms are arranged about the same way but copper atoms are smaller and have more mass than aluminum atoms.



Therefore more can fit into the same volume and each one has more mass. This makes copper more dense than aluminum.

A sample of a substance with a higher density will always have a greater mass than the same size sample of a substance with a lower density. For example, a sample of lead weighs more than the same size sample of wax.

A small sample of a substance with a high density may weigh as much or more than a larger sample of a substance with a lower density. For example, a small piece of iron may weigh as much or more than a larger piece of plastic.

A closer look at mass and volume

In order to find the density of a substance, you need to measure the mass and the volume of a sample of the substance.

Mass is the amount of matter in an object.

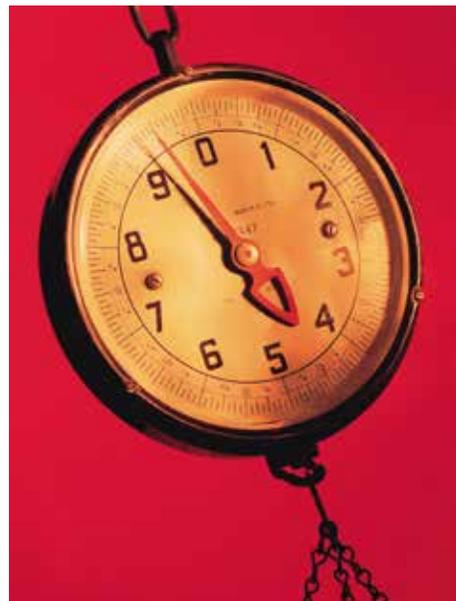
People are often confused between the meaning of “mass” and “weight”. The meaning of mass and weight are different but they are related to each other. Let’s say you have an object like a bowling ball. The bowling ball, like everything else, is made of a certain amount of matter. Let’s call the amount of matter that makes up the bowling ball the *mass* of the bowling ball. You hook the bowling ball to a scale that shows that the amount of mass that makes up the bowling ball weighs 9 pounds.

Then you do something unusual: You fly the bowling ball and the scale to the moon and hook the bowling ball to the same scale again. The moon has less gravity than Earth so the bowling ball is not pulled down as hard as it was on Earth.

Let’s say that on the moon, the scale shows that the mass that makes up the bowling ball weighs only 1.5 pounds. You know that the bowling ball itself didn’t change. It is still made of the same amount of matter so it still has the same mass. The only thing that changed was the force of gravity pulling down on the bowling ball.

So mass is a measure of the amount of matter that makes up an object. Weight is a measure of the force of gravity on a certain mass.

So how can you measure the mass of an object by putting it on a scale? Since gravity is pulling down on the object, why doesn’t a scale always measure weight? That is a great question and the answer has to do with how the scale is made.



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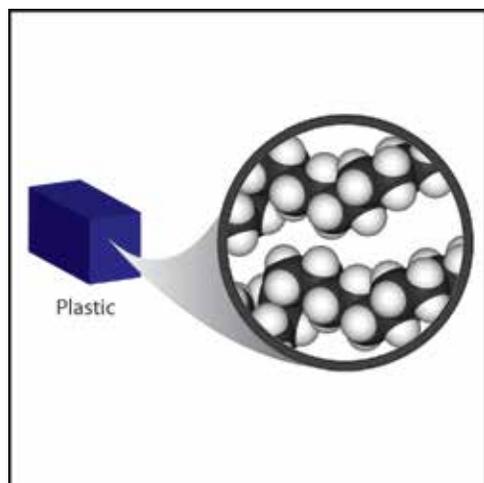
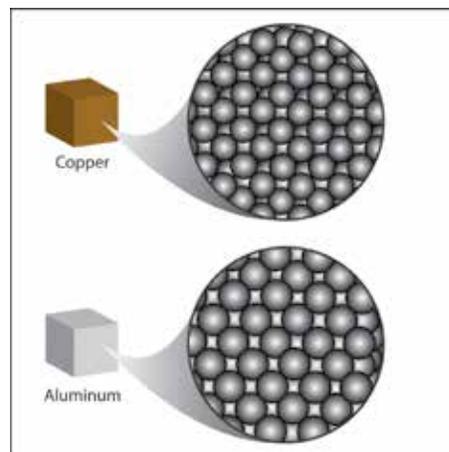
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When you put an object on a scale, of course the scale “feels” the effect of gravity but the scale is programmed or calibrated to do an internal calculation to factor out the effect of gravity and to display only the mass.

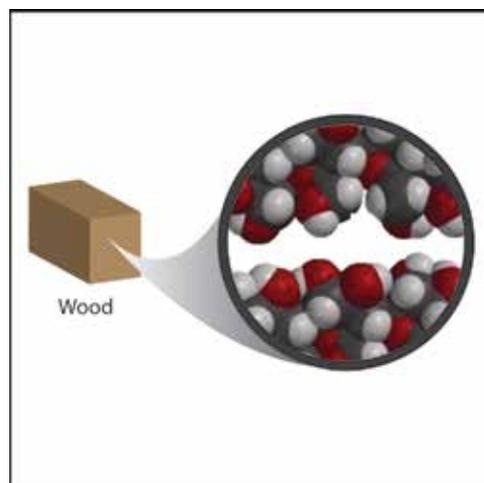
The density of copper is greater than the density of plastic or wood

If you compared the density of a copper cube and a plastic cube of the same volume, the copper is more dense. This is because copper is made from small massive atoms that are packed closely together. This gives copper a fairly high density.

Plastics are made mostly of carbon and hydrogen which are not as massive for their size. They are connected together in long chains and not packed as tightly as the atoms in copper. This makes plastic less dense than copper.



Wood is made of carbon, hydrogen, and oxygen atoms. These are pretty similar in size and mass to the atoms in plastics. Wood is also made mostly of long molecules that are arranged and packed together to make the structure of the wood.



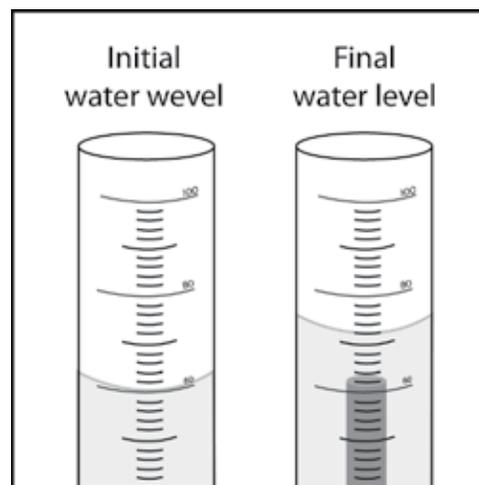
Because of the size, mass, and arrangement of its molecules, the density of wood is more similar to plastic than to copper.

Finding volume using the water displacement method

Sometimes finding the volume of an object is not as easy as simply using a metric ruler to measure its length, width, and height. Another method is the water displacement method. There are two basic ways of doing the water displacement method to find volume.

Fill a graduated cylinder with water to a level that is high enough so that the object placed in the water will be submerged. Record the initial level of the water.

Carefully place an object in the water and let it sink to the bottom. If it doesn't sink, push it down gently with something thin like a pencil point so that it is just under the surface. The water level will rise. Record the final water level. Subtract the initial water level from the final water level to calculate the volume of water that was displaced by the object. The volume of water displaced by an object equals the volume of the object. In this case the final water level (72 mL) - initial water level (60 mL) = 12 mL. Since a milliliter is the same as a cubic centimeter (cm^3), the volume of the object is 12 cm^3 .



Another water displacement method:

Place a cup or beaker in a larger container. Fill the cup as full as possible until it is ready to overflow. Gently place an object in the water and let it sink to the bottom. If it doesn't sink, push it down gently with something thin like a pencil point so that the object is just under the surface.

Some water will flow out of the cup and into the outer container. This water was displaced by the object. Carefully pour this water into a graduated cylinder to measure its volume. The volume of the displaced water equals the volume of the object.



The density of a substance is the same no matter what the size of the sample.

This means that a big piece of wax, for example, has the same density as a small piece of the same wax. Take water for instance. 100 mL of water has a mass of 100 grams. Since density = mass/volume, this sample of water has a density of 100 grams/100 mL = 1 gram/mL = 1 g/cm³. 50 mL of water has a mass of 50 grams. The density of this sample of water is 50 grams/50 mL = 1 gram/cm³.

Water or any substance always has the same density no matter what the size of the sample.

Density in sinking and floating

The density of an object and the density of the liquid it is placed in determine whether an object will sink or float.

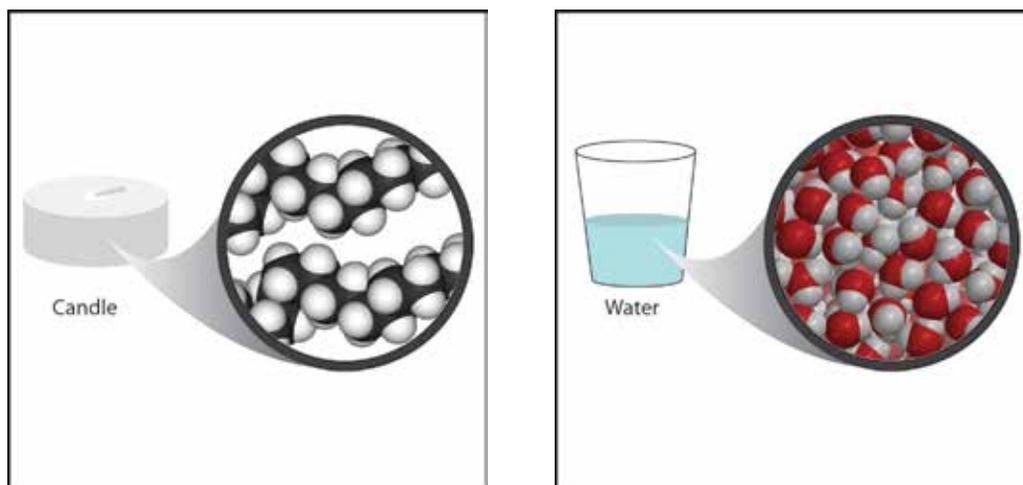
Even though one object might be heavier than another, the heavy one might float and the lighter one might sink. An example is a piece of clay and a heavier wax candle placed in water. Even though the wax is heavier than the clay, the wax floats and the clay sinks.

This is possible because it's not the mass of the object that matters in sinking and floating but its density compared to the density of water.

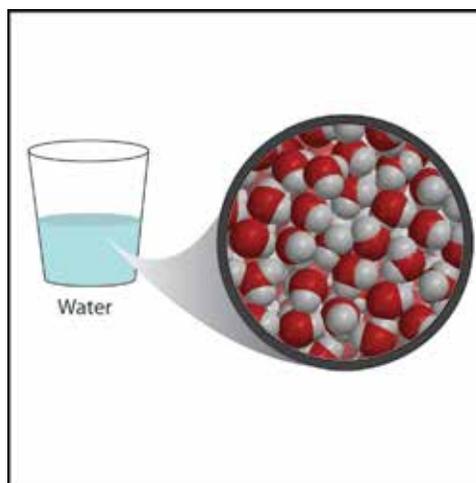
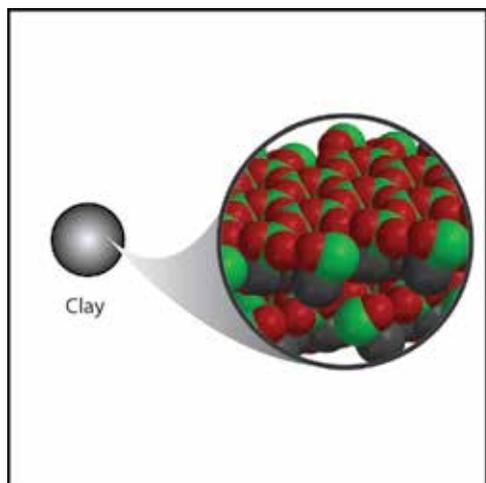
An object that is less dense than water will float. An object that is more dense than water will sink. This is why the clay sinks in water and the wax floats. The clay is more dense than water and the wax is less dense than water.

Since water is more dense than wax, a volume of water has more mass than an equal volume of wax. You can prove this by putting a volume of wax and an equal volume of water on opposite ends of a balance. The balance will show that the water has a greater mass than the wax. This means that the water is more dense than the wax. This is why the wax floats.

Wax is made of carbon and hydrogen atoms linked together into long molecules. These atoms are pretty light and the arrangement of the molecules makes wax less dense than water.



If you put a volume of clay and an equal volume of water on opposite ends of a balance, the balance will show that the clay has more mass. This means that the clay is more dense than the water. This is why the clay sinks. Clay is made from oxygen and heavier atoms like aluminum and silicon. The mass of these atoms and their arrangement make clay more dense than water.



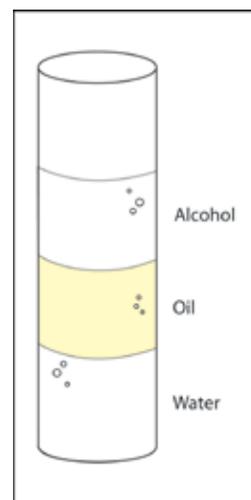
Liquids can sink and float in other liquids

Sinking and floating applies to liquids too. You can test this by carefully placing different liquids together in a graduated cylinder. For example, if you add vegetable oil to water, the oil floats on the water. If you add isopropyl alcohol, the alcohol floats on the oil.

This means that oil is less dense than water and alcohol is less dense than oil.

The density of an object or the water it is placed in can be changed so that an object that normally sinks will float.

If an object sinks in water, this means that the object is more dense than the water. There are two possible ways to make the object float. You can increase the density of the water so that the water becomes more dense than the object. Or you could increase the volume of the object so that the object becomes less dense than the water.



If you put a slice of carrot in water, the carrot sinks. This is because the carrot is more dense than water. But if you dissolve something like salt in the water, the density of the water increases.

This is because the mass of the water increases a lot and the volume does not increase as much. Since $D=m/v$, increasing mass a lot and volume just a little will result in an increase in density. If enough salt dissolves, the density of the water can be increased enough that the salt water will become more dense than the carrot and the carrot will float.

Why do heavy boats float?

Why can boats float when they are made out of material that is more dense than water?

An object made from dense material can float if its volume is made large enough. Here's an example: If you have a cube of clay that is 3 cm on each side, the cube has a volume of $3\text{ cm} \times 3\text{ cm} \times 3\text{ cm} = 27\text{ cm}^3$. Let's say that the cube has a mass of 60 grams. The density of the clay cube is $60\text{ g}/27\text{ cm}^3 = 2.2\text{ g/cm}^3$. Since the density of water is 1 g/cm^3 , the clay cube sinks. [Picture]



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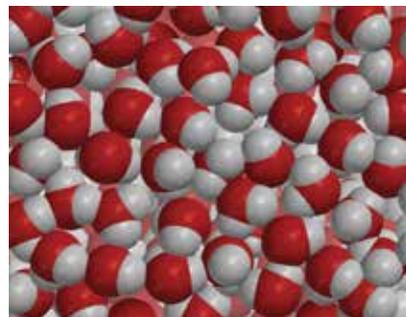
But you can take the same cube of clay and press it down into a thin pancake and form it into a bowl that is like half a sphere. If it has a diameter of 8 cm, the clay bowl will have a volume of about 134 cm^3 or more than four times the volume of the cube. The density of the bowl is now $60\text{ g}/134\text{ cm}^3 = .45\text{ g/cm}^3$. The density of the clay itself doesn't change but the density of the object does. The increase in the volume of the object decreases the density so that the density of the bowl is less than the density of water so the bowl floats.

The temperature of a substance affects its density.

When a substance like a solid or a liquid is heated, its molecules move faster and get slightly further apart. The substance still has the same mass but it takes up a larger volume. Since $D=m/v$, a larger volume results in a decreased density.

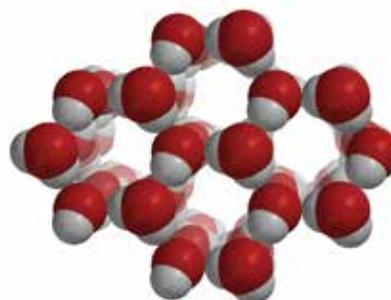
This can be shown pretty easily with liquids. If hot water is carefully placed on cold water, the hot water floats because it is less dense than the cold water.

When water is cooled, its molecules move slower and get a little closer together. The water still has the same mass but it takes up a smaller volume. Since $D=m/v$, a smaller volume results in an increased density. If cold water is placed on hot water, the cold water sinks because it is more dense than the hot water.



Ice is less dense than liquid water.

Normally, when a liquid is cooled, its molecules slow down and the attractions between molecules bring them closer together. But water is different. When water gets below $4\text{ }^\circ\text{C}$, its molecules actually begin getting a little further apart as they orient themselves into the crystal structure of ice. Since the mass does not change but the ice occupies a larger volume than the water, the density of ice is less than the density of water. This is why ice floats in water.



Chapter 4, Lesson 1: Protons, Neutrons, and Electrons

Key Concepts

- Atoms are made of extremely tiny particles called protons, neutrons, and electrons.
- Protons and neutrons are in the center of the atom, making up the nucleus.
- Electrons surround the nucleus.
- Protons have a positive charge.
- Electrons have a negative charge.
- The charge on the proton and electron are exactly the same size but opposite.
- Neutrons have no charge.
- Since opposite charges attract, protons and electrons attract each other.

Summary

Students will put a static charge on a strip of plastic by pulling it between their fingers. They will see that the plastic is attracted to their fingers. Students will be introduced to the idea that rubbing the strip with their fingers caused electrons to move from their skin to the plastic giving the plastic a negative charge and their skin a positive charge. Through these activities, students will be introduced to some of the characteristics of electrons, protons, and neutrons, which make up atoms.

Objective

Students will be able to explain, in terms of electrons and protons, why a charged object is attracted or repelled by another charged object. They will also be able to explain why a charged object can even be attracted to an uncharged object. Students will also be able to explain that the attraction between positive protons and negative electrons holds an atom together.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- Plastic grocery bag
- Scissors
- Inflated balloon
- Small pieces of paper, confetti-size

Materials for the Demonstration

- Sink
- Balloon

ENGAGE

1. Show a picture of a pencil point and how the carbon atoms look at the molecular level.

Project the image *Pencil Zoom*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson1#pencil_zoom

Students should be familiar with the parts of the atom from Chapter 3 but reviewing the main points is probably a good idea.

Ask students questions such as the following:

- **What are the three different tiny particles that make up an atom?**
Protons, neutrons, and electrons.
- **Which of these is in the center of the atom?**

Protons and neutrons are in the center (nucleus) of the atom. You may want to mention that hydrogen is the only atom that usually has no neutrons. The

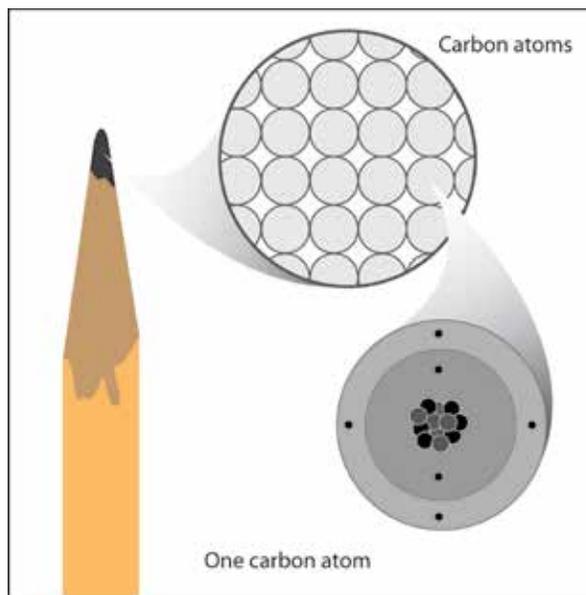
nucleus of most hydrogen atoms is composed of just 1 proton. A small percentage of hydrogen atoms have 1 or even 2 neutrons. Atoms of the same element with different numbers of neutrons are called *isotopes*. These will be discussed in Lesson 2.

- **What zooms around the nucleus of an atom?**

Electrons

- **Which one has a positive charge, a negative charge, and no charge?**

Proton—positive; electron—negative; neutron—no charge. The charge on the proton and electron are exactly the same size but opposite. The same number of protons and electrons exactly cancel one another in a neutral atom.



Note: The picture shows a simple model of the carbon atom. It illustrates some basic information like the number of protons and neutrons in the nucleus. It also shows that the number of electrons is the same as the number of protons. This model also shows that some electrons can be close to the nucleus and others are further away. One problem with this model is that it suggests that electrons orbit around the nucleus in perfect circles on the same plane, but this is not true. The more widely accepted model shows the electrons as a more 3-dimensional “electron cloud” surrounding the nucleus. Students will be introduced to these ideas in a bit more detail in Lesson 3. But for most of our study of chemistry at the middle school level, the model shown in the illustration will be very useful. Also, for most of our uses of this atom model, the nucleus will be shown as a dot in the center of the atom.

2. Show animations and explain that protons and electrons have opposite charges and attract each other.



Project the animation *Protons and Electrons*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson1#protons_and_electrons

Explain to students that two protons repel each other and that two electrons repel each other. But a proton and an electron attract each other. Another way of saying this is that the same or “like” charges repel one another and opposite charges attract one another.

Since opposite charges attract each other, the negatively charged electrons are attracted to the positively charged protons. Tell students that this attraction is what holds the atom together.

Project the animation *Hydrogen Atom*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson1#hydrogen_atom

Explain to students that in a hydrogen atom, the negatively charged electron is attracted to the positively charged proton. This attraction is what holds the atom together.

Tell students that hydrogen is the simplest atom. It has only 1 proton, 1 electron, and 0 neutrons. It is the only atom that does not have any neutrons. Explain that this is a simple model that shows an electron going around the nucleus.

Click on the button “Show cloud” and explain to students that this is a different model. It shows the electron in the space surrounding the nucleus that is called an electron cloud or energy level. It is not possible to know the location of an electron but only the region where it is most likely to be. The electron cloud or energy level shows the region surrounding the nucleus where the electron is most likely to be.

Note: Inquisitive students might ask how the positively charged protons are able to stay so close together in the nucleus: Why don't they repel each other? This is a great question. The answer is well beyond an introduction to chemistry for middle school, but one thing you can say is that there is a force called the “Strong Force,” which holds protons and neutrons together in the nucleus of the atom. This force is much stronger than the force of repulsion of one proton from another.

Another good question: Why doesn't the electron smash into the proton? If they are attracted to each other, why don't they just collide? Again, a detailed answer to this question is beyond the scope of middle school chemistry. But a simplified answer has to do with the energy or speed of the electron. As the electron gets closer to the nucleus, its energy and speed increases. It ends up

moving in a region surrounding the nucleus at a speed that is great enough to balance the attraction that is pulling it in, so the electron does not crash into the nucleus.

Give each student an activity sheet.

Have students answer questions about the illustration on the activity sheet. Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions.



EXPLORE

3. Do an activity to show that electrons and protons attract each other.

Students can see evidence of the charges of protons and electrons by doing an activity with static electricity.

Note: When two materials are rubbed together in a static electricity activity, one material tends to lose electrons while the other material tends to gain electron. In this activity, human skin tends to lose electrons while the plastic bag, made of polyethylene, tends to gain electrons.

Question to investigate

What makes objects attract or repel each other?



Materials for each group

- Plastic grocery bag
- Scissors

Procedure, part 1

Charged plastic and charged skin

1. Cut 2 strips from a plastic grocery bag so that each is about 2–4 cm wide and about 20 cm long.
2. Hold the plastic strip firmly at one end. Then grasp the plastic strip between the thumb and fingers of your other hand as shown.
3. Quickly pull your top hand up so that the plastic strip runs through your fingers. Do this three or four times.
4. Allow the strip to hang down. Then bring your other hand near it.



5. Write “attract” or “repel” in the chart on the activity sheet to describe what happened.

Expected results

The plastic will be attracted to your hand and move toward it. Students may notice that the plastic is also attracted to their arms and sleeves. Let students know that later in this lesson they will investigate why the plastic strip is also attracted to surfaces that have not been charged (neutral).



Note: If students find that their plastic strip does not move toward their hand, it must not have been charged well enough. Have them try charging their plastic strip by holding it down on their pants or shirt and then quickly pulling it with the other hand. Then they should test to see if the plastic is attracted to their clothes. If not, students should try charging the plastic again.

EXPLAIN

4. Show students models comparing the number of protons and electrons in the plastic and skin before and after rubbing them together.

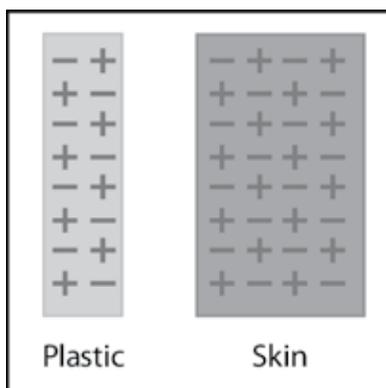
Tell students that the plastic strip and their skin are made of molecules that are made of atoms. Tell students to assume that the plastic and their skin are neutral—that they have the same number of protons as electrons.

Project the image *Charged plastic and hand.*

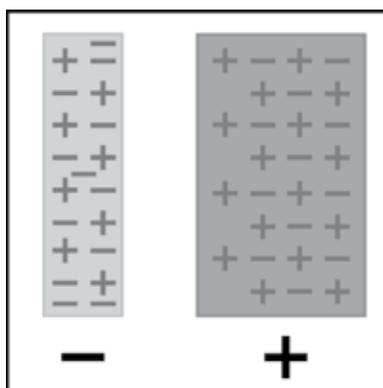
www.middleschoolchemistry.com/multimedia/chapter4/lesson1#charged_plastic_and_hand.jpg

Point out that before the students pulled the plastic between their fingers, the number of protons and electrons in each is the same. Then, when students pulled the plastic through their fingers, electrons from their skin got onto the plastic. Since the plastic has more electrons than protons, it has a negative charge. Since their fingers gave up some electrons, their skin now has more protons than electrons so it has a positive charge. The positive skin and the negative plastic attract each other because positive and negative attract.

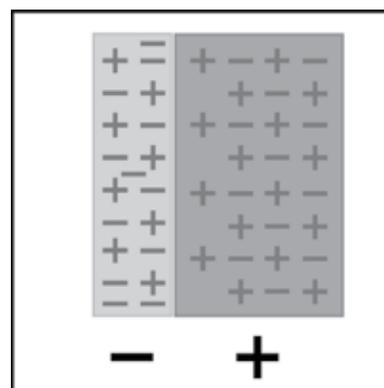
Protons and electrons
before rubbing



Protons and electrons
after rubbing



Opposites attract



EXPLORE

5. Have students investigate what happens when a rubbed plastic strip is held near a desk or chair.



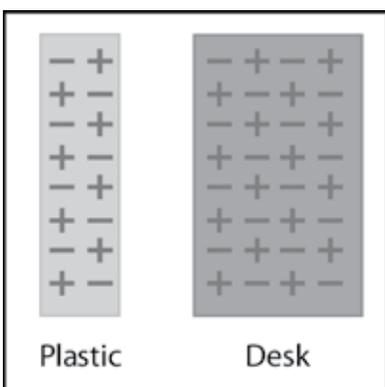
Procedure, part 2

Charged plastic and neutral desk

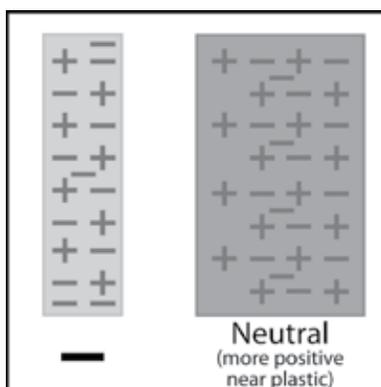
1. Charge one strip of plastic the same way you did previously.
2. This time, bring the plastic strip toward your desk or chair.
3. Write "attract" or "repel" in the chart.



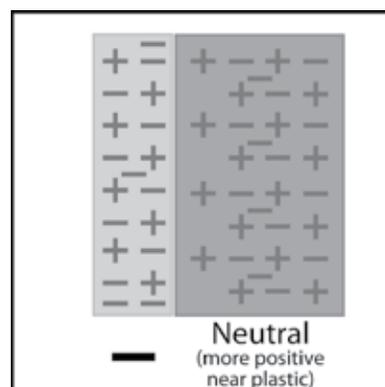
Protons and electrons
before rubbing



Protons and electrons after
rubbing



Opposites attract



Expected results

The plastic moves toward the desk.

Explain to students why the plastic is attracted to the desk. The answer takes a couple of steps, so you can guide students by drawing or projecting a magnified illustration of the plastic and desk.

After pulling the plastic between their fingers, the plastic gains extra electrons and a negative charge. The desk has the same number of protons as electrons and is neutral. When the plastic gets close to the desk, the negatively charged plastic repels electrons on the surface of the desk. This makes the surface of the desk near the plastic slightly positive. The negatively charged plastic is attracted to this positive area, so the plastic moves toward it.

6. Have students charge two pieces of plastic and hold them near each other to see if electrons repel one other.

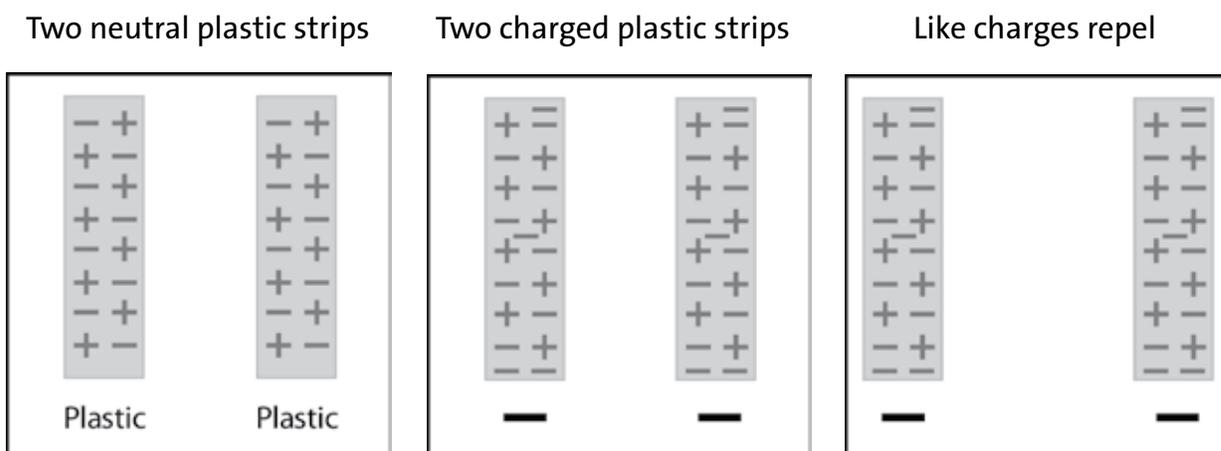
Ask students to make a prediction:

- What do you think will happen if you charge two strips of plastic and bring them near each other?

Procedure, part 3

2 pieces of charged plastic

1. Charge two strips of plastic
2. Slowly bring the two strips of plastic near each other.
3. Write “attract” or “repel” in the chart on the activity sheet.



Expected results

The strips will move away or repel each other. Since both strips have extra electrons on them, they each have extra negative charge. Since the same charges repel one another, the strips move away from each other.

Ask students:

- **What happened when you brought the two pieces of plastic near each other?**
The ends of the strips moved away from each other.
- **Use what you know about electrons and charges to explain why this happens.**
Each strip has extra electrons so they are both negatively charged. Because like charges repel, the pieces of plastic repelled each other.

EXTEND

7. Have students apply their understanding of protons and electrons to explain what happens when a charged balloon is brought near pieces of paper.

Materials for each group

- Inflated balloon
- Small pieces of paper, confetti-size

Procedure

- Rub a balloon on your hair or clothes.
- Bring the balloon slowly toward small pieces of paper.

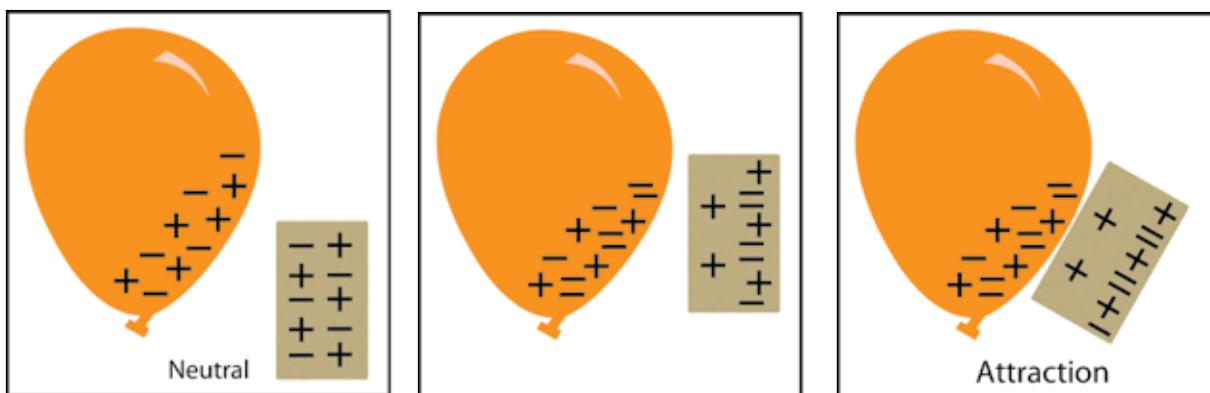
Expected results

The pieces of paper will jump up and stick on the balloon.

Ask students:

- **What did you observe when the charged balloon was held near the pieces of paper?**
The paper pieces moved up and stuck on the balloon.
- **Use what you know about electrons, protons, and charges to explain why this happens.**

When you rub the balloon on your hair or clothes it picks up extra electrons, giving the balloon a negative charge. When you bring the balloon near the paper, the electrons from the balloon repel the electrons in the paper. Since more protons are at the surface of the paper, it has a positive charge. The electrons are still on the paper, just not at the surface, so overall the paper is neutral. Opposites attract, so the paper moves up toward the balloon.



Show the simulation *Balloons and Static Electricity* from the University of Colorado at Boulder's Physics Education Technology site.

http://phet.colorado.edu/simulations/sims.php?sim=Balloons_and_Static_Electricity

In the simulation, check the boxes “show all charges” and “wall”. Uncheck everything else. In this simulation, you can rub the balloon a little bit on the sweater and see that some of the electrons from the sweater move onto the balloon. This gives the balloon a negative charge. Since the sweater lost some electrons, it has more protons than electrons, so it has a positive charge. If you move the balloon toward the sweater, it will be attracted. This is like moving the charged plastic strip toward the cloth it was rubbed on.

You can also move the balloon toward the wall. The excess negative charge on the balloon repels negative charge on the surface of the wall. This leaves more positive charge on the surface of the wall. The negatively charged balloon is attracted to the positive area on the wall. This is like moving the charged plastic strip toward the finger.

EXTRA EXTEND

8. Demonstrate how electrons can attract a stream of water.

Either do the following demonstration or show the video *Balloon and Water*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson1#balloon_and_water

Materials for the demonstration

- Sink
- Balloon

Procedure

1. Rub a balloon on your shirt or pants to give it a static charge.
2. Turn on the faucet so that there is a very thin stream of water.
3. Slowly bring the charged part of the balloon close to the stream of water.

Expected results

The stream of water should bend as it is attracted to the balloon.

Ask students:

- **What did you observe when the charged balloon was held near the stream of water?**

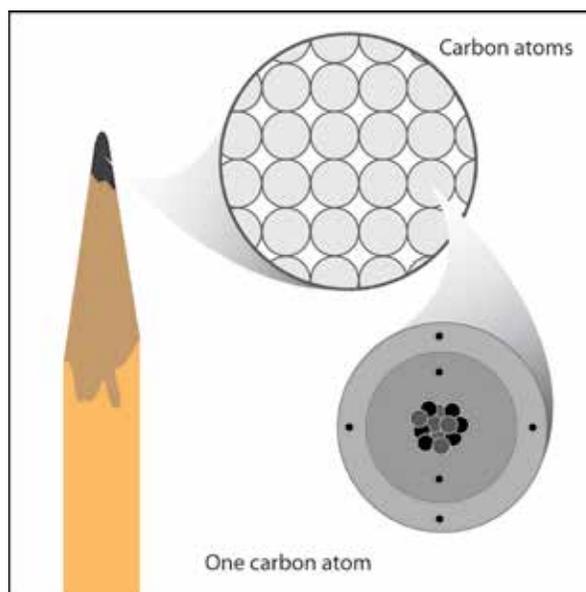
The stream of water bent toward the balloon.

- **Use what you know about electrons, protons, and charges to explain why this happens.**

When you rub the balloon on your hair or clothes it picks up extra electrons, giving the balloon a negative charge. When you bring the balloon near the stream of water, the electrons from the balloon repel the electrons in the water. Since more protons are at the surface of the water, it has a positive charge. Opposites attract, so the water moves toward the balloon.

INTRODUCTION

If you look closely at the tip of a sharpened pencil, you will see that it is made of graphite. Going deeper, graphite is made of carbon atoms. Deeper still, each carbon atom is made of protons, neutrons, and electrons. In this lesson, you will explore these subatomic particles and their charges.



1. Label the nucleus (protons, neutrons) and electrons in the drawing of a carbon atom above.
2. Draw a line between the subatomic particle and its charge.

proton
electron
neutron

no charge
positive charge
negative charge

3. Would the following subatomic particles attract each other or repel one another?

Two protons _____

Two electrons _____

A proton and an electron _____

ACTIVITY

Question to investigate

What makes objects attract or repel each other?

Materials for each group

- Plastic grocery bag
- Scissors

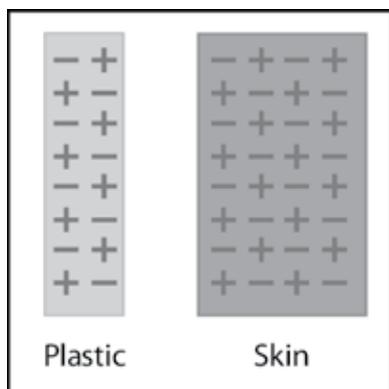
Procedure, part 1

Charged plastic and charged skin

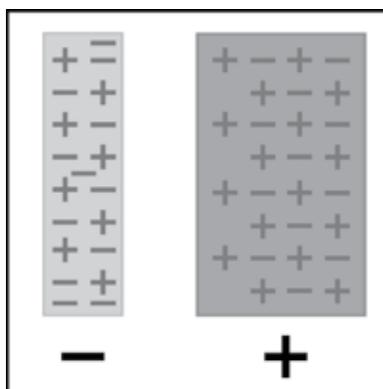
1. Cut 2 strips from a plastic grocery bag so that each is about 2–4 cm wide and about 20 cm long.
2. Hold the plastic strip firmly at one end. Then grasp the plastic strip between the thumb and fingers of your other hand as shown.
3. Quickly pull your top hand up so that the plastic strip runs through your fingers. Do this three or four times.
4. Allow the strip to hang down. Then bring your other hand near it.
5. Write “attract” or “repel” in the chart on page 256 to describe what happened.



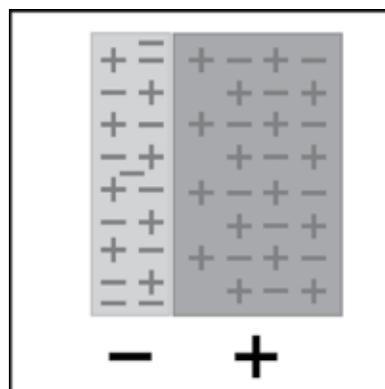
Protons and electrons
before rubbing



Protons and electrons
after rubbing



Opposites attract



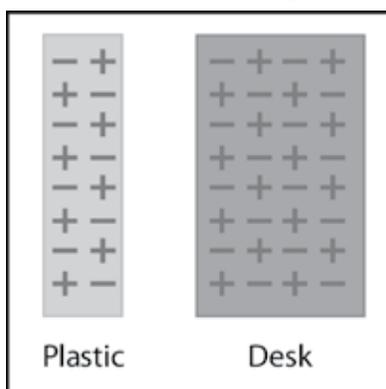
Procedure, part 2

Charged plastic and neutral desk

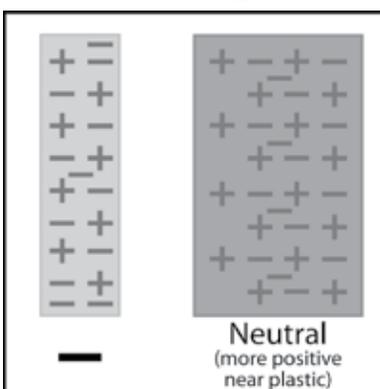
1. Charge one strip of plastic the same way you did previously.
2. This time, bring the plastic strip toward your desk or chair.
3. Write “attract” or “repel” in the chart on the next page.



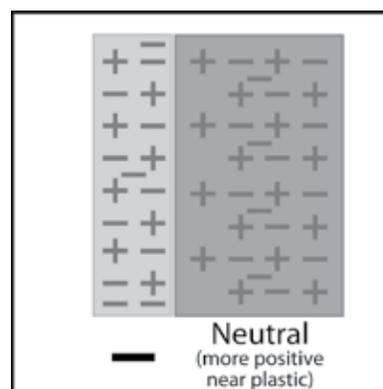
Protons and electrons
before rubbing



Protons and electrons after
rubbing



Opposites attract

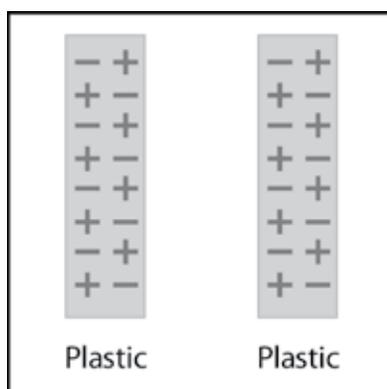


Procedure, part 3

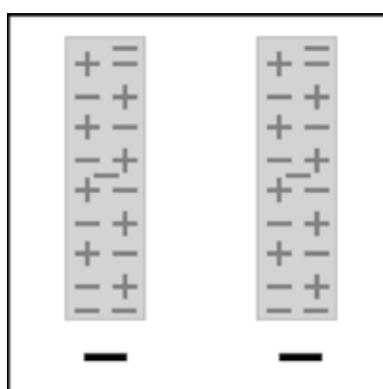
2 pieces of charged plastic

1. Charge two strips of plastic
2. Slowly bring the two strips of plastic near each other.
3. Write “attract” or “repel” in the chart on the next page.

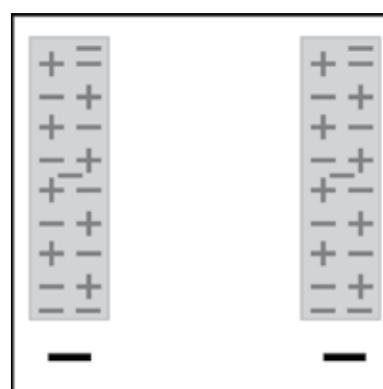
Two neutral plastic strips



Two charged plastic strips



Like charges repel



EXPLAIN IT WITH ATOMS & MOLECULES

What happened when you brought the following materials near each other?		
Materials	Attract or Repel	Use what you know about electrons, protons, and charges to explain your observations
charged plastic + charged skin		
charged plastic + neutral desk		
charged plastic + charged plastic		

TAKE IT FURTHER

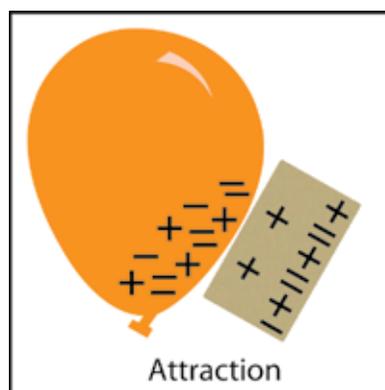
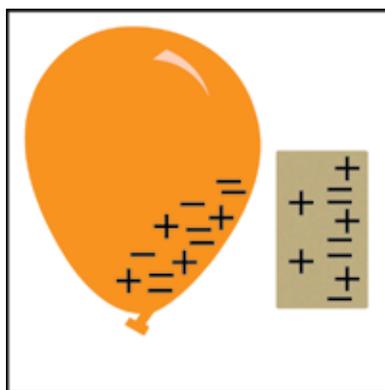
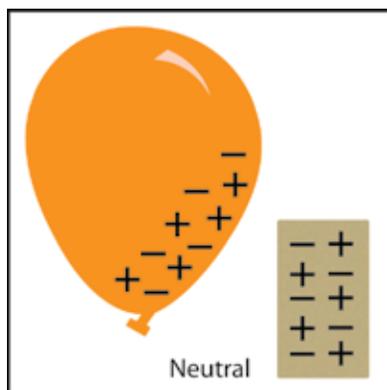
Materials for each group

- Inflated balloon
- Small pieces of paper, confetti-size

Procedure

- Rub a balloon on your hair or clothes.
- Bring the balloon slowly toward small pieces of paper.

4. Write captions beneath each picture explaining what happened between the balloon and your hair and the balloon and the paper in the activity.



Chapter 4, Lesson 2: The Periodic Table

Key Concepts

- The periodic table is a chart containing information about the atoms that make up all matter.
- An element is a substance made up of only one type of atom.
- The atomic number of an atom is equal to the number of protons in its nucleus.
- The number of electrons surrounding the nucleus of an atom is equal to the number of protons in its nucleus.
- Different atoms of the same element can have a different number of neutrons.
- Atoms of the same element with different numbers of neutrons are called “isotopes” of that element.
- The atomic mass of an element is the average mass of the different isotopes of the element.
- The atoms in the periodic table are arranged to show characteristics and relationships between atoms and groups of atoms.

Summary

Students will begin to look closely at the periodic table. They will be introduced to the basic information given for the elements in most periodic tables: the name, symbol, atomic number, and atomic mass for each element. Students will focus on the first 20 elements. They will try to correctly match cards with information about an element to each of the first 20 elements. Students will then watch several videos of some interesting chemical reactions involving some of these elements.

Objective

Students will identify different atoms by the number of protons in the nucleus and realize that the number of electrons equals the number of protons in a neutral atom. They will also be able to explain the meaning of atomic number and atomic mass.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

About this Lesson

Lessons 2 and 3 both use the 20 atom description cards beginning on page 240.

Teacher preparation

Print out the 20 pages of element cards. The first page is shown. Laminate each page and cut out the cards. For Lesson 2, you will need the 5 cards for each element from the left side of each sheet. You will also need the card in the upper right corner. This is the atom name card. Tape each of the 20 atom name cards to a spot in the room where students can place the cards that match that atom nearby. For Lesson 3, you will need the atom name card, taped in the same location in the room, and the four cards beneath it. Divide the class into 10 groups of 2 or 3 students each.

The atom you are looking for has: 1 Proton in its Nucleus.	Atomic Number 1 SYMBOL OF ELEMENT: H Name: H
The atom you are looking for has: 1 Electron surrounding its Nucleus.	The atom you are looking for has: 1 Electron surrounding its Nucleus.
The atom you are looking for has: 0 Neutrons (usually) in its Nucleus.	The atom you are looking for has: 0 Neutrons (usually) in its Nucleus.
The atom you are looking for has: 1 Fewer Proton than Helium (He).	The atom you are looking for has: 1 Fewer Proton than Helium (He).
The atom you are looking for has: 2 Fewer Electrons than Lithium (Li).	The atom you are looking for has: 2 Fewer Electrons than Lithium (Li).

ENGAGE

1. Introduce students to the periodic table.

Project the image *Periodic Table*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson2#periodic_table

Tell students that this is the periodic table. Explain that each box contains information about a different atom. The periodic table shows all the atoms that everything in the known universe is made from. It's kind of like the alphabet in which only 26 letters, in different combinations, make up many thousands of words. The 100 or so atoms of the periodic table, in different combinations, make up millions of different substances.

The Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01															10 Ne Neon 20.18	
11 Na Sodium 22.99	12 Mg Magnesium 24.31															18 Ar Argon 39.95	
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium 178.49	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)						
58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 173.04	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97				
90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)				

Note: It is often confusing for students to see the terms “atom” and “element” used interchangeably as if they are the same thing. Explain to students that an atom is the smallest particle or “building block” of a substance. An element is a substance made up of all the same type of atom. For instance, a piece of pure carbon is made up of only carbon atoms. This piece of pure carbon is a sample of the element carbon. The people who developed the periodic table could have called it the Periodic Table of the Atoms but they did not have a firm understanding of atoms at that time. Since they were working with actual samples of elements such as copper, mercury, sulfur, etc., they called it the periodic table of the elements.

Optional

Play one or both of the following songs.

- *The Elements* by Tom Lehrer with animation by Mike Stanfill
www.privatehand.com/flash/elements.html
- *Meet the Elements* by They Might be Giants
www.youtube.com/watch?v=d0zION8xjbM

2. Explain the meaning of the numbers and letters in the boxes in the periodic table.

Tell students that the class will focus on the first 20 elements over 2 days. On the first day, they will look at the number of protons, electrons, and neutrons in the atoms of each element. On the second day, they will look at the arrangement of electrons in the atoms.

Give each student a copy of the periodic table of the elements, the periodic table of elements 1–20, and the activity sheet.

Students will use the periodic table of elements 1–20, along with the activity sheet, in the lesson they will do today.

Project the image *Periodic Table of the First 20 Elements*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson2#first_twenty

PERIODIC TABLE ELEMENTS 1-20							
HYDROGEN 1 H 1.01							HELIUM 2 He 4.00
LITHIUM 3 Li 6.94	BERYLLIUM 4 Be 9.01	BORON 5 B 10.81	CARBON 6 C 12.01	NITROGEN 7 N 14.01	OXYGEN 8 O 16.00	FLUORINE 9 F 19.00	NEON 10 Ne 20.18
SODIUM 11 Na 22.99	MAGNESIUM 12 Mg 24.31	ALUMINUM 13 Al 26.98	SILICON 14 Si 28.09	PHOSPHORUS 15 P 30.97	SULFUR 16 S 32.07	CHLORINE 17 Cl 35.45	ARGON 18 Ar 39.95
POTASSIUM 19 K 39.10	CALCIUM 20 Ca 40.08						

Project the image *Element explanation*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson2#element_explanation

Explain what the numbers and letters in each box on the periodic table represent.

<p>Atomic number</p> <p>—</p> <p>The number of protons in the nucleus of the atom.</p>	<p>CARBON</p> <p>6</p> <p>C</p> <p>12.01</p>	<p>Element name</p> <p>Usually from a Greek or Latin word for the element or a substance containing the element.</p>
<p>Atomic mass</p> <p>—</p> <p>The average mass of the atoms in an element.</p>		<p>Symbol</p> <p>Short-hand abbreviation for the element name.</p>

Explain atomic mass.

The atomic mass of an element is based on the mass of the protons, neutrons, and electrons of the atoms of that element. The mass of the proton and neutron are about the same, but the mass of the electron is much smaller (about 1/2000 the mass of the proton or neutron). The majority of the atomic mass is contributed by the protons and neutrons.

For any element in the periodic table, the number of electrons in an atom of that element always equals the number of protons in the nucleus. But this is not true for neutrons. Atoms of the same element can have different numbers of neutrons than protons. Atoms of the same element with different numbers of neutrons are called *isotopes* of that element. The atomic mass in the periodic table is an average of the atomic mass of the isotopes of an element. For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

For example, the vast majority of carbon atoms have 6 protons and 6 neutrons, but a small percentage have 6 protons and 7 neutrons, and an even smaller percentage have 6 protons and 8 neutrons. Since the majority of carbon atoms have a mass very close to 12, and only a small percentage are greater than 12, the average atomic mass is slightly greater than 12.

3. Describe the activity students will do to learn about the first 20 elements of the periodic table.

Show students that you have 100 cards (5 for each of the first 20 elements). Explain that each card contains information about one of the first 20 atoms of the periodic table. The students' job is to read the card carefully, figure out which atom the card is describing, and put the card at the spot in the room for that atom.

Review the information about protons, electrons, and neutrons students need to know in order to match the cards with the correct element:

Proton

- Positively charged particle in the nucleus of the atom.
- The number of protons in an atom's nucleus is the atomic number.

Electron

- Negatively charged particle surrounding the nucleus of the atom.
- The number of electrons surrounding the nucleus of an atom is equal to the number of protons in the atom's nucleus.

Neutron

- Particle in the nucleus that has almost the same mass as a proton but has no charge.
- For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

To match the number of neutrons listed on your card to the correct element, look for an element on the periodic table so that if you add the number of neutrons on your card to the protons of the element, you will get close to the atomic mass for that element.

For example, you may have a card that says that the atom you are looking for has 5 neutrons. You would look at the periodic table to find an atom that you could add 5 to its number of protons that would give you a sum close to the atomic mass given for that element. The answer is beryllium (Be), which has 4 protons and an atomic mass of 9.01.

Note: There are a few neutron cards that have two possible correct elements instead of just one:

- 6 Neutrons—Boron or Carbon
- 10 Neutrons—Fluorine or Neon
- 12 Neutrons—Sodium or Magnesium
- 16 Neutrons—Phosphorous or Sulfur
- 20 Neutrons—Potassium or Calcium

EXPLORE

4. Have groups work together to place each card with its correct atom.

Distribute the cards to groups. If you have 10 groups, each group will get 10 cards. Be available to help students who have trouble with the neutrons and atomic mass.



5. Discuss the placement of the cards for two or three atoms.

Select two or three atoms and review whether the cards were placed correctly. This review will help reinforce the concepts about the structure of atoms and help students determine the number of protons, electrons, and neutrons in each type of atom.

Have students begin filling out the activity sheet with the following information:

- Number of protons
- Number of electrons
- Number of neutrons (usually)

PERIODIC TABLE ELEMENTS 1-20

Write the number of protons, electrons, and neutrons in each element.

HYDROGEN 1 # of Protons: 1 # of Electrons: 1 # of Neutrons: 0 1.01							HELIUM 2 # of Protons: 2 # of Electrons: 2 # of Neutrons: 2 4.00
LITHIUM 3 # of Protons: 3 # of Electrons: 3 # of Neutrons: 4 6.94	BERYLLIUM 4 # of Protons: 4 # of Electrons: 4 # of Neutrons: 5 9.01	BORON 5 # of Protons: 5 # of Electrons: 5 # of Neutrons: 6 10.81	CARBON 6 # of Protons: 6 # of Electrons: 6 # of Neutrons: 6 12.01	NITROGEN 7 # of Protons: 7 # of Electrons: 7 # of Neutrons: 7 14.01	OXYGEN 8 # of Protons: 8 # of Electrons: 8 # of Neutrons: 8 16.00	FLUORINE 9 # of Protons: 9 # of Electrons: 9 # of Neutrons: 10 19.00	NEON 10 # of Protons: 10 # of Electrons: 10 # of Neutrons: 10 20.18
SODIUM 11 # of Protons: 11 # of Electrons: 11 # of Neutrons: 12 22.99	MAGNESIUM 12 # of Protons: 12 # of Electrons: 12 # of Neutrons: 12 24.31	ALUMINUM 13 # of Protons: 13 # of Electrons: 13 # of Neutrons: 14 26.98	SILICON 14 # of Protons: 14 # of Electrons: 14 # of Neutrons: 14 28.09	PHOSPHORUS 15 # of Protons: 15 # of Electrons: 15 # of Neutrons: 16 30.97	SULFUR 16 # of Protons: 16 # of Electrons: 16 # of Neutrons: 16 32.07	CHLORINE 17 # of Protons: 17 # of Electrons: 17 # of Neutrons: 18 35.45	ARGON 18 # of Protons: 18 # of Electrons: 18 # of Neutrons: 22 39.95
POTASSIUM 19 # of Protons: 19 # of Electrons: 19 # of Neutrons: 20 39.10	CALCIUM 20 # of Protons: 20 # of Electrons: 20 # of Neutrons: 20 40.08	<p><i>Note: The number of neutrons may be different in the atoms of the same element. The atoms of an element with different numbers of neutrons are called isotopes of that element. The number of neutrons shown in the chart represents the most common isotope for that element.</i></p>					

EXTEND

6. Introduce students to their element project and an online resource that they can use.

Assign each student to an element. Include the first 20 elements and any other elements that you find interesting so that each student can research and present their own.

Each student should find and present some basic information about their element to the class. The presentation can be in the form of a poster, pamphlet, PowerPoint presentation or other form. The presentations should be short and can include: atom name, atomic number, derivation of name, when and where discovered, natural sources of the element, major uses, and any other information you find important.

Some Internet sources for this information can be overwhelming. They can also contain advertising that you may not want students exploring. For basic information about the periodic table, including some images and video, *The Journal of Chemical Education's Periodic Table Live* is an excellent resource.

www.chemeddl.org/collections/ptl/

If there is time available, have students work on this atom project during the week.

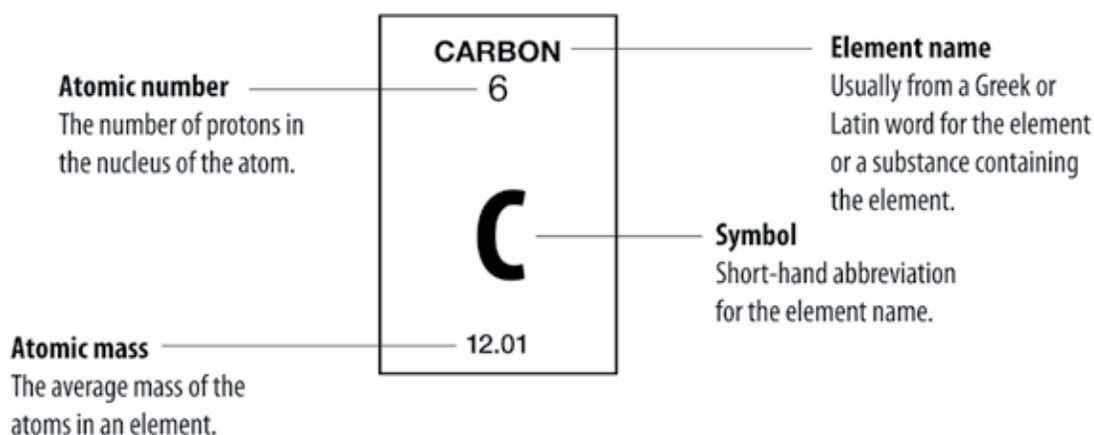
Activity Sheet
Chapter 4, Lesson 2
The Periodic Table

Name _____

Date _____

Your group will receive a set of cards with information that describes a particular atom. Your job is to figure out which atom the card describes and to place it in the area in your classroom for that atom.

You will use the Periodic Table, Elements 1–20 chart to help you determine what atom your card describes. The diagram and information below will help you match your cards to the correct atoms.



Parts of an Atom

Proton

Positively charged particle in the nucleus of the atom.

The number of protons in an atom's nucleus is the atomic number.

Electron

Negatively charged particle surrounding the nucleus of the atom.

The number of electrons surrounding the nucleus of an atom is equal to the number of protons in the atom's nucleus.

Neutron

Particle in the nucleus that has about the same mass as a proton but has no charge. For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

Placing your cards

Once you know what the information in each box on your periodic table stands for and you know the parts of the atom, you will be able to correctly place most of your cards with the atoms they describe. You will need to know the following additional information in order to answer any question having to do with neutrons.



To match the number of neutrons listed on your card to the correct element, look for an element on the periodic table so that if you add the number of neutrons on your card to the protons of the element, you will get close to the atomic mass for that element.

For example, you may have a card that says, “The atom you are looking for has 5 neutrons.” Look at the periodic table to find an atom that you could add 5 to its number of protons that would give you a sum close to the atomic mass given for that element. The answer is beryllium (Be), which has 4 protons and an atomic mass of 9.01.

PERIODIC TABLE ELEMENTS 1-20

Write the number of protons, electrons, and neutrons in each element.

HYDROGEN 1 # of Protons: # of Electrons: # of Neutrons: 1.01	HELIUM 2 # of Protons: # of Electrons: # of Neutrons: 4.00					
LITHIUM 3 # of Protons: # of Electrons: # of Neutrons: 6.94	BORON 5 # of Protons: # of Electrons: # of Neutrons: 10.81	CARBON 6 # of Protons: # of Electrons: # of Neutrons: 12.01	NITROGEN 7 # of Protons: # of Electrons: # of Neutrons: 14.01	OXYGEN 8 # of Protons: # of Electrons: # of Neutrons: 16.00	FLUORINE 9 # of Protons: # of Electrons: # of Neutrons: 19.00	NEON 10 # of Protons: # of Electrons: # of Neutrons: 20.18
SODIUM 11 # of Protons: # of Electrons: # of Neutrons: 22.99	MAGNESIUM 12 # of Protons: # of Electrons: # of Neutrons: 24.31	ALUMINIUM 13 # of Protons: # of Electrons: # of Neutrons: 26.98	SILICON 14 # of Protons: # of Electrons: # of Neutrons: 28.09	SULFUR 16 # of Protons: # of Electrons: # of Neutrons: 32.07	CHLORINE 17 # of Protons: # of Electrons: # of Neutrons: 35.45	ARGON 18 # of Protons: # of Electrons: # of Neutrons: 39.95
POTASSIUM 19 # of Protons: # of Electrons: # of Neutrons: 39.10	CALCIUM 20 # of Protons: # of Electrons: # of Neutrons: 40.08					

Note: Remember that the number of neutrons is not the same for every atom of an element. The number of neutrons you write in this chart will be a number, that when added to the number of protons, gives a sum as close as possible to the atomic mass.

The Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00										
3 Li Lithium 6.94	4 Be Beryllium 9.01															5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18						
11 Na Sodium 22.99	12 Mg Magnesium 24.31															13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95						
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80										
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29										
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)										
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)																
<div style="display: flex; align-items: center; justify-content: center;"> <div style="border: 1px solid black; padding: 5px; margin-right: 10px;"> 3 Li Lithium 6.94 </div> <div style="margin-right: 10px;"> _____ Atomic Number _____ Element Symbol _____ Element Name </div> <div style="margin-right: 10px;"> _____ Average Atomic Mass </div> </div>																											
58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)

The atom you are looking for has

1

Proton in its Nucleus.

Atomic Number **1**
HYDROGEN (H)
Atomic Mass 1.01

H

The atom you are looking for has

1

Electron surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

0

Neutrons (usually)
in its Nucleus.

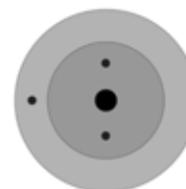
The atom you are looking for has
1 Electron on the **First** Energy Level.

The atom you are looking for has

1

fewer Proton
than Helium (He).

The atom you are looking for is
directly above the atom with this Energy Level.



The atom you are looking for has

2

fewer Electrons
than Lithium (Li).

The atom you are looking for is the only atom with only
1 Electron in the **First** Energy Level.

The atom you are looking for has

2

Protons in its Nucleus.

Atomic Number **2**
HELIUM (He)
Atomic Mass 4.00

He

The atom you are looking for has

2

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

2

Neutrons (usually)
in its Nucleus.

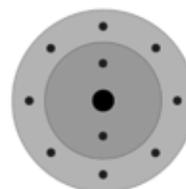
The atom you are looking for has
2 Electrons on the **First** Energy Level and
no other electrons.

The atom you are looking for has

1

more Proton
than Hydrogen (H).

The atom you are looking for is
directly above the atom with this Energy Level.



The atom you are looking for has

2

fewer Electrons
than Beryllium (Be).

The atom you are looking for is the only atom with only
2 Electrons in the First Energy Level and
no other electrons on any other level.

The atom you are looking for has

3

Protons in its Nucleus.

Atomic Number **3**

LITHIUM (Li)

Atomic Mass 6.94

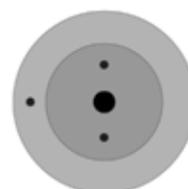
Li

The atom you are looking for has

3

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

4

Neutrons (usually)
in its Nucleus.

The atom you are looking for has
2 Electrons on the **First** Energy Level and
1 Electron on the **Second** Energy Level.

The atom you are looking for has

3

fewer Protons
than Carbon (C).

The atom you are looking for is
directly below the atom with this Energy Level.

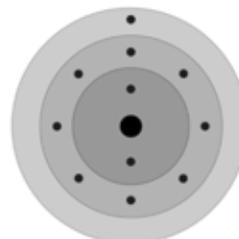


The atom you are looking for has

2

fewer Electrons
than Boron.

The atom you are looking for is
directly above the atom with this Energy Level.



The atom you are looking for has

4

Protons in its Nucleus.

Atomic Number **4**

BERYLLIUM (Be)

Atomic Mass 9.01

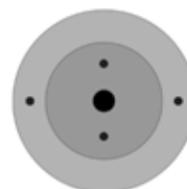
Be

The atom you are looking for has

4

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

5

Neutrons (usually)
in its Nucleus.

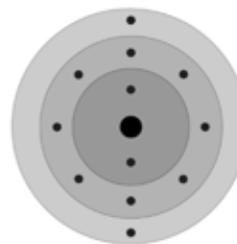
The atom you are looking for has
2 Electrons on the **First** Energy Level and
2 Electrons on the **Second** Energy Level.

The atom you are looking for has

4

fewer Protons
than Oxygen (O).

The atom you are looking for is
directly above the atom with this Energy Level.

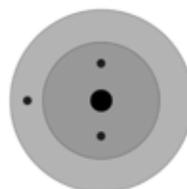


The atom you are looking for has

3

fewer Electrons
than Nitrogen (N).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

5

Protons in its Nucleus.

Atomic Number **5**

Boron (B)

Atomic Mass 10.81

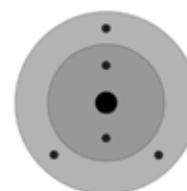
B

The atom you are looking for has

5

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

6

Neutrons (usually)
in its Nucleus.

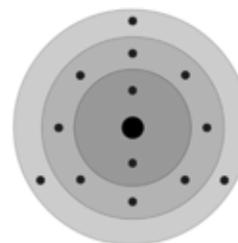
The atom you are looking for has
2 Electrons on the **First** Energy Level and
3 Electrons on the **Second** Energy Level.

The atom you are looking for has

4

more Protons
than Hydrogen (H).

The atom you are looking for is
directly above the atom with this Energy Level.

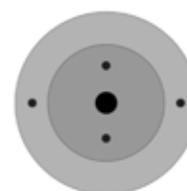


The atom you are looking for has

3

more Electrons
than Helium (He).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

6

Protons in its Nucleus.

Atomic Number **6**

Carbon (C)

Atomic Mass 12.01

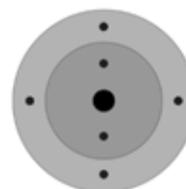
C

The atom you are looking for has

6

Electrons surrounding its Nucleus.

The atom you are looking for has this Energy Level Model:



The atom you are looking for has

6

Neutrons (usually) in its Nucleus.

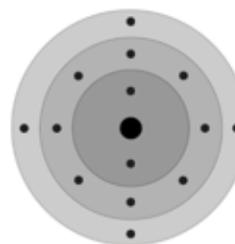
The atom you are looking for has **2** Electrons on the **First** Energy Level and **4** Electrons on the **Second** Energy Level.

The atom you are looking for has

1

more Proton than Boron (B).

The atom you are looking for is **directly above** the atom with this Energy Level.

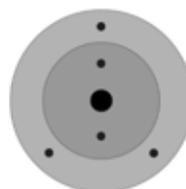


The atom you are looking for has

3

more Electrons than Lithium (Li).

The atom you are looking for is **directly to the right** of the atom with this Energy Level.



The atom you are looking for has

7

Protons in its Nucleus.

Atomic Number **7**
Nitrogen (N)
Atomic Mass 14.01

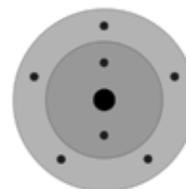
N

The atom you are looking for has

7

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

7

Neutrons (usually)
in its Nucleus.

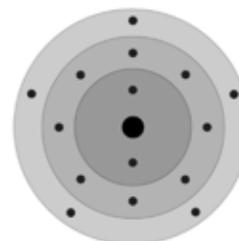
The atom you are looking for has
2 Electrons on the **First** Energy Level and
5 Electrons on the **Second** Energy Level.

The atom you are looking for has

3

fewer Protons
than Neon (Ne).

The atom you are looking for is
directly above the atom with this Energy Level.

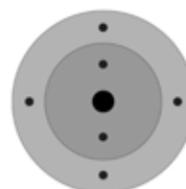


The atom you are looking for has

1

fewer Electron
than Oxygen (O).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

8

Protons in its Nucleus.

Atomic Number **8**

Oxygen (O)

Atomic Mass 16.00

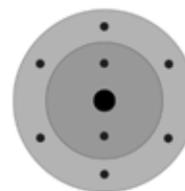
O

The atom you are looking for has

8

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

8

Neutrons (usually)
in its Nucleus.

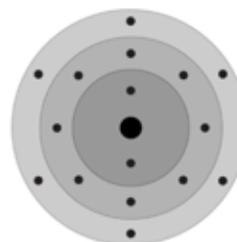
The atom you are looking for has
2 Electrons on the **First** Energy Level and
6 Electrons on the **Second** Energy Level.

The atom you are looking for has

2

more Protons
than Carbon (C).

The atom you are looking for is
directly above the atom with this Energy Level.

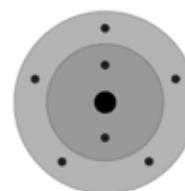


The atom you are looking for has

6

more Electrons
than Helium (He).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

9

Protons in its Nucleus.

Atomic Number **9**
Fluorine (F)
Atomic Mass 18.99

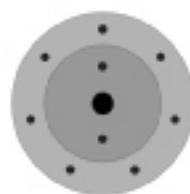
F

The atom you are looking for has

9

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

10

Neutrons (usually)
in its Nucleus.

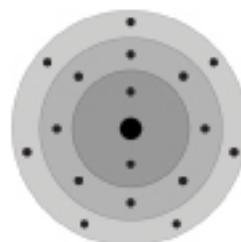
The atom you are looking for has
2 Electrons on the **First** Energy Level and
7 Electrons on the **Second** Energy Level.

The atom you are looking for has

1

fewer Proton
than Neon (Ne).

The atom you are looking for is
directly above the atom with this Energy Level.

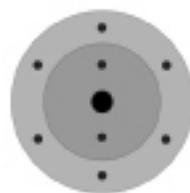


The atom you are looking for has

2

more Electrons
than Nitrogen (N).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

10

Protons in its Nucleus.

Atomic Number **10**
Neon (Ne)
Atomic Mass 20.18

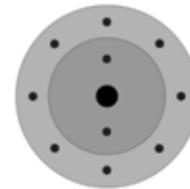
Ne

The atom you are looking for has

10

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

10

Neutrons (usually)
in its Nucleus.

The atom you are looking for has
2 Electrons on the **First** Energy Level and
8 Electrons on the **Second** Energy Level.

The atom you are looking for has

8

more Protons
than Helium (He).

The atom you are looking for is
directly below the atom with this Energy Level.

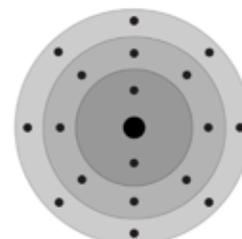


The atom you are looking for has

2

more Electrons
than Oxygen (O).

The atom you are looking for is
directly above
of the atom with this Energy Level.



The atom you are looking for has

11

Protons in its Nucleus.

Atomic Number **11**

Sodium (Na)

Atomic Mass 22.99

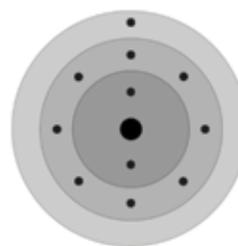
Na

The atom you are looking for has

11

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

12

Neutrons (usually)
in its Nucleus.

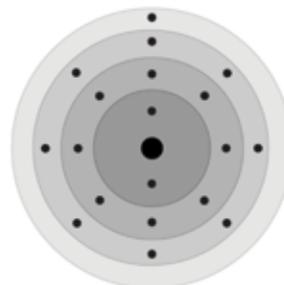
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
1 Electron on the **Third** Energy Level.

The atom you are looking for has

2

fewer Protons
than Aluminum (Al).

The atom you are looking for is
directly above the atom with this Energy Level.

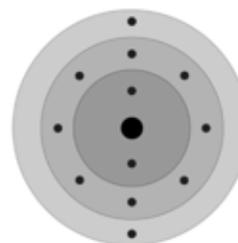


The atom you are looking for has

3

more Electrons
than Oxygen (O).

The atom you are looking for is
directly to the left
of the atom with this Energy Level.



The atom you are looking for has

12

Protons in its Nucleus.

Atomic Number 12
Magnesium (Mg)
Atomic Mass 24.31

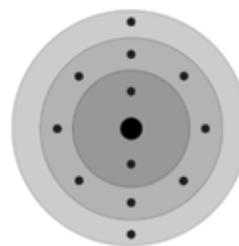
Mg

The atom you are looking for has

12

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

12

Neutrons (usually)
in its Nucleus.

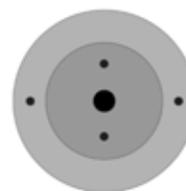
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
2 Electrons on the **Third** Energy Level.

The atom you are looking for has

10

more Protons
than Helium (He).

The atom you are looking for is
directly below the atom with this Energy Level.

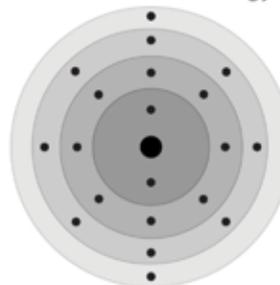


The atom you are looking for has

8

more Electrons
than Beryllium (Be).

The atom you are looking for is
directly above
of the atom with this Energy Level.



The atom you are looking for has

13

Protons in its Nucleus.

Atomic Number **13**
Aluminum (Al)
Atomic Mass 26.98

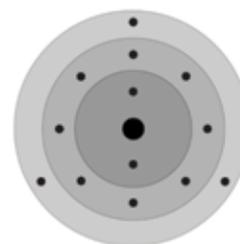
Al

The atom you are looking for has

13

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

14

Neutrons (usually)
in its Nucleus.

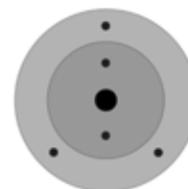
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
3 Electrons on the **Third** Energy Level.

The atom you are looking for has

8

more Protons
than Boron (B).

The atom you are looking for is
directly below the atom with this Energy Level.

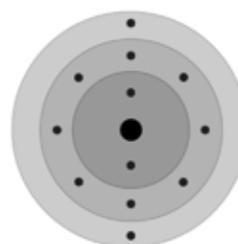


The atom you are looking for has

5

more Electrons
than Oxygen (O).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

14

Protons in its Nucleus.

Atomic Number **14**

Silicon (Si)

Atomic Mass 28.09

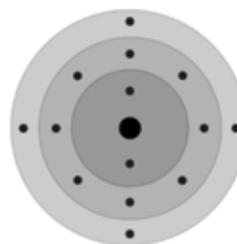
Si

The atom you are looking for has

14

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

14

Neutrons (usually)
in its Nucleus.

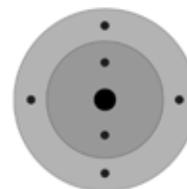
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
4 Electrons on the **Third** Energy Level.

The atom you are looking for has

3

fewer Protons
than Chlorine (Cl).

The atom you are looking for is
directly below the atom with this Energy Level.

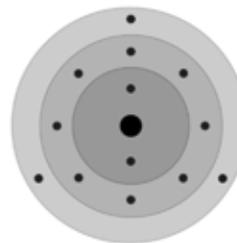


The atom you are looking for has

2

more Electrons
than Magnesium (Mg).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

15

Protons in its Nucleus.

Atomic Number **15**
Phosphorous (P)
Atomic Mass 30.97

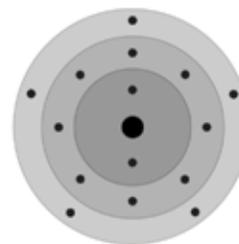
P

The atom you are looking for has

15

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

16

Neutrons (usually)
in its Nucleus.

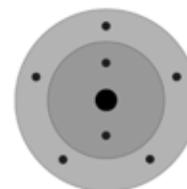
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
5 Electrons on the **Third** Energy Level.

The atom you are looking for has

8

more Protons
than Nitrogen (N).

The atom you are looking for is
directly below the atom with this Energy Level.

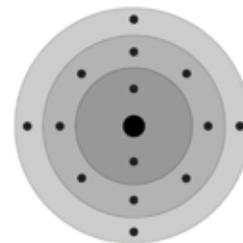


The atom you are looking for has

3

fewer Electrons
than Argon (Ar).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

16

Protons in its Nucleus.

Atomic Number **16**

Sulfur (S)

Atomic Mass 32.07

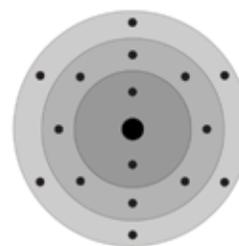
S

The atom you are looking for has

16

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

16

Neutrons (usually)
in its Nucleus.

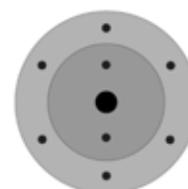
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
6 Electrons on the **Third** Energy Level.

The atom you are looking for has

10

more Protons
than Carbon (C).

The atom you are looking for is
directly below the atom with this Energy Level.

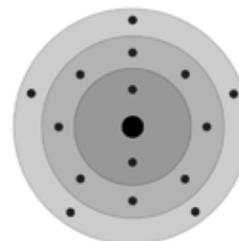


The atom you are looking for has

6

more Electrons
than Neon (Ne).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

17

Protons in its Nucleus.

Atomic Number **17**

Chlorine (Cl)

Atomic Mass 35.45

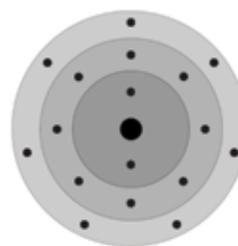
Cl

The atom you are looking for has

17

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

18

Neutrons (usually)
in its Nucleus.

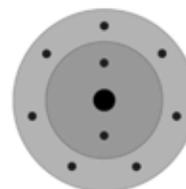
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
7 Electrons on the **Third** Energy Level.

The atom you are looking for has

3

fewer Protons
than Calcium (Ca).

The atom you are looking for is
directly below the atom with this Energy Level.

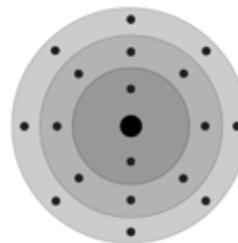


The atom you are looking for has

8

more Electrons
than Fluorine (F).

The atom you are looking for is
directly to the left
of the atom with this Energy Level.



The atom you are looking for has

18

Protons in its Nucleus.

Atomic Number **18**

Argon (Ar)

Atomic Mass 39.95

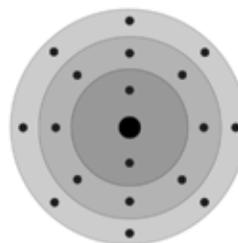
Ar

The atom you are looking for has

18

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

22

Neutrons (usually)
in its Nucleus.

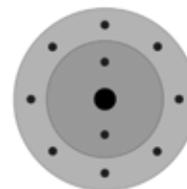
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level, and
8 Electrons on the **Third** Energy Level.

The atom you are looking for has

7

more Protons
than Sodium (Na).

The atom you are looking for is
directly below the atom with this Energy Level.

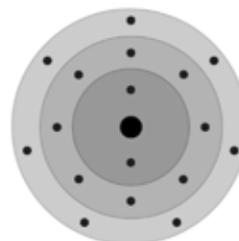


The atom you are looking for has

8

more Electrons
than Neon (Ne).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



The atom you are looking for has

19

Protons in its Nucleus.

Atomic Number **19**

Potassium (K)

Atomic Mass 39.10

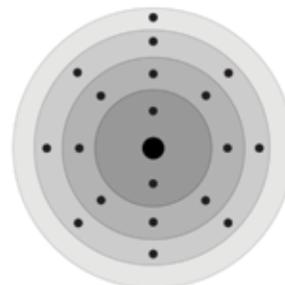
K

The atom you are looking for has

19

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

20

Neutrons (usually)
in its Nucleus.

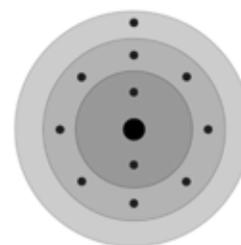
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level,
8 Electrons on the **Third** Energy Level, and
1 Electron on the **Fourth** Energy Level.

The atom you are looking for has

4

more Protons
than Phosphorous (P).

The atom you are looking for is
directly below the atom with this Energy Level.

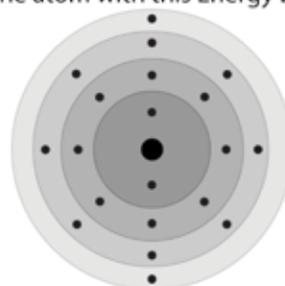


The atom you are looking for has

18

more Electrons
than Hydrogen (H).

The atom you are looking for is
directly to the left
of the atom with this Energy Level.



The atom you are looking for has

20

Protons in its Nucleus.

Atomic Number **20**

Calcium (Ca)

Atomic Mass 40.08

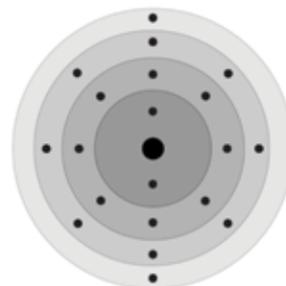
Ca

The atom you are looking for has

20

Electrons surrounding
its Nucleus.

The atom you are looking for has this
Energy Level Model:



The atom you are looking for has

20

Neutrons (usually)
in its Nucleus.

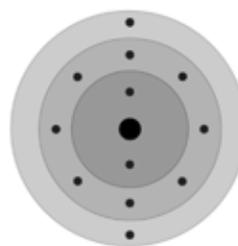
The atom you are looking for has
2 Electrons on the **First** Energy Level,
8 Electrons on the **Second** Energy Level,
8 Electrons on the **Third** Energy Level, and
2 Electrons on the **Fourth** Energy Level.

The atom you are looking for has

8

more Protons
than Magnesium (Mg).

The atom you are looking for is
directly below the atom with this Energy Level.

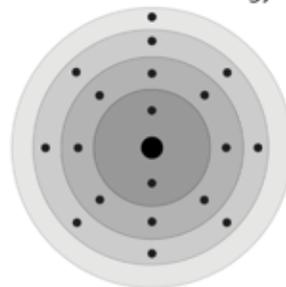


The atom you are looking for has

2

more Electrons
than Argon (Ar).

The atom you are looking for is
directly to the right
of the atom with this Energy Level.



Chapter 4, Lesson 3: The Periodic Table and Energy-Level Models

Key Concepts

- The electrons surrounding an atom are located in regions around the nucleus called “energy levels”.
- An energy level represents the 3-dimensional space surrounding the nucleus where electrons are most likely to be.
- The first energy level is closest to the nucleus. The second energy level is a little farther away than the first. The third is a little farther away than the second, and so on.
- Each energy level can accommodate or “hold” a different number of electrons before additional electrons begin to go into the next level.
- When the first energy level has 2 electrons, the next electrons go into the second energy level until the second level has 8 electrons.
- When the second energy level has 8 electrons, the next electrons go into the third energy level until the third level has 8 electrons.
- When the third energy level has 8 electrons, the next 2 electrons go into the fourth energy level.
- The electrons in the energy level farthest from the nucleus are called *valence* electrons.
- Atoms in the same column (group) in the periodic table have the same number of valence electrons.

Summary

Students will again focus on the first 20 elements. Students will first look at a diagram and animation to understand the basic pattern of the arrangement of electrons on energy levels around an atom. Students will be given cards with information about the electrons and energy levels for each of the first 20 atoms. They will again try to correctly match the cards with each element.

Objective

Students will be able to interpret the information given in the periodic table to describe the arrangement of electrons on the energy levels around an atom.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

About this Lesson

Be sure that the 20 atom name cards are posted around the room. You will need the five cards on the right hand side of each sheet. This lesson is intended as a follow-up to chapter 4, lesson 2.

ENGAGE

1. Introduce students to the idea that electrons surround the nucleus of an atom in regions called energy levels.

Review with students that in lesson two they focused on the number of protons, neutrons, and electrons in the atoms in each element. In this lesson, they will focus on the arrangement of the electrons in each element.

Project the image *Energy level cross-section*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#energy_level_cross_section

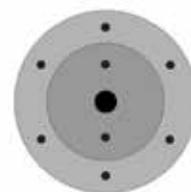
Explain to students that electrons surround the nucleus of an atom in three dimensions, making atoms spherical. They can think of electrons as being in the different energy levels like concentric spheres around the nucleus. Since it is very difficult to show these spheres, the energy levels are typically shown in 2 dimensions.



Project the image *Oxygen atom*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#oxygen_atom

Tell students that this energy level model represents an atom. The nucleus is represented by a dot in the center, which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons in the energy levels. Let students know that they will learn more about electrons and energy levels later in this lesson.



Have students look at the *Periodic table of the elements 1–20* they used in lesson 2 to answer the following question:

- **Can you identify which atom this model represents?**

If students can't answer this question, point out that there are 8 electrons. Because neutral atoms in the periodic table have the same number of electrons as protons, the atom must have 8 protons. The number of protons is the same as the atomic number, so the atom is oxygen.

Read more about energy level models in the additional teacher background section at the end of this lesson.

2. Have groups work together to place each card with its correct atom.

Show students that you have 80 cards (4 for each of the first 20 elements). Before distributing the cards, explain that each card contains information about electrons and energy levels for the first 20 elements of the periodic table. The students' job is to read the card carefully, figure out which element the card is describing, and put the card at the spot in the room for that element. Remind students that they will need to count the electrons in order to identify each atom. Once students understand what their assignment is, distribute the cards to groups.



3. Discuss the placement of the cards for two or three atoms.

After all cards have been placed at the 20 different atoms, select two or three atoms and review whether the cards were placed correctly. This review will help reinforce the concepts about the structure of atoms and help students determine the number of protons and electrons in each atom.

Give each student a *Periodic Table of Energy Levels* activity sheet. This table contains energy level models for the first 20 elements. The electrons are included only for the atoms at the beginning and end of each period.



EXPLORE

4. Project the *Periodic table of energy levels* and discuss the arrangement of electrons as students complete their activity sheet.

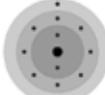
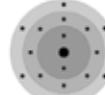
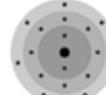
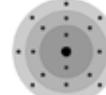
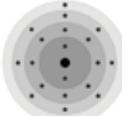
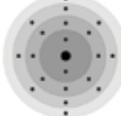
Project the image *Periodic table of energy levels*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#energy_levels

The image you project contains all of the electrons for elements 1–20. However, the periodic table on the activity sheet contains electrons only for the elements at the beginning and end of each period. Discuss the arrangement of electrons within the energy levels for these atoms and have students fill in the electrons for the other atoms.

Note: In the energy level diagrams, the electrons are spread out evenly in the level. Some books show them spread out this way and some show them in pairs. The pairing of electrons is meant to represent that electrons are in separate orbitals within each energy level. At the middle school

level, it is not necessary for students to learn about electron orbitals. This information is offered so that it is clearer to you why electrons are often shown in pairs in energy level diagrams and in the dot diagrams used as an extension at the end of this chapter. An orbital defines a region within an energy level where there is a high probability of finding a pair of electrons. There can be a maximum of two electrons in each orbital. This is why the electrons are often shown in pairs within an energy level.

ENERGY LEVELS ELEMENTS 1-20							
HYDROGEN 1  1.01							HELIUM 2  4.00
LITHIUM 3  6.94	BERYLLIUM 4  9.01	BORON 5  10.81	CARBON 6  12.01	NITROGEN 7  14.01	OXYGEN 8  16.00	FLUORINE 9  19.00	NEON 10  20.18
SODIUM 11  22.99	MAGNESIUM 12  24.31	ALUMINUM 13  26.98	SILICON 14  28.09	PHOSPHORUS 15  30.97	SULFUR 16  32.07	CHLORINE 17  35.45	ARGON 18  39.95
POTASSIUM 19  39.10	CALCIUM 20  40.08						

Tell students that the rows across on the periodic table are called *periods*.

Period 1

- *Hydrogen*
Explain that hydrogen has 1 proton and 1 electron. The 1 electron is on the first energy level.
- *Helium*
Explain that helium has 2 protons and 2 electrons. The 2 electrons are on the first energy level.

Period 2

- *Lithium*

Explain that lithium has 3 protons and 3 electrons. There are 2 electrons on the first energy level and 1 electron on the second. Explain that the first energy level can only have 2 electrons so the next electron in lithium is on the next (second) level.

- *Neon*

Explain that neon has 10 protons and 10 electrons. There are 2 electrons on the first energy level and 8 electrons on the second level.

- *Beryllium–fluorine*

Help students fill in the correct number of electrons in the energy levels for the rest of the atoms in period 2.

Period 3

- *Sodium*

Explain that sodium has 11 protons and 11 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, and 1 electron on the third energy level. Explain that the second energy level can only have 8 electrons so the next electron in sodium has to be on the next (third) level.

- *Argon*

Explain that argon has 18 protons and 18 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, and 8 electrons on the third energy level. Have students complete the energy level model for argon in their periodic table.

- *Magnesium–chlorine*

Help students fill in the correct number of electrons in the energy levels for the rest of the atoms in period 3.

Period 4

- *Potassium*

Explain that potassium has 19 protons and 19 electrons. There are 2 electrons on the first energy level, 8 electrons on the second level, 8 electrons on the third energy level, and 1 on the fourth energy level. Explain that after the third energy level has 8 electrons, the next electron goes into the fourth level.

- *Calcium*

Help students fill in the correct number of electrons in the energy levels for calcium.

Note: Students may wonder why an energy level can hold only a certain number of electrons. The answer to this is far beyond the scope of a middle school chemistry unit. It involves thinking of electrons as 3-dimensional waves and how they would interact with each other and the nucleus.

5. Have students look for patterns in rows and columns of the first 20 elements in the periodic table.

Continue to project the image *Periodic table of energy levels for elements 1–20* and have students look at their activity sheets to find patterns in the number of electrons within each energy level.

Have students look at the periods (rows going across).

Number of energy levels in each period

- The atoms in the first period have electrons in 1 energy level.
- The atoms in the second period have electrons in 2 energy levels.
- The atoms in the third period have electrons in 3 energy levels.
- The atoms in the fourth period have electrons in 4 energy levels.

How the electrons fill in the energy levels

- First energy level = 1, 2
- Second energy level = 1, 2, 3, ...8
- Third energy level = 1, 2, 3, ...8
- Fourth energy level = 1, 2

Read more about the periodic table in the additional teacher background section at the end of this lesson.

A certain number of electrons go into a level before the next level can have electrons in it. After the first energy level contains 2 electrons (helium), the next electrons go into the second energy level. After the second energy level has 8 electrons (neon), the next electrons go into the third energy level. After the third energy level has 8 electrons (argon), the next 2 electrons go into the fourth energy level.

Note: The third energy level can actually hold up to 18 electrons, so it is not really filled when it has 8 electrons in it. But when the third level contains 8 electrons, the next 2 electrons go into the fourth level. Then, believe it or not, 10 more electrons continue to fill up the rest of the third level. Students do not need to know this.

Have students look at the groups (columns going down).

Tell students that the vertical columns in the periodic table are called *groups* or *families*. Ask students to compare the number of electrons in the outermost energy level for the atoms in a group. Students should realize that each atom in a group has the same number of electrons in its outermost energy level. For instance, hydrogen, lithium, sodium, and potassium all have 1 electron on their outer energy level. Let students know that these electrons in the outermost energy level are called *valence* electrons. They are the electrons responsible for bonding, which students will investigate in the next lesson.

EXTEND

6. Compare the way different elements react chemically and relate this to their location on the periodic table.

Tell students that in the periodic table atoms in the same column, called a group, share certain characteristics and can react in a similar way.

Project the video *Sodium in water and potassium in water.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#sodium_in_water

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#potassium_in_water

Students will see that although potassium reacts more vigorously than sodium, the reactions are similar. Have students look at the periodic table to see where sodium and potassium are in relation to one another.

Project the video *Calcium in water.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#calcium_in_water

Students will see that this reaction is different from the sodium and the potassium. Have them locate calcium on the periodic table and point out that it is in a different group than sodium and potassium.

Project the videos *Sodium in acid and potassium in acid.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#sodium_in_acid

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#potassium_in_acid

Show sodium reacting with acid and then potassium reacting with acid. The HCl is hydrochloric acid. The HNO₃ is nitric acid. Each acid is used in two different concentrations. Make sure students realize that the sodium and potassium react in a similar way even though the potassium reacts more vigorously.

Project the video *Calcium in acid.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson3#calcium_in_acid

Point out that calcium reacts differently from the sodium and the potassium.

Ask students:

- **Do elements in the same group have similar properties and react in similar ways?**

Students should realize that sodium and potassium are in the same group and react similarly. Calcium is near them on the periodic table, but is in a different group, so it reacts differently.

Activity Sheet
Chapter 4, Lesson 3
The Periodic Table and Energy Level Models

Name _____

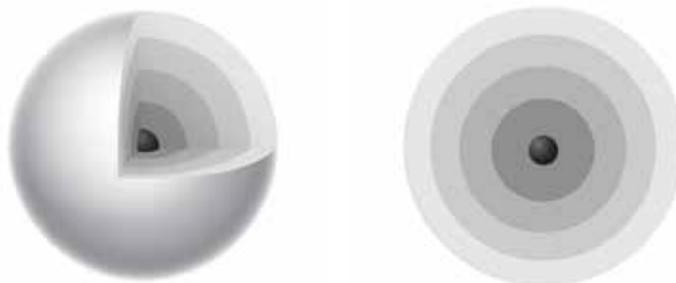
Date _____

Your group will receive a set of cards with information about the energy levels of a particular atom. Your job is to figure out which atom the card describes and to place it in the area in your classroom for that atom. Use the activity sheet from lesson 2 along with this activity sheet as a reference.



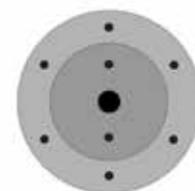
Energy levels

Electrons surround the nucleus of an atom in regions called *energy levels*. Even though atoms are spherical, the energy levels in an atom are more easily shown in concentric circles.



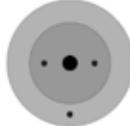
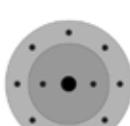
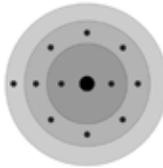
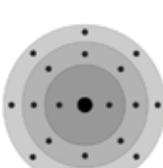
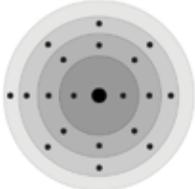
Which atom is this supposed to be?

The larger dot in the center of this atom represents the nucleus, which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons. In order to figure out which atom this represents, count up the number of electrons. There are 8 electrons in this atom. Because the number of electrons and protons is the same in an atom, this atom has 8 protons. Look at the chart Periodic Table, Elements 1–20. The number of protons is the same as the atomic number, so this drawing represents an oxygen atom.



ENERGY LEVELS ELEMENTS 1-20

Complete each energy level model by drawing the correct number of electrons in their corresponding energy levels.

<p>HYDROGEN 1</p>  <p>1.01</p>	<p>HELIUM 2</p>  <p>4.00</p>						
<p>LITHIUM 3</p>  <p>6.94</p>	<p>BERYLLIUM 4</p>  <p>9.01</p>	<p>BORON 5</p>  <p>10.81</p>	<p>CARBON 6</p>  <p>12.01</p>	<p>NITROGEN 7</p>  <p>14.01</p>	<p>OXYGEN 8</p>  <p>16.00</p>	<p>FLUORINE 9</p>  <p>19.00</p>	<p>NEON 10</p>  <p>20.18</p>
<p>SODIUM 11</p>  <p>22.99</p>	<p>MAGNESIUM 12</p>  <p>24.31</p>	<p>ALUMINIUM 13</p>  <p>26.98</p>	<p>SILICON 14</p>  <p>28.09</p>	<p>PHOSPHORUS 15</p>  <p>30.97</p>	<p>SULFUR 16</p>  <p>32.07</p>	<p>CHLORINE 17</p>  <p>35.45</p>	<p>ARGON 18</p>  <p>39.95</p>
<p>POTASSIUM 19</p>  <p>39.10</p>	<p>CALCIUM 20</p>  <p>40.08</p>						

Additional Teacher Background

Chapter 4 Lesson 3, p. 291

As the note on page 292 points out, there are other ways to model the electron energy levels of atoms. Some middle school texts show the electrons in pairs on an energy level. This pairing of electrons is intended to suggest information about the substructure *within* energy levels. This substructure is made up of regions called *orbitals* which comprise each energy level. The shape and size of the orbital is defined by the space around the nucleus where there is a high probability of finding electrons. There can be a maximum of two electrons in any orbital so showing electrons in pairs on an energy level model is an attempt to suggest information about the orbitals within the level.

In Middle School Chemistry, we chose to spread electrons out evenly on energy levels to indicate only the *number* of electrons on a level and not to suggest anything about the substructure of orbitals *within* energy levels. An understanding that the different energy levels can accommodate a certain number of electrons seems enough for students in middle school. They will see more refined models in high school and college when they learn more details about the orbitals within energy levels.

Some teachers might like to use a different model that shows more details of orbitals because it is more complete, even if they do not intend to explain those aspects of the model in much detail. Another argument is that a model showing paired and unpaired electrons may be useful for certain discussions about bonding. Other teachers may be more comfortable showing a less-detailed model even if it leaves out certain aspects of energy levels because they do not intend to discuss those details and they intend to handle bonding in a more general way. No model can be complete and accurate for all purposes and all have limitations. All models involve aspects of judgment and compromise. A good model focuses on the important points without too much to distract from those main features. The model you choose will have a lot to do with how much you think is important to explain and what the students are able to understand.

Some energy level models you might see and what they represent
For helium (atomic number 2), the energy level model in Middle School Chemistry is:



Helium has two electrons on the first energy level.



Some other middle school texts might show an energy level model for helium like this:

The *first* energy level has only one orbital. This is known as the 1s orbital. The “1” means that it is in the first energy level and the “s” stands for an orbital within that energy level with a particular shape. This 1s orbital can hold up to two electrons. So helium has its two electrons in the 1s orbital. The practice of showing the electrons together or *paired* in an energy level is meant to indicate how many orbitals in that level have been completely occupied by two electrons. For the first energy level, the pairing is not very useful for showing which orbitals are full and which aren't because there is only one orbital. But it becomes more useful for atoms that have more orbitals where some orbitals may be filled and others not.

For boron (atomic number 5), the energy level model in Middle School Chemistry is:



Boron has 2 electrons on the first energy level and 3 electrons on the second level.



Some other middle school texts might show an energy level model for boron like this:

The model shows that boron has two electrons in the 1s orbital of the first energy level which are shown as paired. It also has 3 electrons in the second energy level.

The second energy level is made up of four orbitals. There is a spherical orbital called 2s. The “2” means that it is in the second energy level. It is like the 1s orbital but is further from the nucleus. The second energy level also has 3 other orbitals that are all the same shape and distance from the nucleus but oriented in different directions. These orbitals are called 2p. The “p” orbitals are a different shape than the “s” orbitals. The 2s orbital can hold up to two electrons and each of the 2p orbitals can also hold up to 2 electrons. So the second energy level can hold up to eight electrons in its four orbitals. In this model of boron, two electrons are shown as paired in the 2s orbital and the last electron is shown in one of the 2p orbitals.

Another middle school text might show a model of boron like this:

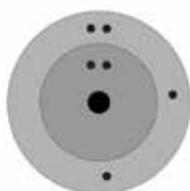


Here, they paired the electrons in the 1s orbital but did not show the detail of pairing the electrons in the 2s orbital of the second energy level. They chose to spread the three electrons out on the second energy level.

For carbon (atomic number 6), the energy level model in Middle School Chemistry is:



Carbon has 2 electrons on the first energy level and 4 on the second. Some other middle school texts might show a model of carbon like this:



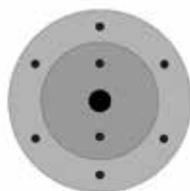
This model shows that carbon has two electrons in the 1s orbital of the first energy level which are shown as paired. It also has 4 electrons in the second energy level. In this model, two electrons are shown as paired in the 2s orbital and the other two electrons are shown separately or unpaired. This is done to indicate that each of the electrons is in a separate 2p orbital. One of the details of orbitals is that an electron goes into an empty available orbital of the same type before it goes into an orbital that already has an electron in it.

Another middle school text might show a model of carbon like this:



This model pairs the 1s electrons but spreads out the four electrons in the second energy level regardless of what orbital they are in. This approach would show electrons being paired on the second energy level for the first time in nitrogen.

For oxygen (atomic number 8), the energy level model in Middle School Chemistry is:



Oxygen has 2 electrons on the first energy level and 6 on the second. Oxygen is an interesting example because the other two types of models come out with the same result which looks like this:



Here, the electrons are paired in the 1s orbital. In the second energy level, whether the electrons are paired in the 2s to begin with or whether they are spread out and only paired after placing 1 electron in each of the four orbitals and then adding the last two electrons to make two pairs, the result is the same.

If the energy level models in Middle School Chemistry are different than those in your text book, you can use either one to teach that energy levels only have a certain number of electrons. You could also use the difference to suggest that there is more detail about energy levels that students may learn about later.

Additional Teacher Background

Chapter 4 Lesson 3, p. 295

What determines the shape of the standard periodic table?

One common question about the periodic table is why it has its distinctive shape. There are actually many different ways to represent the periodic table including circular, spiral, and 3-D. But in most cases, it is shown as a basically horizontal chart with the elements making up a certain number of rows and columns. In this view, the table is not a symmetrical rectangular chart but seems to have steps or pieces missing.

The key to understanding the shape of the periodic table is to recognize that the characteristics of the atoms themselves and their relationships to one another determine the shape and patterns of the table.

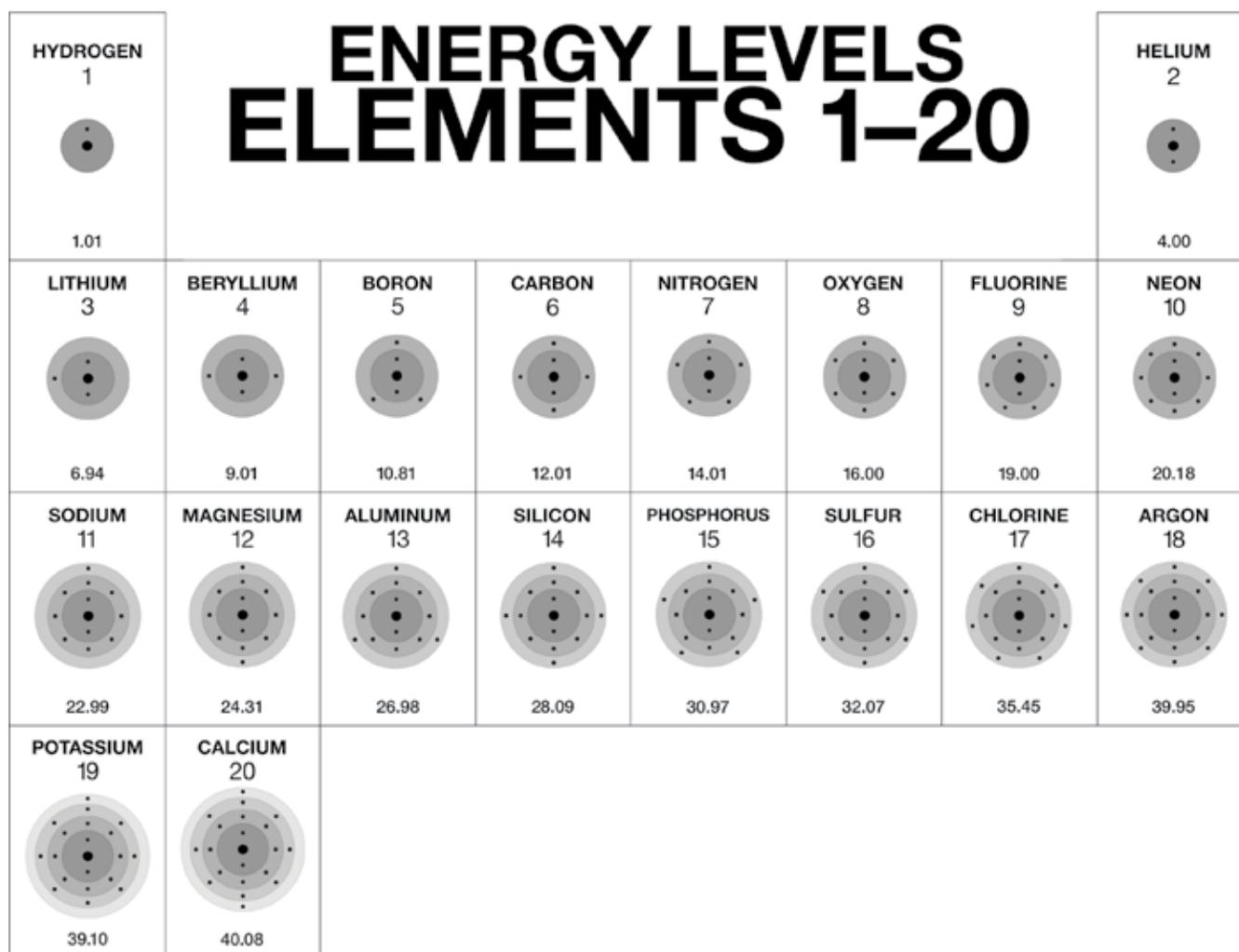
The Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00						
3 Li Lithium 6.94	4 Be Beryllium 9.01																	5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31																	13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80						
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29						
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)						
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium 178.49	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (261)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)												
			58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97							
			90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)							

Legend for the periodic table:

- 3 — Atomic Number
- Li** — Element Symbol
- Lithium — Element Name
- 6.94 — Average Atomic Mass

A helpful starting point for explaining the shape of the periodic table is to look closely at the structure of the atoms themselves. You can see some important characteristics of atoms by looking at the chart of energy level diagrams. Remember that an energy level is a region around an atom's nucleus that can hold a certain number of electrons. The chart shows the number of energy levels for each element as concentric shaded rings. It also shows the number of protons (atomic number) for each element under the element's name. The electrons, which equal the number of protons, are shown as dots within the energy levels. The relationship between atomic number, energy levels, and the way electrons fill these levels determines the shape of the standard periodic table.



What determines the sequence of the elements?

One of the main organizing principles of the periodic table is based on the atomic number (number of protons in the nucleus) of the atoms. If you look at any row, the atoms are arranged in sequence with the atomic number increasing by one from left to right. Since the number of electrons equals the number of protons, the number of electrons also increases by one from left to right across a row.

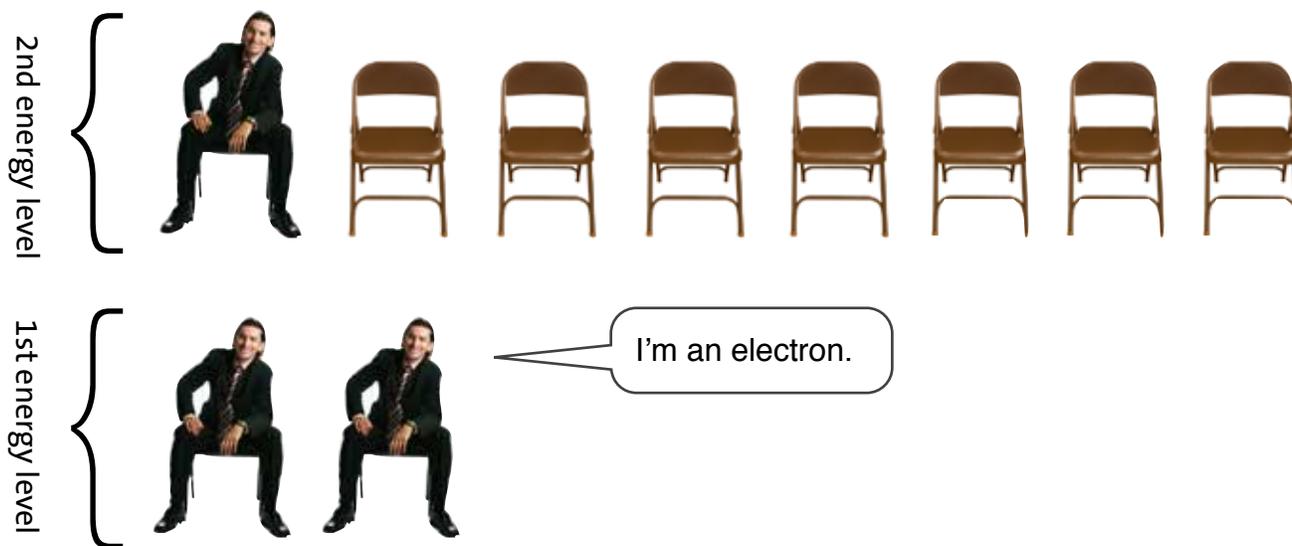
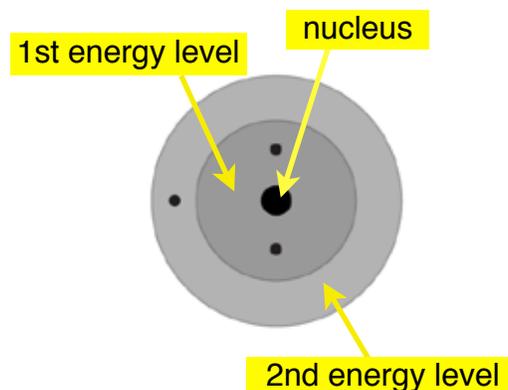
What do the rows represent?

The rows in the periodic table correspond to the number of energy levels of the atoms in that row. If you look at the chart, you can see that the atoms in the first row have one energy level. The atoms in the second row have two energy levels and so on. Understanding how electrons are arranged within the energy levels can help explain why the periodic table has as many rows and columns as it does. Let's take a closer look.

Electrons and Energy Levels

Every atom contains different energy levels that can hold a specific number of electrons. For a moment, let's imagine the simplest possible scenario: once all the positions are occupied within one energy level, any remaining electrons begin filling positions in the next energy level.

To picture this, imagine people filling rows of chairs in an auditorium. If each person sits next to another person until one row is filled, any remaining people must begin taking their seats in the second row, and so on.



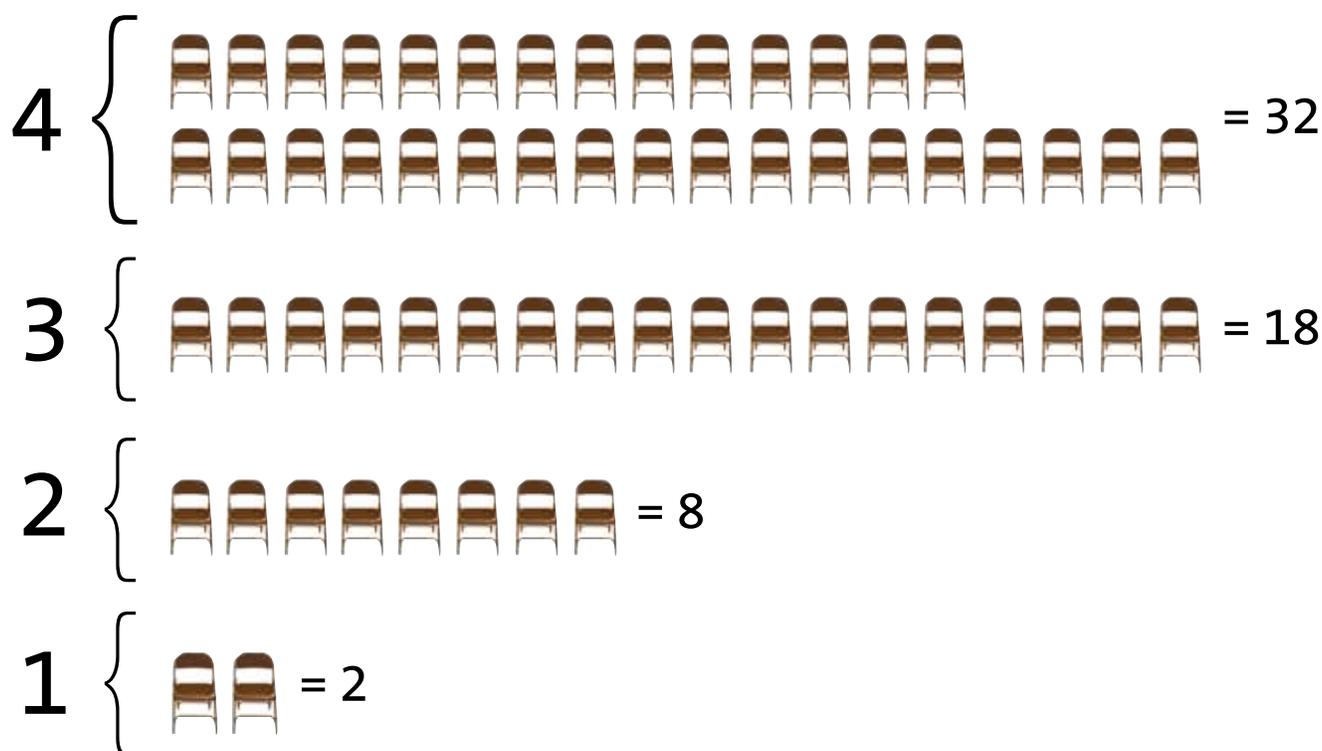
Not so bad, right? In general, this simple case is a helpful analogy. Electrons fill a given section until it is full, and then any more electrons move on to another unoccupied section where they continue filling there. Electrons begin filling the lowest energy level (closest to the nucleus) and then move on to higher energy levels (further from the nucleus). Unfortunately, the actual process is a bit more complicated. Let's see why.

Energy Levels Can Hold Different Numbers of Electrons

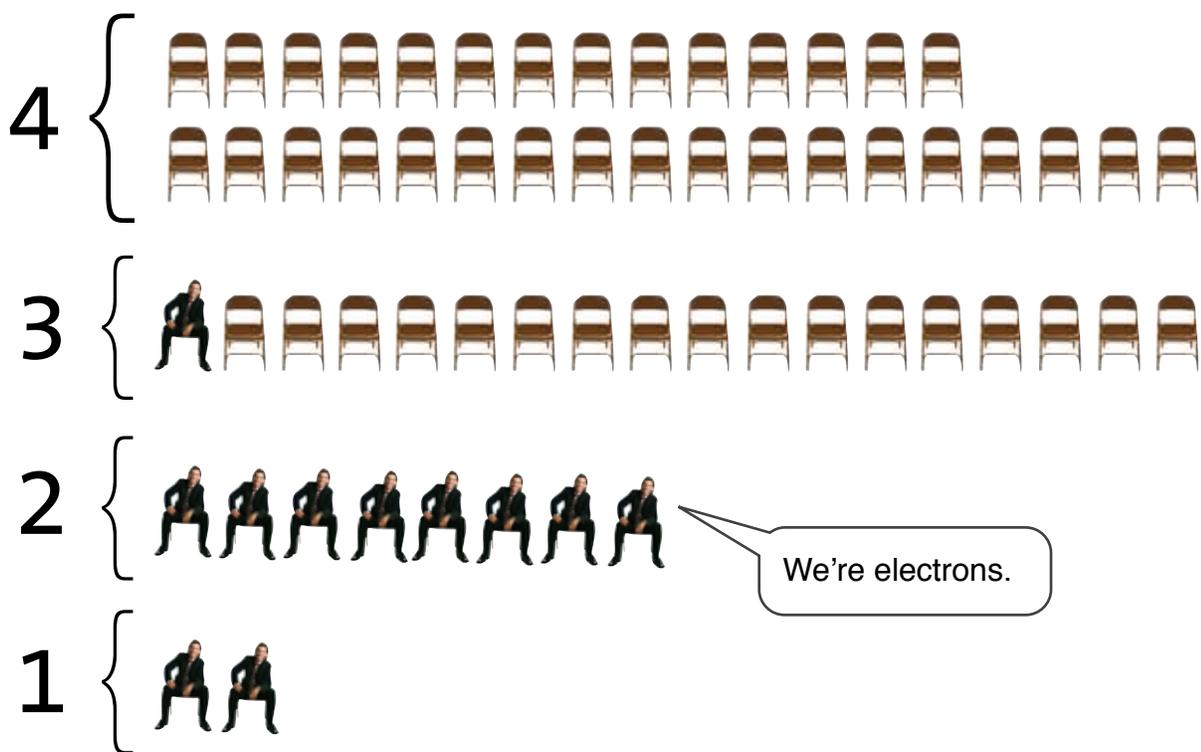
One thing that is slightly tricky about electrons filling these energy levels is that not all the energy levels can hold the same number of electrons. While the first energy level can hold only 2 electrons, the second energy level can hold 8, the third can hold 18, and the fourth can hold 32.

We'll stop there for now.

If we return to our rows of chairs analogy, it would be as if the first row was shorter than the second or third or fourth rows, so that after 2 people, any people remaining would have to begin occupying the second row. Then, if the second row were longer than the first row (but shorter than the third row), after 8 more people had been seated, any remaining individuals would have to begin occupying the third row.



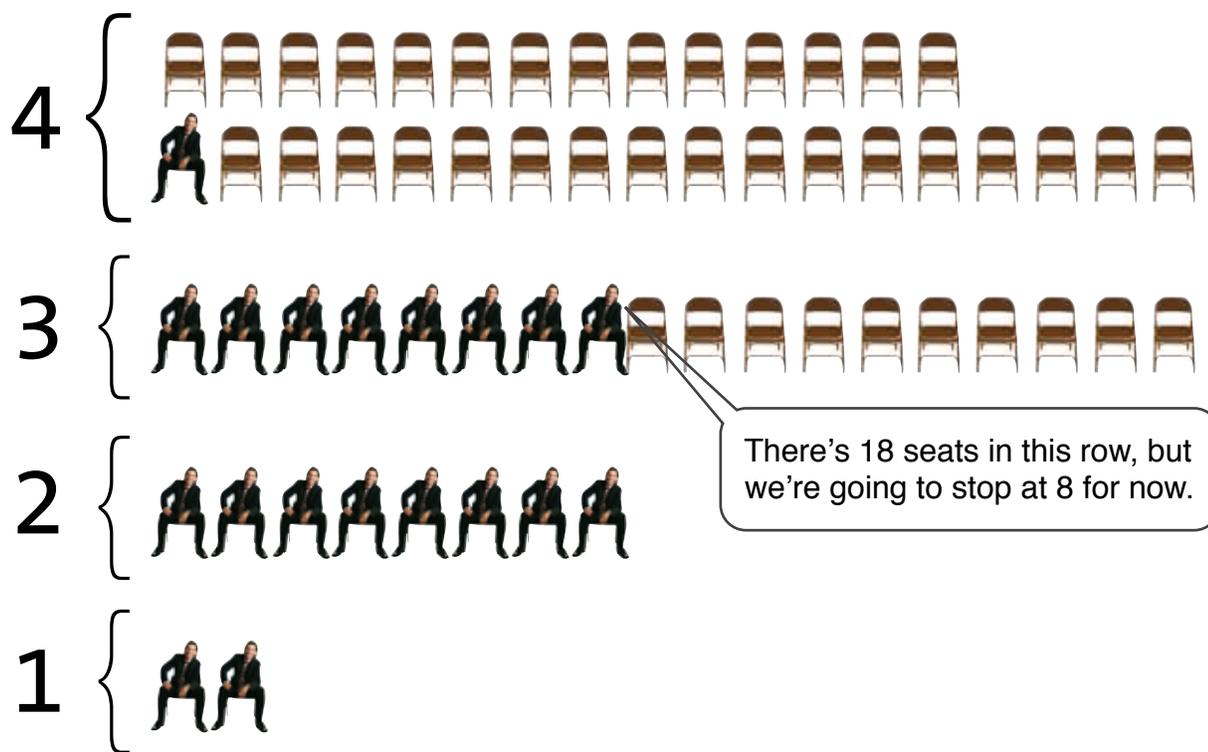
Extending our analogy of theater patrons as electrons, let's look at how the element sodium, with its 11 electrons, might fill these energy levels.



Because sodium has 11 electrons, it fills up the first energy level, which can hold only 2 electrons. It also fills up the second energy level, because it can only hold 8. Together, the first and second energy levels can hold a total of 10 electrons. Sodium has 11 electrons, so that final remaining electron that can't be accommodated by the first and second energy level begins filling in the third energy level. This pattern generally holds for the first 18 elements, up through argon, which has 18 electrons.

Energy Levels are Further Divided into Sections

But something funny happens beginning with potassium. Potassium has 19 electrons. Because the first, second, and third energy levels can hold a total of 28 electrons ($2+8+18=28$) it would seem that all the electrons of potassium could be "seated" within the third energy level. It turns out, however, that even though the third energy level has a total capacity of 18, only 8 "seats" are filled before the electrons begin filling the fourth energy level. So, potassium would fill up the energy levels like this:

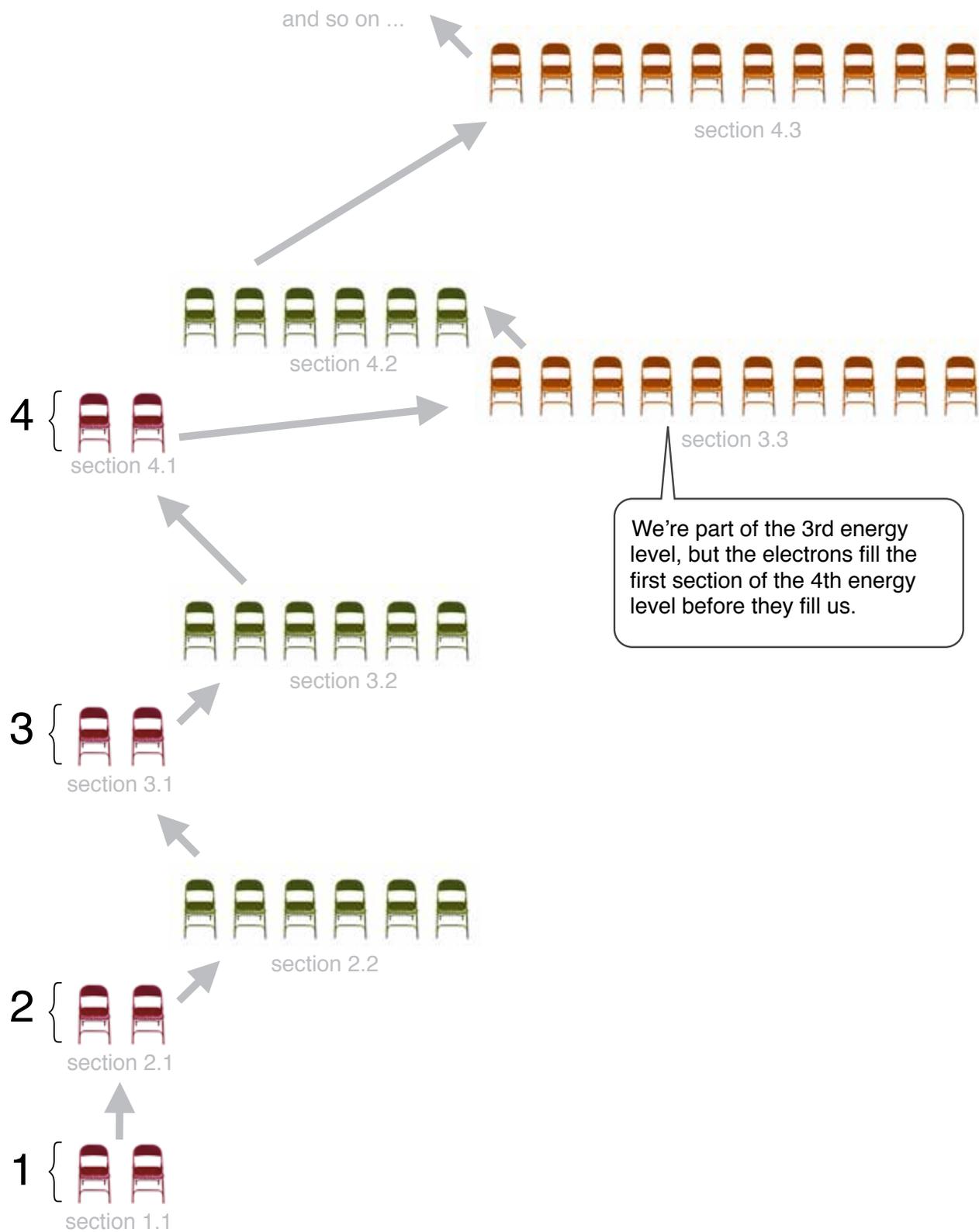


Whoa. *Whoa*. That's crazy. Why does *that* happen?

This is the second complication with our simple chairs analogy. It turns out that in addition to distinct energy levels (first, second, third, etc.) each energy level is further divided into sections where electrons can be found.

In terms of our analogy, the first row would have just one section. The second row would have two sections. The third row would have 3 sections and the fourth row would have four sections. As you can see, the number of sections an energy level has is equal to the number of that energy level.

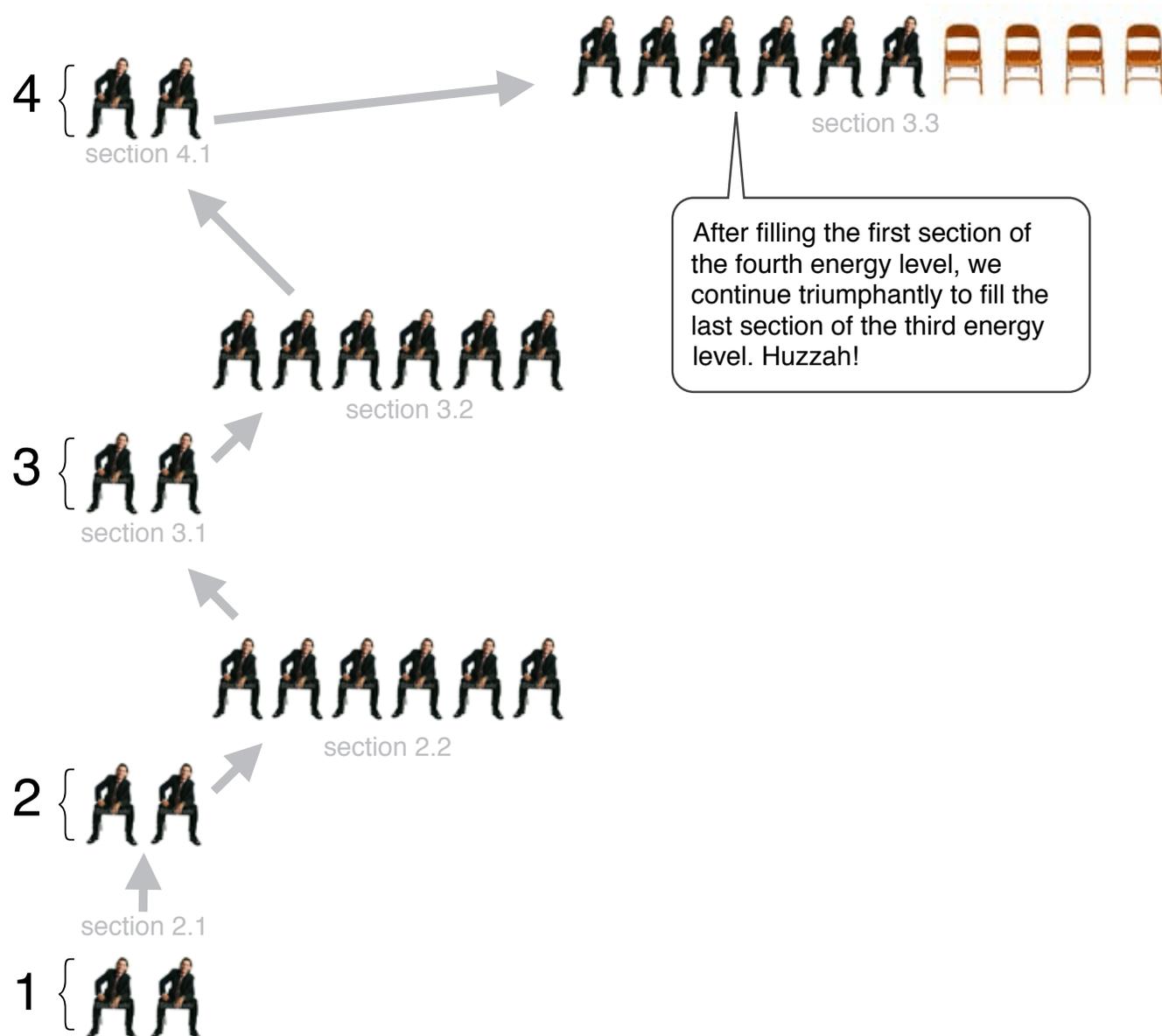
The reason why that last electron from potassium begins filling the fourth energy level rather than continuing to fill the third energy level is that the first section of the fourth energy level is actually closer (or at lower energy) than the last section of the third energy level (the last 10 "seats"). So, really, our chairs would now look something like this:



Admittedly, this doesn't look much like rows of chairs in an auditorium anymore, but the idea is still the same. Electrons will continue filling energy levels, one section at a time, until all the electrons are used up. When one section of the next energy level is actually lower in energy than the next section of the same energy level, the electrons will begin filling there. This is what we depicted in the diagram for potassium. Its last electron filled the first section in the fourth energy level, because that section was actually closer (at lower energy) than the last section of the third energy level.

Eventually, the electrons will continue filling the empty section in the third energy level. The idea is exactly what we've just described. Unusual as it might seem, in some cases, the first section of the next energy level is filled before the electrons continue to fill the last section of the preceding energy level.

Consider, for example, the element Iron. Its 26 electrons would fill energy levels like this:



Whew! So what does all of this mean?

Mainly this: understanding how electrons fill energy levels can help us to understand why the periodic table has as many rows as it does. Each row can roughly be thought of as starting a new energy level. As we proceed across a row, electrons fill energy levels in sections according to where they can be at the lowest energy. So, rather than the row continuing on forever, the periodic table begins a new row which signifies that the electrons in the elements in the next row begin filling a new energy level.

The Periodic Table of the Elements

Legend:

- 3 — Atomic Number
- Li — Element Symbol
- Lithium — Element Name
- 6.94 — Average Atomic Mass

1 H Hydrogen 1.01																	2 He Helium 4.00						
3 Li Lithium 6.94	4 Be Beryllium 9.01																	5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31																	13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80						
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29						
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.84	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.98	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)						
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium 178.49	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (269)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)												
			58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97							
			90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium 168.93	102 No Nobelium (259)	103 Lr Lawrencium (262)							

As we saw, sodium doesn't fall to the right of neon on the periodic table just because sodium has more electrons than neon has. Because sodium begins placing its electrons in a new energy level, it is positioned on the far left side at the beginning of a new row.

If we understand a few of the rules about energy level capacity and filling, we can begin to make sense of the periodic table's unusual shape. Why does the first row only consist of two elements? Well, it's because the first energy level can only hold two electrons, and helium, with an atomic number of two, has exactly two electrons. All elements after it have more than 2 electrons, and so they must continue filling their electrons at higher energy levels.

Why does the second period consist of eight elements? It's because the second energy level can only hold eight electrons. If we add in the two the first energy level can hold, the first and second energy levels combined can hold 10 electrons, and neon, the last element of the second period, has exactly 10 electrons.

Though it's a little tricky, potassium begins placing its electrons in the fourth energy level (even with 10 "seats" still available in the third energy level) because the first section of the fourth energy level is lower in energy than the last section of the third energy level.

The number of rows the periodic table has corresponds to the number of energy levels needed to hold all of the electrons of an atom with the greatest known number of electrons.

And what about these rows hanging out at the bottom? What's their deal?

The other peculiar feature found in most copies of the periodic table are two mysterious rows, often situated below the rest of the table, which seem to have no relation to the rest of the elements. These rows are called the lanthanide series and actinide series, respectively.

These rows are often placed below the rest of table simply as a convenience. In reality, the elements within the lanthanide series, beginning with the element lanthanum, actually belong alongside barium on the periodic table. Because this would make the table very wide, they are usually placed below the rest of the table so that the format of the periodic table fits more easily on a standard size poster or piece of paper. The same is true for the elements that comprise the actinide series. Beginning with the element actinium, these elements actually belong alongside radium, just below where the lanthanide series would be situated. Some periodic tables actually are formatted in this elongated version. As a convention, however, they are placed below for convenience.

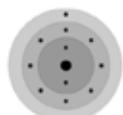
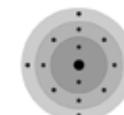
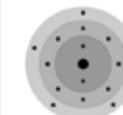
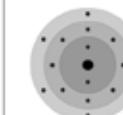
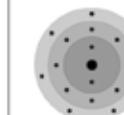
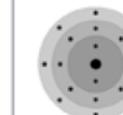
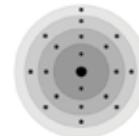
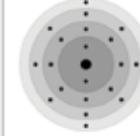
Although alternative forms of the periodic table have been created, some taking unusual shapes like a series of concentric circles in an archery target, the conventional table with the familiar groups and periods is considered the standard.

Why do atoms in the same column have the same number of outer (valence) electrons?

If you think about how the energy levels fill up with electrons, and how the periodic table is designed, you can see how certain atoms end up in the same column. An important point about the columns is that the number of electrons in the outer energy level, called valence electrons, will be the same for all the elements in that column.

The periodic table is designed so that the first electron starting a new energy level starts a new row on the far left. Each new row starts after the outer energy level of the previous row has eight electrons. An exception to this is starting the second level after the first level has two electrons. Let's look again at the energy level chart.

ENERGY LEVELS ELEMENTS 1-20

<p>HYDROGEN 1</p>  <p>1.01</p>							<p>HELIUM 2</p>  <p>4.00</p>
<p>LITHIUM 3</p>  <p>6.94</p>	<p>BERYLLIUM 4</p>  <p>9.01</p>	<p>BORON 5</p>  <p>10.81</p>	<p>CARBON 6</p>  <p>12.01</p>	<p>NITROGEN 7</p>  <p>14.01</p>	<p>OXYGEN 8</p>  <p>16.00</p>	<p>FLUORINE 9</p>  <p>19.00</p>	<p>NEON 10</p>  <p>20.18</p>
<p>SODIUM 11</p>  <p>22.99</p>	<p>MAGNESIUM 12</p>  <p>24.31</p>	<p>ALUMINUM 13</p>  <p>26.98</p>	<p>SILICON 14</p>  <p>28.09</p>	<p>PHOSPHORUS 15</p>  <p>30.97</p>	<p>SULFUR 16</p>  <p>32.07</p>	<p>CHLORINE 17</p>  <p>35.45</p>	<p>ARGON 18</p>  <p>39.95</p>
<p>POTASSIUM 19</p>  <p>39.10</p>	<p>CALCIUM 20</p>  <p>40.08</p>						

With these principles in mind, you can see why the atoms in the first column, which contains hydrogen (H), lithium (L), sodium (Na), and potassium (K), each have one electron in the outer energy level. In the second column, beryllium (Be), magnesium (Mg) and calcium (Ca), all have two valence electrons. The atoms in the column with boron (B) and aluminum (Al) all have three valence electrons. The atoms in the column with carbon (C) and silicon (Si) have four valence electrons. The rest of the columns follow this same pattern. The transition elements in the middle of the periodic table (not shown in the chart) for the most part have two valence electrons.

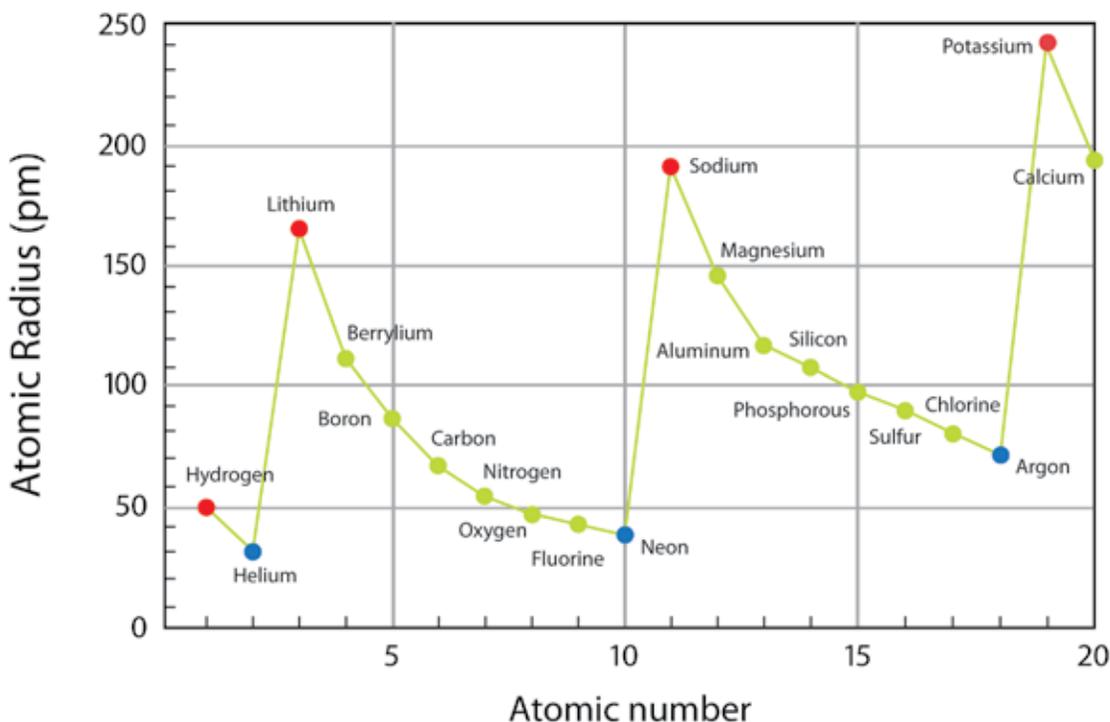
What is the significance of the word “periodic” in the periodic table of the elements?

Because of the way the atoms are organized in the periodic table, a pattern of characteristics or properties that repeat “periodically” can be seen from row to row in the table. This is called *periodicity*. Hence the periodic table.

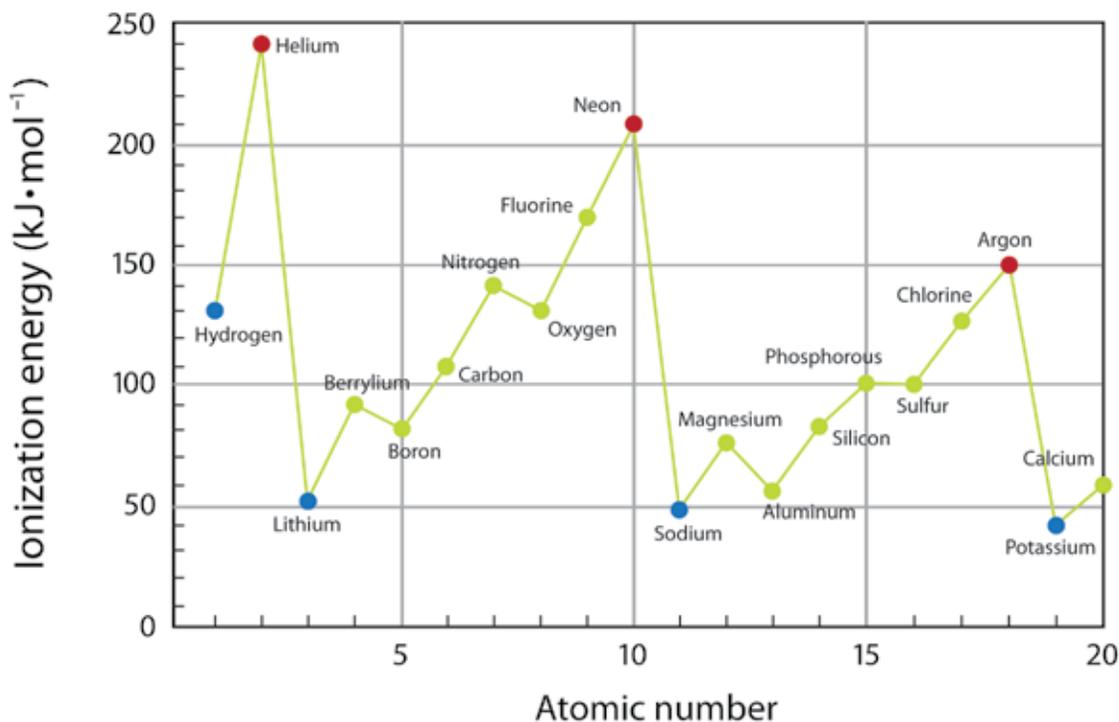
One property that demonstrates the idea of periodicity is atomic radius. Scientists measure atomic radii to tell them how large atoms are. As we proceed across a row (from left to right) we observe that atomic radii decrease. For example, magnesium has a smaller atomic radius than sodium, and aluminum has a smaller atomic radius than magnesium, and so on. This same pattern repeats itself in the next row and the next row in a periodic way.

<h1 style="text-align: center;">ATOMIC SIZE & MASS ELEMENTS 1-20</h1>							
HYDROGEN 1  1.01							HELIUM 2  4.00
LITHIUM 3  6.94	BERYLLIUM 4  9.01	BORON 5  10.81	CARBON 6  12.01	NITROGEN 7  14.01	OXYGEN 8  16.00	FLUORINE 9  19.00	NEON 10  20.18
SODIUM 11  22.99	MAGNESIUM 12  24.31	ALUMINUM 13  26.98	SILICON 14  28.09	PHOSPHORUS 15  30.97	SULFUR 16  32.07	CHLORINE 17  35.45	ARGON 18  39.95
POTASSIUM 19  39.10	CALCIUM 20  40.08						

Note: Some charts of atomic radii show the atoms in the last column on the right (Helium, Neon, Argon ...) having larger radii than one or more of the atoms to their left. This is a result of using a different measuring technique. For our purposes, the trend is decreasing atomic size as you move from left to right along a row.



Another example of periodicity is a property called *ionization energy*. Ionization energy refers to the amount of energy needed to remove an electron from an atom to form an ion. The more difficult it is to remove an electron from an atom, the higher its ionization energy. As a trend, ionization energy increases as you move across a row (from left to right). For example, in the first row, Hydrogen, on the far left, has a low ionization energy and Helium, on the far right, has a high ionization energy. Each row begins with a low value and ends with a high value.



There are other periodic trends in the properties of atoms. You can see some of these by going to <http://acswebcontent.acs.org/games/pt.html>.

Click the “Plot Data” tab to get to a drop-down menu of different properties you can choose. The first one displayed is “Molar Mass” which shows the mass of a mole of each atom. This can be thought of as comparing the atomic mass of one atom to another. This property is not periodic. You can choose atomic radius or ionization energy to see the trends discussed earlier or select another property.

Chapter 4, Lesson 4: Energy Levels, Electrons, and Covalent Bonding

Key Concepts

- The electrons on the outermost energy level of the atom are called valence electrons.
- The valence electrons are involved in bonding one atom to another.
- The attraction of each atom's nucleus for the valence electrons of the other atom pulls the atoms together.
- As the attractions bring the atoms together, electrons from each atom are attracted to the nucleus of both atoms, which "share" the electrons.
- The sharing of electrons between atoms is called a covalent bond, which holds the atoms together as a molecule.
- A covalent bond happens if the attractions are strong enough in both atoms and if each atom has room for an electron in its outer energy level.
- Atoms will covalently bond until their outer energy level is full.
- Atoms covalently bonded as a molecule are more stable than they were as separate atoms.

Summary

Students will look at animations and refer to the energy level models they have been using to make drawings of the process of covalent bonding. Students will consider why atoms bond to form molecules like H₂ (hydrogen), H₂O (water), O₂ (oxygen), CH₄ (methane), and CO₂ (carbon dioxide).

Objective

Students will be able to explain that attraction between the protons and electrons of two atoms cause them to bond. Students will be able to draw a model of the covalent bonds between the atoms in H₂ (hydrogen), H₂O (water), O₂ (oxygen), CH₄ (methane), and CO₂ (carbon dioxide).

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Salt
- Clear plastic cup
- Tape

About this Lesson

This lesson will probably take more than one class period.

ENGAGE

1. Show an animation to introduce the process of covalent bonding.

Introduce the question students will investigate in this lesson:

- If atoms have an equal number of protons and electrons, why do atoms bond to other atoms? Why don't they just stay separate?

Begin to answer this question by using hydrogen as an example.

Project the animation *Covalent bond in hydrogen.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_hydrogen_animation

Make sure students see that each hydrogen atom has 1 proton and 1 electron. Remind students that the electron and its own proton are attracted to each other. Explain that if the atoms get close enough to each other, the electron from each hydrogen atom feels the attraction from the proton of the other hydrogen atom (shown by the double-headed arrow). Point out to students that the attractions are not strong enough to pull the electron completely away from its own proton. But the attractions are strong enough to pull the two atoms close enough together so that the electrons feel the attraction from both protons and are shared by both atoms. At the end of the animation, explain that the individual hydrogen atoms have now bonded to become the molecule H_2 . This type of bond is called a *covalent* bond. In a covalent bond, electrons from each atom are attracted or “shared” by both atoms.



EXPLAIN

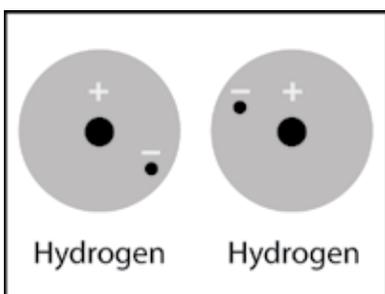
2. Discuss the conditions needed for covalent bonding and the stable molecule that is formed.

Project the image *Covalent bond in hydrogen.*

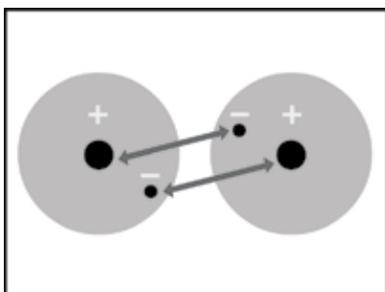
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_hydrogen_illustrations

Note: This model of covalent bonding for the hydrogen molecule (H_2) starts with 2 individual hydrogen atoms. In reality, hydrogen atoms are never separate to start with. They are always bonded with something else. To simplify the process, this model does not show the hydrogen atoms breaking their bonds from other atoms. It only focuses on the process of forming covalent bonds between two hydrogen atoms.

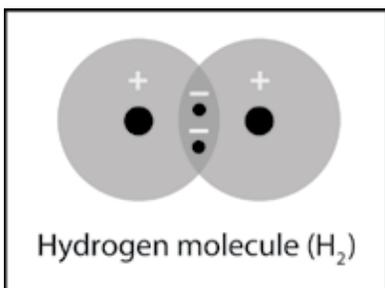
Read more about bonding in the additional teacher background section at the end of this lesson.



Two hydrogen atoms are near each other.



When two hydrogen atoms come close enough to each other, their electrons are attracted to the proton of the other atom.



Because there is both a strong enough attraction between atoms and room for electrons in the outer energy level of both atoms, the atoms share electrons. This forms a covalent bond.

Tell students that there are two main reasons why two hydrogen atoms bond together to make one hydrogen molecule:

- There needs to be a strong enough attraction between the electrons of each atom for the protons of the other atom.
- There needs to be room in the outer energy level of both atoms.

Once bonded, the hydrogen molecule is more stable than the individual hydrogen atoms. Explain to students that by being part of a covalent bond, the electron from each hydrogen atom gets to be near two protons instead of only the one proton it started with. Since the electrons are closer to more protons, the molecule of two bonded hydrogen atoms is more stable than the two individual unbonded hydrogen atoms.

This is why it is very rare to find a hydrogen atom that is not bonded to other atoms. Hydrogen atoms bond with other hydrogen atoms to make hydrogen gas (H_2). Or they can bond with other atoms like oxygen to make water (H_2O) or carbon to make methane (CH_4) or many other atoms.

3. Show students that when two hydrogen atoms bond together, the outer energy level becomes full.

Have students look at their *Periodic table of energy levels for elements 1–20* distributed in lesson 3.

Explain that the two electrons in the hydrogen molecule (H_2) can be thought of as “belonging” to each atom. This means that each hydrogen atom now has two electrons in its first energy level. The first energy level in the outer energy level for hydrogen and can only accommodate or “hold” two electrons. Atoms will continue to covalently bond until their outer energy levels are full. At this point, additional atoms will not covalently bond to the atoms in the H_2 molecule.

4. Have students describe covalent bonding in a hydrogen molecule on their activity sheet and then review their answers.

Give each student an activity sheet.

Have students write a short caption under each picture to describe the process of covalent bonding and answer the first three questions. The rest of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions.

Ask students:

- **What did you write for the second and third pictures of covalent bonding?**
Center drawing: When two hydrogen atoms come close enough, their electrons are attracted to the proton of the other atom.
Last drawing: This brings the atoms close enough together that they share electrons.
- **What are two conditions atoms must have in order to form covalent bonds with one another?**
There is a strong enough attraction between atoms and there is room for electrons in the outer energy level of both atoms.
- **Why is a hydrogen molecule (H_2) more stable than two individual hydrogen atoms?**
In the hydrogen molecule, the electrons from each atom are able to be near two protons instead of only the one proton it started with. Whenever negative electrons are near additional positive protons, the arrangement is more stable.
- **Why doesn't a third hydrogen atom join the H_2 molecule to make H_3 ?**
When two hydrogen atoms share their electrons with each other, their outer energy levels are full.

You could explain to students that when the outer energy levels are full, sharing electrons with another atom would not happen for two main reasons:

1. An electron from a new atom would have to join an atom in the H_2 molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction.
2. An electron from an atom already in the H_2 molecule and close to the nucleus would need to move further away to share with the new atom.

Both of these possibilities would make the molecule less stable and therefore would not happen.

5. Discuss the process of covalent bonding in a water molecule.

Project the animation *Covalent bond in water*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_water



Before hitting the “play” button, point out the oxygen atom and the two hydrogen atoms.

Ask students:

- **Is there anything that might attract these atoms to one another?**

Students should suggest that the electrons from each atom are attracted to the protons of the other atoms.

Play the animation to show the attraction between the protons of oxygen for the electron from each of the hydrogen atoms, the attraction of the proton from the hydrogen atoms for the electrons of oxygen, and the atoms coming together.

Explain that the electrons are shared by the oxygen and hydrogen atoms forming a covalent bond. These bonds hold the oxygen and hydrogen atoms together and form the H_2O molecule. The reason why the atoms are able to bond is that the attractions are strong enough in both directions and there is room for the electrons on the outer energy level of the atoms.

The electron from each hydrogen atom and the electrons from the oxygen atom get to be near more protons when the atoms are bonded together as a molecule than when they are separated as individual atoms. This makes the molecule of bonded oxygen and hydrogen atoms more stable than the individual separated atoms.

Explain to students that the two electrons in the bond between the hydrogen atom and the oxygen atom can be thought of as “belonging” to each atom. This gives each hydrogen atom two electrons in its outer energy level, which is full. It also gives oxygen 8 electrons in its outer energy level, which is also full.

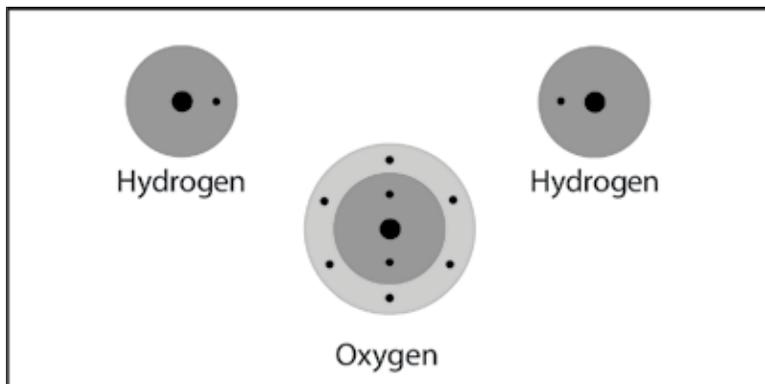
Project the image *Covalent bond in water*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_water_illustrations

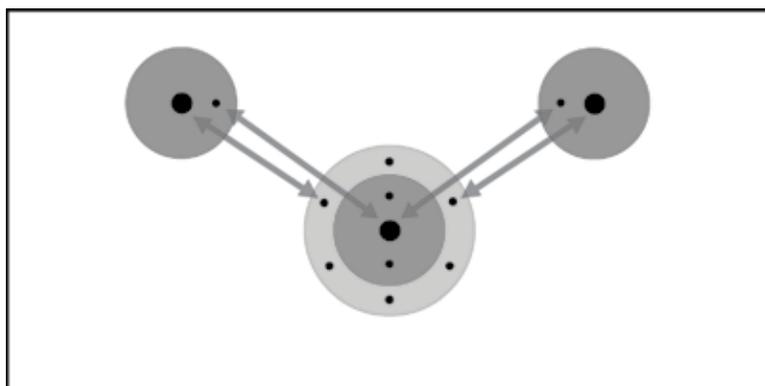
Review with students the process of covalent bonding covered in the animation.

6. Have students describe covalent bonding in a water molecule on their activity sheet.

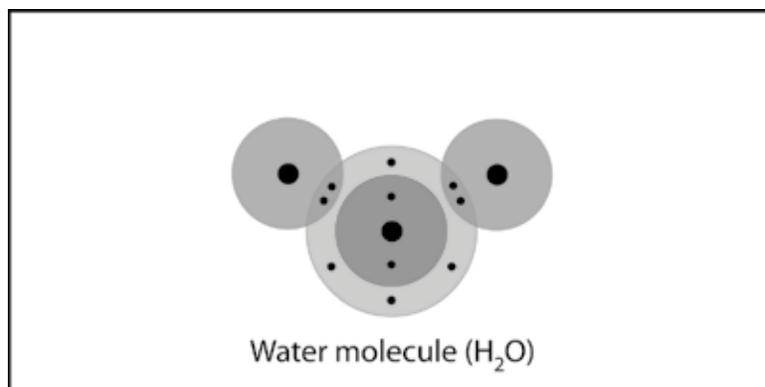
Have students write a short caption beside each picture to describe the process of covalent bonding in the water molecule.



Two hydrogen atoms and one oxygen atom are near each other.



When two hydrogen atoms come close enough to an oxygen atom, their electrons are attracted to the proton of the other atom.



Because there is both a strong enough attraction between atoms and room for electrons in the outer energy levels of the atoms, they share electrons. This forms a covalent bond.

Note: This model of covalent bonding for a water molecule starts with 2 individual hydrogen atoms and 1 oxygen atom. In reality, these atoms are never separate to start with. They are always bonded with something else. To simplify the process, this model does not show the hydrogen and oxygen atoms breaking their bonds from other atoms. It only focuses on the process of forming covalent bonds to make water.

Ask students:

- **Why can't a third hydrogen atom join the water molecule (H_2O) to make H_3O ?**
Once the outer energy levels are full, sharing electrons with another atom would not happen for two main reasons: An electron from a new atom would have to join an atom in the H_2O molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction. An electron from an atom already in the H_2O molecule and close to the nucleus would need to move further away to share with the new atom. Both of these possibilities would make the molecule less stable and would not happen.

EXPLORE

7. Have students use electricity to break the covalent bonds in water molecules.

Tell students that electrical energy can be used to break the covalent bonds in water molecules to produce hydrogen atoms and oxygen atoms. Two hydrogen atoms then bond to form hydrogen gas (H_2) and two oxygen atoms bond to form oxygen gas (O_2).

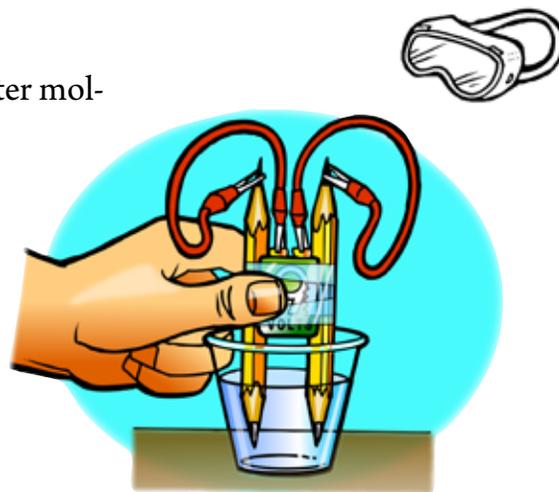
You may choose to do this activity as a demonstration or show the video *Electrolysis*.
www.middleschoolchemistry.com/multimedia/chapter4/lesson4#electrolysis

Question to investigate

What is produced when the covalent bond in water molecules is broken?

Materials for each group

- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Salt
- Clear plastic cup
- Tape



Procedure

1. Place a battery between 2 pencils. Be sure that the battery is more than half-way up.
2. With the help of a partner, wrap tape around the pencils and battery as shown.
3. Add water to a clear plastic cup until it is about $\frac{1}{2}$ -full.
4. Add about $\frac{1}{2}$ teaspoon of salt to the water and stir until the salt dissolves.
5. Connect one alligator clip to one terminal of the battery.
6. Using the other wire, connect one alligator clip to the other terminal of the battery.

7. Connect one end of the pencil lead to the alligator clip at the end of one of the wires.
8. Using the other wire, connect one end of the other pencil lead to the alligator clip at the end of the wire.
9. Place the ends of the pencil into the water as shown.

Expected results

Bubbles will form and rise initially from one pencil lead. Soon, bubbles will form and rise from the other. Students should be able to see that there is more of one gas than the other. The gas that forms the small bubbles that comes off first is hydrogen. The other gas that forms the larger bubbles and lags behind a bit is oxygen.

Note: There will be bubbling when hydrogen and oxygen gas form on the pencil leads. Be sure students do not get the misconception that the bubbles they see mean that the water is boiling. In boiling, the bonds holding the atoms together in water molecules do not come apart. In the process of electrolysis, the bonds holding the atoms together do come apart.

8. Discuss student observations.

Ask students:

- **What are the bubbles made out of in the activity?**
Hydrogen gas (H_2) and oxygen gas (O_2)
- **Why was there more hydrogen gas produced than oxygen gas?**
Each water molecule breaks into 2 hydrogen atoms and 1 oxygen atom. Two hydrogen atoms then bond to form hydrogen gas (H_2) and 2 oxygen atoms bond to form oxygen gas (O_2). Each water molecule has all the atoms needed to make 1 molecule of hydrogen gas. But with only 1 oxygen atom, a water molecule only has half of what is needed to make 1 molecule of oxygen gas. So, 2 water molecules will produce 2 molecules of hydrogen gas but only 1 molecule of oxygen gas.

EXTEND

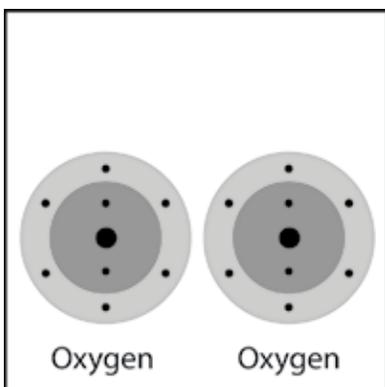
9. Help students understand how atoms combine to form the molecules of oxygen, methane, and carbon dioxide.

Remind students that in this lesson they looked at the covalent bonds in hydrogen molecules and in water molecules. Tell them that they will look at the covalent bonds in three other common substances.

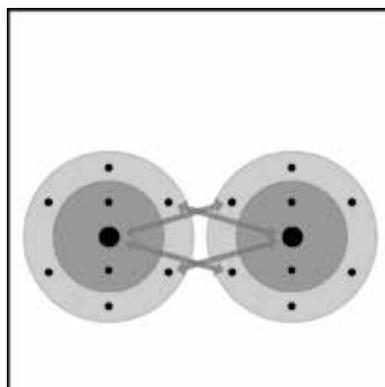
Project the animation *Oxygen's double bond*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#oxygen_double_bond

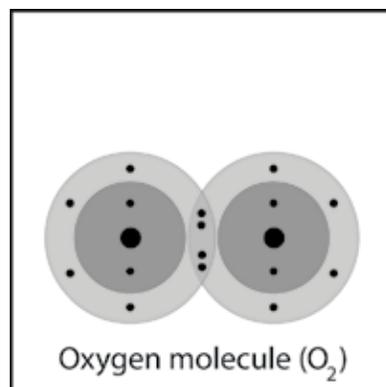
Explain to students that the oxygen molecules that are present in our air are made up of 2 oxygen atoms. This animation will show them what the covalent bond between 2 oxygen atoms is like. Narrate the animation by pointing out that each oxygen atom has 6 valence electrons. When the oxygen atoms get close together, the attractions from the nucleus of both atoms attract the outer electrons. In this case, 2 electrons from each atom are shared. This is called a double bond.



Each oxygen atom has 6 valence electrons in its outer energy level.



When two oxygen atoms get close to each other, the attractions from the nucleus of both atoms attract the outer electrons.



In this case, two electrons from each atom are shared. This is called a double bond.

Project the image *Oxygen's double bond II*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#oxygen_double_bond_illustrations

Review with students the process of covalent bonding covered in the animation.

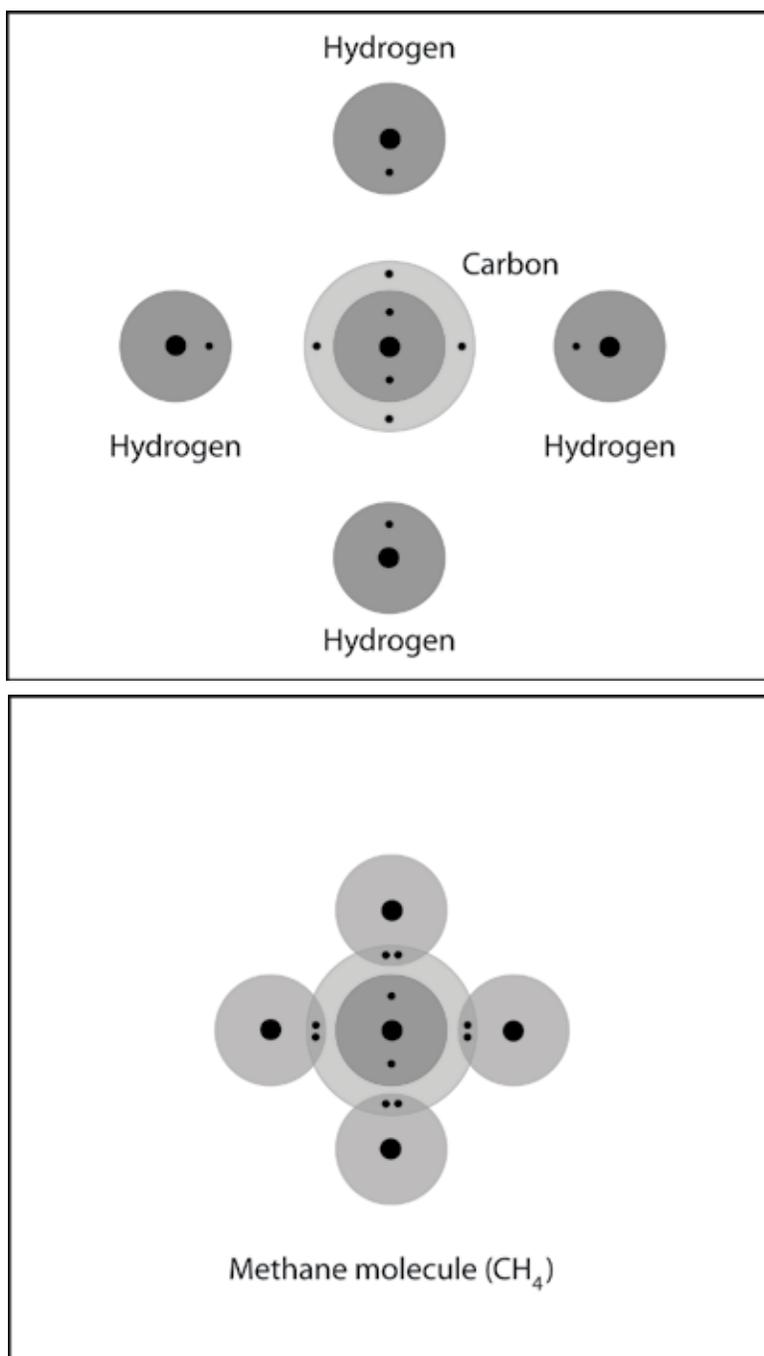
Project the before and after pictures *Covalent bonding of methane.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bonding_methane

Ask students:

- **Briefly describe the process of covalent bonding between the carbon and the four hydrogen atoms to make a methane molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.**

Be sure students realize that the protons of each atom attracts the other atoms electrons, which brings the atoms together. Atoms continue to bond with other atoms until their outer energy levels are full.



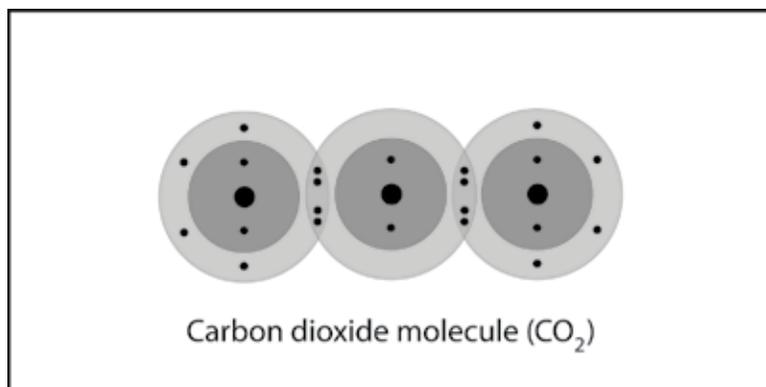
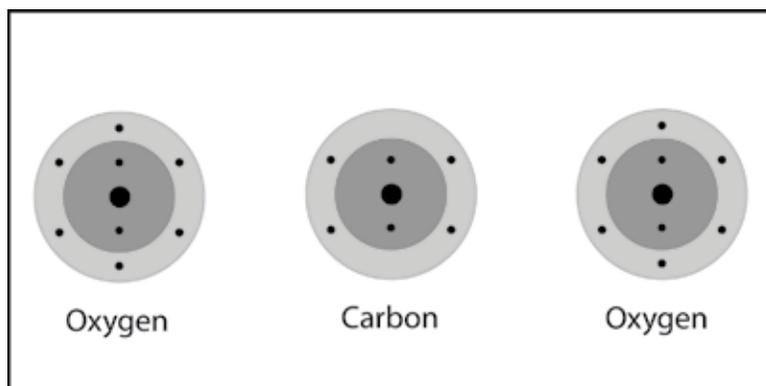
Project the before and after pictures *Covalent bonding of carbon dioxide gas.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson4#covalent_bond_carbon_dioxide

Ask students:

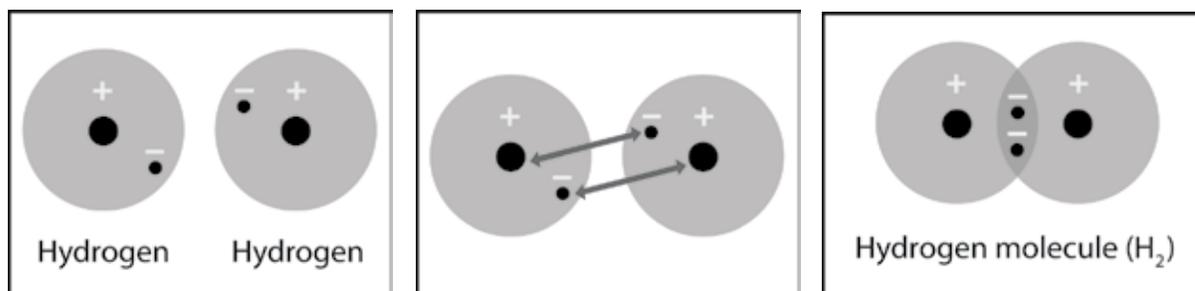
- **Briefly describe the process of covalent bonding between the carbon and the two oxygen atoms to make a carbon dioxide molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.**

Be sure students realize that the protons of each atom attracts the other atoms electrons, which brings the atoms together. Atoms continue to bond with other atoms until their outer energy levels are full.



EXPLAIN IT WITH ATOMS & MOLECULES

1. Write a short caption under each picture to describe the process of covalent bonding.



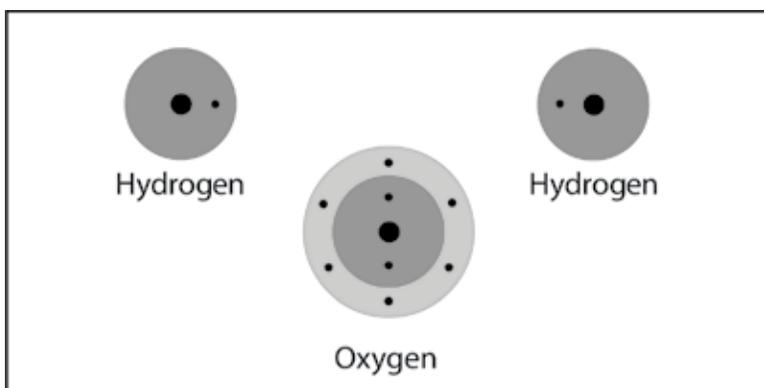
Two hydrogen atoms are near each other.

2. What are two conditions atoms must have in order to form covalent bonds with one another?

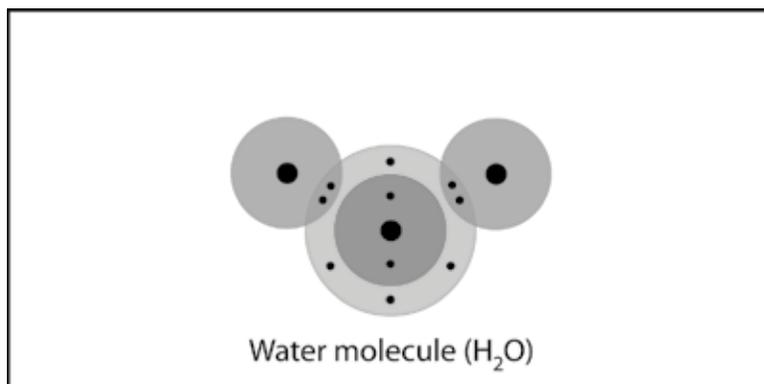
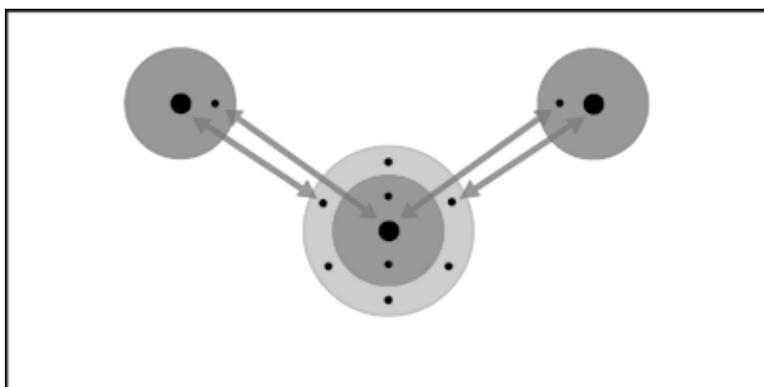
3. Why is a hydrogen molecule (H_2) more stable than two individual hydrogen atoms?

4. Why can't a third hydrogen atom join the H_2 molecule to make H_3 ?

5. Write a short caption beside each picture to describe the process of covalent bonding.



Two hydrogen atoms and one oxygen atom are near each other.



6. Why can't a third hydrogen atom join the water molecule (H_2O) to make H_3O ?

ACTIVITY

Question to investigate

What is produced when the covalent bond in water molecules is broken?

Materials for each group

- 9-volt battery
- 2 wires with alligator clips on both ends
- 2 pencils sharpened at both ends
- Water
- Salt
- Clear plastic cup
- Tape

Procedure

1. Place a battery between 2 pencils. Be sure that the battery is more than half-way up.
2. With the help of a partner, wrap tape around the pencils and battery as shown.
3. Add water to a clear plastic cup until it is about $\frac{1}{2}$ -full.
4. Add about a $\frac{1}{2}$ teaspoon of salt to the water and stir until the salt dissolves.
5. Connect one alligator clip to one terminal of the battery.
6. Using the other wire, connect one alligator clip to the other terminal of the battery.
7. Connect one end of the pencil lead to the alligator clip at the end of one of the wires.
8. Using the other wire, connect one end of the other pencil lead to the alligator clip at the end of the wire.
9. Place the ends of the pencil into the water as shown.

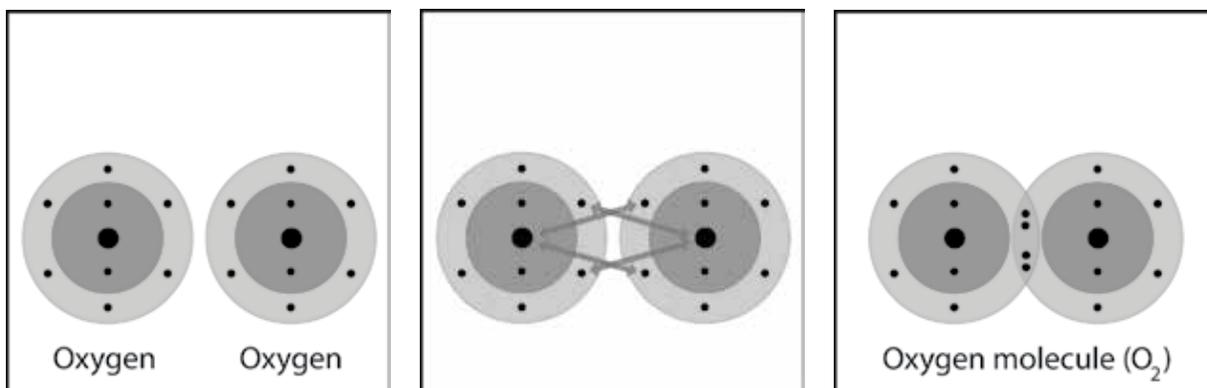


7. What were the bubbles made out of in this activity?

8. Why was there more hydrogen gas produced than oxygen gas?
HINT: Look back at the drawings showing the number of hydrogen and oxygen atoms that bond to form a water molecule.

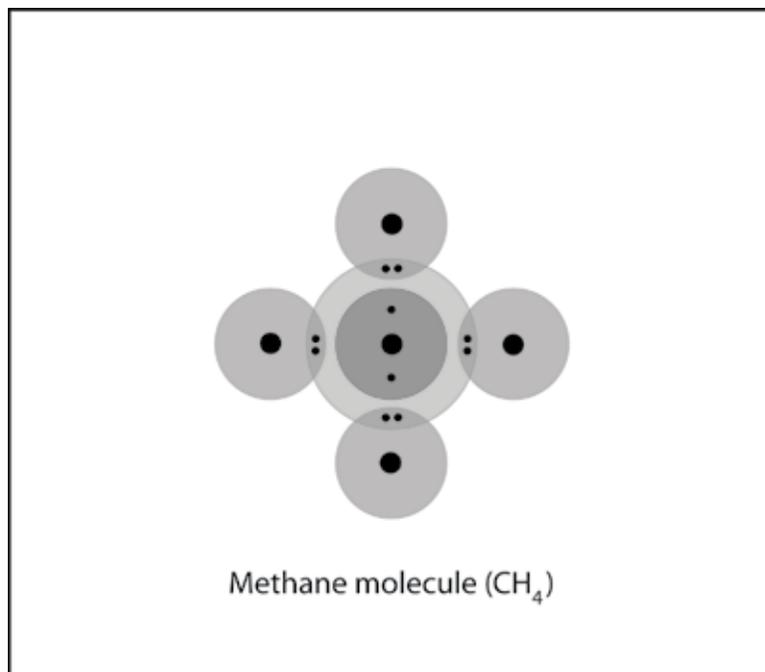
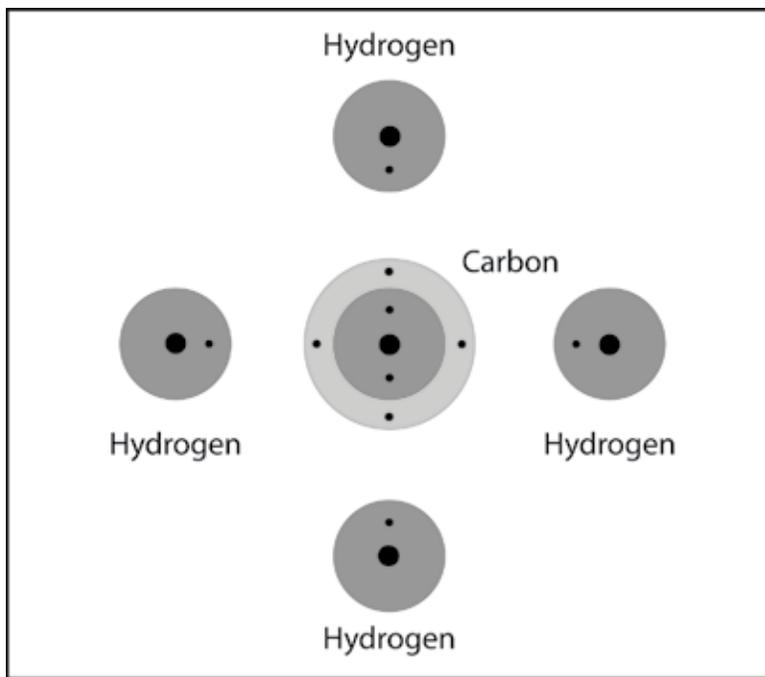
TAKE IT FURTHER

9. Briefly describe the process of covalent bonding between two oxygen atoms to make an oxygen molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.

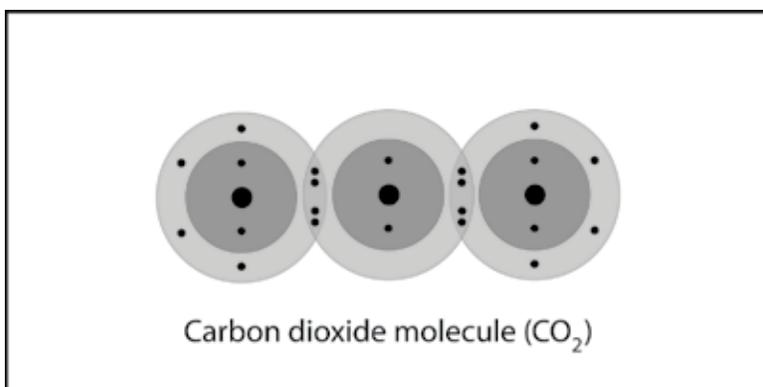
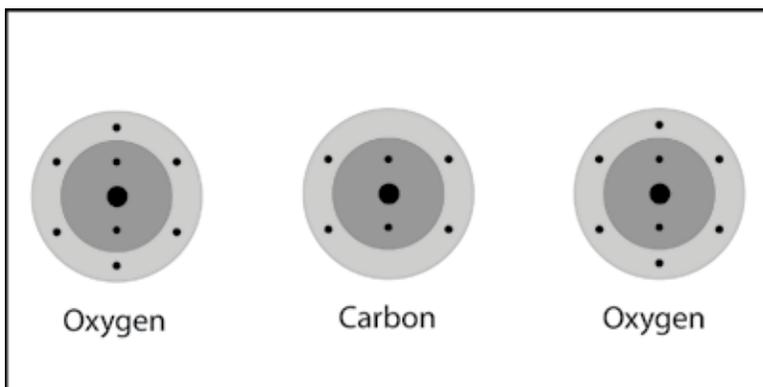


Each oxygen atom has 6 valence electrons in its outer energy level.

10. Briefly describe the process of covalent bonding between the carbon and the four hydrogen atoms to make a methane molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.



11. Briefly describe the process of covalent bonding between the carbon and the two oxygen atoms to make a carbon dioxide molecule. Be sure to mention attractions between electrons and protons and the number of electrons in the outer energy level for the atoms in the final molecule.



Additional Teacher Background

Chapter 4, Lesson 4, p. 318

A common approach to figuring out how atoms bond covalently and ionically is to use the “octet rule”. This rule relies on the fact that atoms bond until they have 8 electrons in their outer energy levels or 2 electrons in the outer level in the case of hydrogen and helium. It is often stated that atoms “want” to have 8 electrons in their outer energy level so they bond until they have 8 as if having 8 electrons is a goal in itself.

The approach taken in Lesson 4 and 5 achieves the same result but it does not use the goal of having 8 electrons or wanting 8 electrons as the reason why atoms bond. Instead the approach emphasizes the fact that, if the attractions are favorable in both directions and there is room to accommodate electrons, atoms continue to bond until it is unfavorable to do so. This occurs when the outer energy levels of the atoms are full.

Chapter 4, Lesson 5: Energy Levels, Electrons, and Ionic Bonding

Key Concepts

- The attractions between the protons and electrons of atoms can cause an electron to move completely from one atom to the other.
- When an atom loses or gains an electron, it is called an ion.
- The atom that loses an electron becomes a positive ion.
- The atom that gains an electron becomes a negative ion.
- A positive and negative ion attract each other and form an ionic bond.

Summary

Students will look at animations and make drawings of the ionic bonding of sodium chloride (NaCl). Students will see that both ionic and covalent bonding start with the attractions of protons and electrons between different atoms. But in ionic bonding, electrons are transferred from one atom to the other and not shared like in covalent bonding. Students will use Styrofoam balls to make models of the ionic bonding in sodium chloride (salt).

Objective

Students will be able to explain the process of the formation of ions and ionic bonds.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for Each Group

- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker

Materials for Each Student

- 2 small Styrofoam balls
- 2 large Styrofoam balls
- 2 toothpicks

Note: In an ionically bonded substance such as NaCl, the smallest ratio of positive and negative ions bonded together is called a “formula unit” rather than a “molecule.” Technically speaking, the term “molecule” refers to two or more atoms that are bonded together covalently, not ionically. For simplicity, you might want to use the term “molecule” for both covalently and ionically bonded substances.

ENGAGE

1. Show a video of sodium metal reacting with chlorine gas.



Project the video *Sodium and chlorine react*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chlorine_react

Before starting the video, tell students that chlorine is a greenish poisonous gas and sodium is a shiny, soft, and very reactive metal. But when they react, they form sodium chloride (table salt). Tell students that in the video, the drop of water helps expose the atoms at the surface of the sodium so that they can react with the chlorine. The formation of the salt crystals releases a lot of energy.

Note: If students ask if the salt they eat is made this way in salt factories, the answer is no. The salt on Earth was produced billions of years ago but probably not from pure chlorine gas and sodium metal. These days, we get salt from mining it from a mineral called halite or from evaporating sea water.

EXPLAIN

2. Show an animation to introduce the process of ionic bonding.

Project the animation *Ionic bond in sodium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#ionic_bond_in_sodium_chloride

Remind students that in covalent bonding, atoms share electrons. But there is another type of bonding where atoms don't share, but instead either take or give up electrons. This is called ionic bonding. This animation shows a very simplified model of how sodium and chloride ions are formed.

Note: In order to simplify the model of ionic bonding, a single atom of sodium and chlorine are shown. In reality, the chlorine atom would be bonded to another chlorine atom as part of the gas Cl_2 . The sodium atom would be one of billions of trillions of sodium atoms bonded together as a solid. The combination of these substances is a complex reaction between the atoms of the two substances. The animation shows single separated atoms to illustrate the idea of how ions and ionic bonds are formed.

Explain what happens during the animation.

Tell students that the attraction of the protons in the sodium and chlorine for the other atom's electrons brings the atoms closer together. Chlorine has a stronger attraction for electrons than sodium (shown by the thicker arrow). At some point during this process, an electron from the sodium is transferred to the chlorine. The sodium loses an electron and the chlorine gains an electron.

Tell students that when an atom gains or loses an electron, it becomes an *ion*.

- Sodium loses an electron, leaving it with 11 protons, but only 10 electrons. Since it has 1 more proton than electrons, sodium has a charge of +1, making it a positive ion.
- Chlorine gains an electron, leaving it with 17 protons and 18 electrons. Since it has 1 more electron than protons, chlorine has a charge of -1, making it a negative ion.
- When ions form, atoms gain or lose electrons until their outer energy level is full.
 - For example, when sodium loses its one outer electron from the third energy level, the second level becomes the new outer energy level and is full. Since these electrons are closer to the nucleus, they are more tightly held and will not leave.
 - When chlorine gains an electron, its third energy level becomes full. An additional electron cannot join, because it would need to come in at the fourth energy level. This far from the nucleus, the electron would not feel enough attraction from the protons to be stable.
- Then the positive sodium ion and negative chloride ion attract each other and form an ionic bond. The ions are more stable when they are bonded than they were as individual atoms.

3. Have students describe the process of ionic bonding in sodium chloride on their activity sheet.

Give each student an activity sheet.

Have students write a short caption under each picture to describe the process of covalent bonding and answer the first three questions. The rest of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions.



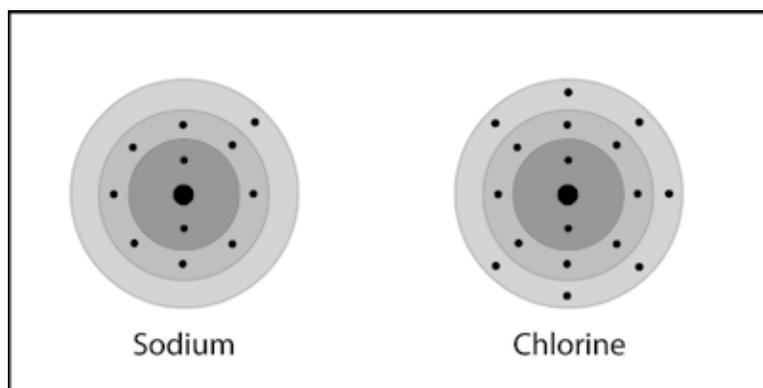
Project the image *Ionic bond in sodium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#_ionic_bond_in_sodium_chloride_2

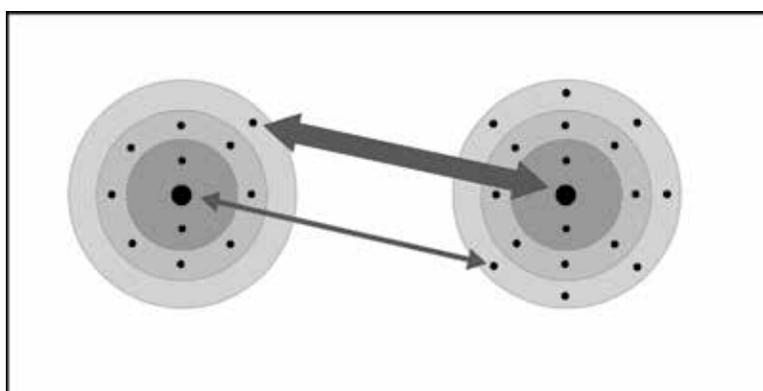
Review with students the process of ionic bonding covered in the animation.



Help students write a short caption beside each picture to describe the process of ionic bonding in sodium and chloride ions.

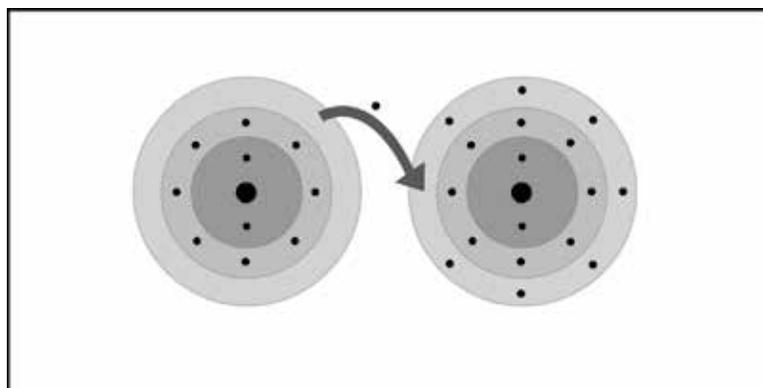


Sodium and chlorine atoms are near each other.

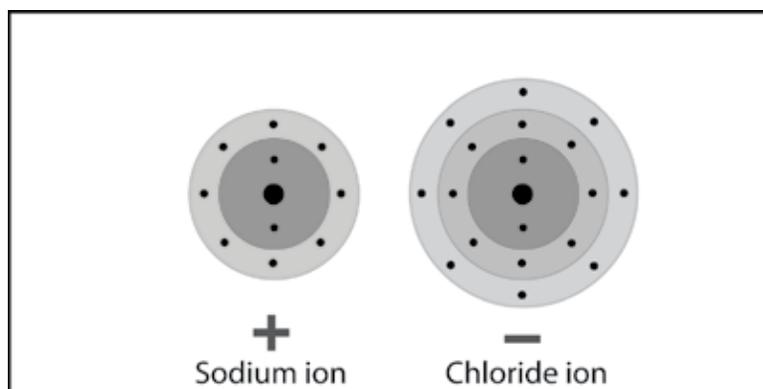


The protons of the two atoms attract the electrons of the other atom.

The thicker arrow shows that chlorine has a stronger attraction for electrons than sodium has.

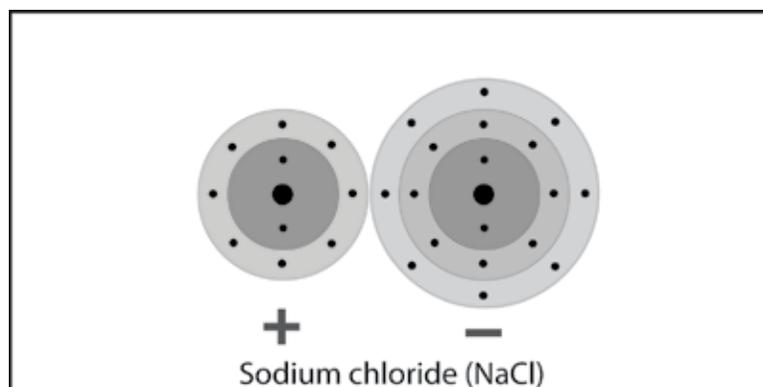


During the interactions between the atoms, the electron in sodium's outer energy level is transferred to the outer energy level of the chlorine atom.



Since sodium *lost* an electron, it has 11 protons, but only 10 electrons. This makes sodium a *positive* ion with a charge of +1.

Since chlorine *gained* an electron it has 17 protons and 18 electrons. This makes chloride a *negative* ion with a charge of -1.



The positive sodium ion and negative chloride ion attract one another. They make an ionic bond and form the ionic compound NaCl.

4. Show students a model of a sodium chloride crystal and have them identify the ions.



Project the image *Sodium chloride crystal*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chloride_crystal

Review with students the process of ionic bonding covered in the animation so that students will understand why the sodium ions are positive and the chloride ions are negative.

Remind students that the scale of any model of atoms, ions, or molecules is enormous compared to the actual size. In a single grain of salt there are billions of trillions of sodium and chloride ions.

Ask students:

- **What ion is the larger ball with the negative charge?**
The chlorine ion.
- **What made it negative?**
It gained an electron.
- **What is the ion with the positive charge?**
The sodium ion.
- **What made it positive?**
It lost an electron.

EXPLORE

5. Have students observe actual sodium chloride crystals and relate their shape to the molecular model.

This two-part activity will help students see the relationship between the arrangement of ions in a model of a sodium chloride crystal and the cubic shape of real sodium chloride crystals.

Teacher preparation

The day before the lesson, dissolve about 10 grams of salt in 50 ml of water. Use Petri dishes or use scissors to cut down 5 or 6 clear plastic cups to make shallow plastic dishes. Pour enough saltwater to just cover the bottom of each dish (1 for each group). Leave the dishes overnight to evaporate so that new salt crystals will be produced.

Materials for each group

- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker



Materials for each student

- 2 small Styrofoam balls
- 2 large Styrofoam balls
- 2 toothpicks

Procedure, Part 1

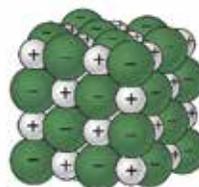
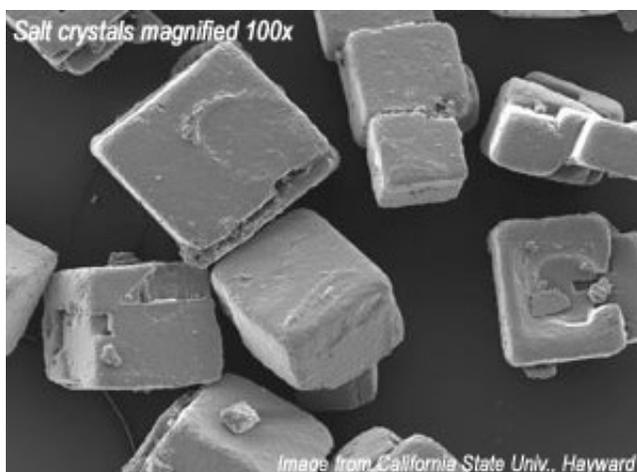
Observe sodium chloride crystals.

1. Place a few grains of salt on a piece of black paper. Use your magnifier to look closely at the salt.
2. Use your magnifier to look at the salt crystals in the cup.

Project the image *Cubic sodium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#cubic_sodium_chloride

The image shows both a magnified view of ordinary table salt and a model of the sodium and chloride ions that make up a salt crystal.



Project the animation *Sodium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#sodium_chloride_crystal

The green spheres represent negatively charged chloride ions and the gray spheres represent positively charged sodium ions.

Ask students:

- **What do the photograph, molecular model, and your observations of real salt crystals tell you about the structure of salt?**

In each case, the salt seems to be shaped like a cube.

6. Have students build a 3-dimensional model of sodium chloride.



Each student will make 1 unit of sodium chloride. Students in each group will put their sodium chloride units together. You can help the groups combine their structures into a class model of a sodium chloride crystal.

Procedure, Part 2

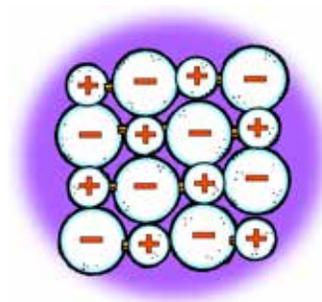
Make NaCl units.

1. Use the marker to put a “-” on the large balls which represent chloride ions.
2. Use the marker to put a “+” on the small balls, which represent sodium ions.
3. Break two toothpicks in half. Use one of the half-toothpicks to connect the centers of the small and large ions together to make a unit of sodium chloride (NaCl). Do the same thing with the other small and large ball.
4. Use another half-toothpick to connect the two NaCl units in a straight line as shown.



Put NaCl ions together to make one layer of ions.

5. Contribute your line of ions to your group and arrange them to make a 4×4 square of ions.
6. Use half-toothpicks to attach the ends of each line to hold the ions together. You only need to place toothpicks in the balls at the end of each line.

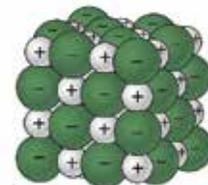


Build a class sodium chloride crystal.

7. Give your group's layer of ions to your teacher. Your teacher will stack these to build a model of a sodium chloride crystal.



Point out that anywhere you look on the crystal, a sodium ion and a chloride ion are always surrounded by the oppositely charged ion. These opposite charges hold the ions together in a crystal.



Ask students

- **Based on the way sodium and chloride ions bond together, why are salt crystals shaped like cubes?**

The size and arrangement of the ions forms a cube on the molecular level. Since the pattern repeats over and over again in the same way, the shape stays the same even when the crystal becomes the normal size that we can see.

EXTEND

7. Show students how calcium and chlorine atoms bond to form the ionic compound calcium chloride.

Tell students that there is another common substance called calcium chloride (CaCl_2). It is the salt that is used on icy sidewalks and roads. Explain that when calcium and chlorine react they produce ions, like sodium and chlorine, but the calcium ion is different from the sodium ion.

Ask students:

- **What ions do you think CaCl_2 is made of?**
One calcium ion and two chloride ions.



Project the animation *Calcium chloride Ionic Bond*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson5#calcium_chloride_ion

Point out that the calcium loses two electrons, becoming a +2 ion. Each of the two chlorine atoms gains one of these electrons, making them each a -1 ion. Help students realize that 1 calcium ion bonds with 2 chloride ions to form calcium chloride (CaCl_2), which is neutral.

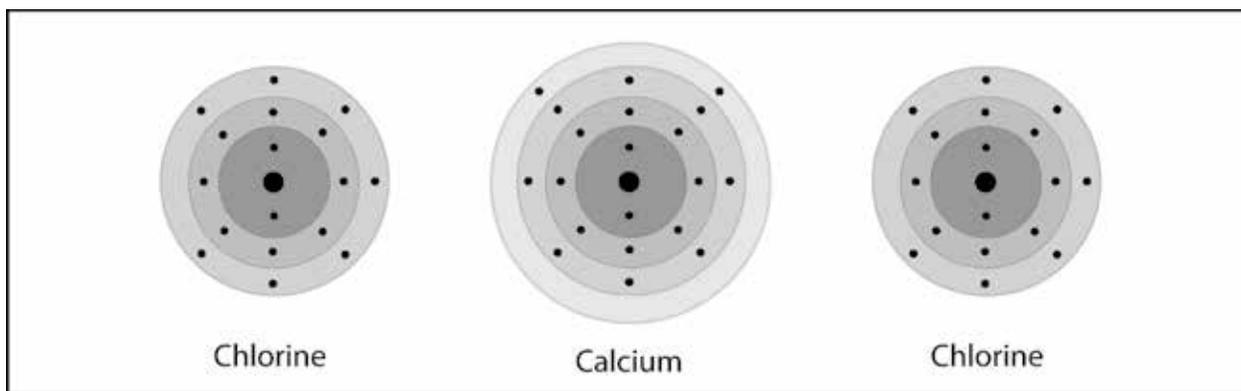
Some atoms gain or lose more than 1 electron. Calcium loses 2 electrons when it becomes an ion. When ions come together to form an ionic bond, they always join in numbers that exactly cancel out the positive and negative charge.

Project the image *Calcium chloride Ionic Bond*.

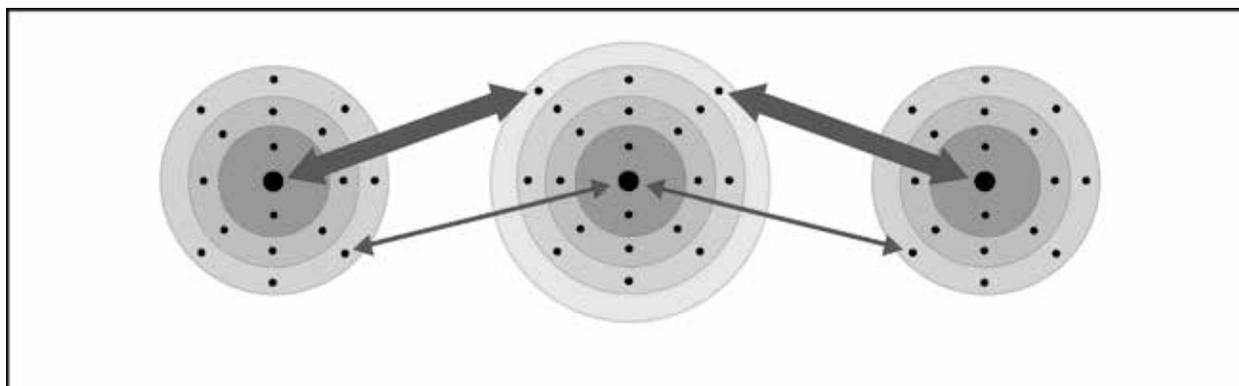
www.middleschoolchemistry.com/multimedia/chapter4/lesson5#calcium_chloride_ion_2

Review with students the process of ionic bonding covered in the animation.

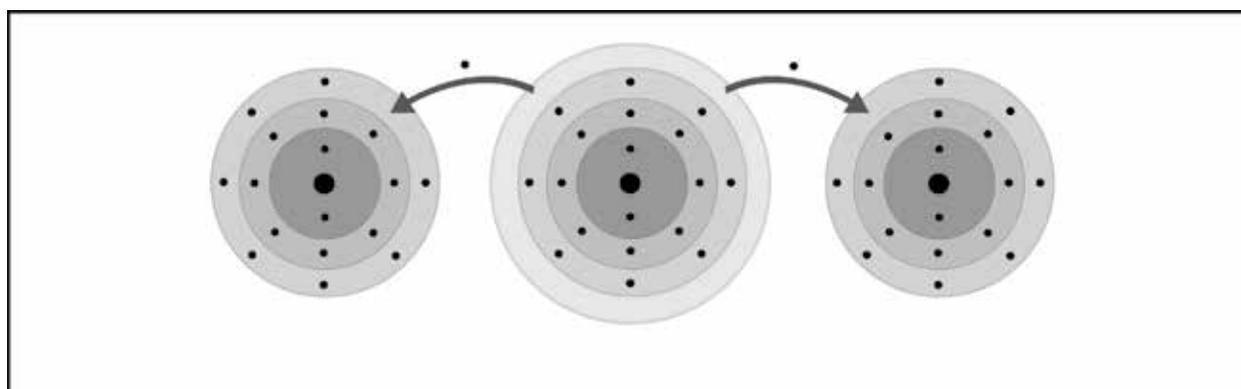
Have students write a short caption beneath each picture to describe the process of ionic bonding in sodium and chloride ions.



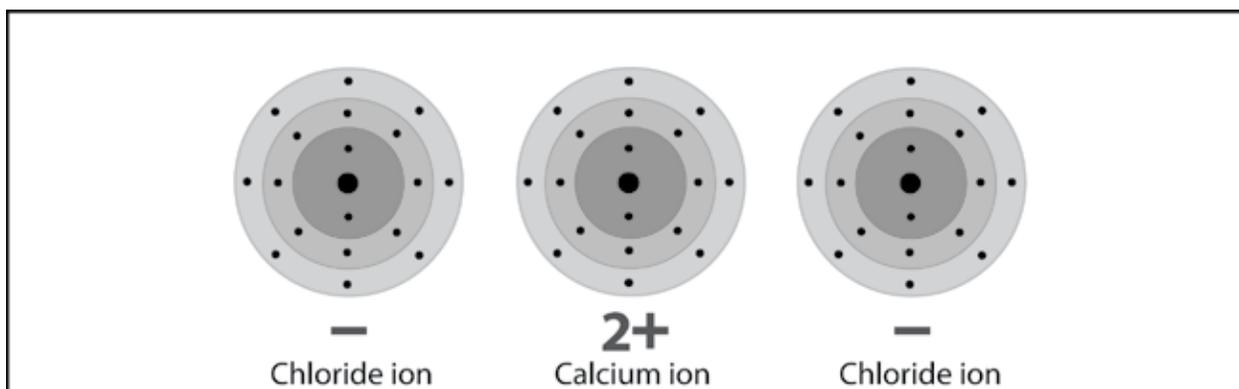
One calcium and two chlorine atoms are near each other.



The protons of the calcium atom attract the electrons from the chlorine atom. The protons of the two chlorine atoms attract the electrons from the calcium atom more strongly as shown by the thicker arrows.

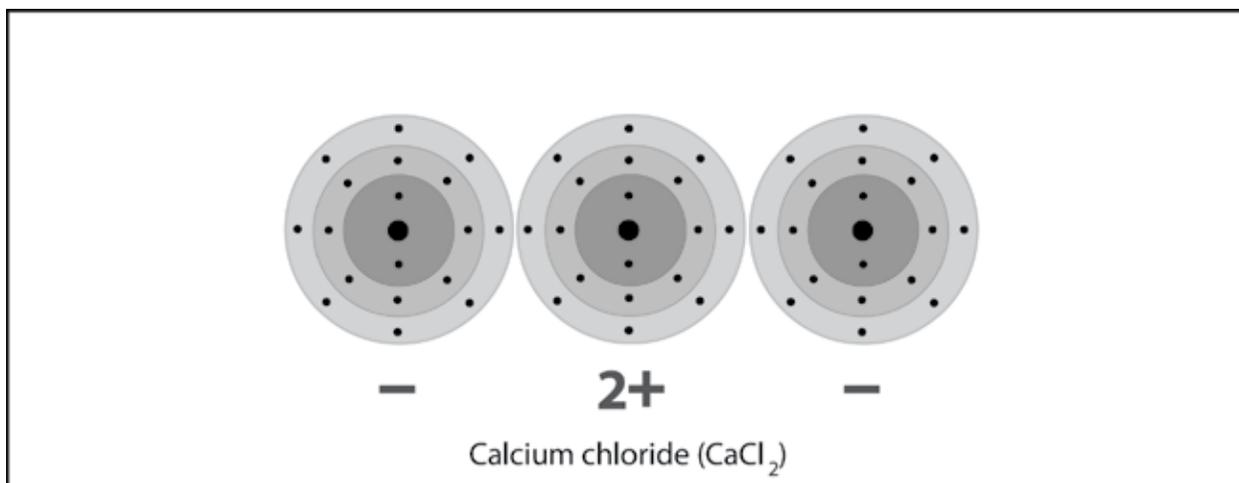


During the interactions between the atoms, the two electrons in calcium's outer energy level are transferred to the outer energy level of each of the chlorine atoms.



Since calcium *lost* two electrons, it has 20 protons, but only 18 electrons. This makes calcium a *positive* ion with a charge of 2+.

Since each chlorine atom *gained* an electron, they each have 17 protons and 18 electrons. This makes each chloride a *negative* ion with a charge of -1.

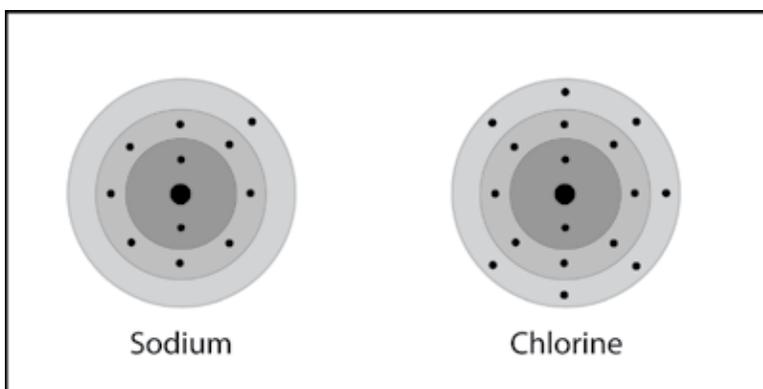


Oppositely charged ions attract each other, forming an ionic bond. The bonded ions are more stable than the individual atoms were.

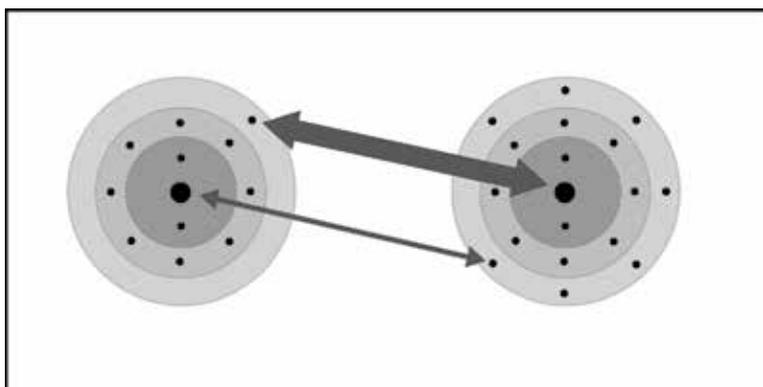
EXPLAIN IT WITH ATOMS & MOLECULES

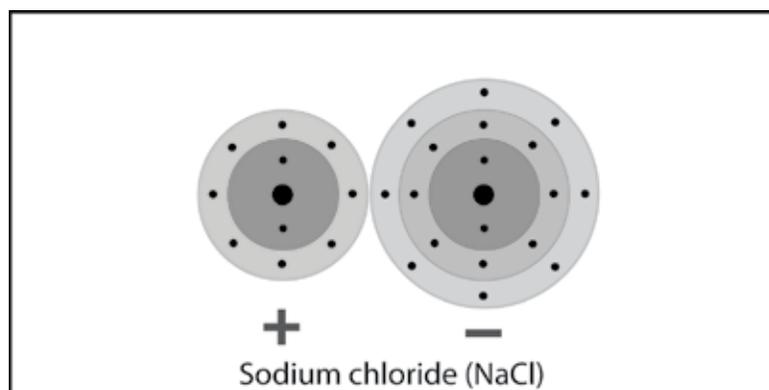
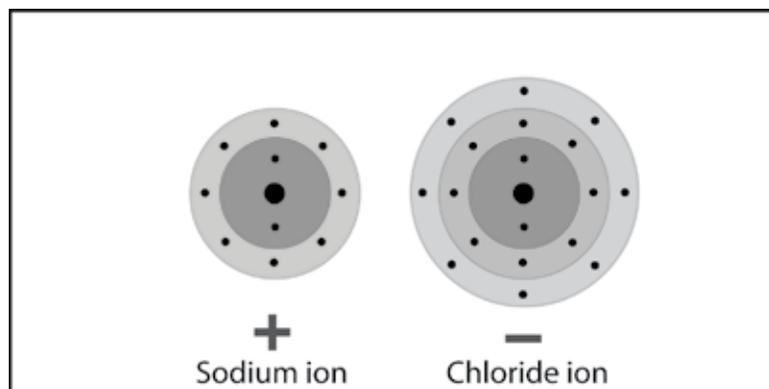
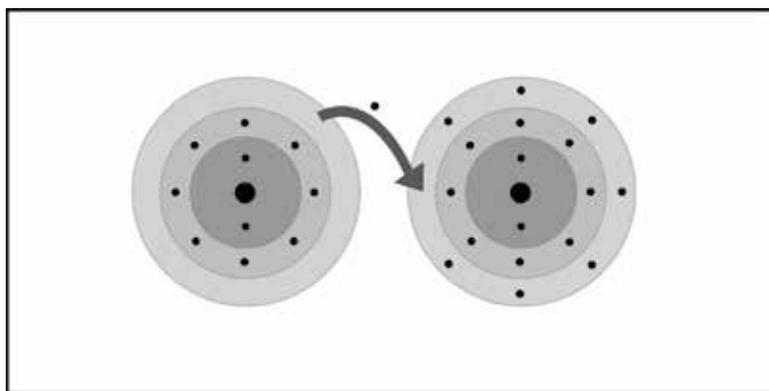
1. What is the basic difference between covalent and ionic bonding?

2. Write a short caption beside each picture to describe the process of ionic bonding.



Sodium and chlorine atoms are near each other.





ACTIVITY



Question to investigate

Why are salt crystals cube-shaped?

Materials for each group

- Black paper
- Salt
- Cup with salt from evaporated saltwater
- Magnifier
- Permanent marker

Materials for each student

- 2 small Styrofoam balls
- 2 large Styrofoam balls
- 2 toothpicks

Procedure, Part 1

Observe sodium chloride crystals.

1. Place a few grains of salt on a piece of black paper. Use your magnifier to look closely at the salt.
2. Use your magnifier to look at the salt crystals in the cup.

Procedure, Part 2

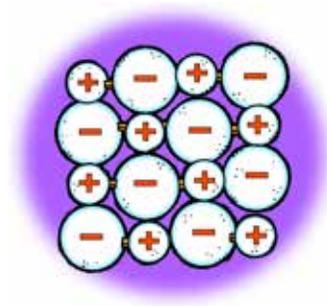
Make NaCl units.

1. Use the marker to put a “-” on the large balls, which represent chloride ions.
2. Use the marker to put a “+” on the small balls, which represent sodium ions.
3. Break two toothpicks in half. Use one of the half-toothpicks to connect the centers of the small and large ions together to make a unit of sodium chloride (NaCl). Do the same thing with the other small and large ball.
4. Use another half-toothpick to connect the two NaCl units in a straight line as shown.



Put NaCl ions together to make one layer of ions.

5. Contribute your line of ions to your group and arrange them to make a 4×4 square of ions.
6. Use half-toothpicks to attach the ends of each line to hold the ions together. You only need to place toothpicks in the balls at the end of each line.



Build a class sodium chloride crystal.

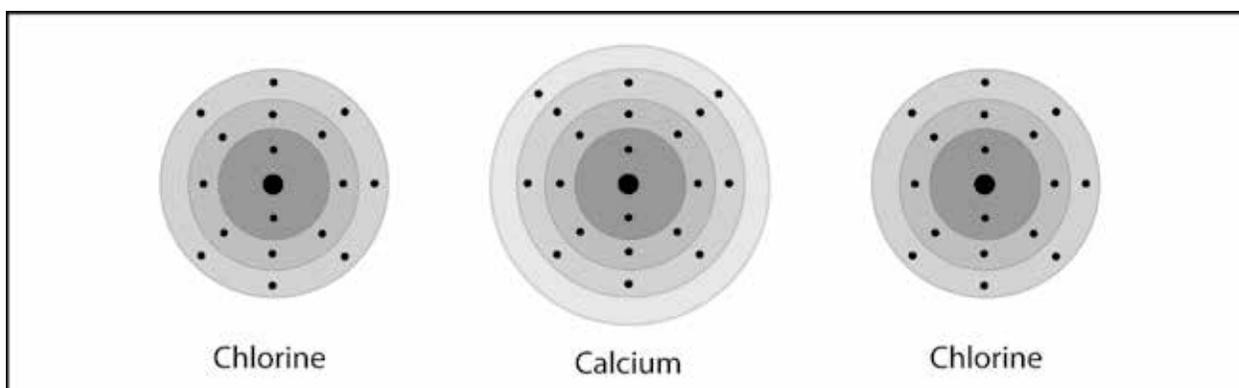
7. Give your group's layer of ions to your teacher. Your teacher will stack these to build a model of a sodium chloride crystal.



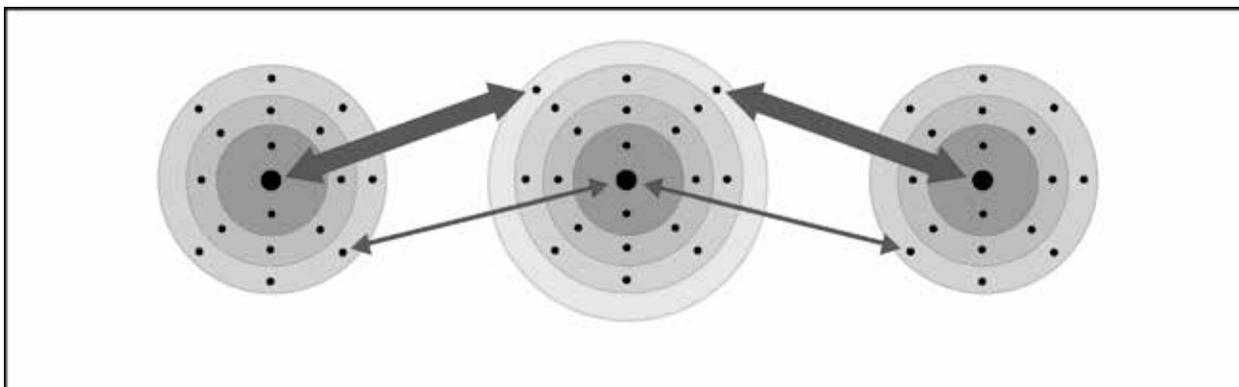
3. Knowing what you do about sodium and chloride ions, why are salt crystals cube-shaped?

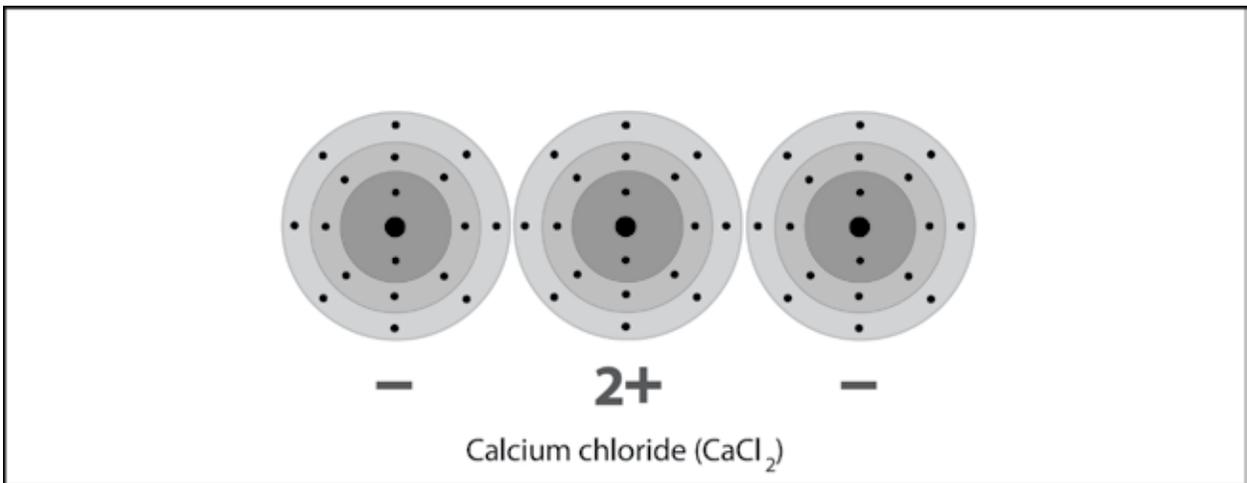
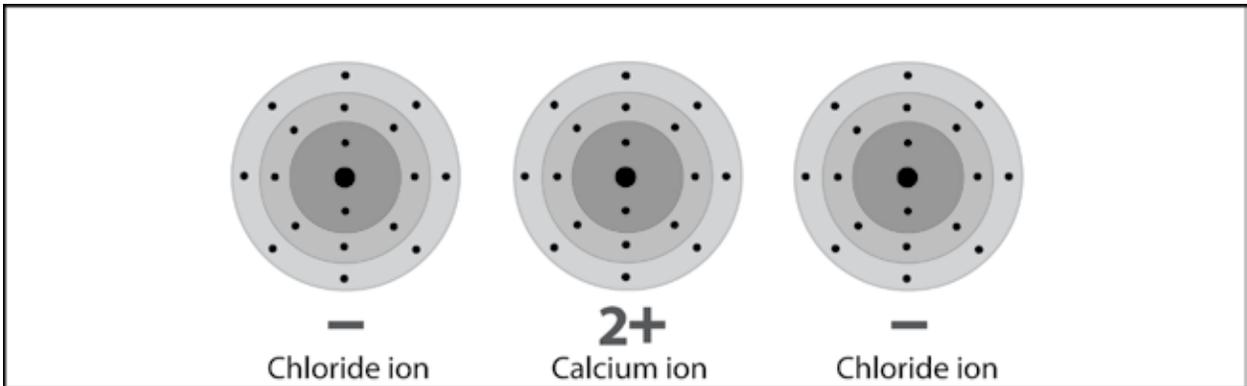
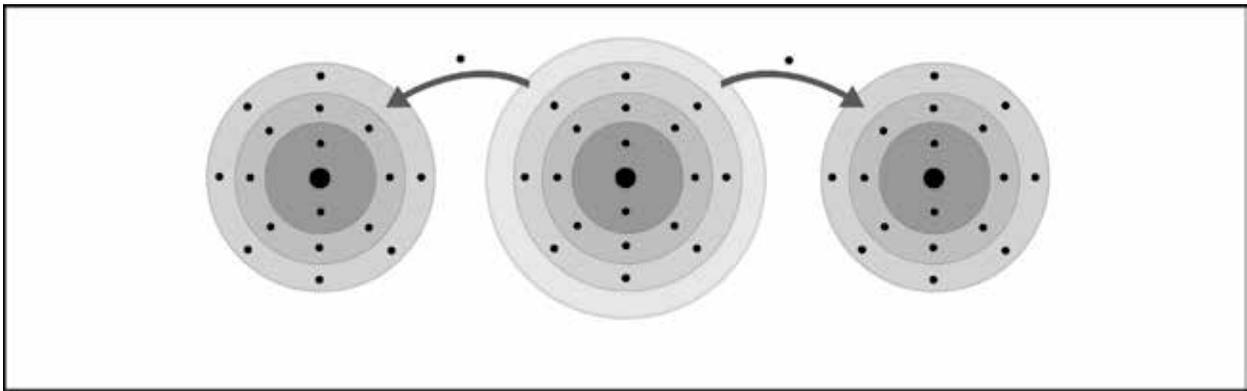
TAKE IT FURTHER

4. Write a short caption beneath each picture below and on the next page to describe the process of ionic bonding. The first one is done for you below.



One calcium and two chlorine atoms are near each other.





Chapter 4, Lesson 6: Represent Bonding with Lewis Dot Diagrams

If you are required to teach Lewis dot structures, this short lesson can help you extend what students have learned about modeling covalent and ionic bonding. Since there is no hands-on activity component, this lesson is not in a 5-E lesson plan format.

Key Concepts

- There are shorthand ways to represent how atoms form covalent or ionic bonds.
- Lewis dot diagrams use dots arranged around the atomic symbol to represent the electrons in the outermost energy level of an atom.
- Single bonds are represented by a pair of dots or one line between atoms.
- Double bonds are represented by two pairs of dots or two lines between atoms.
- Triple bonds are represented by three pairs of dots or three lines between atoms.

Summary

Students will be introduced to the basics of Lewis dot diagrams as they compare the energy level models used in chapter 4 to dot diagrams. Along with the teacher, they will review the Lewis dot diagrams for a few common covalent and ionic compounds.

Objective

Students will be able to interpret and draw Lewis dot diagrams for individual atoms and both covalent and ionic compounds.

Evaluation

The activity sheet serves as a formative assessment and gives students practice interpreting Lewis dot diagrams. A more formal summative assessment is included at the end of each chapter.

About this Lesson

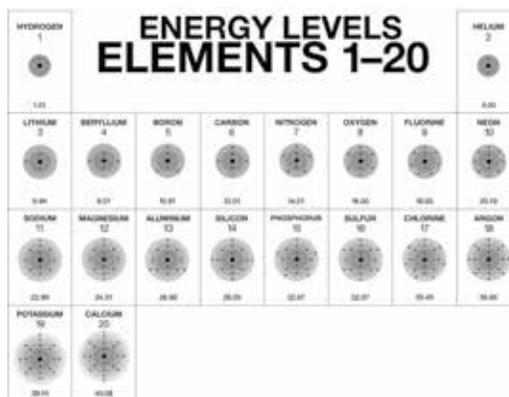
The model of the atom and of covalent and ionic bonding that students have used so far emphasizes the attractions between bonding atoms. The nucleus, electrons, and double-headed arrows show that the protons and electrons from one atom attract the oppositely charged electrons and protons of the other atom, resulting in bonding. The energy levels show that only valence electrons are involved in bonding.

After students understand the important role of attraction of opposite charges, you may introduce them to a common, more symbolic, short-hand way of showing how atoms are bonded together. This information is offered if you feel that showing students these other models of bonding would be useful or if you need to cover basic information about Lewis dot diagrams to satisfy your curriculum.

1. Introduce students to Lewis dot structures.

Tell students that one popular method of representing atoms is through Lewis dot diagrams. In a dot diagram, only the symbol for the element and the electrons in its outermost energy level (valence electrons) are shown.

Note: In the energy level diagrams students have been using, the electrons are spread out evenly in each energy level. Some books show them spread out this way and some show them in pairs. For Lewis dot structures, they are always shown in pairs. This is to indicate that electrons are in separate orbitals within each energy level. It is not necessary for middle school students to learn about electron orbitals. This information is offered so that it is clearer to you why electrons are often shown in pairs in energy level diagrams and in dot diagrams. An orbital is a 3-dimensional space within an energy level where there is a high probability of finding electrons. The further the energy level is from the nucleus, the more orbitals it has. There can be a maximum of two electrons in each orbital. This is why the electrons are shown in pairs.



Have students look at the activity sheet for chapter 4, lesson 3 or distribute the energy level chart at the end of this lesson. They will need to compare the energy levels for elements 1–20 that they completed with the chart you show them.

Project the image *Lewis dot diagrams for elements 1–20.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson6#lewis_dot_diagrams

LEWIS DOT DIAGRAMS ELEMENTS 1–20							
HYDROGEN 1 H 1.01							HELIUM 2 He 4.00
LITHIUM 3 Li 6.94	BERYLLIUM 4 Be 9.01	BORON 5 B 10.81	CARBON 6 C 12.01	NITROGEN 7 N 14.01	OXYGEN 8 O 16.00	FLUORINE 9 F 18.99	NEON 10 Ne 20.18
SODIUM 11 Na 22.99	MAGNESIUM 12 Mg 24.31	ALUMINUM 13 Al 26.98	SILICON 14 Si 28.09	PHOSPHORUS 15 P 30.97	SULFUR 16 S 32.07	CHLORINE 17 Cl 35.45	ARGON 18 Ar 39.95
POTASSIUM 19 K 39.10	CALCIUM 20 Ca 40.08						

Ask students:

- **Compare the dots around each symbol with the energy levels in your chart. What relationship do you notice between the dots in these two charts?**
The dots represent the electrons in the outer energy level (valence electrons) from the energy level models.
- **The number of dots near hydrogen and helium are the same as in the energy level chart. Why?**
The only electrons hydrogen and helium have are valence electrons. All the other dot structures should have fewer electrons than the energy level model because the dot model only shows the outermost electrons.

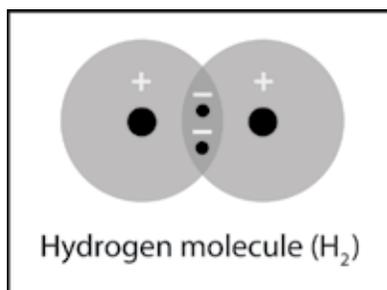
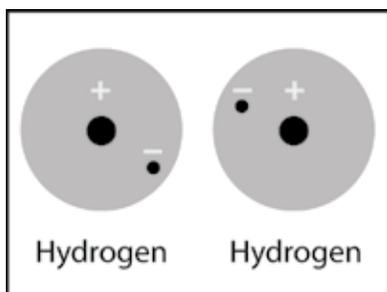
2. Help students recognize the similarities between energy level models and Lewis dot structures that show bonding.

Project the image *Covalent bonding in hydrogen*.

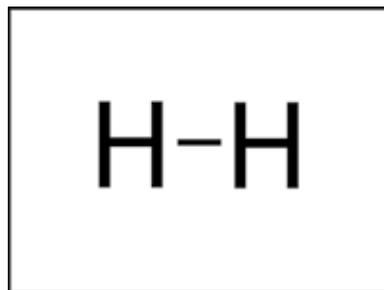
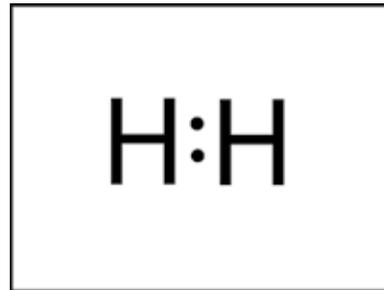
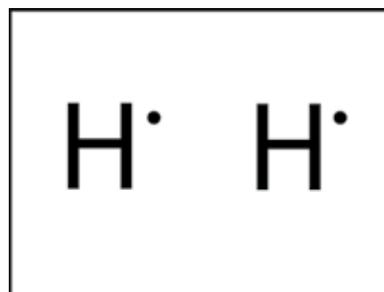
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covalent_bonding_hydrogen

This image shows both the energy level model and Lewis dot structure of two hydrogen atoms before and after bonding.

Energy level model



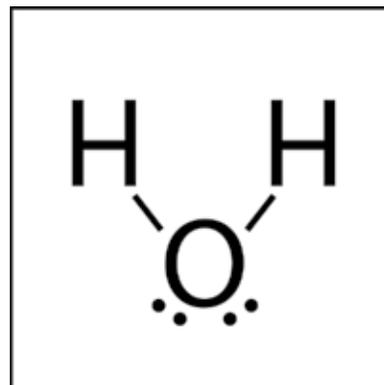
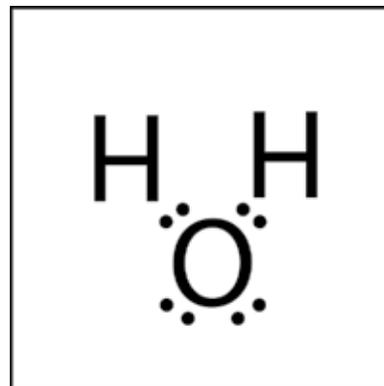
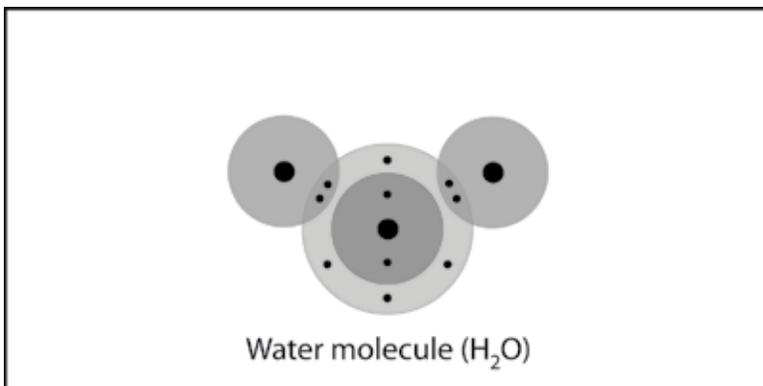
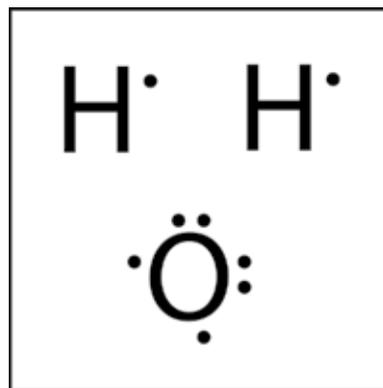
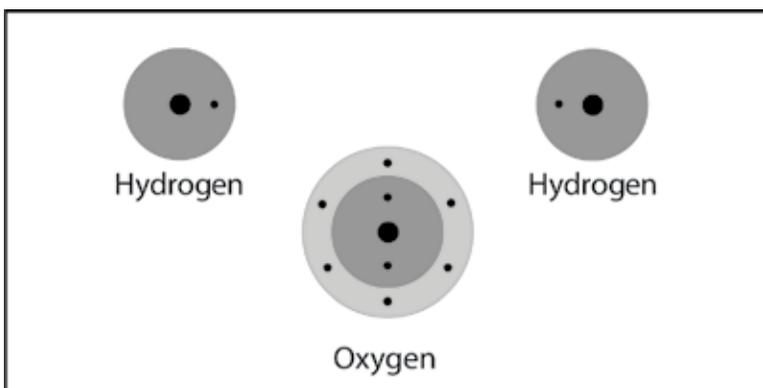
Lewis dot model



Explain to students that in a dot diagram, the electrons that are shared in the bond are placed between the symbol for each atom. Remind students that the electrons between the two atoms are shared and are counted as belonging to each atom. Show students that in the energy level model for the hydrogen molecule, two electrons are shared. The Lewis dot diagram for the hydrogen molecule also shows that two electrons are shared. There is an even more shorthand approach that shows the bond as a line. The line represents one pair of electrons.

Project the image *Covalent bonding in water.*

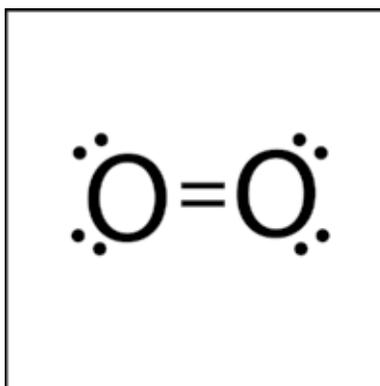
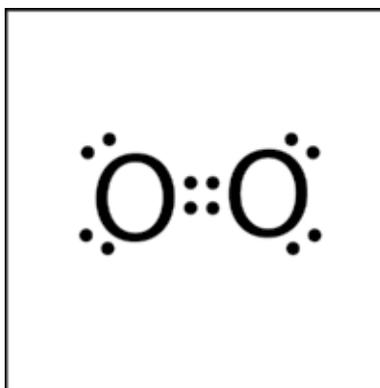
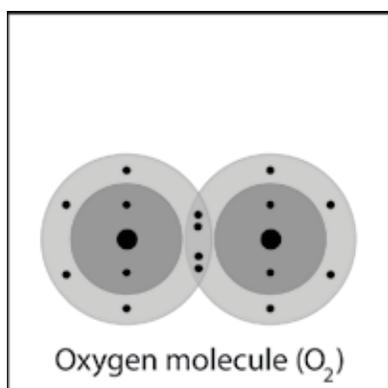
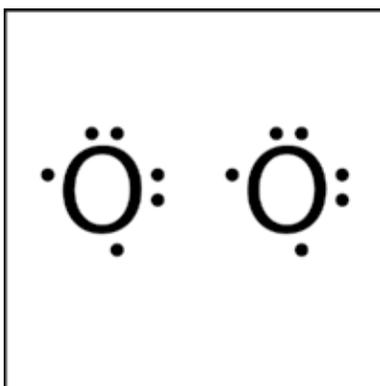
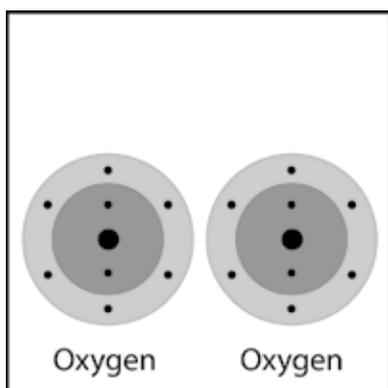
www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covalent_bonding_water



Be sure students notice that the number of dots around the oxygen atom in the Lewis diagram is the same as the number of electrons in the outer energy level of the energy level model. Remind students that the electrons between the atoms are shared and are counted as if they belong to each atom. Show students that in the energy level model for the water molecule, two pairs of electrons are shared. The Lewis dot diagram for the water molecule also shows that two pairs of electrons are shared. The line represents one pair of shared electrons.

Project the image *Covalent bonding in oxygen.*

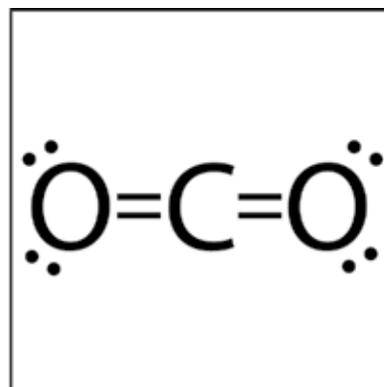
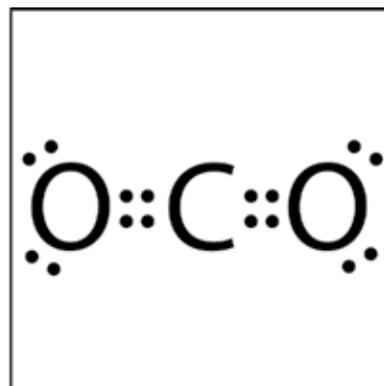
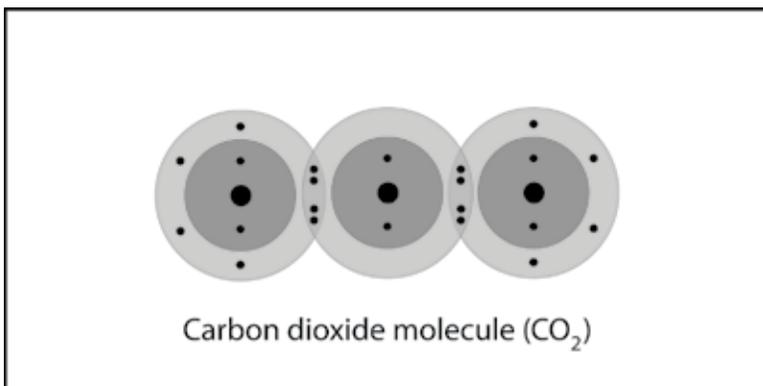
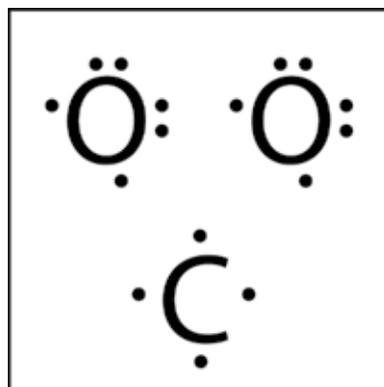
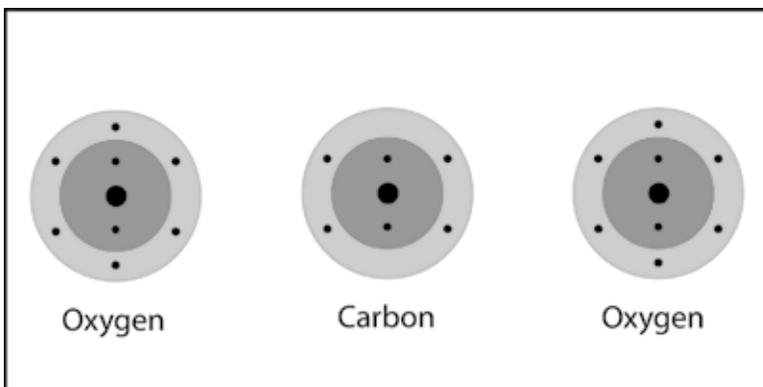
www.middleschoolchemistry.com/multimedia/chapter4/lesson6/#covalent_bond_oxygen



Show students that in the energy level model for the oxygen molecule, two pairs of electrons are shared. The Lewis dot diagram for the water molecule also shows that two pairs of electrons are shared. The remaining electrons are shown paired up around each oxygen atom. In the alternate Lewis dot diagram, there are two lines because there are two pairs of electrons that are shared.

Project the image *Covalent bonding in carbon dioxide.*

www.middleschoolchemistry.com/multimedia/chapter4/lesson6#covlalent_bonding_carbon_dioxide



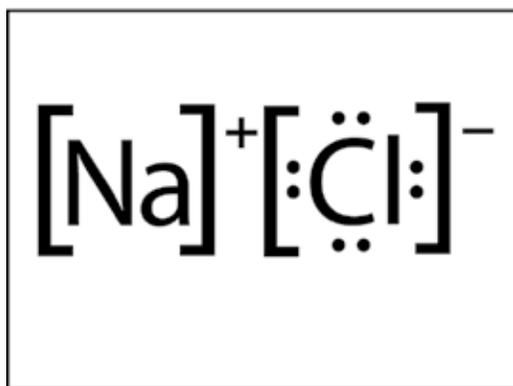
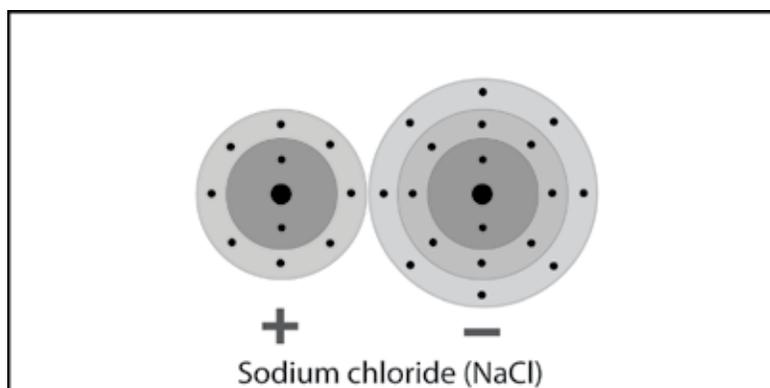
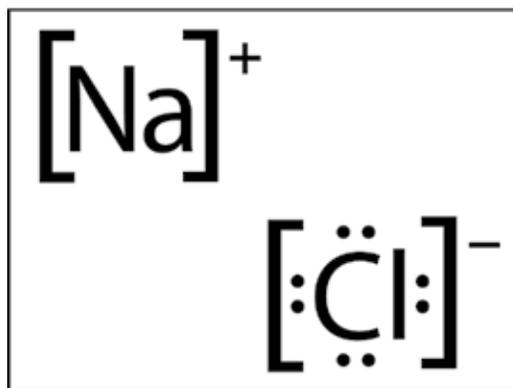
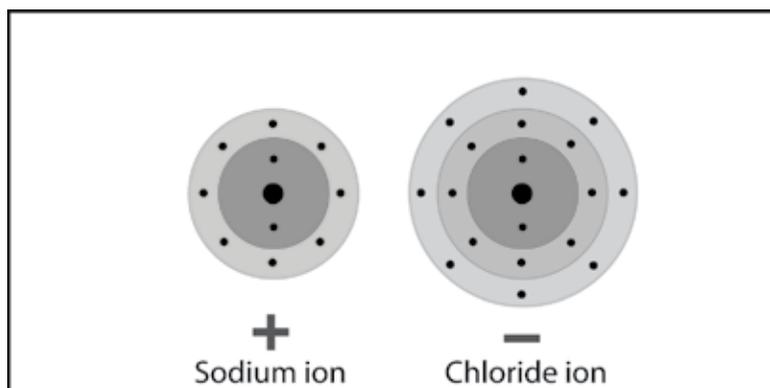
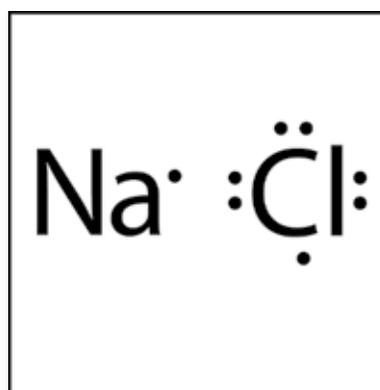
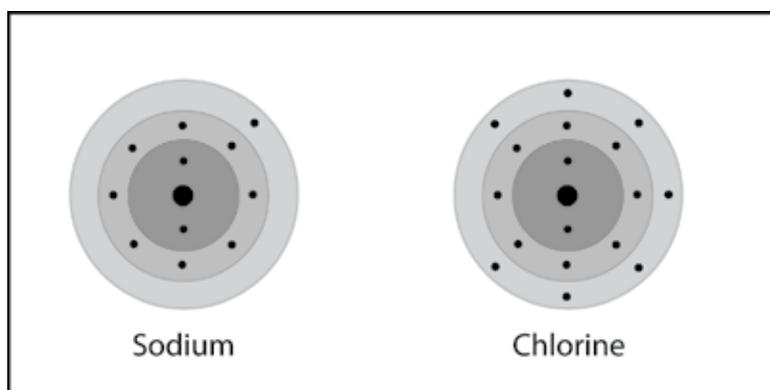
Show students that in the energy level model for carbon dioxide, two pairs of electrons are shared with each oxygen atom. The Lewis dot diagram for carbon dioxide also shows that two pairs of electrons are shared. The remaining electrons are shown paired up around each oxygen atom. In the alternate Lewis dot diagram, there are two lines between each atom to show that two pairs of electrons are shared.

3. Show how Lewis dot diagrams also represent ionic bonding.

Tell students that dot diagrams can also be used to show ionic bonding.

Project the image *Ionic bonding of sodium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson6#ionic_bond_sodium_chloride

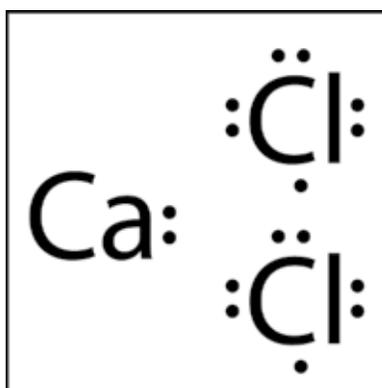
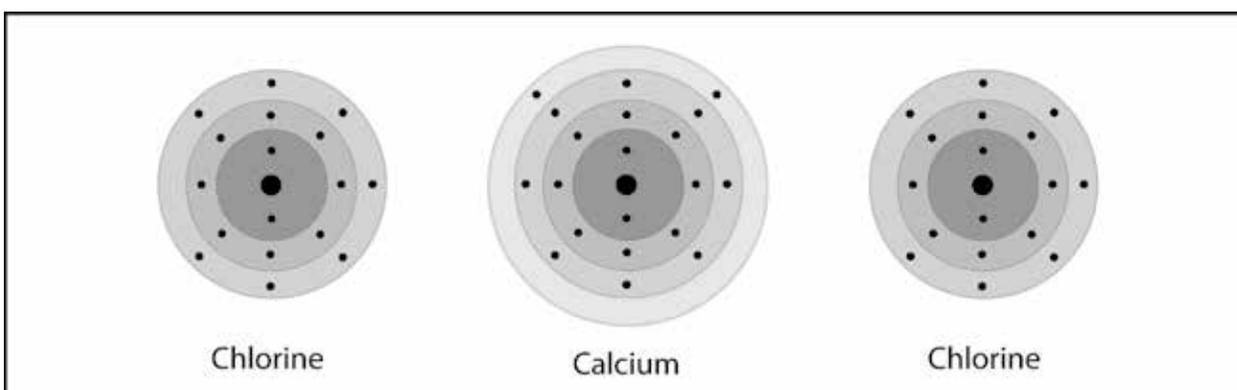


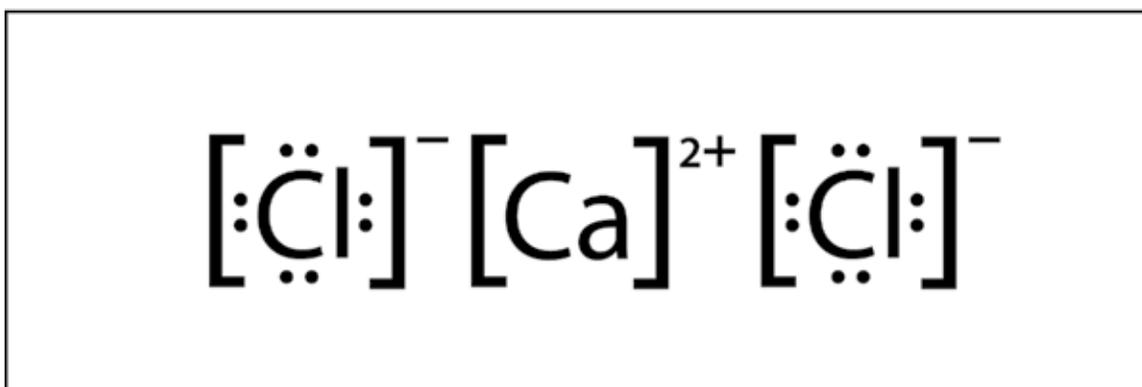
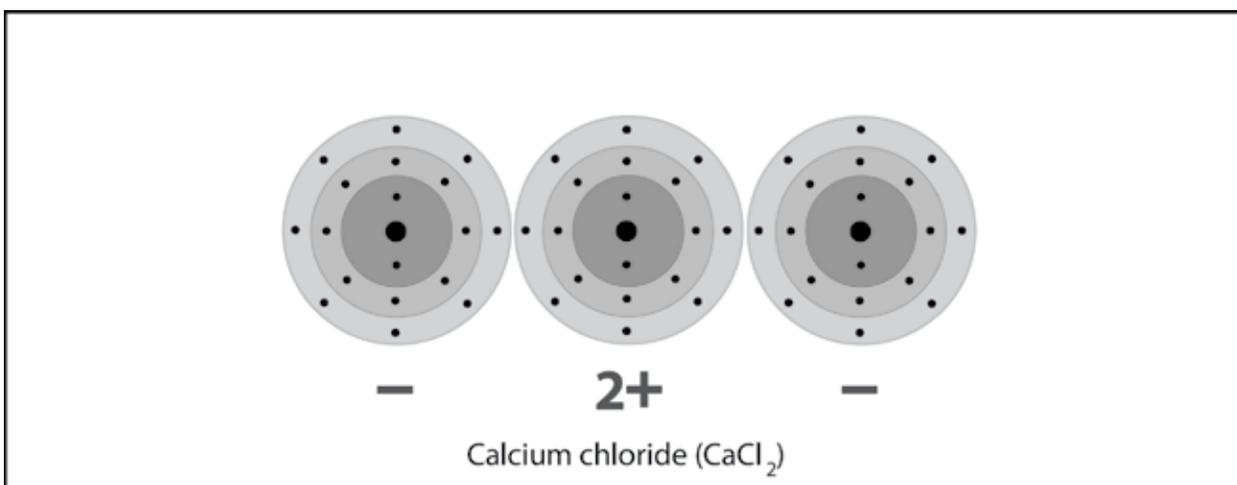
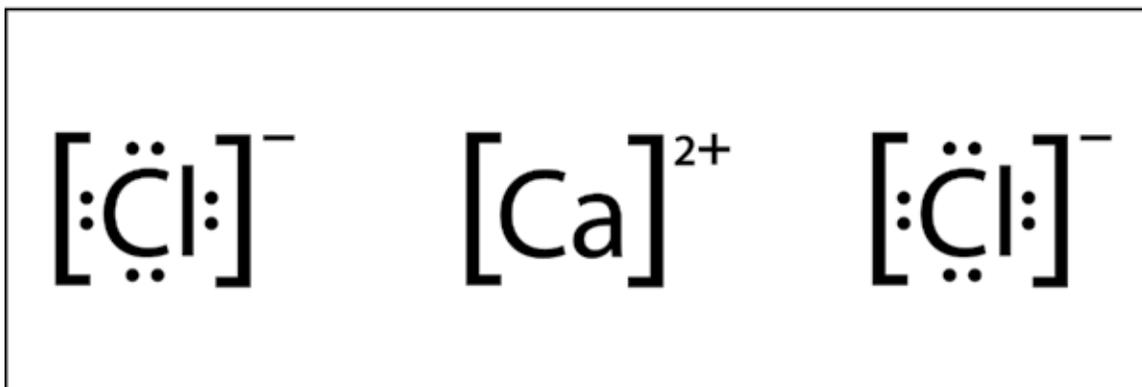
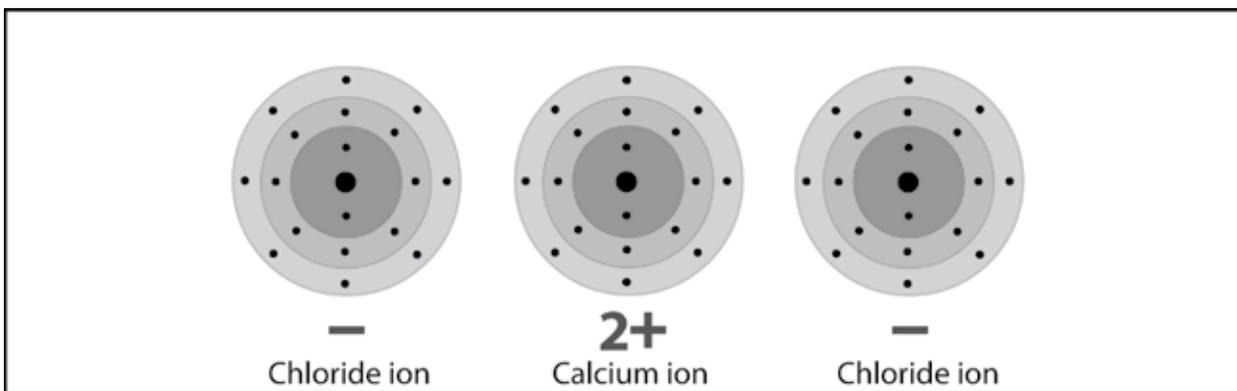
Ask students:

- **In the second dot diagram, why are there no electrons surrounding sodium?**
The electron was transferred to chlorine. The dot diagram only shows electrons in the atom's outermost energy level. No electrons are shown since the only electron in the outermost energy level of sodium was transferred to chlorine.
- **In the final dot diagram of NaCl, the dots between the sodium and chlorine are between the atoms. Are these atoms sharing the electrons?**
No. All electrons shown in the dot diagram belong to chlorine. The pair of electrons is between the two letters only because the symbols are shown close to each other to represent the attraction between the oppositely charged ions.

Project the image *Ionic bonding of calcium chloride*.

www.middleschoolchemistry.com/multimedia/chapter4/lesson6#ionic_bonding_calcium_chloride





Ask students:

- **What do you think the 2+ near the calcium means?**

Calcium had two electrons in its outer energy level. It gave each chlorine one of these electrons. Because the calcium has two more protons than electrons it has a charge of 2+.

- **In the final dot diagram of CaCl_2 , the dots between the calcium and chlorine are between the atoms. Are these atoms sharing the electrons?**

No. All electrons shown in the dot diagram belong to chlorine. The pair of electrons is between the two letters only because the symbols are shown close to each other to represent the attraction between the oppositely charged ions.

Activity Sheet
Chapter 4, Lesson 6
Represent Bonding with Lewis Dot Diagrams

Name _____

Date _____

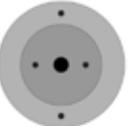
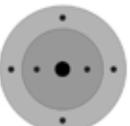
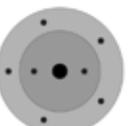
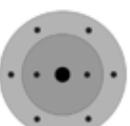
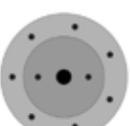
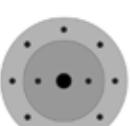
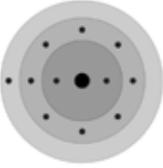
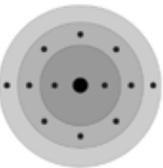
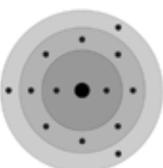
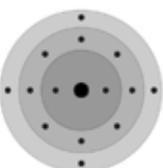
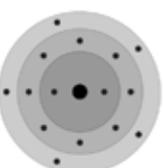
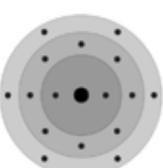
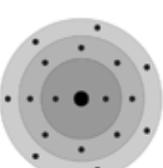
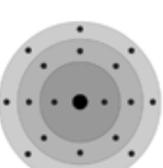
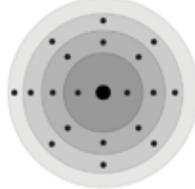
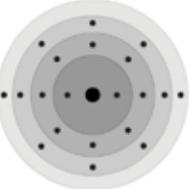
In chapter 4, you saw energy level models for each atom that used concentric circles to represent energy levels and dots for electrons. These diagrams were also used to show what happens to the electrons when different atoms bond. Sometimes electrons were shared (covalent bonding) and sometimes electrons were transferred from one atom to another (ionic bonding).

There is a common, shorthand way to represent bonding called Lewis dot diagrams. Dots still represent electrons, but they are drawn around the symbol for the element. And only the electrons in the outermost energy level are drawn.

1. Compare the periodic table of energy levels to the Lewis dot diagrams. Look at the dots around each symbol and the energy levels in your chart. What relationship do you notice between the dots in these two charts?

2. The number of dots near hydrogen and helium are the same as in the energy level chart. Why?

ENERGY LEVELS ELEMENTS 1-20

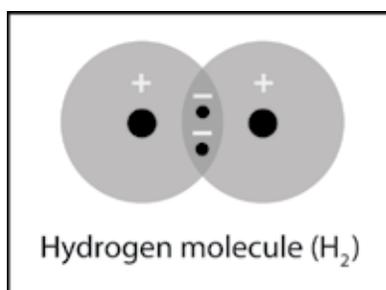
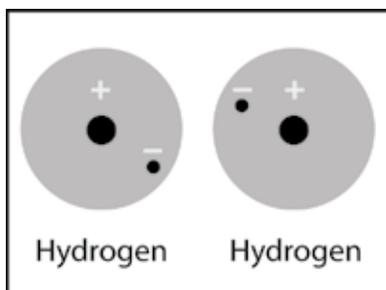
<p>HYDROGEN 1</p>  <p>1.01</p>							<p>HELIUM 2</p>  <p>4.00</p>
<p>LITHIUM 3</p>  <p>6.94</p>	<p>BERYLLIUM 4</p>  <p>9.01</p>	<p>BORON 5</p>  <p>10.81</p>	<p>CARBON 6</p>  <p>12.01</p>	<p>NITROGEN 7</p>  <p>14.01</p>	<p>OXYGEN 8</p>  <p>16.00</p>	<p>FLUORINE 9</p>  <p>19.00</p>	<p>NEON 10</p>  <p>20.18</p>
<p>SODIUM 11</p>  <p>22.99</p>	<p>MAGNESIUM 12</p>  <p>24.31</p>	<p>ALUMINUM 13</p>  <p>26.98</p>	<p>SILICON 14</p>  <p>28.09</p>	<p>PHOSPHORUS 15</p>  <p>30.97</p>	<p>SULFUR 16</p>  <p>32.07</p>	<p>CHLORINE 17</p>  <p>35.45</p>	<p>ARGON 18</p>  <p>39.95</p>
<p>POTASSIUM 19</p>  <p>39.10</p>	<p>CALCIUM 20</p>  <p>40.08</p>						

LEWIS DOT DIAGRAMS ELEMENTS 1-20

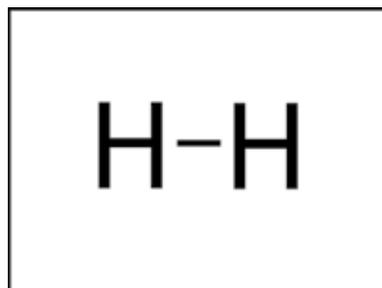
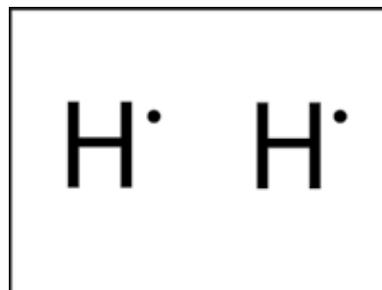
<p>HYDROGEN 1</p> <p>H·</p> <p>1.01</p>	<p>HELIUM 2</p> <p>He·</p> <p>4.00</p>						
<p>LITHIUM 3</p> <p>Li·</p> <p>6.94</p>	<p>BERYLLIUM 4</p> <p>Be·</p> <p>9.01</p>	<p>BORON 5</p> <p>·B·</p> <p>10.81</p>	<p>CARBON 6</p> <p>·C·</p> <p>12.01</p>	<p>NITROGEN 7</p> <p>·N·</p> <p>14.01</p>	<p>OXYGEN 8</p> <p>·Ö·</p> <p>16.00</p>	<p>FLUORINE 9</p> <p>·F·</p> <p>19.00</p>	<p>NEON 10</p> <p>·Ne·</p> <p>20.18</p>
<p>SODIUM 11</p> <p>Na·</p> <p>22.99</p>	<p>MAGNESIUM 12</p> <p>Mg·</p> <p>24.31</p>	<p>ALUMINUM 13</p> <p>·Al·</p> <p>26.98</p>	<p>SILICON 14</p> <p>·Si·</p> <p>28.09</p>	<p>PHOSPHORUS 15</p> <p>·P·</p> <p>30.97</p>	<p>SULFUR 16</p> <p>·S·</p> <p>32.07</p>	<p>CHLORINE 17</p> <p>·Cl·</p> <p>35.45</p>	<p>ARGON 18</p> <p>·Ar·</p> <p>39.95</p>
<p>POTASSIUM 19</p> <p>K·</p> <p>39.10</p>	<p>CALCIUM 20</p> <p>Ca·</p> <p>40.08</p>						

Covalent bonding in the hydrogen molecule, H₂

Energy level model



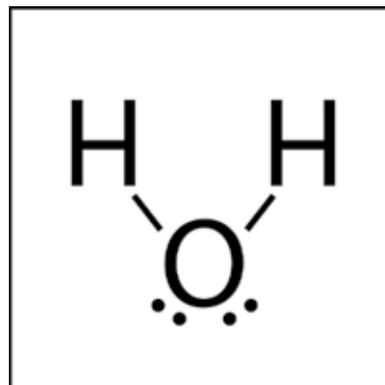
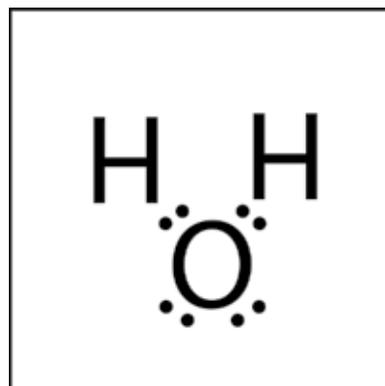
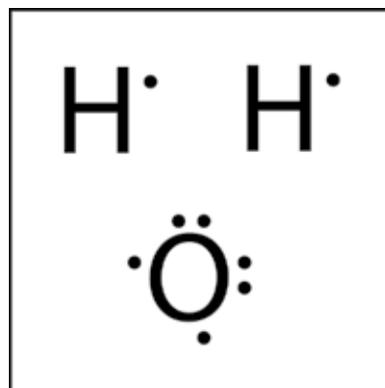
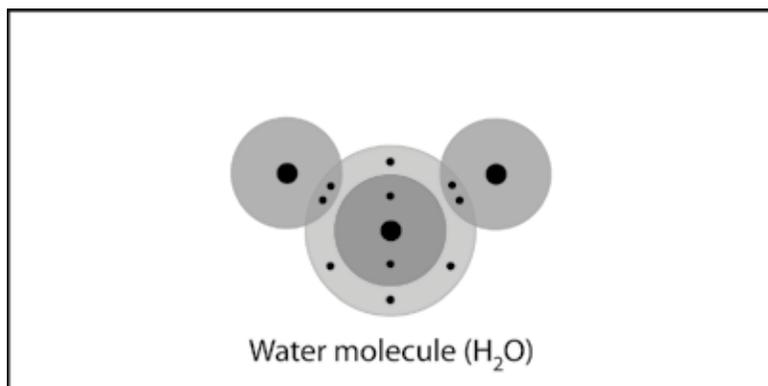
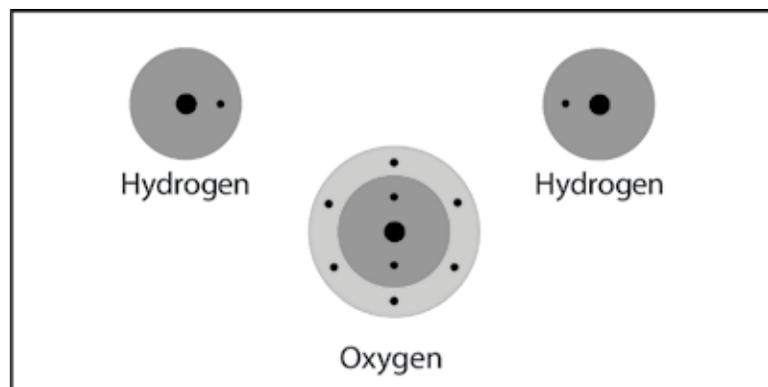
Lewis dot model



3. What do the pair of dots between the two letters "H" represent?

4. What does the line between the two letters "H" represent?

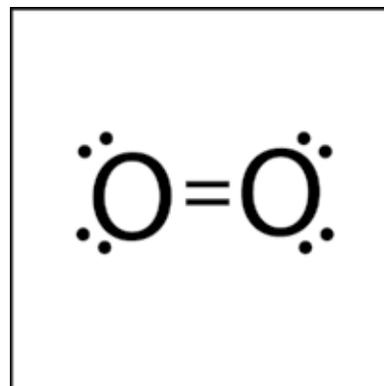
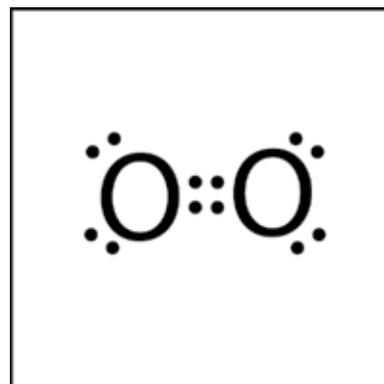
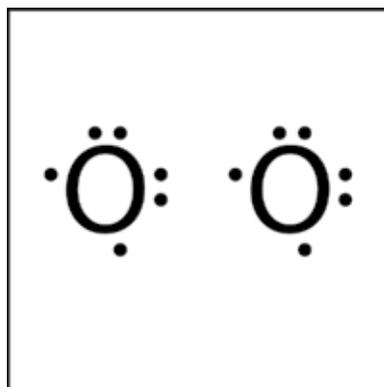
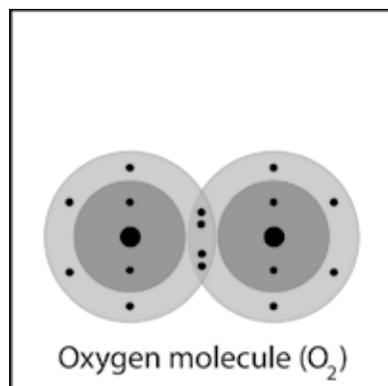
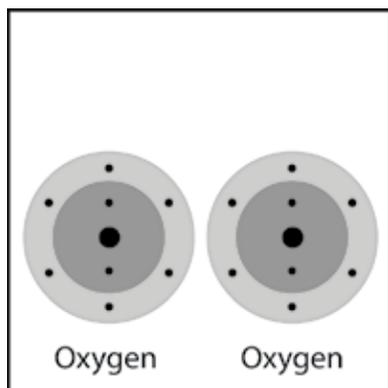
Covalent bonding in the water molecule, H₂O



5. Water has two hydrogen atoms covalently bonded to an oxygen atom. Methane has four hydrogen atoms covalently bonded to a carbon atom.

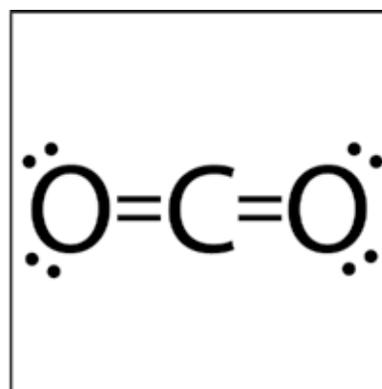
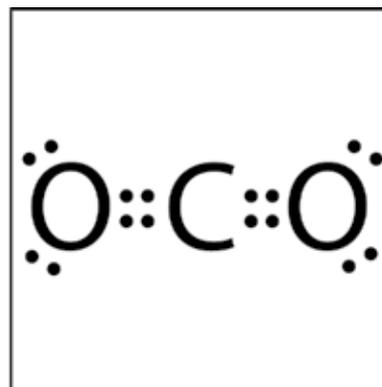
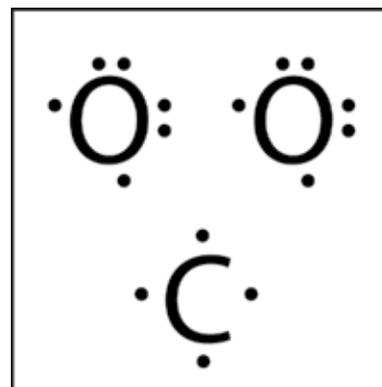
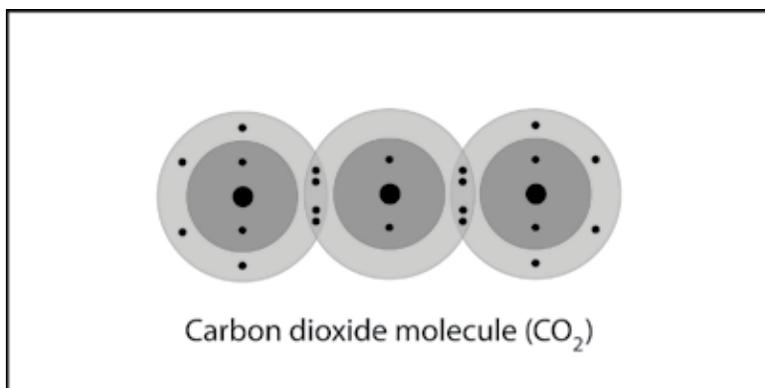
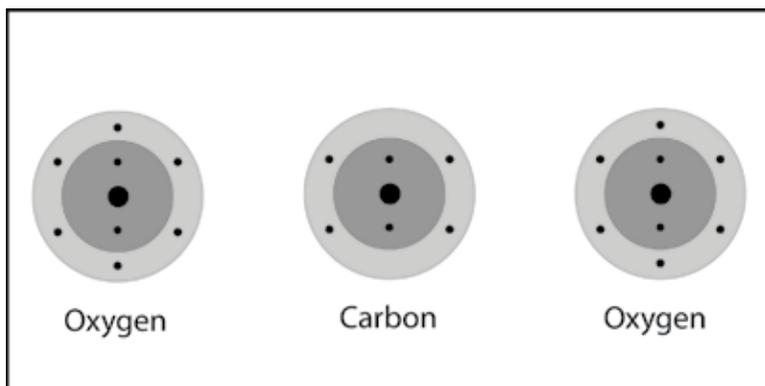
Draw Lewis diagrams for methane using dots for pairs of electrons and then lines for pairs of electrons.

Covalent bonding in the oxygen molecule, O₂



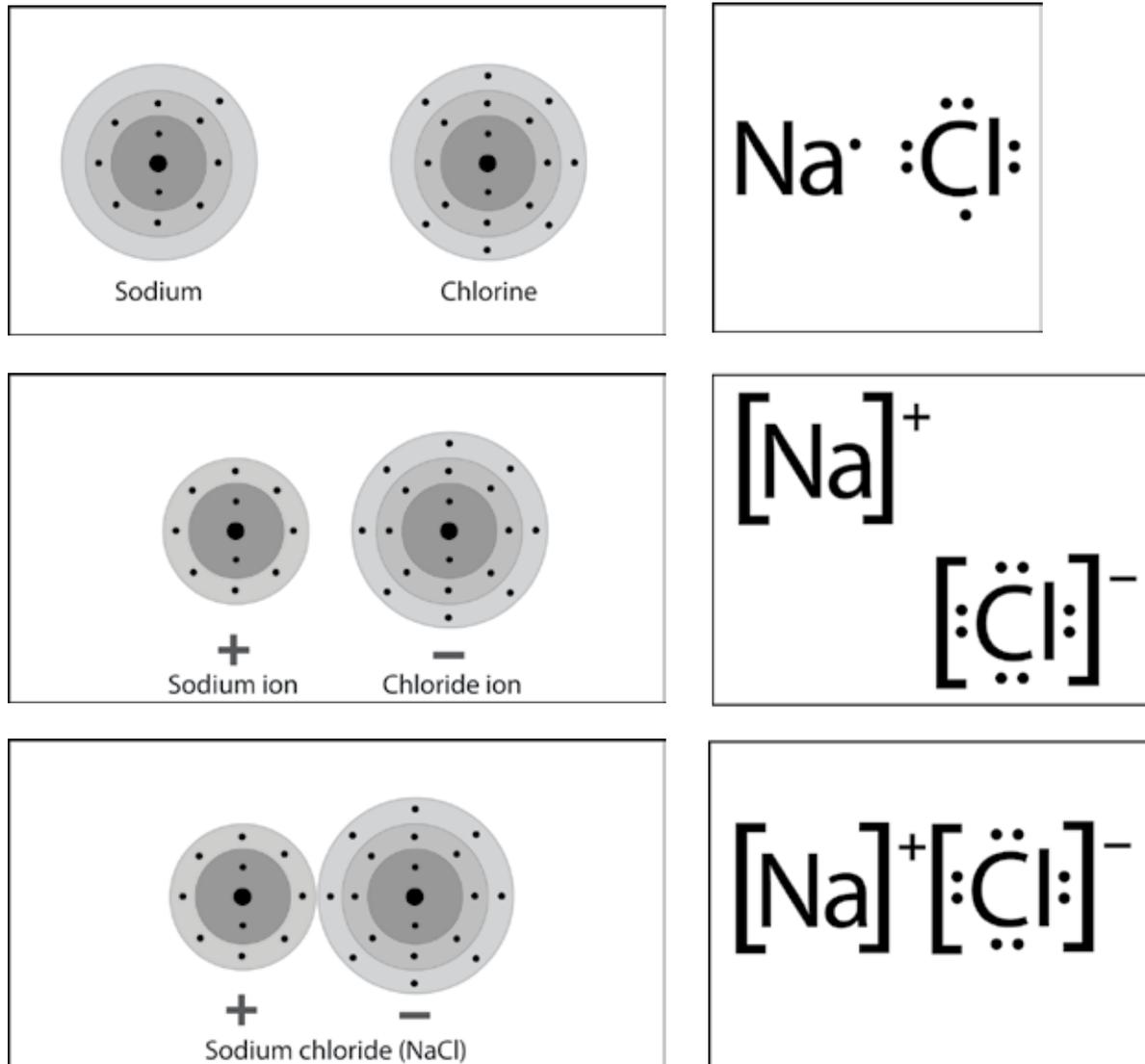
6. Why are there four dots between the two oxygen atoms?

Covalent bonding in the carbon dioxide molecule, CO_2



7. Why are there two sets of lines between the carbon and each oxygen atom?

Ionic bonding of sodium chloride, NaCl



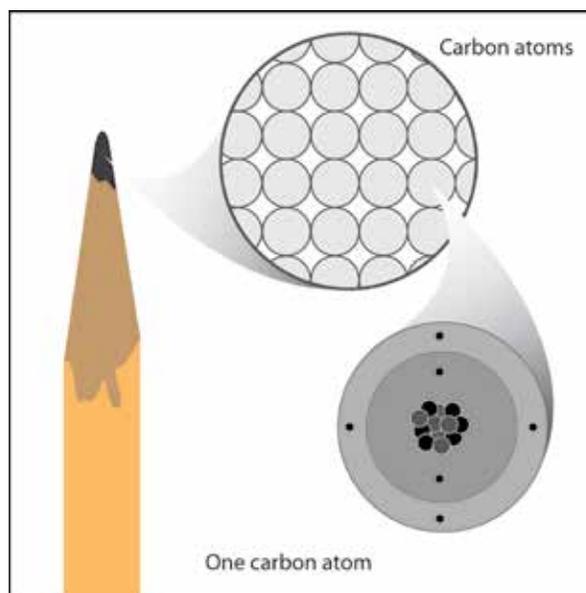
8. In the second dot diagram, why are there no electrons surrounding sodium?
9. In the final dot diagram of NaCl, the dots between the sodium and chlorine are between the atoms. Are these atoms sharing the electrons?

Chapter 4—Student Reading

Parts of the atom

An atom is made up of *protons*, *neutrons*, and *electrons*. Look at the model of a carbon atom from the graphite in the point of a pencil. Protons and neutrons are in the center or *nucleus* of the atom. Electrons are in regions surrounding the nucleus. In the carbon atom, there are six protons, and six electrons. The vast majority of carbon atoms also have six neutrons.

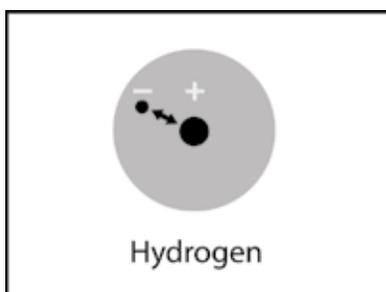
A proton has a positive charge. An electron has a negative charge. A neutron has no charge. The charge on the proton and electron are exactly the same size but opposite. The same number of protons and electrons exactly cancel each other in a neutral atom.



Two protons push each other away or *repel*. Two electrons also repel each other. But a proton and an electron move toward or *attract* each other. Another way of saying this is that the same or “like” charges repel one another and opposite charges attract one another. Since opposite charges attract each other, the negatively charged electrons in an atom are attracted to the positively charged protons. This attraction is what holds an atom together.



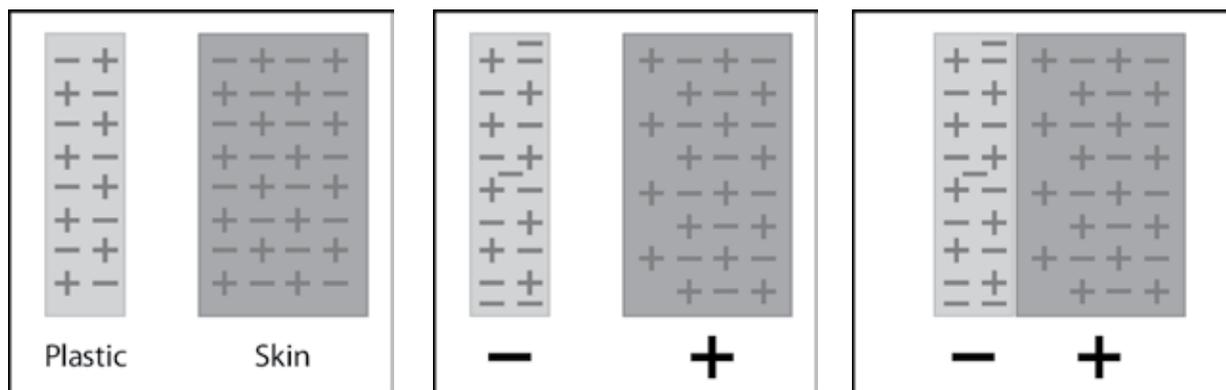
This is a simple model of a hydrogen atom which has one proton and one electron. The arrow shows that the electron is attracted to the proton.



Another model of the hydrogen atom shows a cloudy-looking region in the space surrounding the nucleus. This model represents the electron as a cloud to show that it is not possible to know the exact location of an electron. The electron cloud shows the region surrounding the nucleus where the electron is most likely to be.

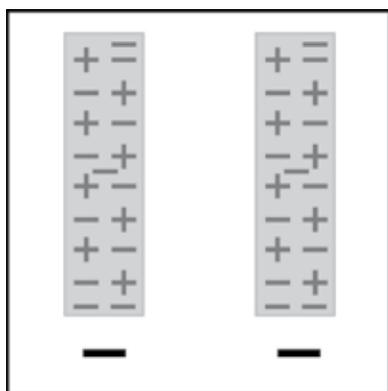
Proton, electrons, and static electricity

You can see evidence of electrons and protons attracting or repelling each other when you make static electricity. For example, when you rub a plastic strip between your fingers, electrons move from your skin to the plastic. If you assume that the plastic and the skin were both neutral before rubbing, the plastic now has more electrons or negative charges than positive.



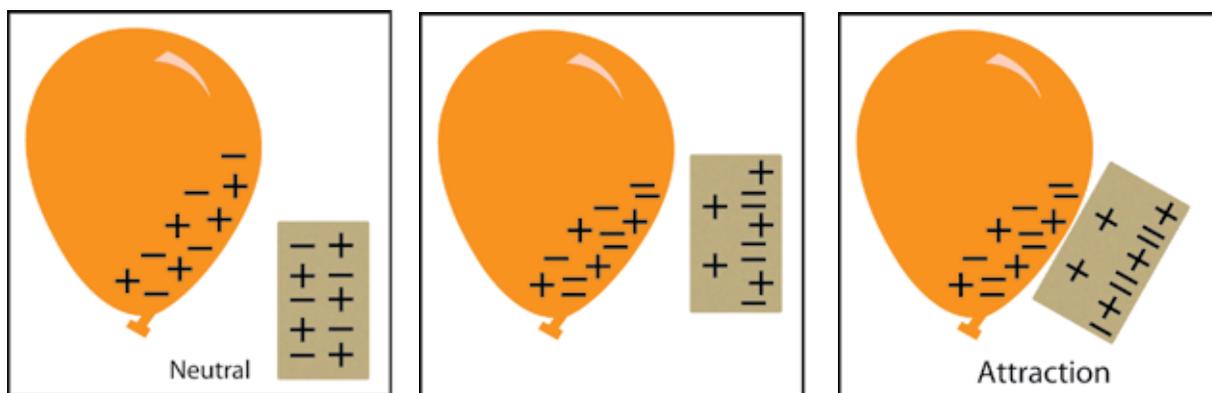
This gives the plastic an overall or *net* negative charge. Since your skin lost some negative charge, it now has more positive charge than negative, so your skin has an overall or net positive charge. When you bring the plastic near your fingers, the plastic is attracted because opposite charges attract.

If you rub *two* plastic strips on your fingers, each strip gains electrons so each one has a net negative charge. If you bring the strips near each other, they repel because like charges repel.



Rubbing a balloon and sticking it to a wall or using it to attract little pieces of paper is also evidence that protons and electrons have opposite charge. When you rub a balloon on your hair or clothes, electrons move onto the balloon. This gives the balloon a negative charge.

When the balloon is brought near a little piece of paper, electrons on the balloon repel electrons in the paper. The electrons in the paper move away from the balloon and leave an area of positive charge near the balloon. The positively charged area of the paper is attracted to the negative balloon and the paper moves to the balloon.



The Periodic Table of the Elements

You have read about protons and electrons, and about the atoms and molecules in different substances. The atoms that make up all solids, liquids, and gases are organized into a chart or table called the periodic table of the elements. The periodic table shows all the atoms that everything in the known universe is made from. Each box contains information about a different atom. It's like the alphabet in which only 26 letters, in different combinations, make up thousands of words. The 100 or so atoms of the periodic table, in different combinations, make up millions of different substances.

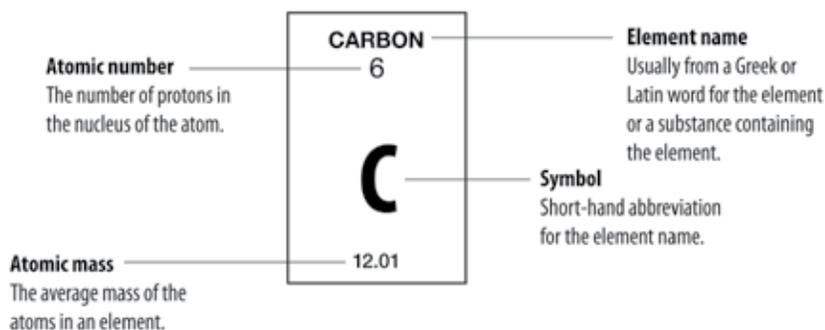
The Periodic Table of the Elements

3
Li
Lithium
6.94

— Atomic Number
— Element Symbol
— Element Name
— Average Atomic Mass

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 18.99	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.39	31 Ga Gallium 69.72	32 Ge Germanium 72.61	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium 98	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.60	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 145	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.97	
87 Fr Francium 223	88 Ra Radium 226	89 Ac Actinium 227	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237	94 Pu Plutonium 244	95 Am Americium 243	96 Cm Curium 247	97 Bk Berkelium 247	98 Cf Californium 251	99 Es Einsteinium 252	100 Fm Fermium 257	101 Md Mendelevium 258	102 No Nobelium 259	103 Lr Lawrencium 262	

Each box in the periodic table contains basic information about an element.



The meaning of the terms “atom” and “element” can be confusing because they are often used as if they are the same thing. They are related to one another but they are not the same. An atom is the smallest particle or “building block” of a substance. An element is a substance made up of all the same type of atom. For instance, a piece of pure carbon is made up of only carbon atoms. The piece of pure carbon is a sample of the element carbon. The people who developed the periodic table could have called it the Periodic Table of the Atoms but they did not have a firm understanding of atoms at that time. Since they were working with actual samples of elements such as copper, mercury, sulfur, etc., they called it the periodic table of the elements.

Atomic mass

The element name, atomic number, and symbol are pretty easy to understand. The atomic mass is a little trickier. The atomic mass of an element is based on the mass of the atoms that make up the element. The mass of the atoms is based on the protons, neutrons, and electrons of the atoms. The mass of the proton and neutron are about the same but the mass of the electron is much smaller (about 1/2000 the mass of the proton or neutron). The vast majority of the atomic mass is contributed by the protons and neutrons.

For any element in the periodic table, the number of electrons in an atom always equals the number of protons in the nucleus. But this is not true for neutrons. Atoms of the same element can have different numbers of neutrons than protons. Atoms of the same element with different numbers of neutrons are called isotopes of that element. The atomic mass given in the periodic table is an *average* of the atomic mass of the isotopes of an element.

For example, the vast majority of carbon atoms have 6 protons and 6 neutrons, but a small percentage of carbon atoms have 6 protons and 7 neutrons, and an even smaller percentage have 6 protons and 8 neutrons. Since the vast majority of carbon atoms have a mass very close to 12, and only a small percentage are greater than 12, the average atomic mass is slightly greater than 12 (12.01). For the atoms of the first 20 elements, the number of neutrons is either equal to or slightly greater than the number of protons.

Hydrogen is an exception to this rule. All hydrogen atoms have one proton but the vast majority have 0 neutrons. There is a small percentage of hydrogen atoms that have 1 neutron and a smaller percentage that have 2 neutrons. When you take the average mass of all the different isotopes of hydrogen, the mass is slightly greater than one (about 1.01).

Electrons are in energy levels surrounding the nucleus

Electrons surround the nucleus of an atom in three dimensions making atoms spherical. Electrons are in different regions around the nucleus like concentric spheres. These regions are called energy levels. Since it is very difficult to draw concentric spheres, the energy levels are usually shown in 2 dimensions.



This energy level model represents an oxygen atom. The nucleus is represented by a dot in the center which contains both protons and neutrons. The smaller dots surrounding the nucleus represent electrons in the energy levels. You can tell that this model is oxygen because there are a total of 8 electrons. Since neutral atoms in the periodic table have the same number of electrons as protons, this atom must have 8 protons. The number of protons is the same as the atomic number, so this atom's atomic number is 8, which is oxygen.

Arrangement of elements in the periodic table

There is a limit to the number of electrons that can go into the different energy levels of an atom. A certain number of electrons go into an energy level before they begin to go into the next level. After the first energy level contains 2 electrons (helium), the next electrons go into the second energy level. After the second energy level has 8 electrons (neon), the next electrons go into the third energy level.

After the third energy level has 8 electrons (argon), the next 2 electrons go into the fourth energy level. An energy level model is shown in the chart below for the first twenty elements in the periodic table.

The rows going across the periodic table are called *periods*. The columns going up and down are called *groups* or *families*.

ENERGY LEVELS ELEMENTS 1-20							
HYDROGEN 1 1.01							HELIUM 2 4.00
LITHIUM 3 6.94	BERYLLIUM 4 9.01	BORON 5 10.81	CARBON 6 12.01	NITROGEN 7 14.01	OXYGEN 8 16.00	FLUORINE 9 19.00	NEON 10 20.18
SODIUM 11 22.99	MAGNESIUM 12 24.31	ALUMINUM 13 26.98	SILICON 14 28.09	PHOSPHORUS 15 30.97	SULFUR 16 32.07	CHLORINE 17 35.45	ARGON 18 39.95
POTASSIUM 19 39.10	CALCIUM 20 40.08						

Number of energy levels in each period

- The atoms in the first period have electrons in 1 energy level.
- The atoms in the second period have electrons in 2 energy levels.
- The atoms in the third period have electrons in 3 energy levels.
- The atoms in the fourth period have electrons in 4 energy levels.

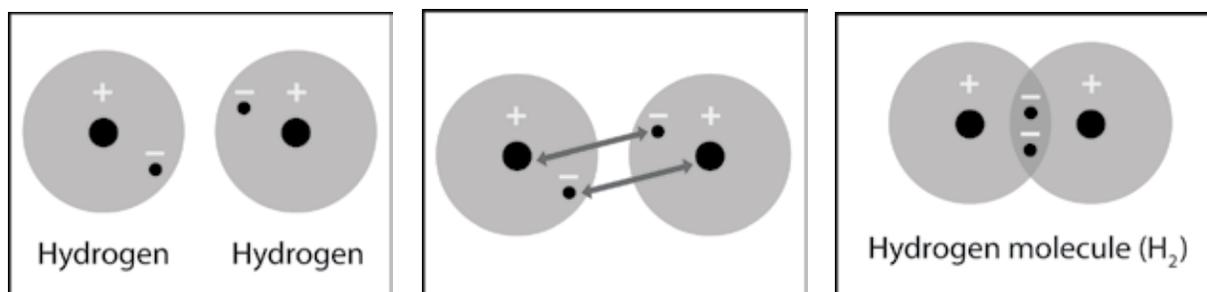
Atoms in a group have the same number of valence electrons

If you look at the atoms in a group, you will see that they each have the same number of electrons in their outermost energy level. Electrons in this level are called valence electrons. For instance, hydrogen, lithium, sodium, and potassium all have 1 valence electron. Valence electrons are important because they interact with other atoms and are responsible for many of the characteristic properties of the atom.

HOW ATOMS BOND TO EACH OTHER

Covalent bonding

Remember that a hydrogen atom has 1 proton and 1 electron and that the electron and the proton are attracted to each other. But if the atoms get close enough to each other, the electron from each hydrogen atom feels the attraction from the proton of the other hydrogen atom (shown by the double headed arrow).



The attractions are not strong enough to pull the electron completely away from its own proton. But the attractions are strong enough to pull the two atoms close enough together so that the electrons feel the attraction from both protons. When the electrons are attracted to and shared by both atoms, the individual hydrogen atoms have bonded to become the *molecule* H₂. This type of bond is called a *covalent* bond. In a covalent bond, electrons from each atom are attracted or “shared” by *both* atoms. Two or more atoms covalently bonded are called a *molecule*.

There are two main requirements for atoms to form a covalent bond and make a molecule:

- There needs to be a strong enough attraction between the electrons in each atom for the protons in the other atom.
- There needs to be room in the outer energy level of both atoms.

Once bonded, the hydrogen molecule is more stable than the individual hydrogen atoms. By being part of a covalent bond, the electron from each hydrogen atom gets to be near two protons instead of only the one proton it started with. Since the electrons are closer to more protons, the molecule of two bonded hydrogen atoms is more stable than the two individual unbonded hydrogen atoms.

Atoms bond until their outer energy levels are full

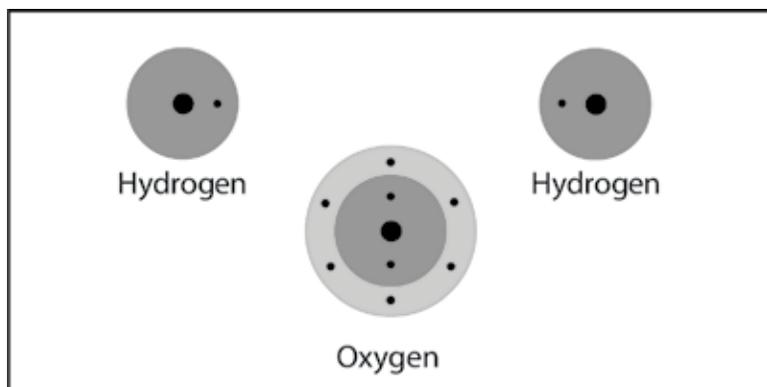
The two electrons in the hydrogen molecule (H_2) can be thought of as “belonging” to each atom. This means that each hydrogen atom now has two electrons in its first energy level. The first energy level is the outer energy level for hydrogen and can only accommodate or “hold” two electrons. This means that the outer energy level is full. Atoms will covalently bond to one another until each atom’s outer energy level is full.

Once the outer energy levels are full, additional atoms will not covalently bond to the atoms in the H_2 molecule. This will not happen for two main reasons:

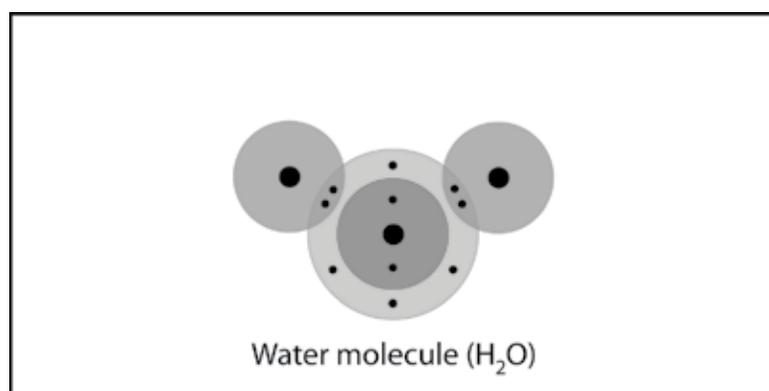
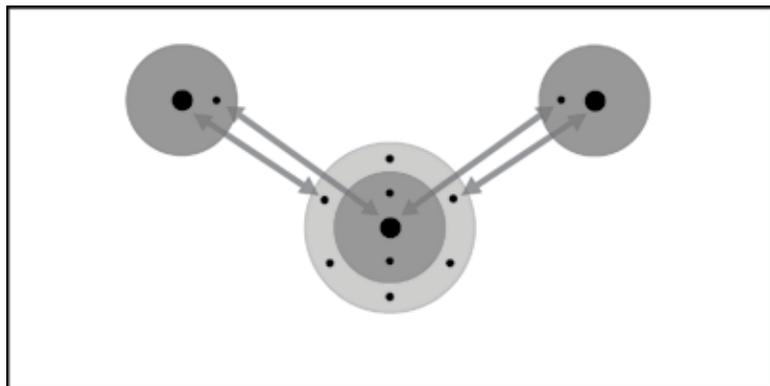
- An electron from a new atom would have to join an atom in the H_2 molecule on the next energy level, further from the nucleus where it would not feel a strong enough attraction.
- An electron from a hydrogen atom already in the H_2 molecule and close to the nucleus would need to move further away to share with the new atom.

Both of these possibilities would make the molecule less stable and would not happen.

Covalent bonding also happens in a water molecule. When hydrogen atoms and an oxygen atom get close enough together, the electrons from the atoms feel the attraction from the other atom’s protons.

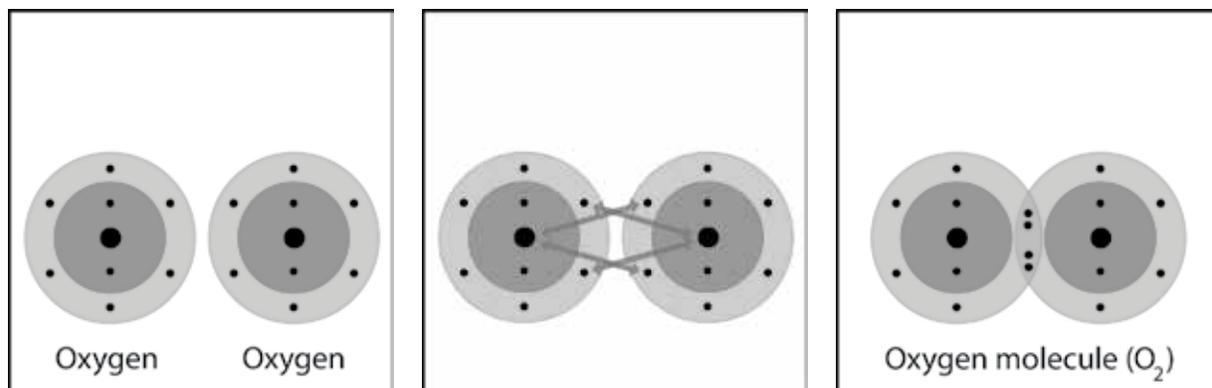


Because there is both a strong enough attraction between the atoms and room for electrons in their outer energy levels, they share electrons. This forms a covalent bond.

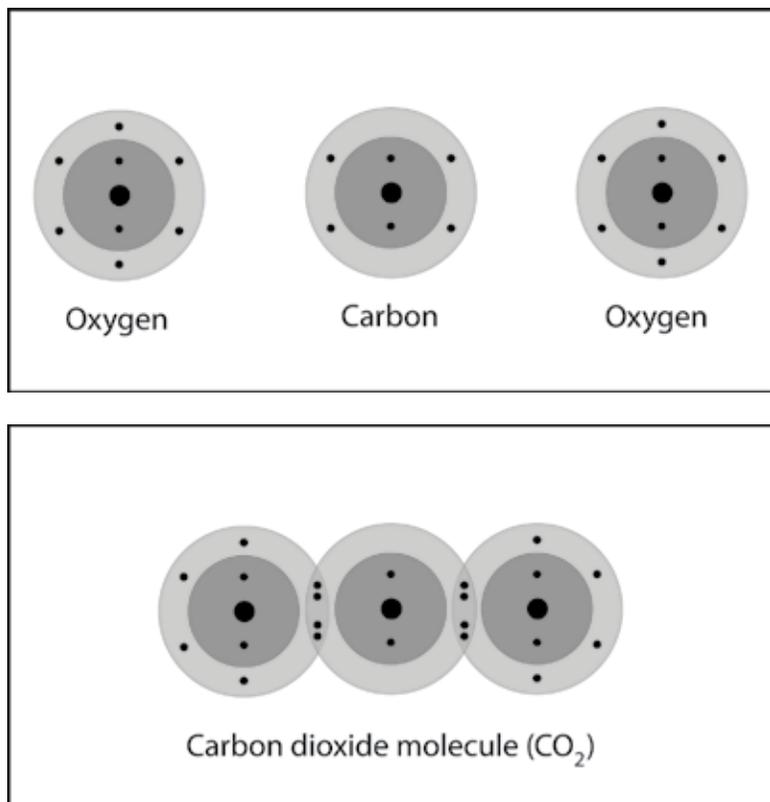


Two oxygen atoms form a double-bond

Oxygen molecules that are present in our air are made up of two oxygen atoms bonded together. Each oxygen atom has 6 valence electrons. When oxygen atoms get close together, the attractions from the nucleus of both atoms attract the outer electrons of the other atom. In this case, 2 electrons from each atom are shared. This is called a double bond.



A carbon atom and two oxygen atoms bond to make carbon dioxide (CO_2)



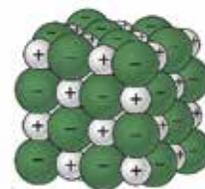
Ionic bonding

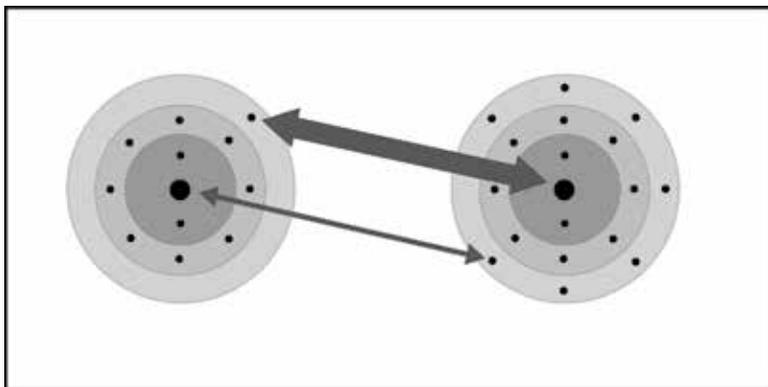
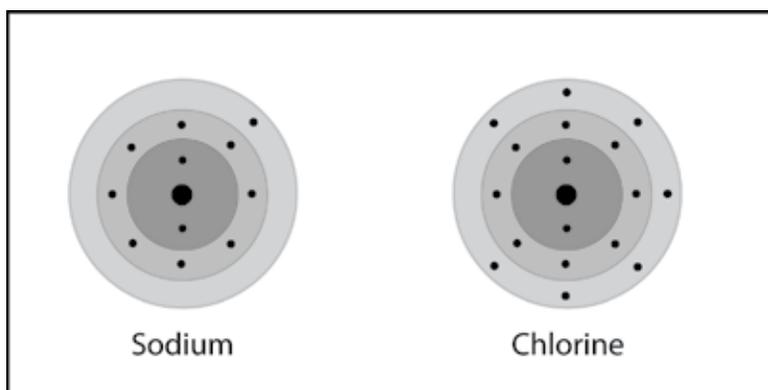
There is another type of bond called an *ionic bond*. One of the most common substances formed by ionic bonding is salt or sodium chloride (NaCl). Look at the model of sodium chloride. The spheres with the “+” and “-” signs on them are called *ions*.

The larger green ones are chloride ions and the smaller gray ones are sodium ions. These ions are formed from chlorine and sodium atoms.

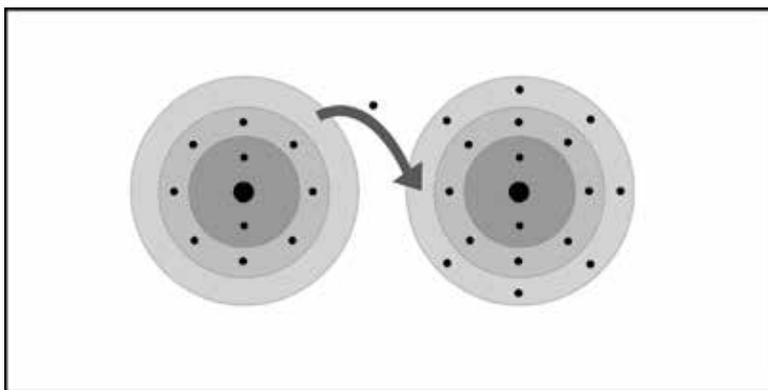
When a sodium and chlorine atom get close enough together, the electrons from the atoms feel the attraction of the protons in the nucleus of the *other* atom.

Chlorine has a stronger attraction for electrons than sodium (shown by the thicker arrow).



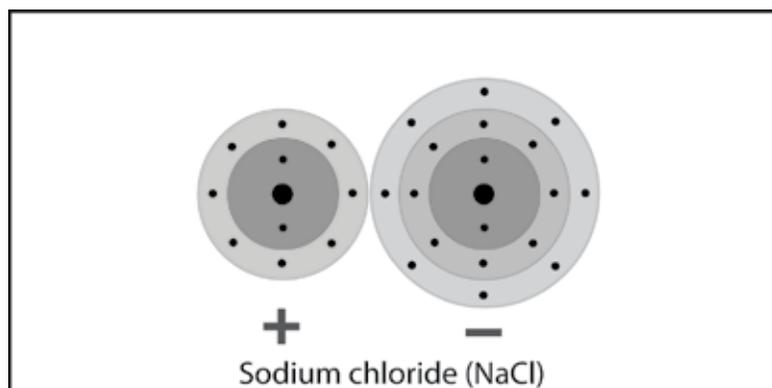


During the interactions between the atoms, the electron in sodium's outer energy level is transferred to the outer energy level of the chlorine atom.



Chlorine *gains* an electron so that the chloride ion has 18 electrons and 17 protons. Since the chloride ion has one more electron than proton, chloride is a *negative* ion with a charge of -1. Sodium loses an electron leaving it with only 10 electrons but 11 protons. This makes sodium a *positive* ion with a charge of +1.

Oppositely charged ions attract each other forming an ionic bond. The bonded ions are more stable than the individual atoms were.

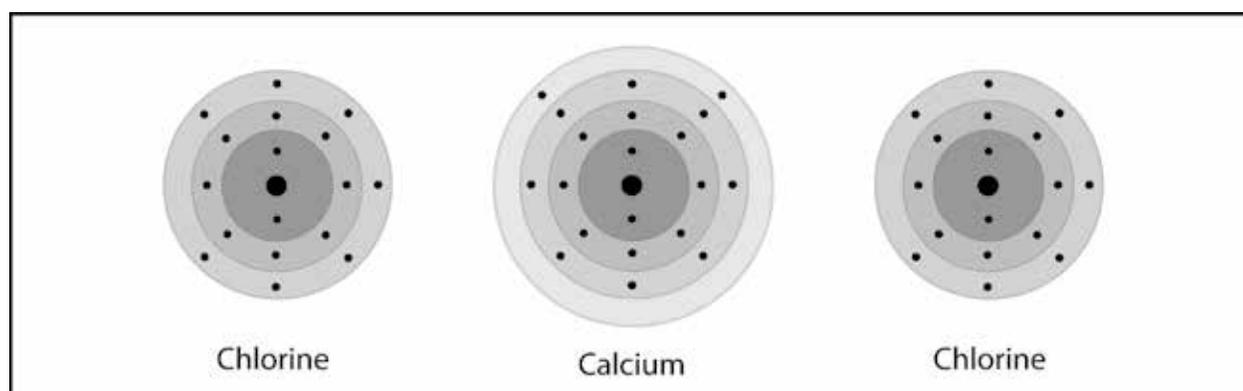


When ions form, atoms gain or lose electrons until their outer energy level is full.

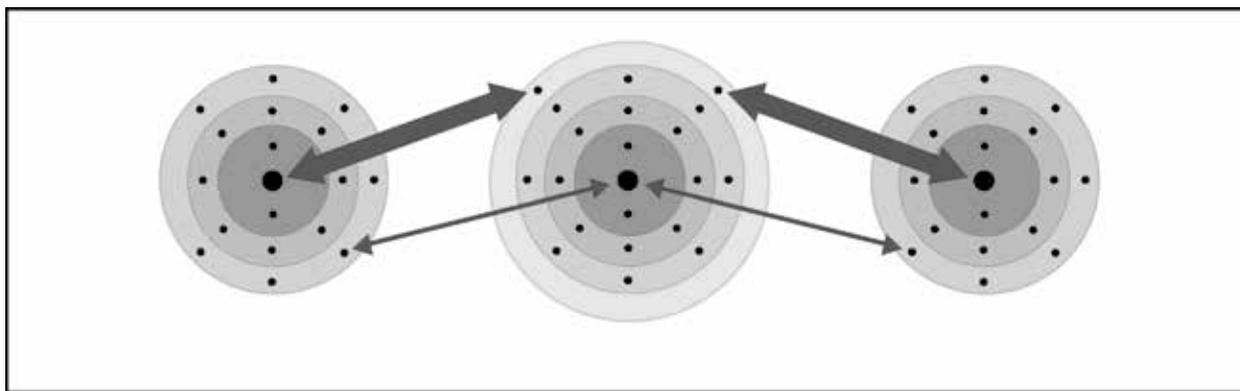
For example, when sodium loses its one outer electron from the third energy level, the second level becomes the new outer energy level and is full. Since these electrons are closer to the nucleus, they are more tightly held and will not leave.

When chlorine gains an electron its third energy level becomes full. An additional electron cannot join because it would need to come in at the fourth energy level. This far from the nucleus, the electron would not feel enough attraction from the protons to be stable.

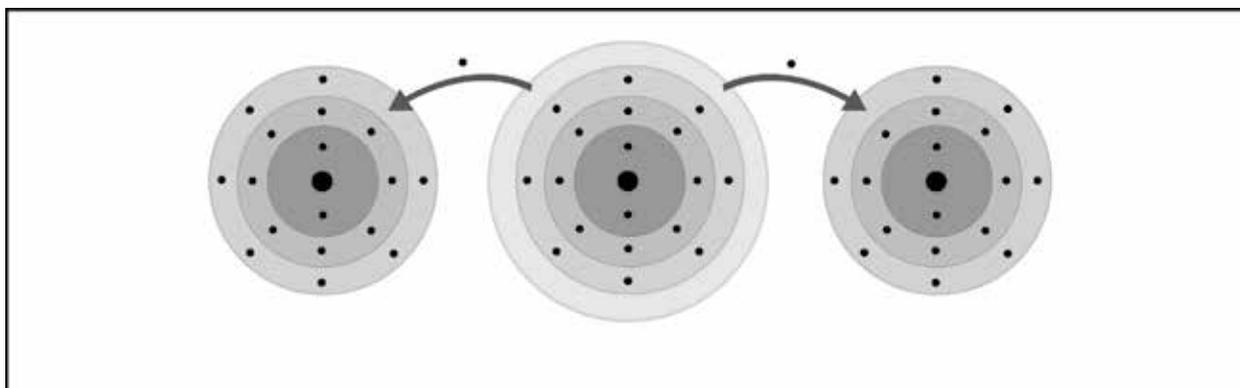
Ionic bonding in calcium chloride (CaCl_2)



The protons of the calcium atom attract the electrons from the chlorine atom. The protons of the two chlorine atoms attract the electrons from the calcium atom more strongly as shown by the thicker arrows.

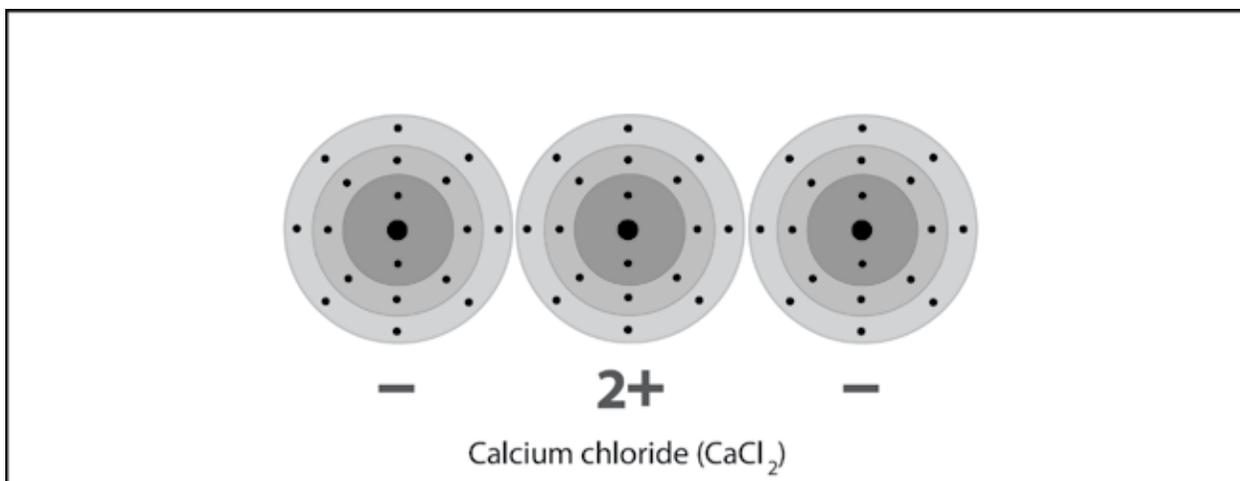


During the interactions between the atoms, the two electrons in calcium's outer energy level are transferred to the outer energy level of each of the chlorine atoms.



Each chlorine atom *gains* an electron so that the chloride ion has 18 electrons and 17 protons. This makes each chloride a *negative* ion with a charge of -1 . Calcium *loses* two electrons leaving it with only 18 electrons and 20 protons. This makes calcium a *positive* ion with a charge of $+2$.

Oppositely charged ions attract each other forming an ionic bond. The bonded ions are more stable than the individual atoms were.



Chapter 5, Lesson 1: Water is a Polar Molecule

Key Concepts

- The water molecule, as a whole, has 10 protons and 10 electrons, so it is neutral.
- In a water molecule, the oxygen atom and hydrogen atoms share electrons in covalent bonds, but the sharing is not equal.
- In the covalent bond between oxygen and hydrogen, the oxygen atom attracts electrons a bit more strongly than the hydrogen atoms.
- The unequal sharing of electrons gives the water molecule a slight negative charge near its oxygen atom and a slight positive charge near its hydrogen atoms.
- When a neutral molecule has a positive area at one end and a negative area at the other, it is a *polar* molecule.
- Water molecules attract one another based on the attraction between the positive end of one water molecule and the negative end of another.

Summary

Students will be introduced to the idea that water has a slight positive charge at one end of the molecule and a slight negative charge at the other (a polar molecule). Students view animations, make illustrations, and use their own water molecule models to develop an understanding of how the polar nature of water molecules can help explain some important characteristics of water.

Objective

Students will be able to explain, on the molecular level, what makes water a polar molecule. Students will also be able to show in a drawing that the polar nature of water can explain some of water's interesting characteristics and help explain its evaporation rate compared to a less polar liquid.

Evaluation

The activity sheet will serve as the "Evaluate" component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Use in well-ventilated room. Dispose of small amounts down the drain or according to local regulations. Have students wash hands after the activity.

Materials for Each Group

- Styrofoam water molecule models from Chapter 2, Lesson 2 (two per student)
- Permanent markers (blue and red)
- Isopropyl alcohol (70% or higher)
- Water
- Brown paper towel
- Droppers

Note about the Materials

Students made molecular models of the water molecule using Styrofoam balls and toothpicks in Chapter 2, Lesson 2. Give each student two of these water molecule models for this activity.

ENGAGE

1. Show students examples of water molecules' attraction for one another.

Remind students that in Chapters 1 and 2, they investigated the behavior of water at different temperatures and explored the state changes of water. Many of the explanations were based on the idea that water molecules are attracted to one another. Remind students that in Chapter 4 they looked at the covalent bonding between oxygen and hydrogen, which creates the water molecule. Now students will look more closely at the details of the covalent bonds in a water molecule to understand why water molecules are attracted to one another.

Project the video *Water Balloon*.

www.middleschoolchemistry.com/multimedia/chapter1/lesson1#water_balloon

This video was shown in Chapter 1, Lesson 1 to show that water molecules are attracted to one another.

Project the video *Water Fountain*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson1#water_fountain

Point out that the water is able to stay together in these arcs because water molecules are very attracted to each other.

EXPLAIN

2. Show molecular model animations that illustrate why water molecules are attracted to each other.



Project the animation *Polar Water Molecule*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson1#polar_water_molecule

First Frame of the Animation

- **Electrons are shared between atoms in a covalent bond.**
Remind students how the shared electrons in a water molecule are attracted to the protons in both the oxygen and the hydrogen atoms. These attractions hold the atoms together.
- **Water molecules are neutral.**
Be sure students realize that no protons or electrons are gained or lost. The water molecule has a total of 10 protons and 10 electrons (8 from the oxygen atom and 1 from each of the two hydrogen atoms). Since it has the same number of protons and electrons, the water molecule is neutral.

Click “Play”

- **The electron cloud model shows where electrons are in a molecule.**
Tell students that another way to see the difference in where the electrons are is by using the electron cloud model. Remind students that it’s impossible to know the exact location of an electron, so sometimes the regions occupied by electrons are shown as “clouds” around the nucleus in an atom or molecule.
- **Unequal sharing of electrons makes water a polar molecule.**
Tell students that the oxygen atom attracts electrons a little more strongly than hydrogen does. So even though the electrons from each atom are attracted by both the oxygen and the hydrogen, the electrons are a bit more attracted to the oxygen. This means that electrons spend a bit more time at the oxygen end of the molecule. This makes the oxygen end of the molecule slightly negative. Since the electrons are not near the hydrogen end as much, that end is slightly positive. When a covalently bonded molecule has more electrons in one area than another, it is called a *polar* molecule.
- **The electron cloud model can show an unequal sharing of electrons.**
Point out that the electron cloud around the oxygen is darker than the electron cloud around the hydrogen. This shows that electrons are more attracted to the oxygen end of the molecule than the hydrogen end, making the water molecule polar.

Click “Next”

- **Color can be added to an electron cloud model to show where electrons are more or less likely to be.**

Tell students that this is another model of a water molecule. In this model, color is used to show the polar areas of the water molecule. The negative area near the oxygen atom is red, and the positive area near the hydrogen atoms is blue.

Project the animation *Attraction between water molecules*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson1#attraction

Ask students:

- **What do you notice about the way water molecules orient themselves?**

The red (oxygen) area of one water molecule is near the blue (hydrogen) end of another water molecule.

- **Why do water molecules attract one another like this?**

Since the oxygen end of a water molecule is slightly negative and the hydrogen end is slightly positive, it makes sense that water molecules attract one another.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



3. Show students that the bonds between atoms in a molecule are different from the polar attractions between molecules.

Project the image *Attractions on different levels*

www.middleschoolchemistry.com/chapter5/lesson1#attractions

Students may be confused about the bonds within a water molecule and the attractions between water molecules.

The bonds *within* molecules and the polar attractions *between* molecules

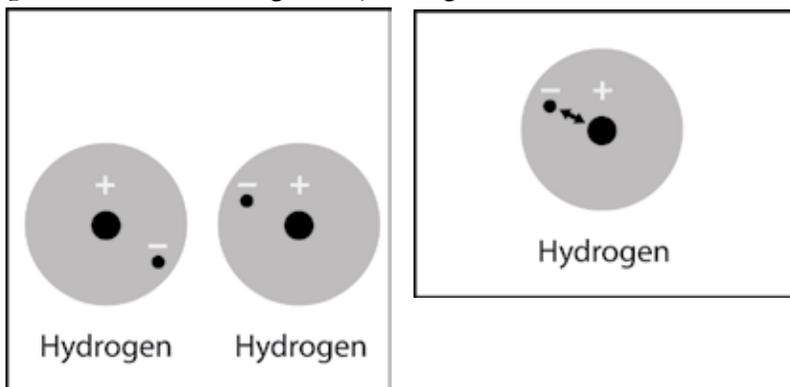
Explain to students that the interaction between the oxygen of one water molecule and the hydrogen of another is different than the sharing of electrons between the oxygen and the hydrogens within the water molecule itself.

It’s all about attractions between positive and negative.

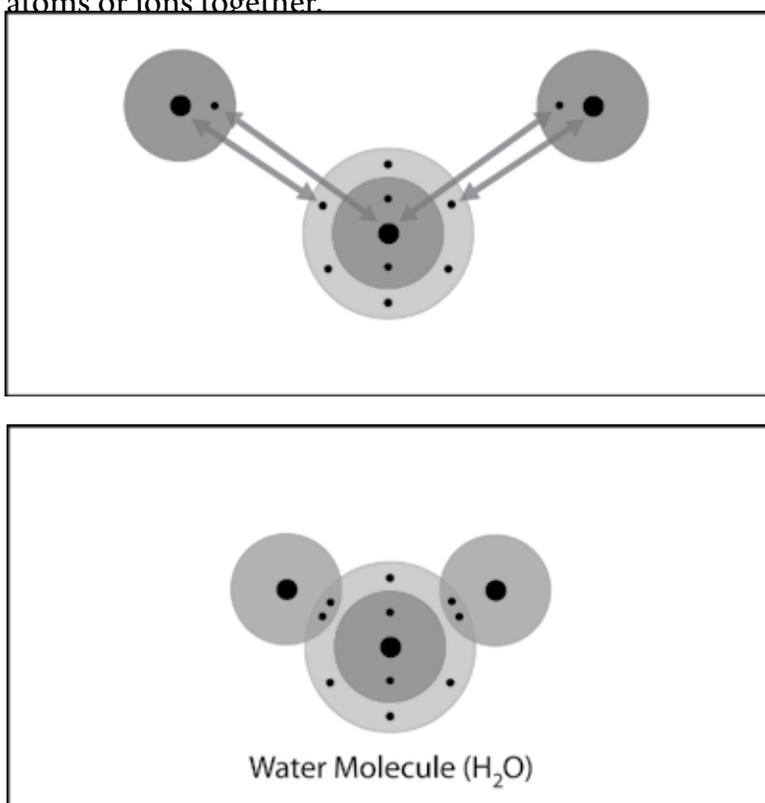
Point out to students that attractions between positive and negative works on three differ-

ent levels.

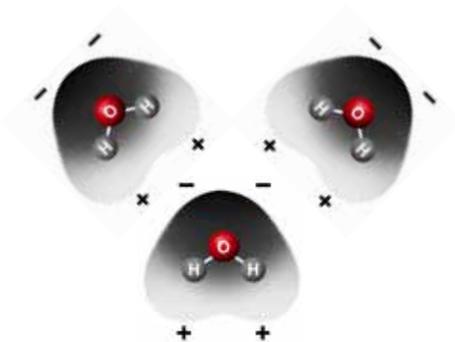
1. A single *atom* stays together because of the attraction between the positively charged protons and the negatively charged electrons.



2. In a molecule, *two or more atoms* stay together because of the mutual attraction between the positively charged protons from one atom and the negatively charged electrons from the other atom. This causes the covalent or ionic bonding that holds atoms or ions together.



3. *Two or more water molecules stay together because of the positive and negative parts of the molecules attracting each other.*



4. **Have students mark the positive and negative areas on a water molecule by color-coding their Styrofoam ball models.**

Materials for Each Group

- Styrofoam water molecule models from Chapter 2, Lesson 2 (two per student)
- Permanent markers (blue and red)

Procedure

1. Draw a blue “+” on each of the hydrogen atoms.
2. Draw two red “-” at the bottom of the oxygen atom.
3. Repeat this for your other water molecule.
4. Position your water molecules so that opposite charges are near each other.

Ask students:

- **How do your Styrofoam ball models of water molecules relate to the color-coded charge density model shown in the animation?**
The different colors show that water is a polar molecule.
- **What do the red “-” signs on the oxygen atom represent?**
The red “-” signs represent the area where there are more electrons.
- **What do the blue “+” signs on the hydrogen atoms represent?**
The blue “+” signs represent the area where there are fewer electrons.
- **Because water molecules are polar, how do they arrange themselves in liquid water?**
The positive area of one water molecule is attracted to the negative area of another water molecule.

EXPLORE

5. Have students design a test to compare the rate of evaporation between water and alcohol.

Remind students that water molecules are very polar. The strong attractions between water molecules affect water's surface tension, boiling point, and rate of evaporation. Tell students that they will do an experiment to compare the evaporation rates of water and another liquid that isn't as polar.

Ask students:

- **Do you think a substance like water with polar molecules would evaporate faster or slower than a substance like alcohol with molecules that are not as polar?**
The more-polar molecules will stick together more and will probably evaporate more slowly than less polar molecules. Less-polar molecules should evaporate faster because they are not as attracted to each other.
- **How could you design a quick and easy evaporation test to compare the rate of evaporation between water and alcohol?**
 - What materials will you need?
 - Should you use the same amount of water and alcohol?
 - How will you know if one evaporates faster than the other?
 - Is there a way to do it so that it will not take a lot of time?

Students should say that they will need the same small amount of water and alcohol. These liquids should be placed at the same time on a surface like a brown paper towel so that students can tell when each liquid evaporates.

6. Have students follow the procedure below to compare the rate of evaporation between water and alcohol.



Question to Investigate

Does water evaporate faster or slower than less-polar alcohol?

Materials for Each Group

- Isopropyl alcohol (70% or higher)
- Water
- Brown paper towel
- Droppers



Procedure

1. At the same time, place 1 drop of water and 1 drop of alcohol on a brown paper towel. Observe.

Expected Results

The dark spot on the paper towel made by the alcohol will turn lighter faster than the dark spot made by the water. This indicates that the alcohol evaporates more quickly than the water.

Note: This test is fine for middle school students but there is something about the test that does not make it completely fair. There are many more water molecules in a drop of water than alcohol molecules in a drop of alcohol. The test would be more fair if the same number of water and alcohol molecules are placed on the paper towel. This requires a way to “count” molecules. Determining the number of particles in a sample is a basic concept in chemistry, but is beyond the scope of a middle school chemistry unit. Even if the same number of water and alcohol molecules were used in this activity, the alcohol would evaporate faster.

Read more about counting molecules in the additional teacher background section at the end of this lesson.

EXPLAIN

7. Discuss student observations and describe the differences in polarity between water and alcohol molecules.

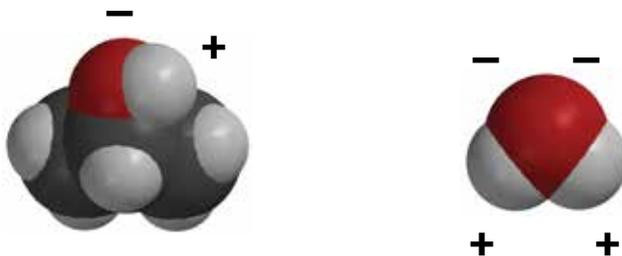
Ask students:

- Which evaporated faster, water or alcohol?
The alcohol evaporated faster.

Project the image *Water and Alcohol Molecules*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson1#water_and_alcohol

Tell students that understanding about polarity can help explain why water evaporates more slowly than alcohol.



Remind students that the oxygen-hydrogen (O–H) bonds in water make it a polar molecule. This polarity makes water molecules attracted to each other.

Explain that the oxygen–hydrogen (O–H) bond in the alcohol molecule is also polar. But, the carbon–hydrogen (C–H) bonds in the rest of the alcohol molecule are nonpolar. In these bonds, the electrons are shared more-or-less evenly.

Because there are both polar and nonpolar areas on the alcohol molecule, they are somewhat less attracted to each other than water molecules are to each other. This makes it easier for alcohol molecules to come apart and move into the air as a gas. This is why alcohol evaporates faster than water.

EXTEND

8. Have students consider how polarity might affect the temperature at which water and alcohol boil.

You know that water and alcohol have different characteristics because of the molecules they are made of and how these molecules interact with each other.

Project the image *Water and Alcohol Boiling*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson1#water_and_alcohol_boil

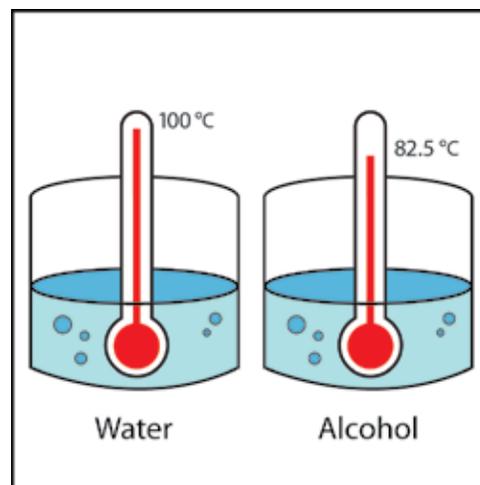
This illustration shows that alcohol boils at a lower temperature than water.

- Water boils at 100 °C
- Alcohol boils at 82.5 °C

Ask students:

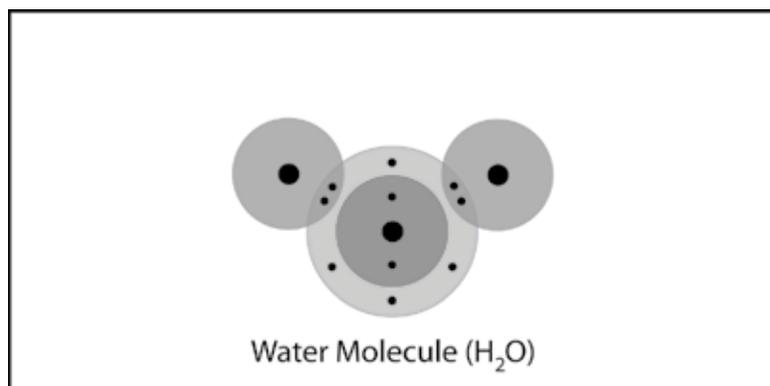
- **Knowing what you do about the polarity of water and alcohol, explain why alcohol boils at a lower temperature than water.**

The polar characteristic of water molecules causes them to attract each other well. The less polar alcohol molecules do not attract one another as strongly as water molecules do. It takes more energy to make water boil than it does to make alcohol boil. In other words, alcohol boils at a lower temperature than water.



EXPLAIN IT WITH ATOMS & MOLECULES

This model of a water molecule shows the number of electrons that can be found in each energy level. It also shows that oxygen and hydrogen share electrons in a covalent bond. But it doesn't show where electrons are most likely to be at any given moment.



1. Do the shared electrons in the water molecule spend more time near the oxygen atom or the hydrogen atoms?

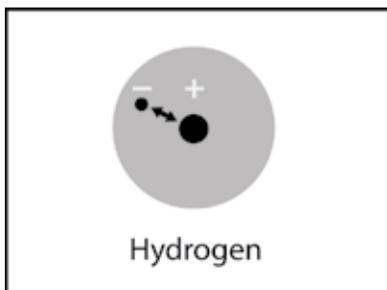
Why?

2. What do the colors and positive and negative signs on the electron cloud model represent?

3. Why are water molecules so attracted to each other?

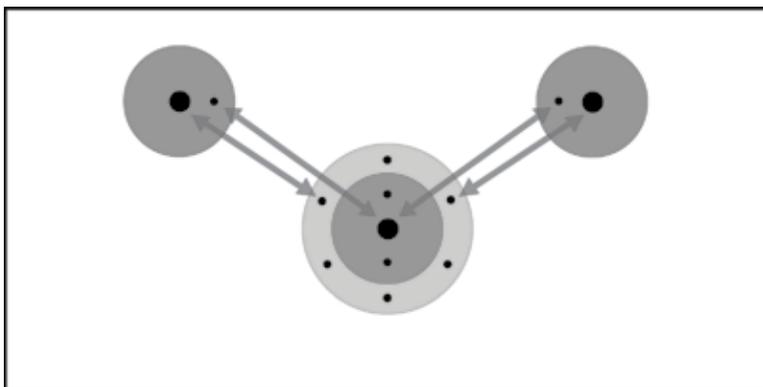
4. Attractions are important in three different ways. Draw a line between the picture and the description of the attractions.

Within an atom



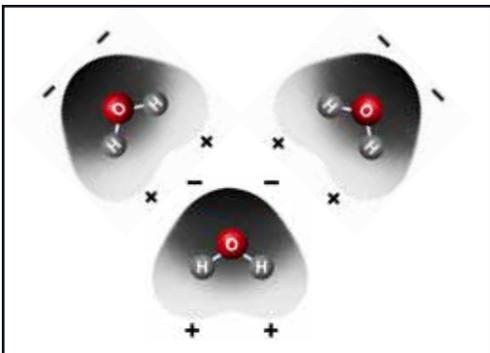
The electrons of each atom are attracted to the protons in the other atom. These mutual attractions keep two or more atoms together as a covalently bonded molecule.

Between the atoms in a molecule



The positive areas of one molecule are attracted to the negative areas of another molecule. These mutual attractions keep a substance together

Between molecules



The electrons are attracted to the protons in an atom. These attractions keep an atom together.

ACTIVITY

Mark the positive and negative areas on a water molecule.

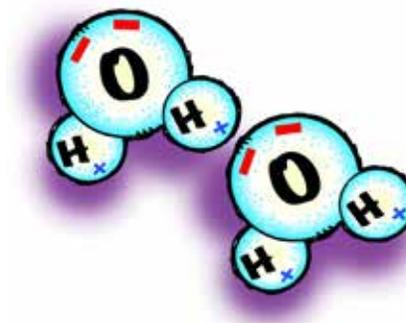


Materials for Each Group

- Styrofoam water molecule models from Chapter 2, Lesson 2 (two per student)
- Permanent markers (blue and red)

Procedure

1. Draw a blue “+” on each of the hydrogen atoms.
2. Draw two red “-” at the bottom of the oxygen atom.
3. Repeat this for your other water molecule.
4. Position your water molecules so that opposite charges are near each other.



5. What do the red “-” signs on the oxygen atom represent?
6. What do the blue “+” signs on the hydrogen atoms represent?
7. Because water molecules are polar, how do they arrange themselves in liquid water?
8. How would you design an experiment to find out which evaporates faster, alcohol or water? Be sure to explain how you would control variables.

ACTIVITY



Question to Investigate

Does water evaporate faster or slower than less-polar alcohol?

Materials for Each Group

- Isopropyl alcohol (70% or higher)
- Water
- Brown paper towel
- Droppers

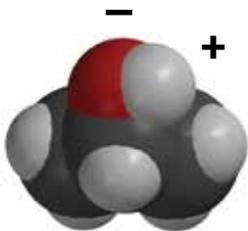
Procedure

1. At the same time, place 1 drop of water and 1 drop of alcohol on a brown paper towel. Observe.



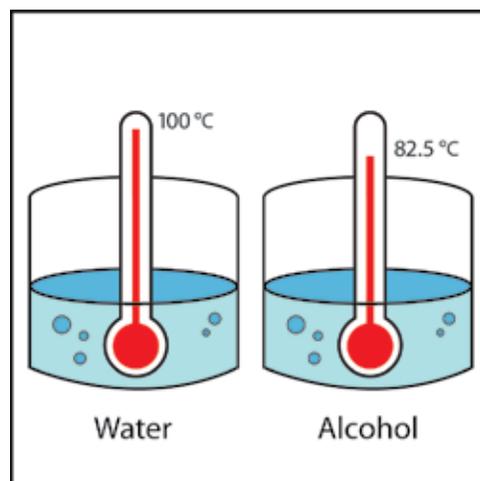
9. Which evaporated faster, water or alcohol?

10. The following molecular models show the polar regions of alcohol and water. Why does alcohol evaporate faster?



TAKE IT FURTHER

11. This illustration shows that alcohol boils at a lower temperature than water. Knowing what you do about the polarity of water and alcohol, explain why alcohol boils at a lower temperature than water.



Additional Teacher Background

Chapter 5, Lesson 1, p. 388

The idea of measuring an “equal amount” of two substances can have different meanings depending on how the amounts will be used. For example, let’s say you wanted an equal amount of carbon and calcium. If you measured out 5 grams of each, you would have the same amount of matter of each substance but you would not have the same number of atoms. Since carbon is lighter (lower atomic mass) than calcium, there would be more carbon atoms in 5 grams of carbon than calcium atoms in 5 grams of calcium.

Counting atoms

But what if you needed to measure the same number of carbon and calcium atoms? There is a way of doing this but you need to know the atomic mass of carbon and calcium. You also need to apply a concept related to the mole or Avogadro’s number.

A mole is a huge number used to count the enormous number of atoms or molecules in even a small sample of matter. A mole is equal to 6.02×10^{23} . The mole concept and the atomic mass work together in an interesting way to allow you to count the atoms or molecules in a sample of a substance.

Here’s how it works: Carbon has an atomic mass of 12. That means that 12 grams of carbon contains 6.02×10^{23} carbon atoms. Another way of saying this is that one mole of carbon atoms has a mass of 12 grams.

Calcium has an atomic mass of 40. This means that 40 grams of calcium contains 6.02×10^{23} calcium atoms. Another way of saying this is that one mole of calcium atoms has a mass of 40 grams.

So to get an equal number of carbon and calcium atoms, you could weigh out 12 grams of carbon and 40 grams of calcium. You could also weigh out $\frac{1}{2}$ the amount of each such as 6 grams of carbon and 20 grams of calcium. As long as you use the same fraction or ratio for both, you could measure larger or smaller amounts and you would always have the same number of atoms.

Counting molecules

This same process works for molecules just like it does for atoms but you add the atomic mass for each atom in the molecule. For example, water is made of one oxygen atom and two hydrogen atoms. The atomic mass for oxygen is 16. The atomic mass for hydrogen is 1. So the mass of the molecule is $16 + 1 + 1 = 18$. This means that 18 grams of water contains 6.02×10^{23} water molecules.

How about isopropyl alcohol? Isopropyl alcohol is C_3H_8O . This means that the mass of the molecule is $3(12) + 8(1) + 16 = 36 + 8 + 16 = 60$. This means that 60 grams of isopropyl alcohol contains 6.02×10^{23} isopropyl alcohol molecules. Another way of saying it is that one mole of isopropyl alcohol molecules has a mass of 60 grams.

So you need 60 grams of isopropyl alcohol to have the same number of molecules as there are water molecules in only 18 grams of water.

So if you wanted to do an evaporation test between water and isopropyl alcohol in which you compared the same number of molecules of each, you could do it. You would just need to use a mass of alcohol and a mass of water that are in a ratio of 60 to 18. The activity does not call for this but if you compared the evaporation of 6 grams of isopropyl alcohol to 1.8 grams of water, the evaporation of the alcohol would be faster.

Chapter 5, Lesson 2—Surface Tension

Key Concepts

- The attraction of molecules at the surface of a liquid is called *surface tension*.
- The polarity of water molecules can help explain why water has a strong surface tension.

Summary

Students will observe several phenomena related to the polarity of water molecules. They will observe a demonstration of a paper clip being placed on the surface of water. Students will place drops of water in an already-filled test tube and on the surface of a penny. They will compare the way water behaves with the less polar liquid isopropyl alcohol and will see how detergent affects water's surface tension. Students will relate these observations to an explanation of surface tension at the molecular level.

Objective

Students will be able to explain, on the molecular level, the effects of polarity on water's surface tension.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Materials for the Demonstration

- 1 clear plastic cup
- Water
- 1 standard size paper clip
- 1 large paper clip

Safety

- Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Use in a well-ventilated room.
- Paper towels wet with alcohol should be allowed to evaporate. Dry paper towel can then be placed in the trash.
- Small amounts of isopropyl alcohol can be disposed of down the drain or according to local regulations.
- Have students wash hands after the activity.

Materials for Each Group

- Water
- Isopropyl alcohol (70% or higher)
- Dish detergent in cup
- Test tube
- 2 pennies
- 2 droppers
- 2 toothpicks
- 2 paper towels

ENGAGE

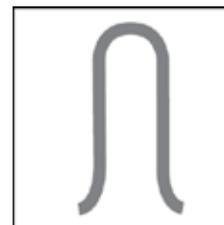
1. Do a demonstration to show water's surface tension.

Either do the following demonstration for students or show them the video *Water's surface tension* at www.middleschoolchemistry.com/multimedia/chapter5/lesson2#water_surface_tension



Materials

- 1 clear plastic cup
- Water
- 1 standard size paper clip
- 1 large paper clip



Teacher Preparation

Unbend the large paper clip until it is straight. Then bend it into a “U” shape. Bend the bottom of each end out a little bit as shown. This will be your device for picking up and placing the smaller paper clip onto the surface of the water. It works like tweezers, but in reverse: allow it to spread it apart in order to pick up the paper clip and squeeze it to release the paper clip.

Procedure

1. Place water in one cup until it is about $\frac{3}{4}$ full.
2. Use your device to pick up a paper clip. Do this by squeezing the ends of the device together a bit and placing them inside the paper clip. Then allow the ends to spread apart so that the tension of the ends pushes against the inside of the paper clip and holds it in place.
3. Very carefully lower the paper clip so that it lies flat on the surface of the water. Slowly squeeze the device to release the paper clip.



Expected Results

The paper clip should rest on the surface of the water. This may take a couple of tries.

Ask students:

- **Why do you think a paper clip, which is more dense than water, can stay on the surface of water?**

Remind students that paper clips are more dense than water and would normally sink. Help students realize that the paper clip in the demo stayed on the surface of the water because of something having to do with the water molecules at the water's surface, called *surface tension*.

2. Have students relate their observations in the demo to a water strider standing on the surface of water.

Project the image *Water Strider and Molecule*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#water_strider

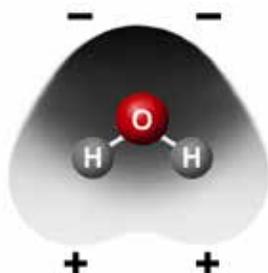
Point out how the surface of the water seems to bend but not break under the water strider's legs. Tell students that what they see is another example of water's surface tension.



Ask students:

- **Why do you think water has such a strong surface tension?**

Encourage students to think about what they already know about the strong attraction between water molecules. Students should remember that water molecules are very attracted to each other and this attraction is the basis for surface tension. This will be explained in more detail later in this lesson.



Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

3. Have students carefully add single drops of water to a filled test tube.

Tell students that there is another phenomenon that is caused by water's surface tension. It is water's ability to fill beyond the top of a container.

Question to Investigate

How much water can you add to a full test tube?



Materials for Each Group

- Water
- Dropper
- Test tube
- Penny
- 2 paper towels

Procedure

1. Pour water into a test tube so that the water is very near the top of the test tube.
2. Hold the test tube up to eye level and use a dropper to carefully add drops of water, one at a time, to the test tube.
3. Watch the water at the top of the test tube while you add the drops. Continue adding drops until the water spills.
4. Place a penny on a paper towel.
5. While watching from the side, add single drops of water to the penny. Continue adding drops until the water spills.



Expected Results

While looking from the side, students will see that the water forms a dome on the top of the test tube and on the penny.

Ask students:

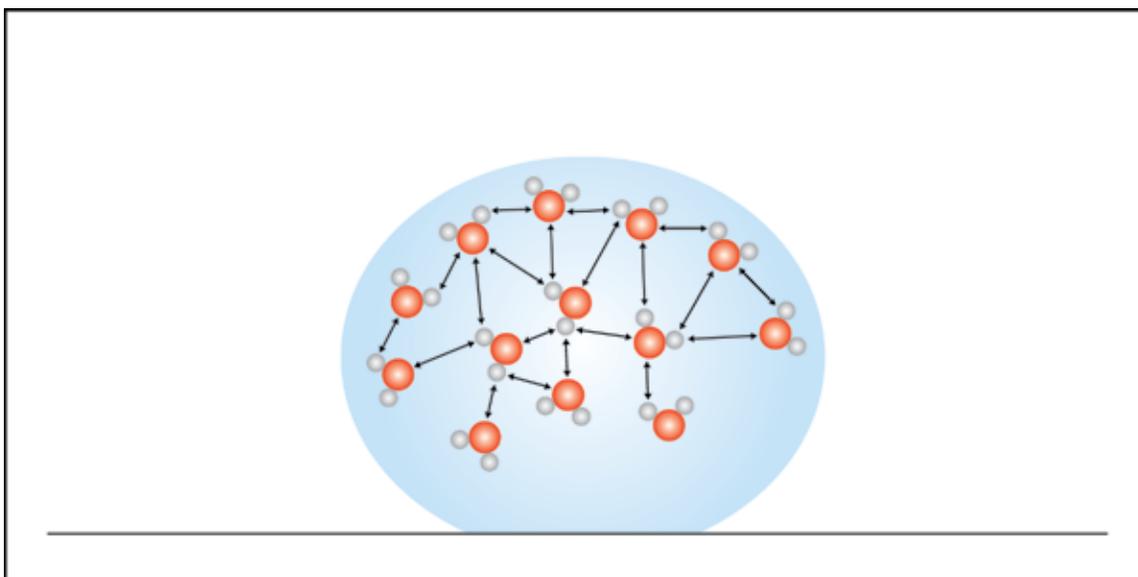
- **What did the water look like as you added it to the top of the test tube and the penny?**
The water makes a “dome” or “hill” of water above the top of the test tube and penny.

EXPLAIN

4. Explain how attractions between water molecules give water its strong surface tension.

Project the image *Why Water Beads*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#water_beads



Explain to students that water's surface tension is based on the attractions between water molecules at the surface and the water molecules in the rest of the water. A water molecule beneath the surface feels attractions from all the molecules around it. But the molecules at the surface only feel attractions from the molecules next to them and beneath them. These surface molecules are pulled together and inward by these attractions. This inward pull has the effect of compressing the surface molecules which form a tight arrangement over the water's surface. This tight arrangement at the surface is called *surface tension*.

The inward pull from the attractions of the molecules results in the smallest possible surface for a volume of water, which is a sphere. This is why water forms a round drop or dome at the top of the filled test tube and on the surface of a penny.

EXPLORE

5. Have students compare the surface tension of water and alcohol.

Ask students:

- **How could we compare the surface tension of water and alcohol?**

Students may suggest placing an equal number of drops of each liquid on wax paper, overflowing a test tube, or comparing the number of drops that can be added to the top of a penny.

Note: Even though there are many ways to compare the surface tension of water and alcohol, the procedure written below compares each liquid on the surface of a penny. In order to make this as fair a test as possible, students should place each liquid on two similar pennies that are either both "heads" or "tails." They should also be sure to add single drops of each liquid slowly and carefully.

Question to Investigate

Which has a greater surface tension, water or alcohol?

Materials for Each Group

- 2 pennies
- 2 droppers
- Water
- Isopropyl alcohol (70% or higher)
- Paper towel



Procedure

1. Place two pennies on a paper towel.
2. Use a dropper to add drops of water to the surface of a penny. Count the drops until the water overflows.
3. Use a dropper to add drops of alcohol to the surface of the other penny. Count the drops until the alcohol overflows.

Expected Results

The water beads up on the penny and the alcohol spreads out flat. Many more drops of water can be added to the penny than drops of alcohol.

EXPLAIN

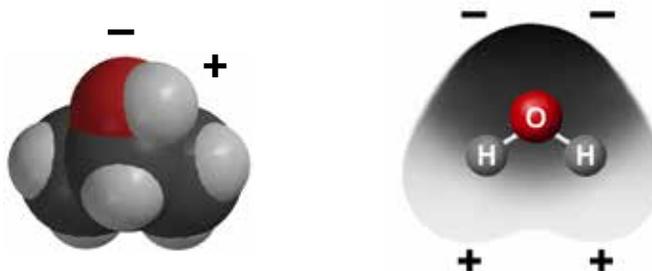
6. Discuss student observations and why students were not able to get as many drops of alcohol on a penny.



Project the image *Water and Alcohol*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#water_and_alcohol

Review that water molecules are polar and that they are very attracted to each other. Point out that alcohol molecules are polar in only one area, making them somewhat attracted to each other. They are not as attracted to other alcohol molecules as water is to other water molecules.



Water and Alcohol on Pennies

Explain that water's attraction pulls itself together into a tight arrangement. Explain that alcohol molecules don't have a structure that is as good as water's for attraction to itself. Alcohol molecules only have 1 O–H bond and they have some C–H bonds that are pretty non polar. There is not as strong an attraction between them as there is between water molecules.

The shape of the water molecule and its polarity at the top and the bottom give water molecules lots of opportunities to attract. Almost anywhere two water molecules meet they can be attracted to each other.

But alcohol has a different size and shape and has its polar part on one end. Alcohol molecules can meet at areas where they would not attract as strongly. The water is more attracted to itself than to the metal of the penny. The alcohol is a bit less attracted to itself so it spreads more on the penny.

EXPLORE

7. Have students add detergent to the water on a penny.



Question to Investigate

How does detergent affect water's surface tension?

Materials

- Dish detergent in cup
- 2 pennies
- Dropper
- 2 tooth picks
- Paper towel

Procedure

1. Place 2 clean dry pennies on a flat surface like a table or desk.
2. Use a dropper to add water to both pennies. Add the same number of drops to each penny so that the water stacks up in a dome shape about the same height on both.
3. Gently touch the water on one penny with a toothpick. Watch the surface of the water as you touch it.



- Dip the toothpick in liquid detergent and then touch the water on the other penny with the toothpick.

Note: This activity works best if the dome of water on the pennies is pretty high.

Expected Results

Touching the water with the toothpick causes the surface of the water to be pressed down and bend. Touching the water with the toothpick and detergent causes the water to collapse and spill off the penny.

EXPLAIN

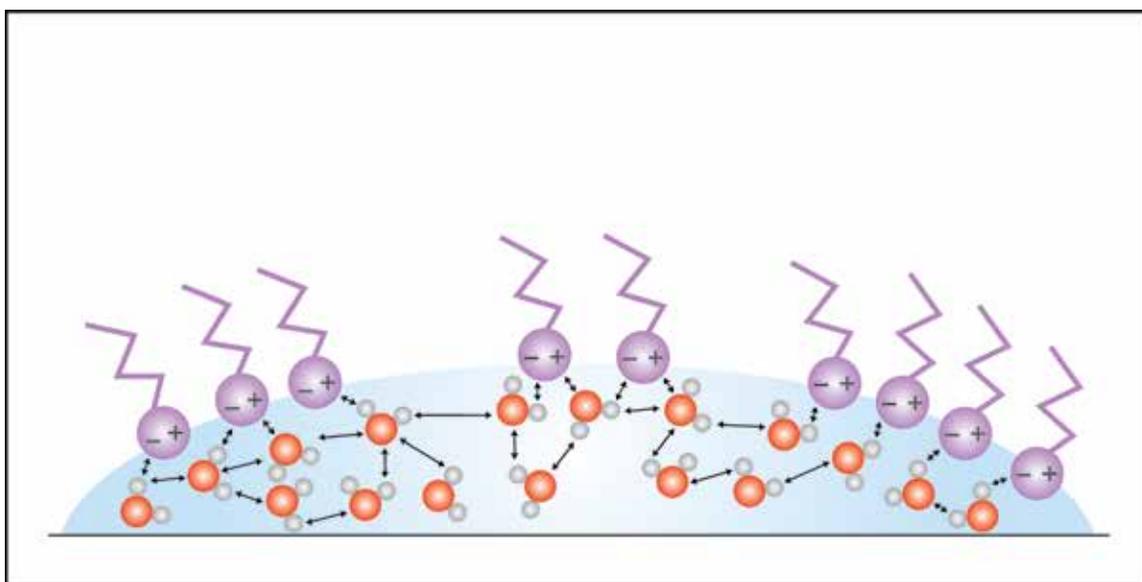
- Explain how detergent interferes with water's surface tension.



Project the image *Water and Detergent*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#water_and_detergent

Explain that detergent is made from molecules that have a charged end and a longer uncharged end. The detergent molecules spread out over the surface of the water with the charged end in the water and the uncharged end sticking out. The water molecules at the surface are attracted to the charged end of the detergent molecules. As the surface water molecules are attracted outward, this acts against their inward attraction that was creating the surface tension. This reduces the surface tension, and the water does not hold its round shape and thus spills.



EXTEND

9. Discuss how the polarity of the material that the water is placed on affects how the water absorbs or beads up.

Ask students:

- If water absorbs into a paper towel but does not absorb into wax paper, what does that say about the polarity of paper and wax paper?

The molecules that make up paper are probably polar, and the molecules that make up wax are probably nonpolar.

Project the animation *Water on Paper Towel*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#paper_towel

Explain that paper towel and other paper is made from cellulose. Cellulose is made from repeating molecules of glucose that are bonded together. The glucose molecule has many O–H bonds, which are polar. Polar water molecules are attracted to polar cellulose.

Project the animation *Water on Wax Paper*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson2#wax_paper

Tell students that wax is made from paraffin, which is repeating carbon–hydrogen bonds. The C–H bond is not very polar, so water is more attracted to itself than to the wax. This causes the water to bead up on wax paper.

EXPLAIN IT WITH ATOMS & MOLECULES

1. You saw a demonstration of a paper clip floating at the surface of water. Paper clips are more dense than water and usually sink. Why do you think the paperclip was able to stay on the surface of the water?



ACTIVITY

Question to Investigate

How much water can you add to a full test tube?



Materials for Each Group

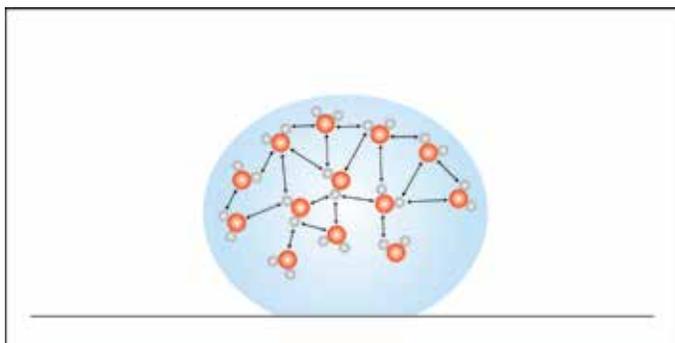
- Water
- Dropper
- Test tube
- Penny
- 2 paper towels

Procedure

1. Pour water into a test tube so that the water is very near the top of the test tube.
2. Hold the test tube up to eye level and use a dropper to carefully add drops of water, one at a time to the test tube.
3. Watch the water at the top of the test tube while you add the drops. Continue adding drops until the water spills.
4. Place a penny on a paper towel.
5. While watching from the side, add single drops of water to the penny. Continue adding drops until the water spills.



2. What did the water look like as you added it to the top of the test tube and the penny?
3. Use the illustration to explain why water has a strong surface tension.



ACTIVITY

Question to Investigate

Which has a greater surface tension, water or alcohol?

Materials for Each Group

- 2 pennies
- 2 droppers
- Water
- Isopropyl alcohol (70% or higher)
- Paper towel

Procedure

1. Place two pennies on a paper towel.
2. Use a dropper to add drops of water to the surface of a penny. Count the drops until the water overflows.
3. Use a dropper to add drops of alcohol to the surface of the other penny. Count the drops until the alcohol overflows.



4. How many drops of each liquid were you able to get on a penny?

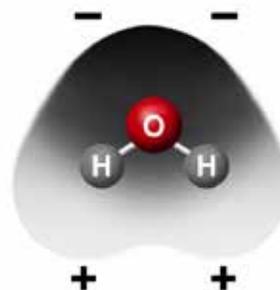
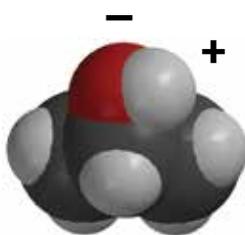
Alcohol

Water

5. Does alcohol or water have a greater surface tension?

How do you know?

6. How does the polarity of alcohol and water molecules affect the surface tension of each liquid?



ACTIVITY

Question to Investigate

How does detergent affect water's surface tension?



Materials

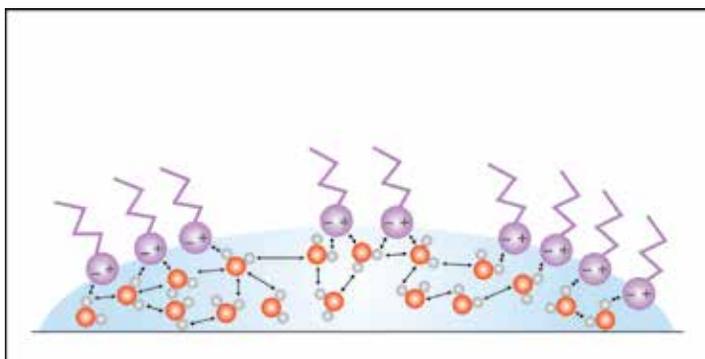
- Dish detergent in cup
- 2 pennies
- Dropper
- 2 tooth picks
- Paper towel

Procedure

1. Place 2 clean, dry pennies on a flat surface like a table or desk.
2. Use a dropper to add water to both pennies. Add the same number of drops to each penny so that the water stacks up in a dome shape about the same height on both.



3. Gently touch the water on one penny with a toothpick. Watch the surface of the water as you touch it.
 4. Dip the toothpick in liquid detergent and then touch the water on the other penny with the toothpick.
7. What happens when you add a small amount of detergent to a large drop of water?
8. Use the illustration to explain how detergent interferes with water's surface tension.



TAKE IT FURTHER

9. If water absorbs into a paper towel but does not absorb into wax paper, what does that say about the polarity of paper and the polarity of wax paper?

Chapter 5, Lesson 3—Why Does Water Dissolve Salt?

Key Concepts

- The polarity of water molecules enables water to dissolve many ionically bonded substances.
- Salt (sodium chloride) is made from positive sodium ions bonded to negative chloride ions.
- Water can dissolve salt because the positive part of water molecules attracts the negative chloride ions and the negative part of water molecules attracts the positive sodium ions.
- The amount of a substance that can dissolve in a liquid (at a particular temperature) is called the *solubility* of the substance.
- The substance being dissolved is called the solute, and the substance doing the dissolving is called the solvent.

Summary

Students will make a 2-D model of a salt crystal and use water molecule cut-outs to show how water dissolves salt. After seeing an animation of water dissolving salt, students will compare how well water and alcohol dissolve salt. They will relate their observations to the structure of salt, water, and alcohol on the molecular level.

Objective

Students will be able to explain, on the molecular level, why water can dissolve salt. Students will be able to identify the variables in their experiment. Students will also be able to explain why a less polar liquid, such as alcohol, is not good at dissolving salt.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Alcohol should be disposed of according to local regulations. Have students wash hands after the activity.

Materials for Each Group

- Construction paper, any color
- Scissors
- Tape or glue
- Water
- Isopropyl alcohol (70% or higher)
- Salt
- Balance
- 2 clear plastic cups
- 2 small plastic cups
- Graduated cylinder

Materials Note:

You may choose to laminate the water molecules, sodium ions, and chloride ions located on the last page of the activity sheet so that you can reuse them with your students next year.

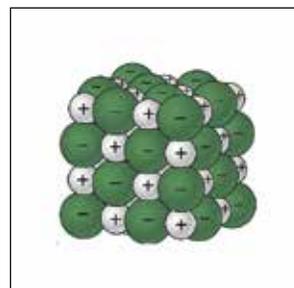
ENGAGE

1. Make a model of a salt crystal.

Project the image *Sodium Chloride Crystal*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson3#sodium_chloride_crystal

Remind students that the green balls represent negative chloride ions and the gray balls represent positive sodium ions.



Ask students:

- **What is it about water molecules and the ions in salt that might make water able to dissolve salt?**

The positive and negative polar ends of a water molecule are attracted to the negative chloride ions and positive sodium ions in the salt.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

Question to Investigate

How does salt dissolve in water?

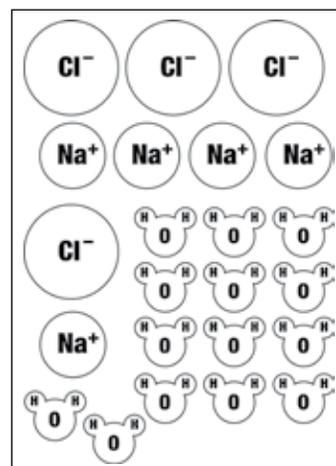
Materials

- Activity sheet with sodium and chloride ions and water molecules
- Construction paper, any color
- Scissors
- Tape or glue

Procedure

Make a model of a salt crystal

1. Cut out the ions and water molecules.
2. Arrange the ions on a piece of construction paper to represent a 2-D salt crystal. Do not tape these pieces down yet.



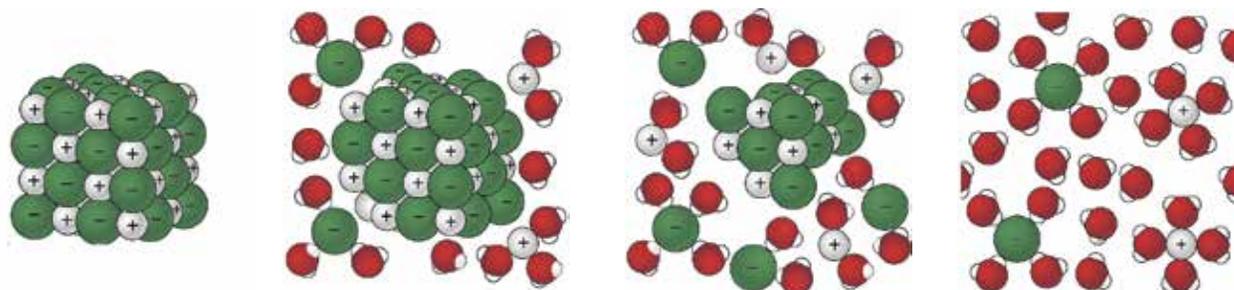
2. Project an image and have students model what happens when salt dissolves in water.

Show students a series of four pictures to help explain the process of water dissolving salt.

Project the image *Sodium Chloride Dissolving in Water*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson3#sodium_chloride_dissolving

Point out that several water molecules can arrange themselves near an ion and help remove it from the crystal. Show students that the positive area of a water molecule will be attracted to the negative chloride ion and that the negative area of a water molecule will be attracted to the positive sodium ion.



Model how water dissolves salt

1. Look at the pictures showing how water molecules dissolve salt. Then arrange the water molecules around the sodium and chloride ions in the correct orientation. The positive part of the water molecules should be near the negative chloride ion. The negative part of the water molecules should be near the the positive sodium ion.
2. Move the water molecules and sodium and chloride ions to model how water dissolves salt.
3. Tape the molecules and ions to the paper to represent water dissolving salt.



Project the animation *Sodium Chloride Dissolving in Water*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson3#sodium_chloride_dissolving_video

Point out that the water molecules are attracted to the sodium and chloride ions of the salt crystal. Explain that the positive area of a water molecule is attracted to a negative chloride ion. The negative area of a water molecule is attracted to a positive sodium ion. Dissolving happens when the attractions between the water molecules and the sodium and chloride ions overcome the attractions of the ions to each other. This causes the ions to separate from one another and become thoroughly mixed into the water.

Tell students that the amount of a substance that can dissolve in a liquid (at a particular temperature) is called *solubility*. Point out the similarity in the words dissolve and solubility. Also tell them that the substance that is dissolved is called the solute. The substance that does the dissolving is called the solvent.

EXPLORE

3. Have students conduct an experiment to find out whether water or isopropyl alcohol would be better at dissolving salt.

Ask students to make a prediction:

- **Think about the polarity of water molecules and alcohol molecules. Do you think alcohol would be just as good, better, or worse than water at dissolving salt?**

Discuss how to set up a test to compare how water and alcohol dissolve salt. Be sure students identify variables such as:

- Amount of water and alcohol used
- Amount of salt added to each liquid
- Temperature of each liquid
- Amount of stirring

Question to Investigate

Is alcohol just as good, better, or worse than water at dissolving salt?

Materials for Each Group

- Water
- Isopropyl alcohol (70% or higher)
- Salt
- Balance
- 2 clear plastic cups
- 2 small plastic cups
- Graduated cylinder



Procedure

1. In separate cups, measure two samples of salt that weigh 5 g each.
2. Place 15 mL of water and alcohol into separate cups.
3. At the same time, add the water and alcohol to the samples of salt.
4. Swirl both cups the same way for about 20 seconds and check for the amount of salt dissolved.
5. Swirl for another 20 seconds and check. Swirl for the last 20 seconds and check.
6. Carefully pour off the water and alcohol from the cups and compare the amount of undissolved salt left in each cup.

Expected Results

There will be less undissolved salt in the cup with the water than the alcohol. This means that more salt dissolved in the water than in the alcohol.

EXPLAIN

4. Discuss how differences in the polarity of alcohol and water explain why water dissolves salt better than alcohol.

Ask students:

- **Is alcohol just as good, better, or worse than water at dissolving salt?**

Alcohol does not dissolve salt as well as water does.

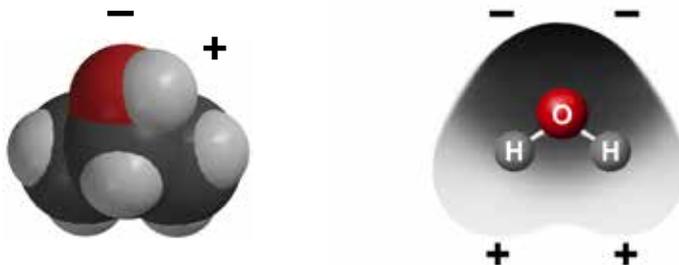
- **How do you know?**

There was more salt left behind in the cup with the alcohol.

- **Think about the polarity of water and alcohol to explain why water dissolves more salt than alcohol.**

Have students look at the models of water and alcohol molecules on their activity sheet.

Read more about polarity in the additional teacher background section at the end of this lesson.



Remind students that isopropyl alcohol has an oxygen atom bonded to a hydrogen atom so it does have some polarity but not as much as water. Since water is more polar than alcohol, it attracts the positive sodium and negative chloride ions better than alcohol. This is why water dissolves more salt than alcohol does. Another way of saying this is that the solubility of salt is greater in water than in alcohol.

EXTEND

5. Have students compare the solubility of two different ionic substances in water.

Compare the solubility of the ionic substances calcium carbonate (CaCO_3) and sodium carbonate (Na_2CO_3) in water.

Ask students:

- **How could you compare the solubility of calcium chloride and calcium carbonate?**

Students should suggest measuring equal amounts of each substance and adding equal amounts of water at the same temperature.

Question to Investigate

Do all ionic substances dissolve in water?

Materials for Each Group

- Sodium carbonate
- Calcium carbonate
- Water
- 2 clear plastic cups
- 2 small plastic cups
- Balance

Procedure

1. Label two clear plastic cups Sodium Carbonate and Calcium Carbonate.
2. Measure 2 g each of sodium carbonate and calcium carbonate and put them in their labeled cups.
3. Measure 15 mL of water into each of two empty cups.
4. At the same time, pour the water into the sodium carbonate and calcium carbonate cups.
5. Gently swirl both cups.



Expected Results

The sodium carbonate will dissolve, but the calcium carbonate will not. Explain that not all ionically bonded solids dissolve in water.

6. Discuss student observations.

Ask students:

- **Do all ionic substances dissolve in water? How do you know?**

Because calcium carbonate does not dissolve in water, students should realize that not all ionic substances dissolve in water.

Explain that on the molecular level, the ions that make up calcium carbonate are attracted so strongly to each other that the attraction by water molecules cannot pull them apart. That is a good thing because calcium carbonate is the material that sea shells and bird eggs are made of. Calcium phosphate is another ionic solid that does not dissolve in water. This is also good because it is the material that bones and teeth are made of.

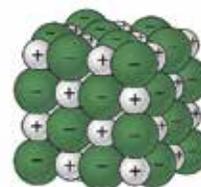
Sodium carbonate breaks apart completely into ions that are incorporated throughout the water, forming a solution. The sodium and carbonate ions will not settle to the bottom and cannot be filtered out of the water.

But calcium carbonate does not break up into its ions. Instead it is just mixed in with the water. If given enough time, the calcium carbonate will settle to the bottom or can be filtered out of the water. Sodium carbonate dissolved in water is a good example of a solution, and undissolved calcium carbonate is a mixture, not a solution.

Note: The carbonate ion is different from the single-atom ions such as sodium (Na^+) and chloride (Cl^-) that students have seen so far. The carbonate ion (CO_3^{2-}) is composed of more than one atom. These types of ions, called polyatomic ions, are made up of a group of covalently bonded atoms that act as a unit. They commonly gain or lose one or more electrons and act as an ion. Another common polyatomic ion is the sulfate ion (SO_4^{2-}). This ion is part of Epsom salt as magnesium sulfate (MgSO_4) and many fertilizers as potassium sulfate (K_2SO_4). You can decide if you would like to introduce students to these two common polyatomic ions.

INTRODUCTION

1. What is it about water molecules and the ions in salt that might make water able to dissolve salt?



Question to Investigate

How does salt dissolve in water?



Materials

- Activity sheet with sodium and chloride ions and water molecules
- Construction paper, any color
- Scissors
- Tape or glue

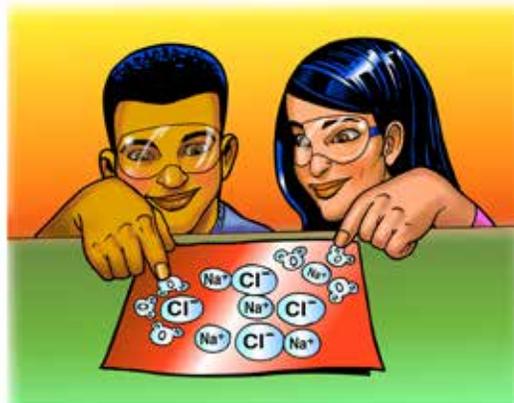
Procedure

Make a model of a salt crystal

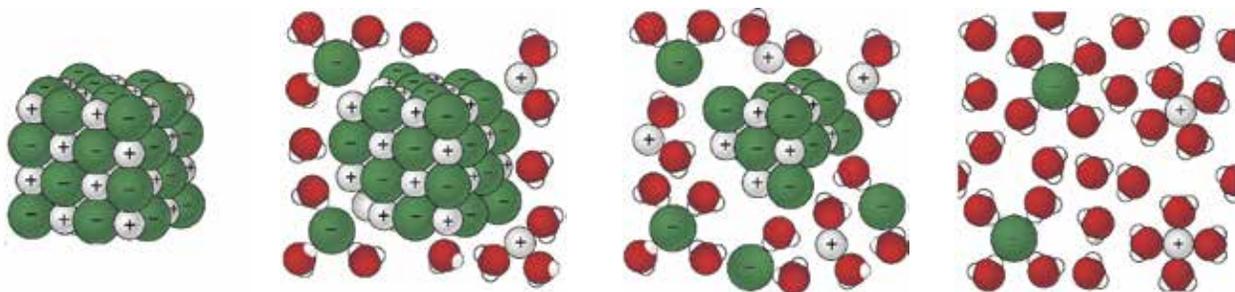
1. Cut out the ions and water molecules.
2. Arrange the ions on a piece of construction paper to represent a 2-dimensional salt crystal. Do not tape these pieces down yet.

Model how water dissolves salt

1. Look at an image and animation showing how water molecules dissolve salt. Then arrange the water molecules around the sodium and chloride ions in the correct orientation. The positive part of the water molecules should be near the negative chloride ion. The negative part of the water molecules should be near the positive sodium ion.
2. Move the water molecules and sodium and chloride ions to model how water dissolves salt.
3. Tape the molecules and ions to the paper to represent water dissolving salt.



2. Describe what happens when water dissolves salt.



ACTIVITY

Question to Investigate

Is alcohol just as good, better, or worse than water at dissolving salt?

Materials for Each Group

- Water
- Isopropyl alcohol (70% or higher)
- Salt
- Balance
- 2 clear plastic cups
- 2 small plastic cups
- Graduated cylinder

Procedure

1. In separate cups, measure two samples of salt that weigh 5 grams each.
2. Place 15 mL of water and alcohol into separate cups.
3. At the same time, add the water and alcohol to the samples of salt.
4. Swirl both cups the same way for about 20 seconds and check for the amount of salt dissolved.
5. Swirl for another 20 seconds and check. Swirl for the last 20 seconds and check.
6. Carefully pour off the water and alcohol from the cups and compare the amount of undissolved salt left in each cup.



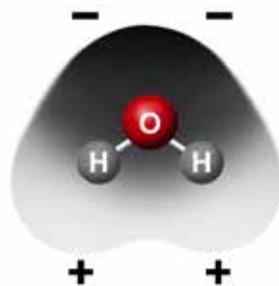
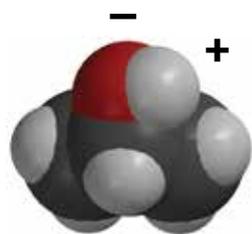
3. Select two variables and explain how they are controlled in this procedure.

4. Is alcohol just as good, better, or worse than water at dissolving salt?

How do you know?

EXPLAIN IT WITH ATOMS & MOLECULES

5. Think about the polarity of water and alcohol to explain why water dissolves salt better than alcohol does.



TAKE IT FURTHER

Question to Investigate

Do all ionic substances dissolve in water?

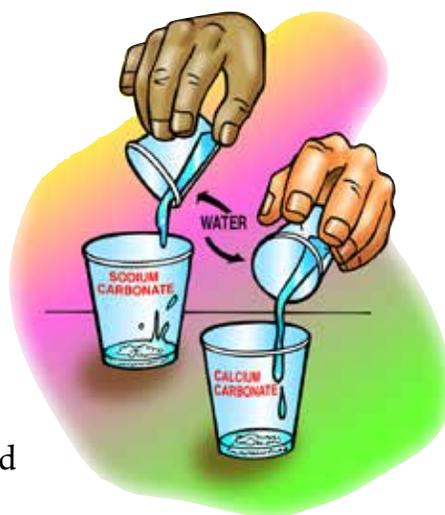
Materials for Each Group

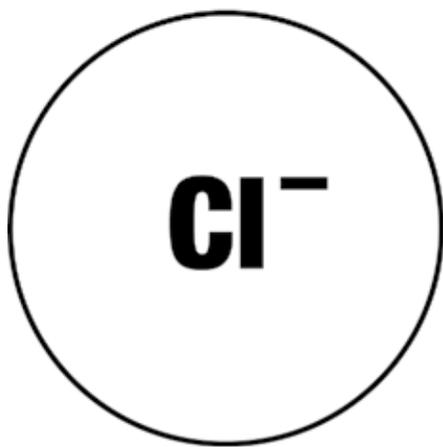
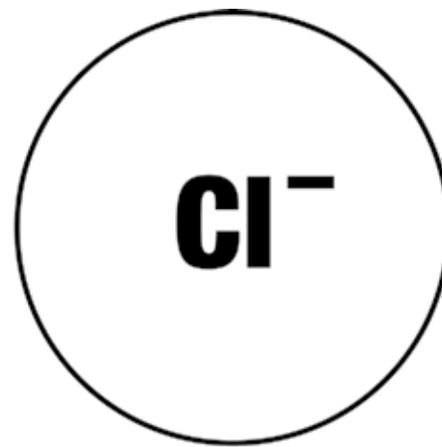
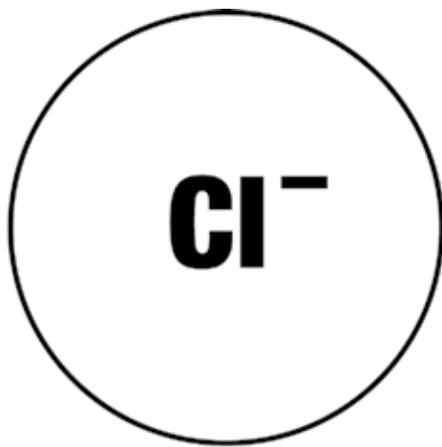
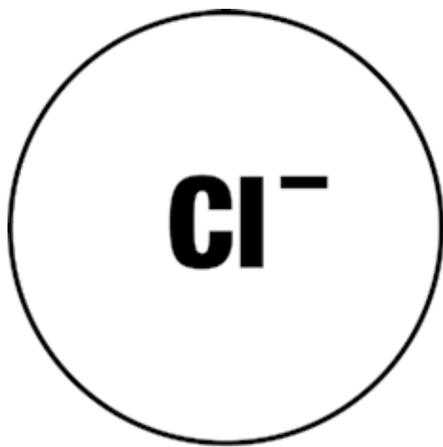
- Sodium carbonate
- Calcium carbonate
- Water
- 2 clear plastic cups
- 2 small plastic cups
- Balance

Procedure

1. Label two clear plastic cups sodium carbonate and calcium carbonate.
 2. Measure 2 g each of sodium carbonate and calcium carbonate and put them in their labeled cups.
 3. Measure 15 mL of water into each of two empty cups.
 4. At the same time, pour the water into the sodium carbonate and calcium carbonate cups.
 5. Gently swirl both cups.
6. Do all ionic substances dissolve in water?

How do you know?





Chapter 5, Lesson 3

Teacher Background, p. 414

Non-polar molecules can attract each other

Lessons 1–3 focus on the principle that molecules attract one another based on their polarity. This principle is true but polarity is not the only cause of attraction between molecules. It may seem strange but even non-polar molecules attract one another. Non-polar molecules such as the molecules in mineral oil or gasoline attract each other and stay together as a liquid. The non-polar molecules of paraffin attract each other and stay together as solid wax.

Non-polar molecules can attract each other based on the movement of electrons within their atoms. The negatively charged electrons are in constant motion in regions around the positively charged protons. But at any given instant, there can be a slight and fleeting imbalance in the electrons on one side of an atom than another. This makes one side of an atom temporarily slightly positive and the other side temporarily slightly negative.

The slightly negative side can repel electrons from a nearby atom in another molecule creating a temporary positive area in that atom. These negative and positive areas attract each other and help hold the molecules together. These types of attractions are called dispersion forces. Non-polar molecules containing a greater number of atoms tend to be held together more strongly than non-polar molecules with fewer atoms. This is because there are more opportunities for *dispersion forces* to attract the molecules.

Chapter 5, Lesson 4: Why Does Water Dissolve Sugar?

Key Concepts

- For a liquid to dissolve a solid, the molecules of the liquid and solid must attract one another.
- The bond between the oxygen and hydrogen atoms (O–H bond) in sugar (sucrose) gives the oxygen a slight negative charge and the hydrogen a slight positive charge. Sucrose is a polar molecule.
- The polar water molecules attract the negative and positive areas on the polar sucrose molecules which makes sucrose dissolve in water.
- A nonpolar substance like mineral oil does not dissolve a polar substance like sucrose.

Summary

Students will observe the dissolving of the sugar coating from an M&M when it is placed in water. Students will then help design an experiment to see if the type of liquid the M&M is placed in affects how much of the coating dissolves.

Objective

Students will be able to explain, on the molecular level, how the polar characteristic of water and sugar interact so that water dissolves sugar. Students will be able to identify and control the variables in their experiment. Students will also be able to explain why a nonpolar liquid, such as mineral oil, is not good at dissolving sugar.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Dispose of isopropyl alcohol and mineral oil according to local regulations. Warn students not to eat the M&M's. Have students wash hands after the activity.

Materials for Each Group

- M&M's
- Water
- Mineral oil
- Isopropyl alcohol (70%)
- Small white plastic plate
- 3 clear plastic cups
- White paper

ENGAGE

1. Help students realize that the candy coating of an M&M is made mostly of sugar and a bit of coloring.

Distribute M&M's to students and have them look at the outside candy coating. Then have students break an M&M to look closely at the coating from the inside.

Ask students:

- **What do you think the coating of an M&M is made from?**
Students will see the layer of color with a layer of white beneath it and suggest that the coating is made of sugar and coloring. Explain that the coating is mostly sugar.
- **Have you ever noticed what happens to the coating of an M&M when it gets wet?**
The color comes off and if it gets wet enough, the entire coating comes off, leaving the chocolate behind.

Tell students that in this activity, they will see what happens to the sugar and color coating of an M&M when it is placed in water.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.

2. Have students place an M&M in a cup of water and observe.

Question to Investigate

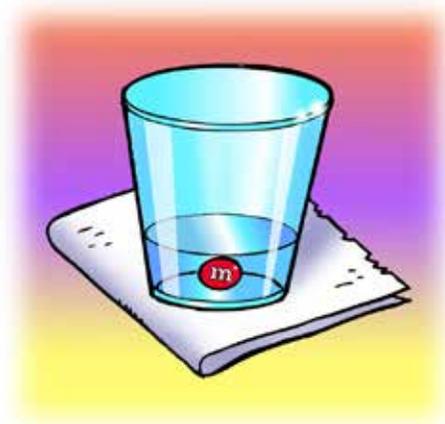
What happens to the sugar and color coating of an M&M when it is placed in water?

Materials

- Clear plastic cup
- Water
- M&M
- White paper

Procedure

1. Pour enough room-temperature water into clear plastic cup so that the water is deep enough to completely cover an M&M and place this cup on a piece of white paper.



2. Once the water has settled, place 1 M&M in the center of the cup. Be careful to keep the water and M&M as still as possible. Observe for about 1 minute.

Expected Results

The coating will dissolve from the M&M, revealing a white layer under the color and then the brown chocolate underneath. The colored coating of the M&Ms will collect in a circular pattern around the M&M. Students may also mention the white streaks in the water from the sugar coating.

3. Discuss student observations.

Ask students:

- **What do you notice about the M&M and the water?**
The color comes off and moves through the water in a circular pattern.
- **What do you think is happening when the color and sugar come off the M&M?**
Point out to students that because the water makes the colored coating come off the M&M and mix into the water, the water is dissolving the sugar and color.

Note: There are actually two processes happening in this activity. The color and sugar are dissolving in the water but they are also diffusing. In order to focus on the amount that dissolves from an M&M, students should look at the amount of coating missing from the M&M, instead of the size of the circle of color in the water.

- **Knowing what you do about the polarity of water, why do you think water dissolves sugar?**
Students may think that sugar is made of ionic bonds like salt. Or they might think that sugar has positive and negative areas and this is why water is attracted to it.

EXPLAIN

4. Show students how the polar areas of a sucrose molecule cause it to dissolve in water.

Explain to students that sugar is made of large molecules called sucrose. Each sucrose molecule is made of atoms that are covalently bonded.

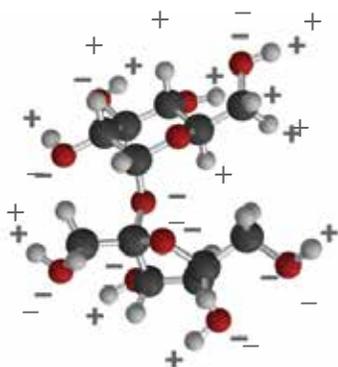
Ask students:

- **The chemical formula for sucrose is $C_{12}H_{22}O_{11}$. What do these letters and numbers mean?**
Sucrose is made up of 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.

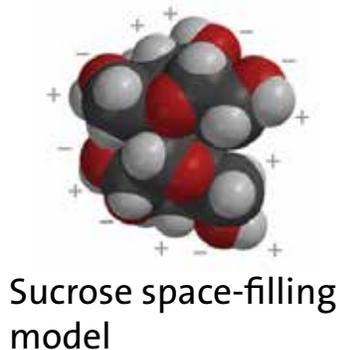
Project the image *Sucrose*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#sucrose

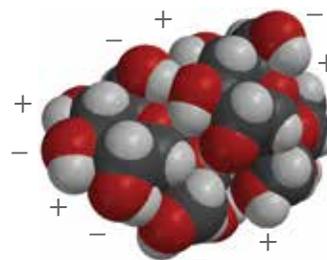
Explain that the first picture is a ball-and-stick model of a single sucrose molecule. The next picture is a space-filling model of a single sucrose molecule. The last picture is two sucrose molecules attracted to each other. These two molecules will separate from each other when sugar dissolves. Point out that in the areas on a sucrose molecule where oxygen is bonded to hydrogen (O–H bond), the oxygen is slightly negative and the hydrogen is slightly positive. This makes sucrose a polar molecule.



Sucrose ball-and-stick model



Sucrose space-filling model

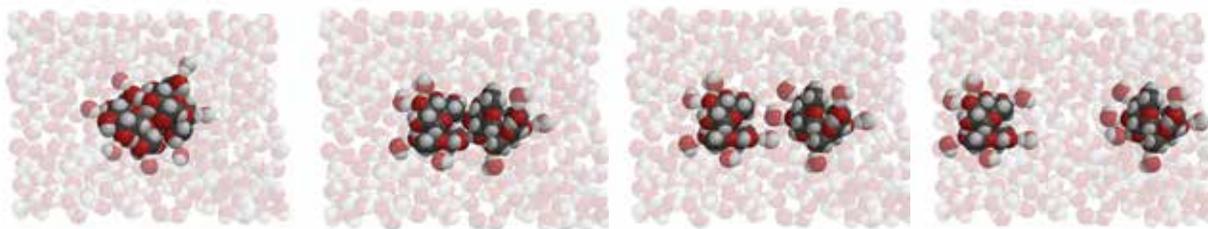


Two sucrose molecules closely associated due to oppositely charged polar areas

Project the image *Water Dissolves Sucrose*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#water_dissolves_sucrose

Explain that a sugar cube (about a half a teaspoon of sugar) is made up of at least one billion trillion sucrose molecules. When sugar dissolves, these whole sucrose molecules separate from one another. The molecule itself doesn't come apart: The atoms that make up each molecule stay together as a sucrose molecule.



Project the animation *Sucrose*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#sucrose_animation

Explain that sucrose has polar areas caused by the same type of oxygen–hydrogen covalent bonds as in the water molecule. Point out the O–H bonds on the outer edges of the molecule. Point out that the bonding of the oxygen and hydrogen in the sucrose makes parts of the sucrose molecule polar in a similar way as in a water molecule. The area near the hydrogen is positive (blue) and the area near the oxygen is negative (red).

Project the animation *Water Dissolves Sucrose*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#water_dissolves_sucrose_video

Tell students that sugar molecules are attracted to each other and held together by the attraction between these polar areas of the molecules.

Help students notice how the positive (blue) area of a water molecule is attracted to the negative (red) area of a sucrose molecule. It also works the other way around. The negative (red) area of a water molecule is attracted to the positive (blue) area of the sucrose molecule.

Explain that the positive and negative areas on water molecules interact with these negative and positive parts of sucrose molecules. When the attraction between water molecules and sucrose molecules overcomes the attraction the sucrose molecules have to other sucrose molecules, they will separate from one another and dissolve.

Point out that one whole sucrose molecule breaks away from another whole sucrose molecule. The molecule itself does not come apart into individual atoms.

Note: Although the focus is on dissolving the polar sugar molecules, the food coloring used to color the M&M is also made from polar molecules. This helps explain why the coloring also dissolves.

EXPLORE

5. Have students conduct an experiment to compare how well water, alcohol, and oil dissolve the sugar and color coating of an M&M.

Ask students to make a prediction:

- Do you think water, alcohol, or oil would be better at dissolving the sugar and color coating of an M&M?

Discuss with students how to design an experiment to compare how well water, alcohol, and vegetable oil dissolve the color and sugar coating from an M&M. Be sure students identify variables such as:

- Amount of water, alcohol, and oil used
- Temperature of each liquid
- Same color of M&M
- Time and location the M&M's are placed in each liquid

Question to Investigate

Is water, alcohol, or oil better at dissolving the color and sugar coating from an M&M?

Materials

- 3 M&Ms (same color)
- Water
- Mineral oil
- Isopropyl alcohol (70%)
- 3 clear plastic cups
- White paper



Procedure

1. Label 3 cups Water, Alcohol, and Oil. Add 15 mL of water, alcohol, and mineral oil to their labeled cups.
2. Place the three cups on a white sheet of paper.
3. At the same time, add 1 M&M to each liquid. Then gently swirl the liquid and M&M in each cup for about 30 seconds.

Expected Results

Water—The sugar and color dissolve from the M&M.

Alcohol—The color dissolves only slightly and the sugar coating doesn't seem to dissolve.

Oil—Neither the color nor the sugar dissolves

EXPLAIN

6. Show students the molecular structures for water, alcohol, and oil and discuss how this relates to their observations.

Project the image *Polarity of Water, Alcohol, and Oil*.

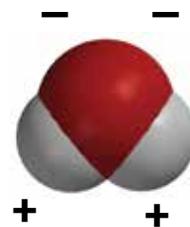
www.middleschoolchemistry.com/multimedia/chapter5/lesson4#polarity

Show students the polar areas on a water molecule, isopropyl alcohol molecule, and an oil molecule.



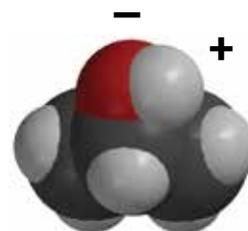
Water

Water molecules are polar. The polar water dissolves the polar coloring and the polar sugar.



Isopropyl Alcohol (70%)

The alcohol is 30% water and 70% alcohol and is not a good dissolver. Alcohol molecules have only one polar area and also have a larger nonpolar area. This makes alcohol not a good dissolver of polar substances. Also, the water and alcohol interact, which means the water doesn't even dissolve the sugar or color as well as it normally would.



Oil

Oil molecules are not polar so they cannot dissolve either the coloring or the sugar.



EXTEND

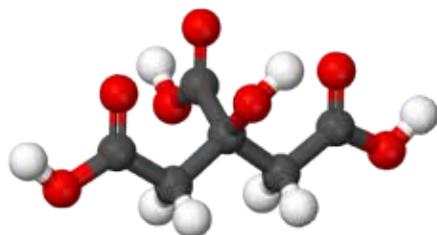
Have students explain on the molecular level why citric acid dissolves so well in water.



Project the image *Citric Acid*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#citric_acid

Explain that the projected image is a model of a citric acid molecule. Tell students that citric acid is the substance that gives lemons, limes, grapefruit, and oranges their tangy sour taste. Citric acid is very soluble in water and is dissolved in the water in the fruit.



Ask students:

- **Why do you think citric acid is so soluble in water?**

HINT: The chemical formula for citric acid is $C_6H_8O_7$.

Every place there is an O–H bond, there is an uneven sharing of electrons. The oxygen atoms in an O–H bond have a slightly negative charge and the hydrogen atoms in the bond have a slightly positive charge. Because water molecules are also polar, the positive ends of water molecules are attracted to the negative areas of the citric acid molecules. The negative ends of water molecules are attracted to the positive areas of the citric acid molecules. These mutual attractions will overcome the attractions citric acid molecules have for other citric acid molecules, causing them to mix thoroughly in the water and dissolve.

INTRODUCTION

Question to Investigate

What happens to the sugar and color coating of an M&M when it is placed in water?

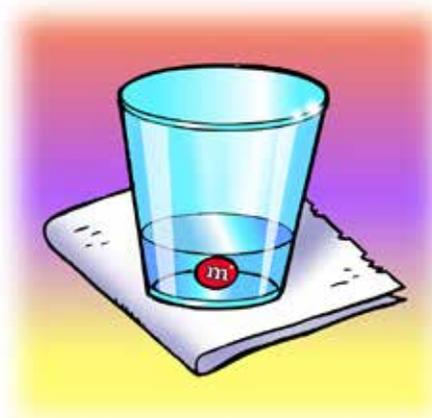


Materials

- Clear plastic cup
- Water
- M&M
- White paper

Procedure

1. Pour enough room-temperature water into a clear plastic cup so that the water is deep enough to completely cover an M&M and place this cup on a piece of white paper.
2. Once the water has settled, place 1 M&M in the center of the cup. Be careful to keep the water and M&M as still as possible. Observe for about 1 minute.



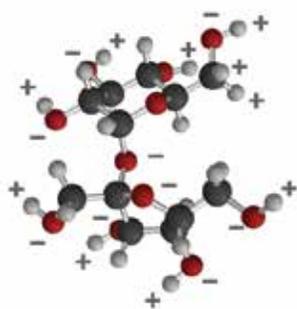
1. What happens to the sugar and color coating when an M&M is placed in water?

2. Knowing what you do about the polarity of water, why do you think water dissolves sugar?

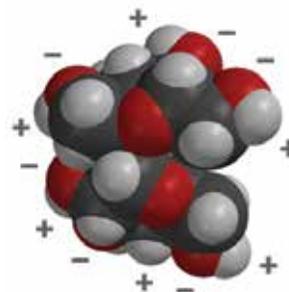
EXPLAIN IT WITH ATOMS & MOLECULES

3. Sucrose makes up the sugar we commonly use. The chemical formula for sucrose is $C_{12}H_{22}O_{11}$. What do these letters and numbers mean?

4. What do the + and – signs around certain parts of the sucrose molecule mean?

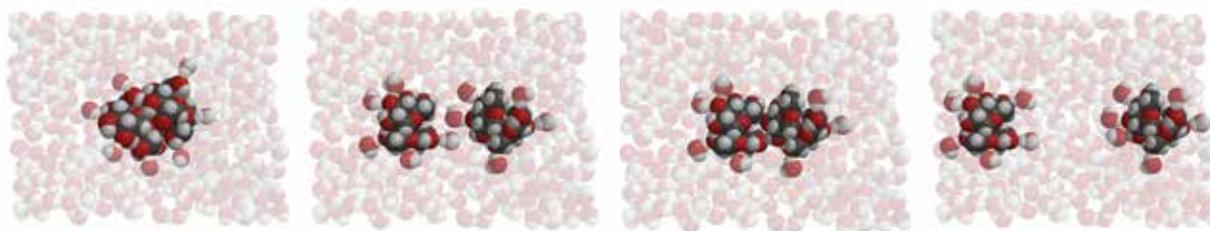


Sucrose ball-and-stick model



Sucrose space-filling model

5. Look at the pictures below and describe what happens when water dissolves sucrose. Be sure to discuss the polarity of both water and sucrose.



ACTIVITY

Question to Investigate

Is water, alcohol, or oil better at dissolving the color and sugar coating from an M&M?

Materials

- 3 M&M's (same color)
- Water
- Mineral oil
- Isopropyl alcohol (70%)
- 3 clear plastic cups
- White paper



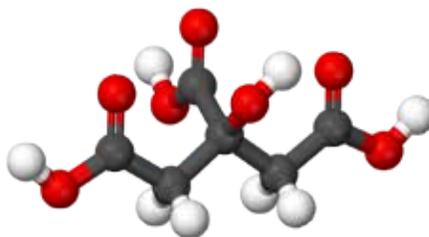
Procedure

1. Label 3 cups Water, Alcohol, and Oil. Add 15 mL of water, alcohol, and mineral oil to their labeled cups.
2. Place the three cups on a white sheet of paper.
3. At the same time, add 1 M&M to each liquid. Then gently swirl the liquid and M&M in each cup for about 30 seconds.

TAKE IT FURTHER

7. Citric acid occurs naturally in fruits like oranges, lemons, and limes. It is dissolved in the water within the fruit and contributes to the fruit's sour taste. When it isn't dissolved in water, citric acid molecules are attracted to other citric acid molecules within a crystal.

The chemical formula for citric acid is $C_6H_8O_7$ and it is very soluble in water. Why do you think citric acid is so soluble in water?



Ball-and-stick model of a citric acid molecule

Chapter 5, Lesson 5—Using Dissolving to Identify an Unknown

Key Concepts

- Different substances are made from different atoms, ions, or molecules, which interact with water in different ways.
- Since dissolving depends on the interaction between water and the substance being dissolved, each substance has a characteristic solubility.

Summary

Students will observe a solubility test between salt and sugar. They will then be presented with four known crystals and an unknown. Based on the solubility demonstration, the class will design a solubility test to discover the identity of the unknown.

Objective

Students will be able to identify and control variables when designing a solubility test. Students will be able to explain why different substances dissolve to different extents in water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for the Demonstrations

- Gram balance
- Simple balance
- Graduated cylinder
- Water
- 4 clear plastic cups
- 2 small plastic cups
- Salt
- Sugar
- Cereal balls (Kix work well)
- Zip-closing plastic bag (quart-size, storage-grade)

Materials for Each Group

- Salt (sodium chloride)
- Epsom salt (magnesium sulfate)
- MSG (monosodium glutamate)
- Sugar (sucrose)
- Coarse kosher salt (sodium chloride)
- Water
- Black construction paper
- Masking tape
- Pen or permanent marker
- Magnifier
- Gram balance
- 5 small plastic cups
- 5 clear plastic cups
- Graduated cylinder
- Paper towel

About this Lesson

This lesson will take two or three class periods.

ENGAGE

1. Do a demonstration to show that different substances have different solubilities.

Tell students that in this demonstration, you will pour salt and sugar into water to find out which dissolves better. In order to make this test fair, you will use the same amount (mass) of salt and sugar, the same amount of water at the same temperature, and you will swirl each in the same way for the same length of time.

Question to Investigate

Which dissolves in water better, salt or sugar?



Materials for the Demonstration

- Balance that measures in grams
- Graduated cylinder
- Water
- 2 clear plastic cups
- 2 small plastic cups
- Salt
- Sugar

Teacher Preparation

- Label 1 clear plastic cup and 1 small cup Salt.
- Label the other clear plastic cup and another small cup Sugar.
- Measure 5 grams of salt and 5 grams of sugar and place them in the pair of small labeled cups.
- Pour 5 mL of room-temperature water into the pair of larger empty cups.



Procedure

1. At the same time, pour the salt and sugar into the water in the corresponding cups. Swirl each cup at the same time and in the same way for about 20 seconds.
2. Walk around the room to show students the amount of salt and sugar left in the bottom of each cup. If you have an overhead projector, place the cups on the projector so that the entire class can compare what is left undissolved in each cup. Ask students whether one substance seems to dissolve better than the other.
3. Swirl again for 20 seconds and observe. Then swirl for 20 more seconds and have students make their final observations.

4. Slowly and carefully pour the solution from each cup back into its empty labeled cup. Try not to let any undissolved crystal go into these cups. Show students the cups so that they can compare the amount of undissolved crystal remaining.

Expected Results

Much more sugar will dissolve than salt. There will be more undissolved salt than sugar left in the cups.

Note: Solubility is normally measured by the number of grams of a substance that dissolves in a certain volume of water at a given temperature. The previous demonstration uses this conventional way to measure solubility. Another approach could be to compare the number of molecules or ions of each substance that dissolves in water. This would require a way to “count” the molecules or ions in each substance.

Read more about counting molecules or ions in the additional teacher background on page 369.

2. **Discuss the results of the demonstration and introduce the idea that each substance has its own characteristic solubility.**

Ask students:

- **Was more salt or sugar left in the bottom of the cup?**
There was more salt left undissolved in the bottom of the cup.
- **Which dissolved better, salt or sugar?**
Because there was little to no sugar in the bottom of the cup, more of it must have dissolved in the water.
- **Do you think we would get similar results if we tried dissolving salt and sugar again?**
We probably would get similar results because the amount of salt or sugar that dissolves has something to do with how each substance interacts with water.
- **How well a substance dissolves in water is called its *solubility*. Would you expect different substances to have the same or different solubility?**
Each substance is made up of its own kind of molecules that will interact with water differently. Different substances should have different solubilities.

Tell students that they will compare the solubility of four different household crystals—salt (sodium chloride), Epsom salt (magnesium sulfate), MSG (monosodium glutamate), and sugar (sucrose). Explain that they will also test an unknown crystal that is chemically the same as one of the other crystals. Because it is chemically the same, it should have the same solubility as one of the crystals they will test. By the end of the activity, students should be able to identify the unknown.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

3. Have students try to identify the unknown based on appearance.

Let students know that before doing a solubility test, they will look closely at the crystals to see if they might be able to get some clues about the identity of the unknown by appearance alone. Have students follow the procedure below and record their observations about the crystals on the activity sheet. Let students know that they can look at the crystals and touch them but they should not taste them.

Question to Investigate

Can you identify the unknown crystal by the way it looks?



Materials for Each Group

- Black construction paper
- Masking tape
- Pen or permanent marker
- 5 small plastic cups
- Salt
- Epsom salt
- MSG
- Sugar
- Unknown (Coarse kosher salt)
- Magnifier

Teacher Preparation

- Label the 5 small plastic cups *Salt*, *Epsom salt*, *MSG*, *Sugar*, and *Unknown*.
- Add at least two teaspoons of each crystal to its labeled cup.

Procedure

1. Use masking tape and a pen to label four corners of a piece of black construction paper *Sugar*, *Salt*, *Epsom salt*, and *MSG*. Label the center *Unknown*.
2. Place small samples of Epsom salt, table salt, sugar, MSG, and the unknown on the labeled areas of the construction paper.
3. Use a magnifier to look carefully at each type of crystal.

Expected Results

All of the crystals are white, but some are more transparent or opaque than others. Each type of crystal is also a different size and shape.



4. Discuss student observations and have groups plan how they might conduct a solubility test to identify the unknown.

Ask students:

- **What do you notice about each crystal? Include any similarities or differences you notice among them.**

Students should describe physical properties such as the size, shape, color and texture. They should also describe whether the crystals are shiny, dull, transparent, or opaque.

- **Can you identify the unknown yet?**

Students should not have enough evidence to correctly identify the unknown at this point. Don't tell students yet that the unknown is coarse kosher salt. They will discover this by the end of this lesson.

Explain that looking at the crystals is not enough to identify the unknown. But a solubility test will provide useful information, if it controls variables well. Ask students to think about how they might conduct a solubility test on salt, Epsom salt, MSG, sugar, and the unknown. Have students work in groups to discuss their ideas and record a simple plan on their activity sheet.

5. Have student groups share their ideas for a solubility test and consider how each plan controls variables.



As each group presents their plans, have the class identify how each solubility test controls variables. All groups will likely suggest that they use the same volume of water at the same temperature in the same type of containers, and the same amount of each crystal. But there may be some disagreement in how to measure the same amount of each crystal. Some students may suggest they use the same volume of each crystal while others may suggest the same mass of each crystal. If no one suggests using mass, explain that in the demonstration, you used an equal mass of salt and sugar—5 g of each.

Ask students:

- **Is it better to use the same volume (like a teaspoon or 5 mL) or the same mass (like 5 g) of each crystal? Why?**

Tell students that you will do a demonstration that will help them see whether they should use a volume or mass measure so that they can dissolve the same amount of each crystal in water.

6. Do a demonstration to show that mass is better than volume in measuring equal amounts for a solubility test.

Question to Investigate

Is it better to measure the same volume or same mass of each crystal when conducting a solubility test to identify an unknown?

Materials for the Demonstration

- 2 clear plastic cups
- Cereal balls (Kix work well)
- Zip-closing plastic bag (quart-size storage-grade)
- Balance

Teacher Preparation

Fill two clear plastic cups completely with cereal balls. Both cups should be identical and contain the same amount of cereal balls. Test these cups on a balance to make sure that these cups have the same mass.

Procedure

1. Hold the cups filled with cereal up so that students can see that both have about the same amount of cereal in them.
2. Place the cups in the center of each end of a simple balance to prove to your students that both contain the same amount of cereal.

Ask students to make a prediction:

- **I am going to crush the cereal balls in one cup. Do you expect the height of cereal in this cup to be higher, lower, or the same as in the other cup?**

Students will probably say that the crushed cereal will not take up as much room in the cup.

3. Pour the cereal from one of the cups into a storage-grade, zip-closing plastic bag. Get as much air out as possible and seal the bag.
4. Place the bag on the ground, and crush the cereal thoroughly with your foot. Once the cereal is pulverized, open the bag, and pour the crushed cereal back into the cup.

Ask students:

- **Which cup contains more cereal?**

Students will realize that both cups contain the same amount of cereal, but some may have an urge to say that the cup with the cereal balls contains more cereal.

- **Was any cereal added or removed from either cup?**

Point out that even though the crushed cereal takes up less space, it is still the same amount of matter (cereal) as was in the cup before it was crushed.

- **How could you prove that these two cups contain the same amount of matter?**

Students should suggest placing the cups on a balance as you did before.



5. Place the cups on opposite ends of a balance to prove that the mass of cereal in each cup is the same.

Expected Results

Even though the volume of cereal balls is greater than the volume of crushed cereal balls, the cups will balance on the scale.

7. Relate student observations in the demonstration to the five crystals they will dissolve in water.

Tell students to imagine that the large cereal balls represent large crystals and the crushed cereal represents small crystals. Explain that the size and shape of the crystals may be different, but the balance shows that their mass is the same. Remind students that mass is a measure of the amount of matter. Because the large and small crystals (cereal) have the same mass, both cups contain the same amount of matter. Conclude that, in order to measure equal amounts, it is better to measure the mass of substances than the volume.

Ask students:

- **In the solubility test you will do, you will need to measure equal amounts of the five crystals. How will you measure equal amounts?**

After this demonstration, students should realize that measuring mass for a solubility test is better than measuring by volume.

Based on what students observed when they looked closely at the different crystals along with evidence from the demonstration, they should realize that different crystals have a slightly different size or shape. This will cause them to pack differently in the spoon so that more granules of one will be in the spoon than the other.

Note: The following explanation may be too difficult for students but is included here for you to think about and discuss with students if you think it is appropriate. Even if salt and sugar granules were exactly the same size and shape and packed exactly the same way in a spoon, it still would not be a good idea to use a teaspoon to measure equal amounts for a solubility test. Here's why: Salt is about 25% more dense than sugar. Therefore a teaspoon of salt weighs more than a teaspoon of sugar by almost 25%. Your dissolving test would not be accurate because you would be starting out with a larger mass of salt than sugar.

8. Have students weigh five grams of each of the crystals for the solubility test.

Materials for Each Group

- Gram balance
- 5 small plastic cups
- 5 clear plastic cups
- Masking tape and pen or permanent marker
- Salt
- Epsom salt
- MSG
- Sugar
- Unknown (coarse kosher salt)
- Water

Procedure

1. Use masking tape and a pen to label the 5 small plastic cups Salt, Epsom salt, MSG, Sugar, and Unknown.
2. Label the 5 larger clear plastic cups the same way.
3. Weigh 5 g of each crystal and place each in its small labeled cup.

If you do not have enough time, you can stop here and have students store the crystals and conduct the test another day. If you do have time to conduct the test, the procedure follows.

9. Have students dissolve the four known crystals and the unknown in room-temperature water.

The amount of water used in the procedure is specific and should be used because it gives clear results. Swirling the crystals in water is a good way of mixing them to help them dissolve. Lead the class so that all groups pour their crystal samples into the water at the same time. Also tell students when to swirl the water and crystals and when to stop and observe. There will be three 20-second intervals.

Question to Investigate

Can you identify an unknown using a solubility test?

Materials for Each Group

- Graduated cylinder
- 5 g each of salt, Epsom salt, MSG, sugar, and unknown (coarse kosher salt)
- 5 clear plastic cups
- Water

Procedure

1. Use a graduated cylinder to add 5 mL of room temperature water to each empty clear plastic cup.
2. Match up each pair of labeled cups so that each cup of crystal is near its corresponding cup of water.
3. When your teacher tells you to, work with your lab partners to pour the weighed amount of each crystal into its cup of water at the same time.
4. With the help of your lab partners, swirl each cup at the same time and in the same way for about 20 seconds and observe. Swirl again for another 20 seconds and observe. Swirl again for the last 20 seconds and make your final observations.
5. Slowly and carefully pour the solution from each cup back into its small empty cup. Try not to let any undissolved crystal go into the small cup. Compare the amount of crystal remaining in each clear plastic cup.



Expected Results

Results may vary. However, sugar should dissolve the most, followed by Epsom salt. MSG should appear to dissolve a bit more than salt and the unknown. The salt and the unknown should appear to dissolve to a similar degree.

10. Discuss student observations and the possible identity of the unknown.

Ask students:

- **Are there any crystals that you could rule out as probably not the unknown?**
Based on their observations, students are most likely to eliminate sugar and Epsom salt as the unknown.
- **Which cup or cups seem to have about the same amount of crystal left undissolved as the unknown?**
The unknown, salt, and MSG appear to have similar amounts of crystal that did not dissolve.
- **What do you think is the identity of the unknown?**
Students might conclude that the unknown is salt, but in some cases might think it could also be MSG.
- **What evidence do you have to support your conclusion?**
Students should cite the amount of crystal left behind in each cup as evidence that the unknown is either salt or MSG.
- **If someone in the class had a very different conclusion and had very different observations, what do you think may have led to these differences?**
Students should mention possible errors in weighing the crystals, in measuring the amount of water used, stirring in a different way, or accidentally pouring the crystals into the wrong cups.

Tell students that their test showed that different substances have different solubilities. In fact, solubility is a characteristic property of a substance. Explain to students that this type of solubility test can help eliminate some of the crystals, but may not be accurate enough to identify the unknown. Since they may have some doubt about the identity of the unknown, students will do a recrystallization test with the crystal solutions made during the solubility test.

11. Have students conduct another test to confirm the identity of the unknown.

Explain to students that they might be able to get more clues about the identity of the unknown if they allow the solutions of the dissolved crystals to recrystallize. Allowing the substances to recrystallize from their solutions might show similarities and differences that were not as easily seen in the original crystals.

***Materials notes:** The recrystallization test should be done immediately after the solubility test with the solutions made during the solubility test. Students will reuse the large clear plastic cups and solutions from the dissolving part of the activity, according to the procedure.*

Question to Investigate

Will the crystals that form when the solutions evaporate help identify the unknown?

Materials for Each Group

- Five solutions made in the activity, each in a small plastic cup
- 5 clear plastic cups from the activity
- Magnifier
- Water
- Paper towel

Procedure

1. Rinse each large, clear plastic cup with water to remove any remaining crystal. Dry each with a paper towel.
2. Carefully pour the solution from each small cup into its corresponding large, clear plastic cup.
3. Allow the solutions to sit overnight.
4. The next day, use a magnifier to carefully observe the crystals from both the top and bottom of the cup.

Expected Results

Salt and the unknown look very similar. Epsom salt, MSG, and sugar look different from each other and different from salt and the unknown. The sugar may not have recrystallized yet, but given more time it will form crystals.

Ask students:

- **Describe the crystals in each cup. What do you think is the identity of the unknown?**

Students should discuss the shape and size of the different crystals and notice that both salt and the unknown look very similar.

Tell students that the unknown is coarse kosher salt. It is chemically the same as regular salt, but the process for making ordinary table salt and kosher salt is different and this is why they look different.

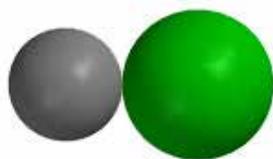
EXPLAIN

12. Show molecular models of salt, Epsom salt, sugar, and MSG.

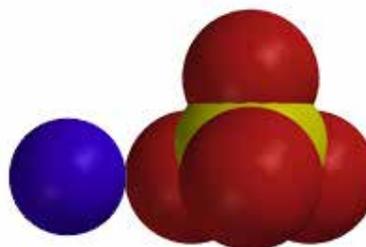
Project the image *All Four Crystals*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson5#all_four_crystals

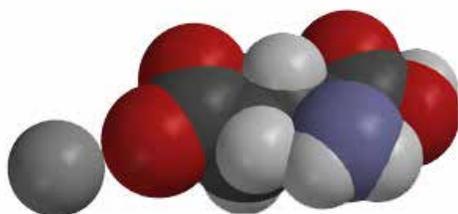




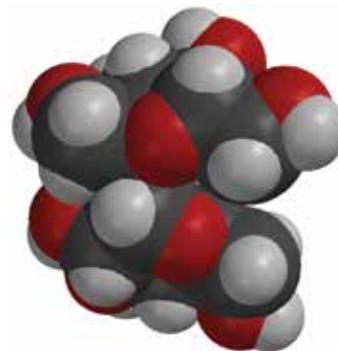
Sodium chloride



Magnesium sulfate



Monosodium glutamate



Sucrose

Explain that because these substances are made up of different atoms and ions bonded together differently, they interact with water differently, giving them each their own characteristic solubility.

- **Salt**

Remind students that sodium chloride is an ionic compound. There is a positive sodium ion (Na^+) and a negative chloride ion (Cl^-). Polar water interacts with these oppositely charged ions to dissolve the salt.

- **Epsom salt**

Tell students that Epsom salt is an ionic compound. There is a positive magnesium ion (Mg^{2+}) and a negative sulfate ion (SO_4^{2-}). Polar water interacts with these oppositely charged ions to dissolve the Epsom salt.

- **MSG**

MSG is made of a positive sodium ion (Na^+) and a negative glutamate ion, which has the molecular formula ($\text{C}_5\text{H}_8\text{NO}_4^-$). Polar water interacts with these oppositely charged ions to dissolve the MSG.

- **Sugar**

Sucrose is not an ionic compound. Sucrose has many O–H bonds, which give it positive and negative polar areas. These areas attract other sucrose molecules and hold them together in a crystal. These polar areas interact with water and cause entire sucrose molecules to separate from one another and dissolve.

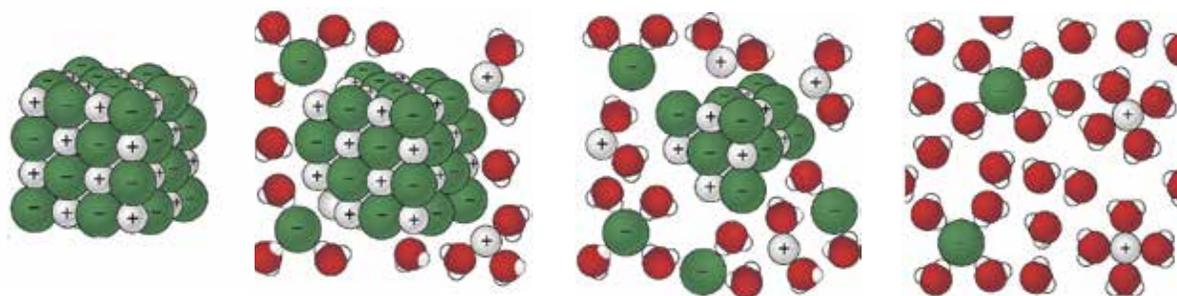
EXTEND

13. Help students review the similarities and differences in the way salt and sugar dissolve in water.

Tell students that depending on the substance being dissolved, ions are separated from each other, or molecules are separated from each other. Salt and sugar are common examples of dissolving both types of solids.

Project the image *Water Dissolves Salt.*

www.middleschoolchemistry.com/multimedia/chapter5/lesson3#sodium_chloride_dissolving



Ask students:

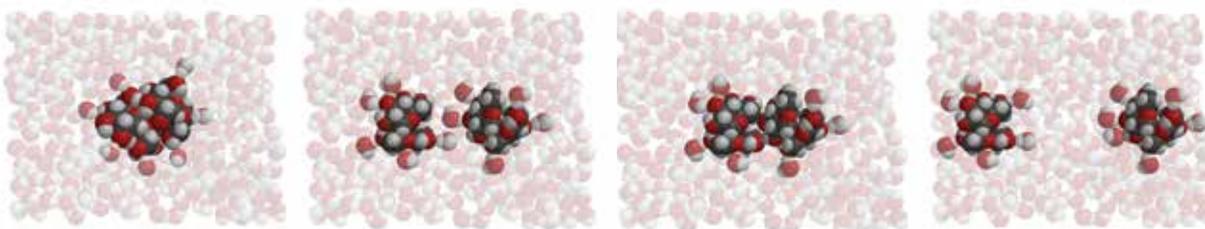
- **When salt dissolves, why are water molecules attracted to the sodium and chloride ions?**

Sodium chloride is an ionic compound with a positive sodium ion (Na^+) and a negative chloride ion (Cl^-). Polar water interacts with these oppositely charged ions to get it to dissolve.



Project the image *Water Dissolves Sugar.*

www.middleschoolchemistry.com/multimedia/chapter5/lesson4#water_dissolves_sucrose



Ask students:

- **When sugar dissolves, why are water molecules attracted to sucrose molecules?**
Sucrose has many O–H bonds, which give it positive and negative polar areas. These areas attract other sucrose molecules and hold them together in a crystal. These polar areas interact with water and cause entire sucrose molecules to separate from one another and dissolve.
- **What are the similarities and differences between water dissolving salt and water dissolving sugar?**
The sodium and chloride ions separate from one another and become surrounded by water molecules as they dissolve. Entire sucrose molecules separate from other sucrose molecules. The covalent bonds holding the atoms in the sucrose molecule do not come apart.

DEMONSTRATION

1. Your teacher did a demonstration comparing the amount of salt and sugar that dissolved in a small amount of water.

Was more salt or sugar left in the bottom of the cup?

Which dissolved better, salt or sugar?



2. How well a substance dissolves in water is called its solubility. Would you expect different substances to have the same or different solubility?

Why?

ACTIVITY

Question to Investigate

Can you identify the unknown crystal by the way it looks?



Materials for Each Group

- Black construction paper
- Masking tape
- Pen or permanent marker
- 5 small plastic cups
- Salt
- Epsom salt

- MSG
- Sugar
- Unknown
- Magnifier

Procedure

1. Use masking tape and a pen to label four corners of a piece of black construction paper Sugar, Salt, Epsom salt, and MSG. Label the center unknown.
2. Place small samples of Epsom salt, table salt, sugar, MSG, and the unknown on the labeled areas of the construction paper.
3. Use a magnifier to look carefully at each type of crystal.



3. What do you notice about each crystal? Include any similarities or differences you notice among them.
4. What do you think might be the identity of the unknown from what you have seen so far?

5. Your teacher did a demonstration with cereal balls. Look at the cups of cereal on the balance in the picture. Which cup contains more cereal?

Why?



6. In the solubility test you will do, you will need to measure equal amounts of the five crystals. How will you measure equal amounts?

GET READY FOR THE NEXT ACTIVITY



Materials for Each Group

- Gram balance
- 5 small plastic cups
- 5 clear plastic cups
- Masking tape and pen or permanent marker
- Salt
- Epsom salt
- MSG
- Sugar
- Unknown
- Water

Procedure

1. Use masking tape and a pen to label the 5 small plastic cups Salt, Epsom Salt, MSG, Sugar, and Unknown.
2. Label the 5 larger clear plastic cups the same way.
3. Weigh 5 g of each crystal and place each in its small labeled cup.

ACTIVITY

Question to Investigate

Can you identify an unknown using a solubility test?



Materials for Each Group

- Graduated cylinder
- 5 g each of salt, Epsom salt, MSG, sugar, and unknown
- 5 clear plastic cups
- Water

Procedure

1. Use a graduated cylinder to add 5 mL of room-temperature water to each empty clear plastic cup.
2. Match up each pair of labeled cups so that each cup of crystal is near its corresponding cup of water.
3. When your teacher tells you to, work with your lab partners to pour the weighed amount of each crystal into its cup of water at the same time.



4. With the help of your lab partners, swirl each cup at the same time and in the same way for about 20 seconds and observe. Swirl again for another 20 seconds and observe. Swirl again for the last 20 seconds and make your final observations.
5. Slowly and carefully pour the solution from each cup back into its small empty cup. Try not to let any undissolved crystal go into the small cup. Compare the amount of crystal remaining in each clear plastic cup.



What did you learn about the unknown from the solubility test?	
Unknown might be:	Unknown is probably not:
What evidence do you have to support your conclusion?	What evidence do you have to support your conclusion?

Question to Investigate

Will the crystals that form when the solutions evaporate help identify the unknown?



Materials for Each Group

- Five solutions made in the activity, each in a small plastic cup
- 5 clear plastic cups from the activity
- Magnifier
- Water
- Paper towel

Procedure

1. Rinse each large clear plastic cup with water to remove any remaining crystal. Dry each with a paper towel.
2. Carefully pour the solution from each small cup into its corresponding large clear plastic cup.
3. Allow the solutions to sit overnight.
4. The next day, use a magnifier to carefully observe the crystals from both the top and bottom of the cup.

Describe each type of re-crystallized crystal. Circle the one that you think is the unknown.	
Salt	Epsom Salt
MSG	Sugar

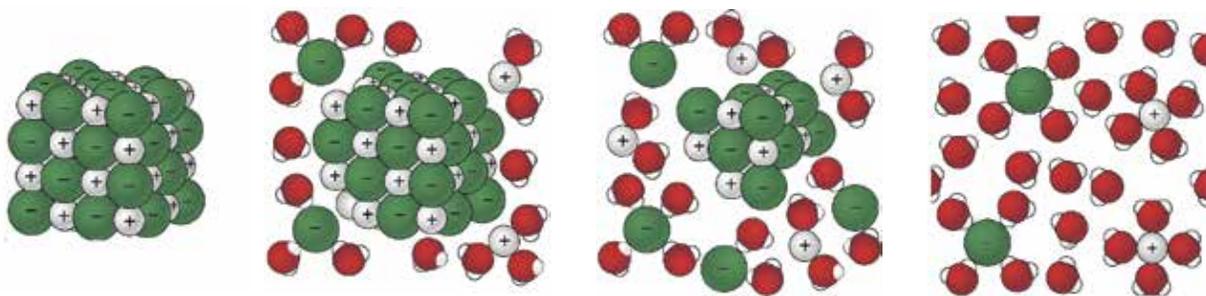
EXPLAIN IT WITH ATOMS & MOLECULES

7. On the molecular level, why do different substances have different solubilities?

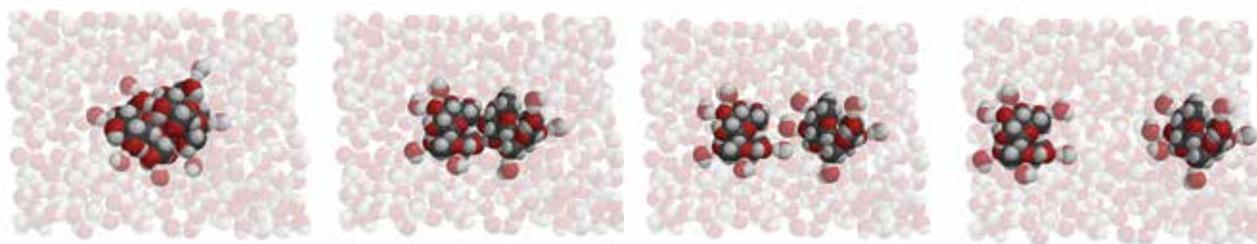
8. Why does the solubility test help you identify the unknown?

TAKE IT FURTHER

Sodium Chloride dissolving in water



Sucrose dissolving in water



9. What are the similarities and differences between salt dissolving in water and sugar dissolving in water?

Chapter 5, Lesson 6—Does Temperature Affect Dissolving?

Key Concepts

- Adding energy (heating) increases molecular motion.
- Increased molecular motion competes with the attraction between solute molecules and tends to make them come apart more easily.
- Increased molecular motion causes more solvent molecules to contact solute molecules and pull on them with more force, usually resulting in more dissolving.
- Since different substances are made from different atoms, ions, or molecules, increased temperature will affect their dissolving to different extents.

Summary

Students revisit the dissolving M&M activity from Lesson 4. They will design an experiment to see if temperature affects the amount of dissolving of the sugar coating of an M&M.

Objective

Students will be able to identify and control variables to design an experiment to see whether the temperature of a solvent affects the speed at which a solute dissolves. Students will be able to explain, on the molecular level, why increasing temperature increases the rate of dissolving.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Use caution when handling hot water.

Materials for Each Group

- M&M's (3 same color)
- 3 clear plastic cups
- 1 sheet of white paper
- Room-temperature water
- Hot water (about 50 °C)
- Cold water (about 5 °C)

Materials for the Demonstration

- 4 graduated cylinders (50 mL)
- Hot water (about 50 °C)
- Cold water (about 5 °C)
- Salt
- Sugar
- Tablespoon
- 2 small cups

ENGAGE

1. Have students work in groups to design an experiment to investigate whether the temperature of water affects the amount of M&M coating that dissolves.

Remind students of the experiment they did in Lesson 4 in which they placed an M&M in water and watched the colored sugar coating dissolve. In that experiment, students used water and then alcohol and oil to see if the solvent used affects the dissolving of the M&M coating.

Ask students:

- **What could you investigate about M&M's dissolving in water?**

If students do not suggest changing the temperature of the water, ask whether they think the temperature of the water affects the amount of coating that dissolves.

Give each student an activity sheet.

Students will describe their experimental design, record their observations, and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



Have student groups discuss the following questions:

- **How could you investigate whether the temperature of water affects the amount of coating that dissolves from an M&M?**
- **What are the variables in this experiment, and how will you control them?**

As you visit the groups and listen to their discussions, check to see if students are thinking about variables such as:

- Kind of container
- Amount of water
- Color of the M&M's
- When the M&M's are placed in the water
- Location of the M&M's in each cup

All these variables should be kept the same. Students should realize that the only variable that should be changed is the temperature of the water.

2. Have student groups share their ideas for comparing the effect of temperature on dissolving and consider how each plan controls variables.

As each group presents their plans, have the class identify how each plan controls variables. Some groups may have planned to test M&M's in hot and cold water but didn't consider using room-temperature water, too. Encourage all groups to test an M&M in all three temperatures of water. The room-temperature water serves as a control and can help students see the difference in how temperature affects dissolving.

EXPLORE

3. Have groups place an M&M in cold, room-temperature, and hot water at the same time to investigate the effect of temperature on dissolving.

Question to Investigate

Does the temperature of water affect the amount of coating that dissolves from an M&M?

Materials for Each Group

- 3 same-colored M&M's
- 3 clear plastic cups
- 1 sheet of white paper
- Room-temperature water
- Hot water (about 50 °C)
- Cold water (about 5 °C)



Procedure

1. Pour cold, room-temperature, and hot water into the cups so that the water is deep enough to cover an M&M.
2. Place the three cups on the white paper. Write Cold, Room-temp, and Hot near its cup.
3. With the help of your partners, place a same-colored M&M in the center of each cup at the same time. Observe for about 1 minute.

Expected Results

More color and sugar dissolve from the M&M in the hot water and less in the room-temperature and cold water. This means there is more chocolate visible on the M&M in the hot water than there is in the room-temperature and cold water. The color and sugar in the room-temperature water dissolve somewhere between the cold and hot water, but are more similar to the cold than the hot.

Note: There are actually two processes happening in this activity. The color and sugar are dis-

solving in the water but they are also diffusing in the water. The temperature of the water affects the amount of dissolving but it also affects the rate of diffusion. Students should focus on the surface of the M&M to judge the amount of color and sugar that dissolves.

4. Record and discuss student observations.

Ask students:

- **Does the temperature of the water affect the amount of colored coating that dissolves from an M&M? How do you know?**

Based on their observations, students should conclude that the hotter water causes more dissolving. Students may have noticed a greater difference in the amount of dissolving between the hot and the room-temperature water than between the room-temperature and the cold water. If no one comments on this, suggest that there is a difference.

EXPLAIN

5. Discuss how differences in molecular motion caused more of the sugar coating to dissolve in hot than in cold water.

Ask students:

- **What are the differences in the way water molecules move in cold, room-temperature, and hot water?**

Students should remember that water molecules move faster in hot water than in cold.

- **Why do you think sugar dissolves better in hot water than in cold water?**

The reason why sugar dissolves at a faster rate in hot water has to do with increased molecular motion. The added energy in the hot water causes water molecules to move faster and sucrose molecules to vibrate faster. This added movement tends to make the bonds between sucrose molecules easier to overcome. When faster-moving water molecules attach to sucrose molecules, a higher proportion of these sucrose–water interactions have enough energy to pull sucrose molecules away from other sucrose molecules, so the rate of dissolving increases.

- **Why do you think there is a greater difference in the amount of dissolving between the hot and room-temperature water than between the room-temperature and cold water?**

There is a greater difference in the rate of dissolving because there is a greater difference in temperature between hot and room-temperature water (about 30 degrees) than between room-temperature and cold water (about 15 degrees).

EXPLORE

6. Either do a demonstration or show a video to investigate whether all substances dissolve much better in hot water than in cold water.

Ask students to make a prediction:

- In the activity, you have seen that hot water dissolves sugar better than cold water. Do you think that salt will dissolve much better in hot water than in cold, like sugar does?

Project the image *Hot and Cold Water Dissolve Salt vs. Sugar*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson6#dissolving_temperature

If you choose to do the demonstration, follow the procedure below.

Question to Investigate

Will salt dissolve much better in hot water than in cold, like sugar does?

Materials for the Demonstration

- 4 graduated cylinders (50 or 100 mL)
- Hot water (about 50 °C)
- Cold water (about 5 °C)
- Salt
- Sugar
- Tablespoon
- 4 small cups

Teacher Preparation

- Label two small cups salt and two cups sugar.
- Place 1 tablespoon of salt and 1 tablespoon of sugar in each of their labeled cups.

Procedure

Salt in Hot and Cold Water

1. Place 25 mL of hot water and 25 mL of cold water in two separate graduated cylinders.
2. At the same time, pour one tablespoon of salt into each graduated cylinder. Do not swirl, shake, or stir. Set these graduated cylinders aside.

Sugar in Hot and Cold Water

- Place 25 mL of hot water and 25 mL of cold water in two separate graduated cylinders.
- At the same time, pour one tablespoon of sugar into each graduated cylinder. Do not swirl, shake, or stir. Set these graduated cylinders aside.



Compare the amounts of salt and sugar left undissolved

- Show students the graduated cylinders with the salt.
- Show students the graduated cylinders with the sugar.

Expected Results

Less sugar is visible in the hot water than in the cold, meaning that more sugar dissolves in the hot water than in the cold water. There is no obvious difference between the amount of salt that dissolves in the hot water compared to the cold water. This shows that temperature affects the dissolving of sugar more than it affects the dissolving of salt.



- Show students the table and graph for the solubility of sugar and salt at different temperatures.**

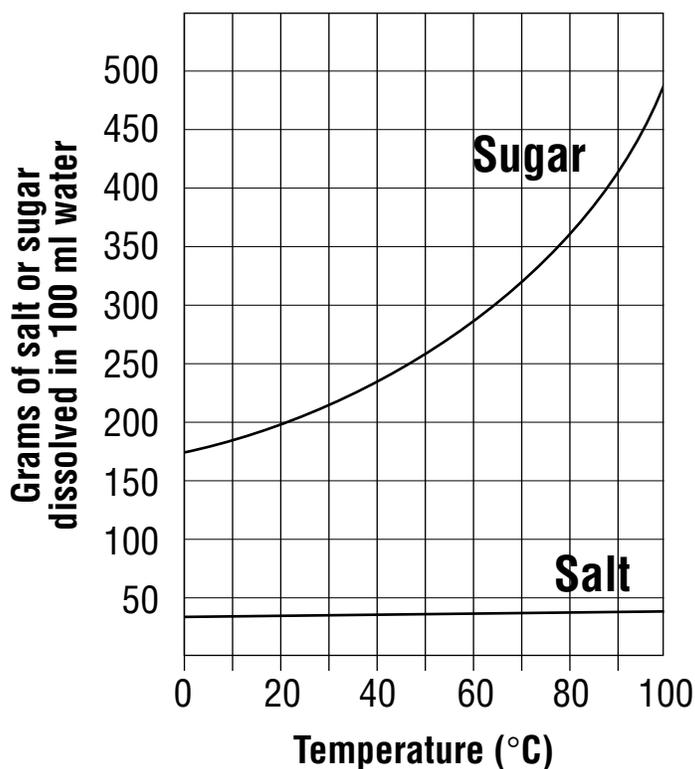
Project the graph *Solubility of Salt and Sugar*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson6#solubility_curve

Help students understand that the graph shows that more sugar dissolves in water as the temperature of the water increases. Also help them to see that the dissolving of salt also increases as the temperature of the water increases. But the dissolving of salt does not increase nearly as much as sugar.

Comparing the solubility of sodium chloride and sucrose Measured in grams of solute dissolved in 100 mL of water						
Temperature °C	0	20	40	60	80	100
Sodium chloride	35.5	36	36.5	37.5	38	39
Sucrose	179	204	241	288	363	487

Solubility of salt and sugar



Note: Students may ask why so much more sugar dissolves at higher temperatures compared to salt. This is not easy to explain on the molecular level at a middle school, high school, or even college level. Tell students that since substances are composed of different atoms, ions, and molecules, they are held together differently and interact with water differently. Changing temperature also affects the motion of the atoms, ions, or molecules of the substance and affects the interaction between water molecules and the particles of the substance. There are so many factors involved that it is difficult to explain why the solubility of one substance is affected more than another by an increase in temperature.

Ask students:

- **The demonstration showed that temperature affects the dissolving of sugar more than it affects the dissolving of salt. Explain how the graph shows this.**
As the temperature increases, more and more sugar can be dissolved, but only slightly more salt can be dissolved.
- **How much sugar dissolves in 100 mL of water at 50 °C?**
About 260 g of sugar will dissolve.
- **How much salt dissolves in 100 mL of water at 50 °C?**

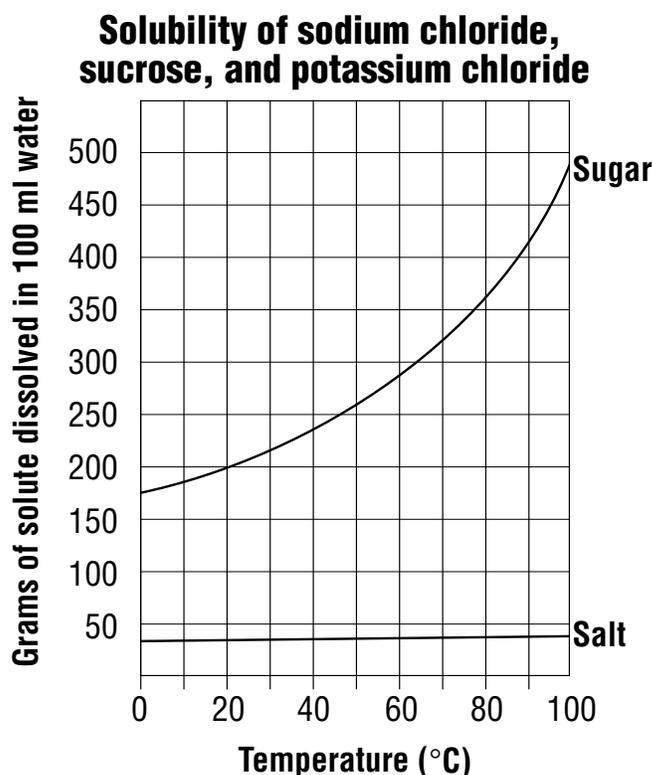
About 37 g of salt will dissolve.

EXTEND

8. Plot a solubility curve for potassium chloride and compare it to the solubility of sugar and salt.

Tell students that they will plot the solubility of potassium chloride on a graph provided on the activity sheet. Explain that potassium chloride is used as a salt substitute for people who should not eat regular salt (sodium chloride.)

The solubility of potassium chloride Measured in grams of solute dissolved in 100 mL of water						
Temperature °C	0	20	40	60	80	100
Potassium chloride	28	33	38	44	50	55



Ask students:

- At what temperature would you say that the solubility of sodium chloride and potassium chloride are about the same?
- At 0 °C, which substance is the least soluble?
- At 0 °C, which substance is the most soluble?

INTRODUCTION

1. Does the temperature of water affect the amount of coating that dissolves from an M&M? Talk with your group and design an experiment to find out.

List three variables and how you will control them.

What is the only variable that should be changed?

ACTIVITY

Question to Investigate

Does the temperature of water affect the amount of coating that dissolves from an M&M?

Materials for Each Group

- 3 same-colored M&Ms
- 3 clear plastic cups
- 1 sheet of white paper
- Room-temperature water
- Hot water (about 50 °C)
- Cold water (about 5 °C)

Procedure

1. Pour cold, room-temperature, and hot water into the cups so that the water is deep enough to cover an M&M.
2. Place the three cups on the white paper. Write *cold*, *room-temp*, and *hot* near its cup.
3. With the help of your partners, place a same-colored M&M in the center of each cup at the same time. Observe for about 1 minute.



2. Does the temperature of the water affect the amount of coating that dissolves from an M&M?

How do you know?

EXPLAIN IT WITH ATOMS & MOLECULES

3. What are the differences in the way water molecules move in cold, room-temperature, and hot water?
4. On the molecular level, why do you think sugar dissolves better in hot water than in cold water?

5. Why do you think there is a greater difference in the amount of dissolving between the *hot and room-temperature water* than between the *room-temperature and cold water*?

DEMONSTRATION

6. Your teacher showed you an activity comparing the amount of salt that can dissolve in hot and cold water and the amount of sugar that can dissolve in hot and cold water. Just like in your M&M experiment, much more sugar dissolved in hot water. Does much more salt dissolve in hot water than in cold?

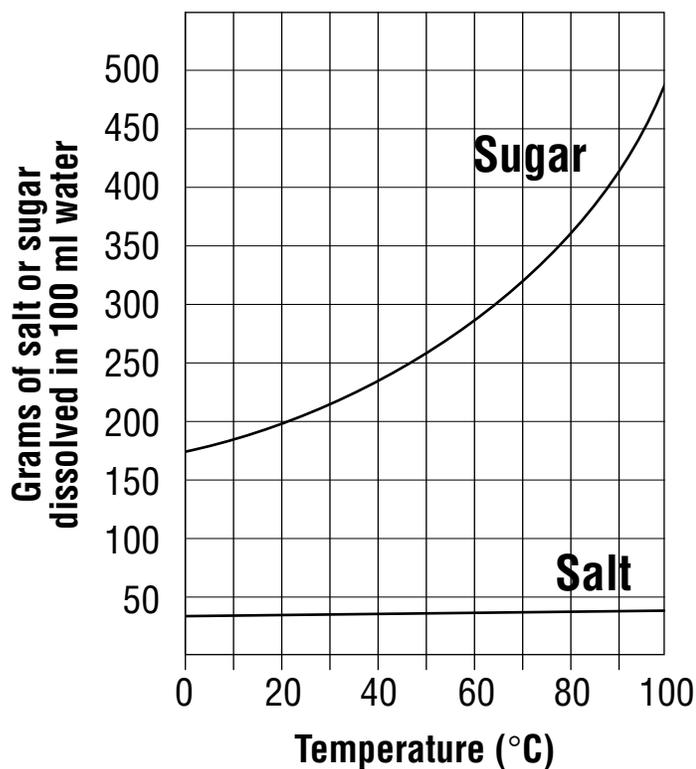
How do you know?



7. The following data table and graph compare how much sugar and salt can dissolve in water over the temperature range 0 °C to 100 °C. Use the table and graph to answer the following questions.

Comparing the solubility of sodium chloride and sucrose						
Measured in grams of solute dissolved in 100 mL of water						
Temperature °C	0	20	40	60	80	100
Sodium chloride	35.5	36	36.5	37.5	38	39
Sucrose	179	204	241	288	363	487

Solubility of salt and sugar



The activity showed that temperature affects the dissolving of sugar more than it affects the dissolving of salt. Explain how the graph shows this.

How much sugar dissolves in 100 mL of water at 50 °C?

How much salt dissolves in 100 mL of water at 50 °C?

TAKE IT FURTHER

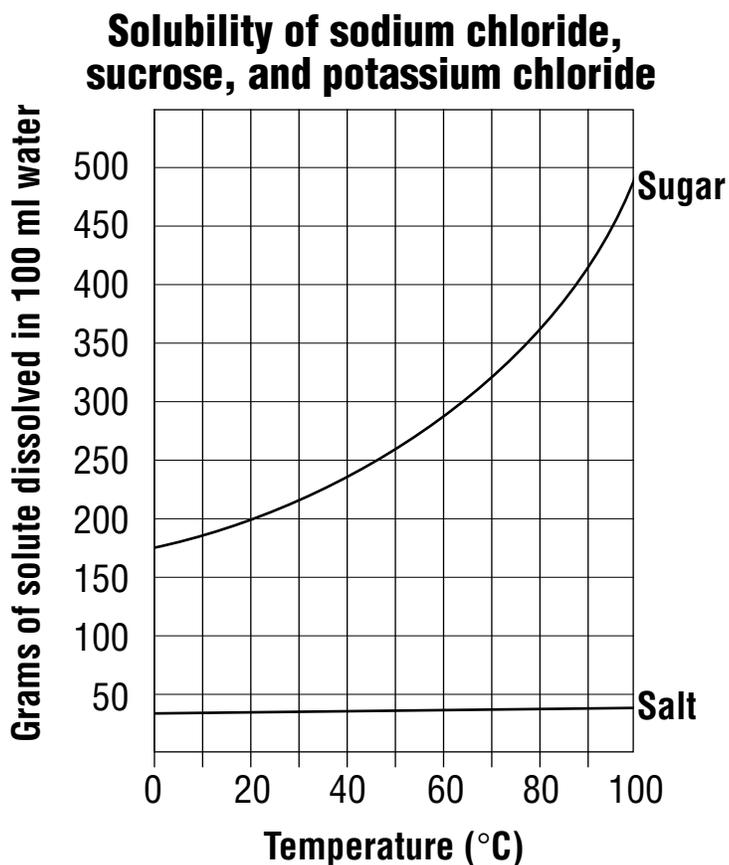
8. Potassium chloride is a salt substitute sold in grocery stores for people who should limit their intake of table salt (sodium chloride). Use the data table to plot the solubility curve for potassium chloride on the graph below. Then use the graph to answer the following questions.

The solubility of potassium chloride Measured in grams of solute dissolved in 100 mL of water						
Temperature °C	0	20	40	60	80	100
Potassium chloride	28	33	38	44	50	55

At what temperature would you say that the solubility of sodium chloride and potassium chloride are about the same?

At 0 °C, which substance is the least soluble?

At 0 °C, which substance is the most soluble?



Chapter 5, Lesson 7—Can Liquids Dissolve in Water?

Key Concepts

- Liquids have characteristic properties based on the molecules they are made of.
- The properties of liquids depend on the attractions the molecules of the liquid have for each other and for other substances.
- Liquids can dissolve certain other liquids, depending on the attractions between the molecules of both liquids.
- Polar liquids, like water, dissolve other liquids which are polar or somewhat polar.
- Polar liquids, like water, do not dissolve nonpolar liquids like oil.

Summary

Students will place isopropyl alcohol, mineral oil, and corn syrup in water to see if any of these liquids dissolve in water. Students will extend their understanding and definition of “dissolving” and see that certain, but not all, liquids can dissolve in water.

Objective

Students will identify and control variables to help design a solubility test for different liquids in water. Students will be able to explain, on the molecular level, why certain liquids, but not all, will dissolve in water. They will also be able to explain that the solubility of a liquid is a characteristic property of that liquid.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Isopropyl alcohol is flammable. Keep it away from flames or spark sources. Read and follow all warnings on the label. Isopropyl alcohol and mineral oil should be disposed of according to local regulations. Have students wash hands after the activity.

Materials for the Demonstrations

- Clear plastic cup
- Water
- Food coloring
- Straw or popsicle stick
- Isopropyl alcohol (70% or higher)
- 2 identical 100-mL graduated cylinders

Materials for Each Group

- Water
- Mineral oil
- Isopropyl alcohol (70% or higher)
- Corn syrup
- 3 clear plastic cups
- 5 small cups
- Permanent marker or masking tape and a pen for labeling cups
- 3 straws or popsicle sticks (for stirring)
- Laminated index card or card covered with wax paper
- Blue water
- Yellow isopropyl alcohol (70% or higher)
- 2 droppers
- Toothpick or popsicle stick

ENGAGE

1. Do a demonstration to introduce the idea that solids aren't the only substances that can dissolve—liquids can also dissolve in liquids.

Ask students to make a prediction:

- Solids, like salt or sugar, can dissolve in water. Do you think that liquids can dissolve in water?



Question to Investigate

Does liquid food coloring dissolve in water?

Materials for the Demonstration

- Clear plastic cup
- Water
- Food coloring
- Straw or popsicle stick

Note: A similar demonstration was conducted back in Chapter 1 as evidence that molecules are in motion. Here, the demonstration is used to show the attraction between water molecules and the liquid coloring molecules that allows the color to dissolve in the water.

Procedure

1. Hold up a clear plastic cup or other clear container of room-temperature water. Add 1 or 2 drops of food coloring and allow the coloring to drift and spread in the water a bit.
2. Stir with a straw or popsicle stick.



Expected results

The food coloring will drift and slowly mix throughout the water. When stirred, the water will be evenly colored throughout.

Ask students:

- **Does food coloring dissolve in water?**
Yes.
- **How do you know when a solute, like food coloring, has dissolved in a solvent, like water?**

As part of the answer to this question, review the definition of dissolving using the food coloring and water as an example. The solute (food coloring) is dissolved in the solvent (water) when the molecules of the solute are so thoroughly intermixed within the molecules of the solvent that they do not settle out or separate.

- **This demonstration showed that food coloring can dissolve in water. Describe an experiment you could do to compare how isopropyl alcohol, mineral oil, and corn syrup dissolve in water.**

Students should agree that they will need three cups filled with the same amount of water. They should also realize that it's important that the same temperature water is used in each cup and that the temperature of each of the three liquids being tested is the same, too.

Give each student an activity sheet.

Students will describe their experimental design, record their observations, and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Have students conduct an activity to see how well isopropyl alcohol, mineral oil, and corn syrup dissolve in water.

Question to Investigate

Do isopropyl alcohol, mineral oil, and corn syrup dissolve in water?

Teacher Preparation for Each Group

- Label 3 small cups Alcohol, Oil, and Syrup for each group.
- Pour about 1 tablespoon of isopropyl alcohol, mineral oil, and corn syrup into their labeled cups.

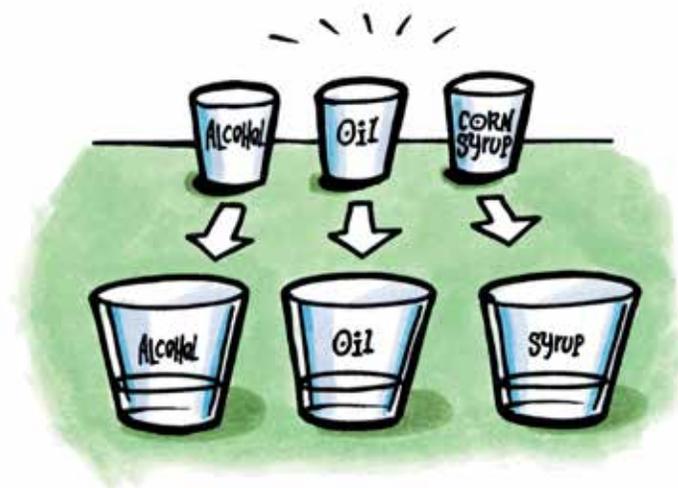


Materials for Each Group

- Water
- Isopropyl alcohol (70% or higher) in small labeled cup
- Mineral oil in small labeled cup
- Corn syrup in small labeled cup
- 3 clear plastic cups
- Permanent marker or masking tape and a pen for labeling cups
- 3 straws or popsicle sticks for stirring

Procedure

1. Label 3 clear plastic cups Alcohol, Oil, and Syrup.
2. Pour water into all three labeled cups until each is about half-full.
3. While looking at the water from the side, slowly pour the alcohol into its labeled cup.
4. Without stirring, watch to see if the alcohol dissolves in the water on its own. Record your observations in the chart.
5. After waiting about 10 seconds, stir to see if the alcohol dissolves. Record your observations.
6. Repeat Steps 2–5 for oil and corn syrup.



Expected Results

- The alcohol looks kind of gray and swirly as it goes into the water. The alcohol tends to stay on the surface of the water because it is less dense than water. It does not seem to dissolve immediately but dissolves when stirred.
- The oil stays on the surface of the water because it is less dense than water but it does not appear to mix much at all with the water. When stirred, the oil breaks apart a bit and then forms a layer again on the surface of the water. The oil does not dissolve.
- The corn syrup sinks in the water because it is more dense than water. It seems to stay there without much initial dissolving. After stirring, the corn syrup dissolves into the water and the solution turns clear.

3. Discuss student observations.

Have students describe what the alcohol, oil, and corn syrup looked like in the water and whether or not they dissolved.

Ask students:

- **Can a liquid dissolve another liquid?**
Students should realize that some liquids, but not all, can dissolve in water.
- **Based on your observations of the way isopropyl alcohol, mineral oil, and corn syrup dissolve in water, would you say that solubility is a characteristic property of a liquid? Why?**
Yes. Solubility is a characteristic property because each liquid interacted with the water differently.

EXPLAIN

4. Discuss how the molecular structure of isopropyl alcohol, mineral oil, and glucose (in corn syrup) determines whether or not each liquid will dissolve in water.

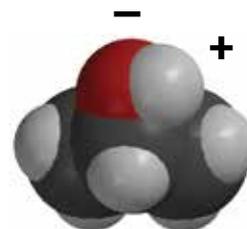
Project the image *Isopropyl Alcohol*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson7#isopropyl_alcohol

Point out the bond between oxygen and hydrogen in one area of the alcohol molecule.

Ask students:

- **What do you already know about the O–H bond?**
It is polar. The oxygen has a slight negative charge, and the hydrogen has a slight positive charge.



- **How do you think this polar part of the molecule affects the solubility of alcohol?**

Even though alcohol has one polar area (O–H bond) and a larger nonpolar area (C–H bonds), polar water molecules and the polar area on alcohol molecules are attracted to each other, causing alcohol to dissolve in water.

Project the image *Mineral Oil*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson7#mineral_oil

Remind students that the carbon (darker gray) and hydrogen atoms share electrons rather evenly. This means that the bonds in mineral oil are nonpolar, so water molecules and mineral oil molecules are not attracted to each other.

- **Why do you think the oil does not dissolve in water?**

The mineral oil molecule is made of carbon atoms bonded to hydrogen atoms. The bond between these atoms creates very little polarity. Water is not very attracted to the oil and so does not dissolve it.



- **In some salad dressings a layer of oil, like canola or olive oil, floats on top of a layer of vinegar, which is mostly water. If you shake a bottle of this kind of salad dressing, the liquids will temporarily combine. But the oil and vinegar do not dissolve in one another because eventually the two liquids will separate out again. Knowing what you do about molecules and dissolving, why doesn't the oil in these salad dressings dissolve in vinegar?**

Oil is nonpolar and is not attracted to the water in vinegar, so it will not dissolve.

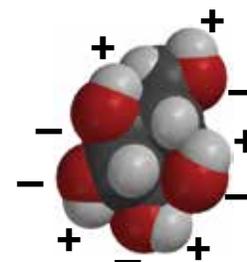
Note: Students should understand that polar molecules, like water, attract other polar molecules but they do not attract nonpolar molecules, like oil. At the middle school level, this understanding is sufficient but could lead to the impression that nonpolar molecules have no attractions at all. This is not true.

Read more about the attractions between nonpolar molecules in the additional teacher background section at the end of this lesson.

Project the image *Glucose*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson7#glucose

Explain to students that corn syrup is mostly glucose but also contains a similar sugar, fructose. Show students the glucose molecule and point out the bonds between oxygen and hydrogen.



Note: Students may have heard of “high fructose corn syrup” used in carbonated beverages and some prepared foods. This type of corn syrup contains a higher percentage of fructose than the corn syrup commonly sold in the baking aisle at grocery stores.

Ask students:

- **Why do you think glucose molecules dissolve well in water?**

Glucose has many areas where oxygen is bonded to hydrogen. These O–H bonds are polar. Polar water molecules and the polar areas of glucose molecules are attracted to each other, causing the corn syrup to dissolve.

- **Some people with diabetes may accidentally let their sugar level get too low. There are glucose tablets to help them with this problem. When a person eats one, do you think it will act quickly to increase his/her blood sugar level? Why or why not?**

Yes. The tablet will act quickly because the water in a person’s saliva and stomach will easily dissolve the glucose.

EXTEND

5. Look more closely at the way water and alcohol mix.

Water and alcohol do some pretty interesting things when they mix. Tell students that you colored water blue and isopropyl alcohol yellow so that they can see the mixing better.

Question to Investigate

What happens as drops of water and alcohol combine?

Teacher Preparation

- Either laminate an index card for each group or cover an index card with a piece of wax paper and secure the wax paper with tape.
- Place about $\frac{1}{4}$ cup water in 1 cup
- Add 2 drops of blue food coloring to the water.
- Place $\frac{1}{4}$ cup isopropyl alcohol in another cup.
- Add 2 drops of yellow color food coloring to the alcohol.
- Label 2 small cups water and alcohol for each group. You may reuse the small cup labeled alcohol from the activity in this lesson.
- Place about 1 teaspoon of each colored liquid into its pair of labeled cups for each group.

Materials for Each Group

- Water (colored blue)
- Isopropyl alcohol (70% or higher and colored yellow)

- Laminated index card or card covered with wax paper
- 2 droppers
- Toothpick or popsicle stick

Procedure

1. Use a dropper to place about 5 drops of blue water together to make 1 large drop on your index card.
2. Use another dropper to make a similar large drop of yellow alcohol close to, but not touching, the blue drop.
3. Use a toothpick to drag the blue water toward the yellow alcohol until they touch. As soon as the drops touch, lift the toothpick away and do not stir.
4. Watch closely as the alcohol and water mix.



Expected Results

The alcohol and water will kind of “shake” or “jiggle” right at the area where they are mixing.

6. Discuss student observations.

Ask students:

- **What do you observe when the drop of alcohol and drop of water combine?**
Students will notice that the alcohol and water seem to shake at the area where they are mixing. As the liquids mix, the yellow and blue colors combine to make green.

Tell students that the mixing of alcohol and water is not completely understood on the molecular level. One reason for the shaky appearance might be that alcohol is less dense than water so as it mixes with water, the density of the overall liquid changes. Alcohol (which alone floats on water) sinks as it mixes. Maybe lots of little “sinkings” make the mixing look shaky. It could be that the changing density causes light to refract differently to cause the shaky look. Maybe the alcohol interferes with water’s surface tension and causes a shaky look on the surface. It is also true that when water and alcohol are mixed, the solution gets warmer. Maybe the heat increases molecular motion at the surface, which somehow contributes to the shaky look.

7. Do a demonstration to show that when water and alcohol combine, the volume of the resulting solution is less than expected.

Tell students that the characteristic way water and alcohol interact with each other causes another interesting phenomenon.

Materials for the Demonstration

- Isopropyl alcohol (90% or higher)
- Water
- 2 identical 100-mL graduated cylinders



Procedure

1. Measure 50 mL of isopropyl alcohol and pour it into a 100-mL graduated cylinder.
2. Measure 50 mL of water and add it to the alcohol in the 100-mL graduated cylinder.

Expected Results

The total volume of the liquid will be about 97 or 98 mL. This is surprising because 50 mL of water + 50 mL of water equals 100 mL.

Explain to students that when the water and alcohol molecules interact, they rearrange and actually take up less room than if you add up their individual volumes.



Note: You may have heard the explanation that the water molecules are in the spaces between the alcohol molecules or the alcohol molecules are in the spaces between the water molecules. This is too passive an explanation. It's not like marbles falling in the spaces between golf balls. There is an active rearranging of molecules that are attracted to one another that results in the final volume. In fact, adding any two liquids that can dissolve in one another will result in a volume that's different from the sum of the separate liquid volumes. Alcohol and water are often used as an example of this phenomenon because they are easy to get and show a particularly big difference.

DEMONSTRATION

1. Your teacher placed some food coloring in water. Did the food coloring dissolve in the water?

How do you know when a *solute*, like food coloring, has dissolved in a *solvent*, like water?



ACTIVITY

Question to Investigate

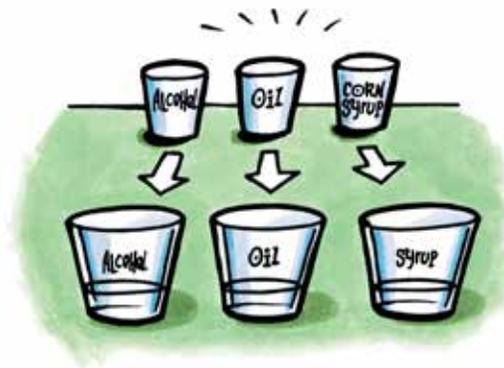
Do isopropyl alcohol, mineral oil, and corn syrup dissolve in water?

Materials for Each Group

- Water
- Isopropyl alcohol (70% or higher) in small labeled cup
- Mineral oil in small labeled cup
- Corn syrup in small labeled cup
- 3 clear plastic cups
- Permanent marker or masking tape and a pen for labeling cups
- 3 straws or popsicle sticks for stirring

Procedure

1. Label 3 clear plastic cups Alcohol, Oil, and Syrup.
2. Pour water into all three labeled cups until each is about half-full.



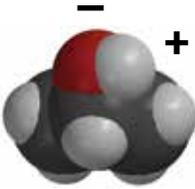
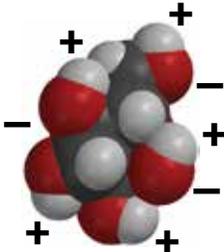
3. While looking at the water from the side, slowly pour the alcohol into its labeled cup.
4. Without stirring, watch to see if the alcohol dissolves in the water on its own. Record your observations in the chart.
5. After waiting about 10 seconds, stir to see if the alcohol dissolves. Record your observations.
6. Repeat Steps 2–5 for oil and corn syrup.

Do alcohol, oil, and syrup dissolve in water?			
Liquid	What does this liquid do when you first place it in water?	Does this liquid dissolve in water after you stir?	How do you know whether this liquid dissolves or does not dissolve in water?
Isopropyl alcohol			
Mineral oil			
Corn syrup			

2. Based on your observations of the way isopropyl alcohol, mineral oil and corn syrup dissolve in water, would you say that solubility is a characteristic property of a liquid? Why?

EXPLAIN IT WITH ATOMS & MOLECULES

Look at the structure of the molecules in isopropyl alcohol, corn syrup, and mineral oil. Explain why either dissolves or does not dissolve in water.

<p>Isopropyl alcohol</p> 	
<p>Mineral oil</p> 	
<p>Glucose in corn syrup</p> 	

3. In some salad dressings a layer of oil, like canola or olive oil, floats on top of a layer of vinegar, which is mostly water. If you shake the bottle of salad dressing, the liquids will temporarily combine. The oil and vinegar do not dissolve in one another because eventually the two liquids will separate out again. Knowing what you do about molecules and dissolving, why doesn't the oil in these salad dressings dissolve in vinegar?

4. Some people with diabetes may accidentally let their sugar level get too low. There are glucose tablets to help them with this problem. When a person eats one, do you think it will act quickly to increase his/her blood sugar level? Why or why not?

TAKE IT FURTHER

Question to Investigate

What happens as drops of water and alcohol combine?

Materials for Each Group

- Water (colored blue)
- Isopropyl alcohol (70% or higher and colored yellow)
- Laminated index card or card covered with wax paper
- 2 droppers
- Toothpick or popsicle stick

Procedure

1. Use a dropper to place about 5 drops of blue water together to make 1 large drop on your index card.
2. Use another dropper to make a similar large drop of yellow alcohol close to, but not touching, the blue drop.
3. Use a toothpick to drag the blue water toward the yellow alcohol until they touch. As soon as the drops touch, lift the toothpick away and do not stir.
4. Watch closely as the alcohol and water mix.



5. What do you observe when the drop of alcohol and drop of water combine?

6. Your teacher combined 50 mL of isopropyl alcohol and 50 mL of water. What is surprising about the result?



Chapter 5, Lesson 8—Can Gases Dissolve in Water?

Key Concepts

- Gases can dissolve in water.
- The dissolving of a gas in water depends on the interaction between the molecules of the gas and the water molecules.
- The amount of gas that can be dissolved in water depends on the temperature of the water.
- More gas can dissolve in cold water than in hot water.

Summary

Students will observe the dissolved carbon dioxide (CO_2) in a bottle of club soda. They will help design an experiment to compare the amount of CO_2 that stays in cold club soda compared to warmer club soda.

Objective

Students will be able to explain, on the molecular level, how a gas dissolves in water. They will also be able to explain why the gas comes out of solution faster in warm water than in cold water.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Warn students not to eat the M&M. Use caution when handling hot water.

Materials for the Demonstration

Unopened 1-liter bottle of club soda, which will also be used in the activity.

Materials for each group

- Club soda
- 3 clear plastic cups
- 1 M&M
- Pipe cleaner
- Hot water
- Cold water
- 2 deli containers (that cups easily fit inside)

ENGAGE

1. Show students the bubbles that appear when a new bottle of soda is opened.

Remind students that they have seen that some solids and liquids can dissolve in water (Chapter 5, Lessons 5 and 7).

Ask students:

- **Do you think that gases can dissolve in water?**
The idea of a gas dissolving may seem strange to students, but this demonstration will help them realize that gases can dissolve in water.

Materials

Unopened 1-liter bottle of club soda



Teacher Preparation

Remove the label from a 1-liter bottle of carbonated water.

Ask students:

- **How is a bottle of carbonated water different from a regular bottle of water?**
Students will probably say that carbonated water has bubbles.
- **Do you see any bubbles in the carbonated water?**
They shouldn't see any, yet.

Procedure

1. Very slowly unscrew the bottle cap.
2. Wait a few seconds to allow students to observe the bubbles.
3. Tighten the cap on the bottle.



Expected Results

When the cap is loosened, many bubbles will appear throughout the soda and rise up through the water to the surface, where they pop. When the cap is tightened, fewer bubbles will form.

Ask students:

- **What did you observe when I opened and then closed the bottle of soda?**
Bubbles appeared only when the bottle was opened. The bubbles stopped forming when the bottle cap was tightened.
- **What is the gas that makes these bubbles?**
Carbon dioxide (CO_2)
- **Where was the CO_2 before the bottle was opened?**
The carbon dioxide was dissolved in the water.

4. Explain that carbonated water is made of carbon dioxide gas dissolved in water.

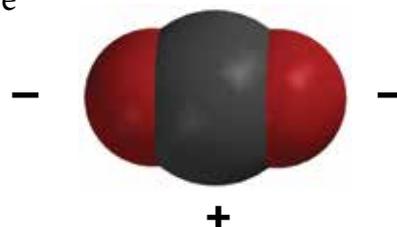
Tell students that at a soda factory, carbon dioxide gas is added to cold water under high pressure to make carbonated water. The pressure forces more gas to dissolve than ordinarily would.



Project the image CO₂ Molecule.

www.middleschoolchemistry.com/multimedia/chapter5/lesson8#carbon_dioxide +

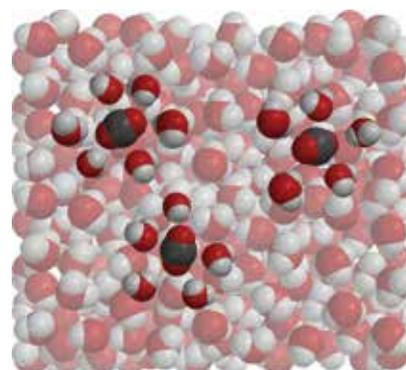
Point out that a molecule of carbon dioxide has a slight negative charge near the oxygen and a slight positive charge near the carbon. CO₂ is soluble because water molecules are attracted to these polar areas. The bond between carbon and oxygen is not as polar as the bond between hydrogen and oxygen, but it is polar enough that carbon dioxide can dissolve in water.



Project the image CO₂ Dissolved in Water.

www.middleschoolchemistry.com/multimedia/chapter5/lesson8#carbon_dioxide_dissolved_in_water

Explain that in carbonated water, molecules of carbon dioxide are thoroughly mixed and dissolved in water. This is similar to molecules of sucrose, sodium and chloride ions from salt, or the molecules of isopropyl alcohol, which students dissolved in water in previous activities in this chapter. Point out that when dissolved, the molecules of CO₂ are not like tiny little bubbles of gas mixed in the water. Instead single molecules of CO₂ are surrounded by water molecules.



Let students know that although the CO₂ dissolves, the molecules are not attracted as strongly by the water molecules as substances like salt or sugar. Due to these weaker attractions, the molecules of CO₂ come out of solution relatively easily. This is why soda becomes flat if it is left uncapped for too long.

Give each student an activity sheet.

Students will describe their experimental design, record their observations, and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

5. Have students add objects to carbonated water to see if they can get carbon dioxide gas to come out of solution.

Ask students:

- **Aside from shaking soda, or leaving it uncovered, are there other ways to make carbon dioxide gas come out of carbonated water?**

Tell students that objects can be placed in the soda that can cause the carbon dioxide to bubble out of the soda.

Question to Investigate

How can you make carbon dioxide gas come out of solution?



Materials for Each Group

- Club soda in clear plastic cup
- 2 clear plastic cups
- M&M
- Pipe cleaner

Teacher Preparation

Immediately before the activity, use the bottle of carbonated water from the demonstration to pour about $\frac{3}{4}$ cup of carbonated water into a clear plastic cup for each group.

Procedure

1. Evenly divide the club soda among the 3 clear plastic cups. Push two of these cups aside to use later.
2. Place a pipe cleaner in the soda and observe.
3. Place an M&M in the soda and observe.



Expected Results

Bubbles form on the pipe cleaner. Bubbles also form on the M&M and rise to the surface.

Ask students:

- **Where did the gas bubbles that you observed come from?**
There were molecules of carbon dioxide dissolved in the water.
- **Where did the carbon dioxide gas that was dissolved in the water go?**
The carbon dioxide bubbles rose to the surface and popped, releasing carbon dioxide into the air.

Explain that the objects placed in the soda had tiny bumpy areas where the carbon dioxide molecules collected. When enough of the molecules were together in a certain area, they became a bubble. When this bubble, which is less dense than the water around it, became big enough, it floated to the surface and popped, releasing carbon dioxide gas into the air.

Ask students:

- **While drinking soda pop with a straw, you may have noticed that bubbles form on the outside of the straw. Now that you have done this activity, why do you think these bubbles form on the straw?**

Even though the straw looks smooth, it also has tiny bumpy areas, where molecules of carbon dioxide collect. When enough of them collect in an area, they become a bubble of carbon dioxide gas.

4. Discuss how to find out whether temperature affects how quickly gas escapes from carbonated water.

Ask students:

- **Would you expect carbon dioxide to stay dissolved better in hot or cold water?**
Hint: Soda pop is kept in the refrigerator after it is opened.

Students will discover the answer to this question as they do the following activity.

- **How could you set up an experiment to find out whether carbon dioxide stays dissolved better in water that is warmed or water that is cooled?**

Students should realize that they will need two cups of carbonated water. They will then need to heat one and cool the other. One simple way to heat and cool the cups is to use a hot or cold water bath like the one described in the procedure.

- **If more carbon dioxide gas escapes from the soda, will there be more or fewer bubbles popping at the surface?**

There will be more bubbles popping at the surface.

- **If more carbon dioxide gas stays dissolved, will there be more or fewer bubbles popping at the surface?**

- There will be fewer bubbles popping at the surface.

Conclude that students can compare how fast carbon dioxide gas escapes or stays in solution by comparing the quantity of bubbles they see rise to the surface and pop. More bubbles rising and popping means more gas is escaping from the solution. Fewer bubbles rising and popping means more gas is remaining in the solution.

5. Have students warm and cool 2 cups of carbonated water to find out if temperature affects the solubility of carbon dioxide.

Question to Investigate

Does carbon dioxide stay dissolved better in water that is warmed or water that is cooled?

Materials for Each Group

- Carbonated water in 2 clear plastic cups
- Hot water (about 50 °C)
- Cold water (about 5 °C)
- 2 deli containers (that cups easily fit in)

Procedure

1. Get the two cups of carbonated water that you moved aside earlier.
2. Fill one empty deli container about $\frac{1}{3}$ of the way with ice cold water and another about $\frac{1}{3}$ of the way with hot tap water.
3. Place each of the cups of carbonated water into the cold and hot water, as shown.
4. Watch the surface of the soda in each cup of carbonated water.

Expected Results

More bubbles form and rise to the surface in the carbonated water that is placed in hot water.



6. Discuss student observations.

Students should realize that dissolved gas comes out of solution faster when soda is warm than when it is cold. The reverse is also true: Dissolved gas tends to stay dissolved better in cold soda.

Ask students:

- **Does carbon dioxide stay dissolved better in hot water or in cold water?**
Carbon dioxide stays dissolved better in cold water.
How do you know?
More gas escapes from the soda placed in hot water.
- **Based on what you observed in this experiment, why do you think people store soda pop that has been opened in the refrigerator?**
Since less carbon dioxide escapes when soda is colder, keeping soda in the refrigerator prevents the soda from going flat.

EXPLAIN

7. Explain why carbon dioxide gas escapes from hot water faster than it does from cold water.

Remind students that carbon dioxide gas is ready to come out of carbonated water no matter the temperature of the water. This is because gas molecules and water molecules are somewhat attracted to each other, but not very much. All you have to do is leave a bottle of soda pop open and the carbon dioxide will come out on its own, making your soda pop flat. Warming the soda increases the motion of the water and carbon dioxide molecules, making their attachments even looser and allowing the gas to escape even faster.

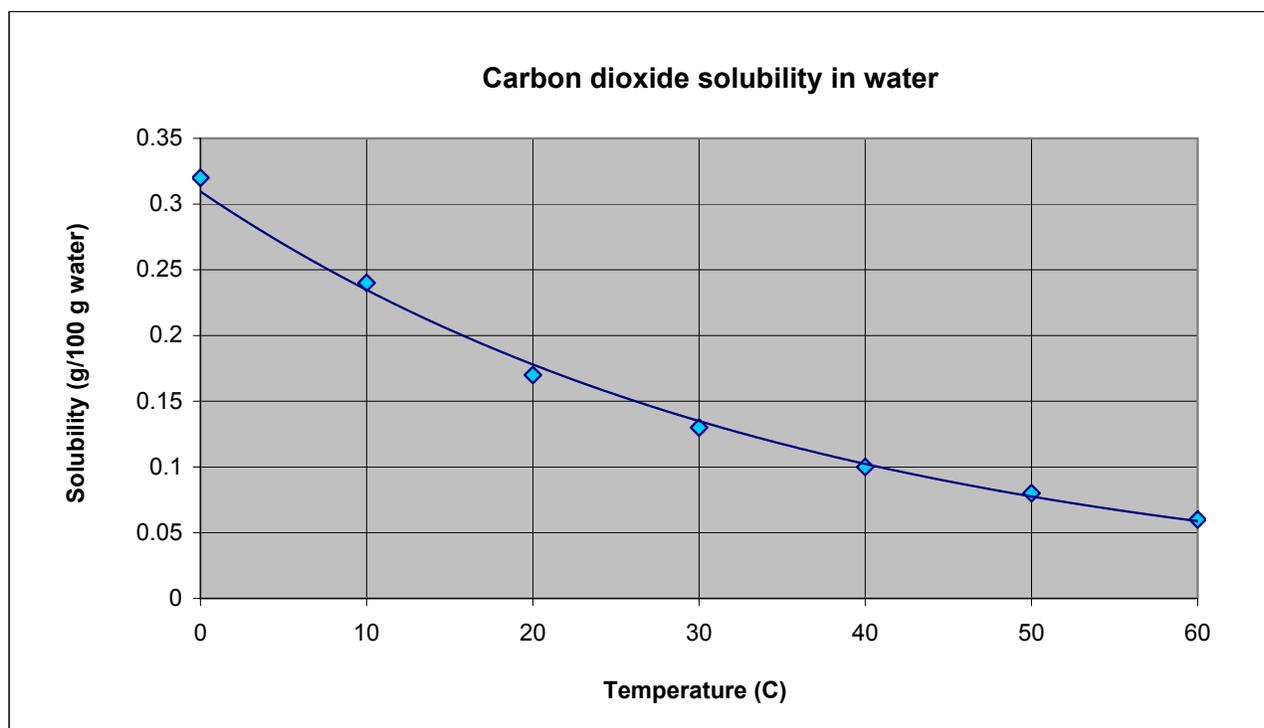
Note: Even a soda that we call “flat” has a little carbon dioxide in it because some CO₂ from the air dissolves into the water.

8. Help students relate their observations to the graph of the solubility of carbon dioxide in water.

Project the image *Solubility Graph for CO₂*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson8#solubility_curve_carbon_dioxide

Look at the graph to see how the concentration of carbon dioxide in water changes with temperature.



Ask students:

- **As the temperature increases, is carbon dioxide more soluble in water or less soluble in water?**

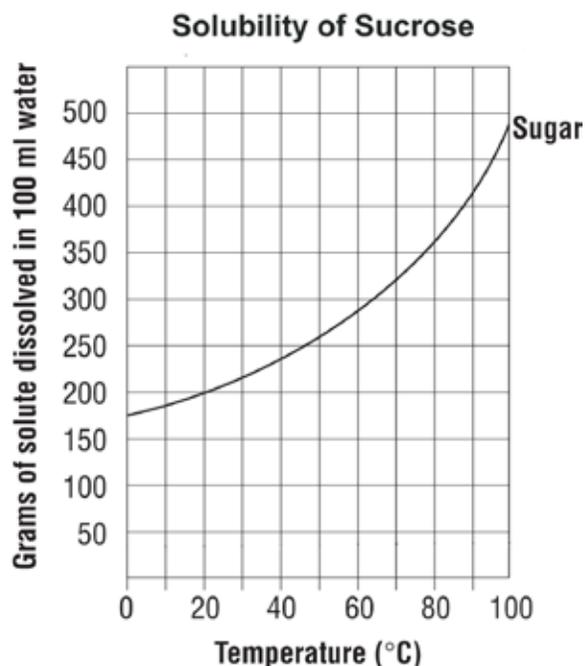
Carbon dioxide is less soluble as the temperature of the water increases.

- **Does this graph match or not match your observations? Explain.**

This graph matches student observations in the activity. As the carbonated water was warmed, more CO_2 left the solution. Since more gas left, less was dissolved in the warmer water.

- **What do the graphs tell you about the solubility of carbon dioxide compared to sucrose, as temperature increases?**

The curve showing the solubility of carbon dioxide goes down as the temperature of the water increases, while the curve showing the solubility of sucrose goes up as the temperature of the water increases. More sucrose can dissolve in hot water than in cold. But for carbon dioxide, more can dissolve in cold water than in hot.



EXTEND

9. **Relate students' experiences with the solubility of carbon dioxide to the solubility of oxygen.**

Tell students that there is another common example of a gas dissolved in water. The water in which fish and other aquatic creatures live contains dissolved oxygen gas. These creatures use their gills to get the oxygen from the water in order to stay alive. Like the solubility of carbon dioxide in water, the solubility of oxygen decreases as the temperature increases.

Ask students:

- **During a long hot summer, you may notice fish gulping air at the surface of a pond. Why do you think the fish come to the surface like this, instead of breathing dissolved oxygen in the water the way they normally do?**

Like carbon dioxide, the concentration of dissolved oxygen is also affected by temperature. Cold water can hold more dissolved oxygen than warm water. In winter and early spring, when the water temperature is low, the dissolved oxygen concentration is high. In summer and early fall, when the water temperature is high, the dissolved-oxygen concentration is lower.

- **Coal-burning power plants heat water to turn turbines to make electricity. After using the water, it is cooled and then returned to the river or lake it came from. Why is it important to cool the water before returning it to the river?**

Dissolved gases, like oxygen for fish and carbon dioxide for aquatic plants, would escape if the returned water were hot. Cool water helps keep gases dissolved, which fish, other aquatic creatures, and underwater plants need.

10. Have students observe and explain what happens when Mentos candies are dropped in a bottle of Diet Coke.



Ask students:

- **Has anyone ever seen the Diet Coke and Mentos demonstration?**

If students have seen it, ask them to describe the activity. An entire packet of Mentos mints is dropped into a 2-liter bottle of carbonated beverage, usually Diet Coke. The soda pop shoots out of the bottle with a lot of force and goes high into the air.

Project the video *Mentos and Diet Coke Demo*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson8#mentos

If you are willing to do this demonstration, it must be done outside. Instructions can be found at <http://crazysciencedemos.com>.

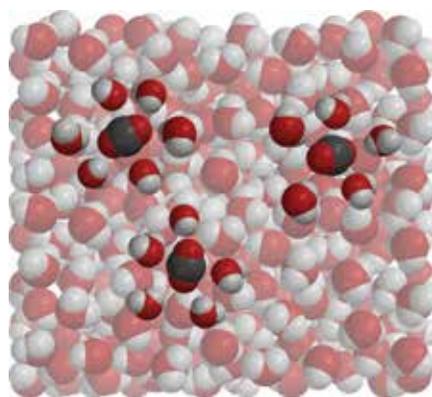
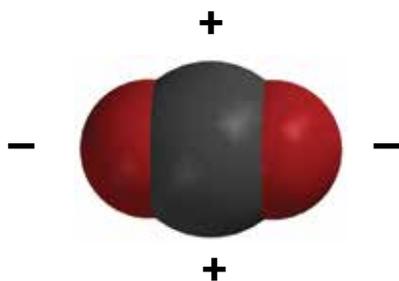
Remind students that the pipe cleaner and the M&M they added to carbonated water caused CO₂ to escape from the solution. Mentos and Diet Coke work in the same way. On the microscopic level, the surface of the mint is rough with many tiny bumps and pits. When the candy is added to the soda pop, carbon dioxide molecules adhere to these tiny spots called *nucleation points*. More carbon dioxide molecules collect in these areas, forming bubbles. The bubbles of carbon dioxide form quickly and grow in all directions but can only escape from the top of the bottle. Because many bubbles are forming and rising to the surface all at once, they bring a large amount of the soda pop with them as they come out of the soda, creating a “fountain” of soda.

DEMONSTRATION

1. What gas is inside the bubbles you saw when your teacher opened a bottle of carbonated water?
2. Where was this gas before the bottle was opened?



EXPLAIN IT WITH ATOMS & MOLECULES



3. Why does carbon dioxide dissolve in water?
4. Why does carbon dioxide gas come out of solution (opposite of dissolving) so easily?

ACTIVITY



Question to Investigate

How can you make carbon dioxide gas come out of solution?

Materials for Each Group

- Club soda in clear plastic cup
- 2 clear plastic cups
- M&M
- Pipe cleaner

Procedure

1. Evenly divide the club soda among the 3 clear plastic cups. Push two of these cups aside to use later.
 2. Place a pipe cleaner in the soda and observe.
 3. Place an M&M in the soda and observe.
5. What was it about the pipe cleaner and M&M that caused bubbles to form?
6. While drinking soda pop with a straw, you may have noticed that bubbles form on the outside of the straw. Now that you have done this activity, why do you think these bubbles form on the straw?



ACTIVITY

Question to Investigate

Does carbon dioxide stay dissolved better in water that is warmed or water that is cooled?



Materials for Each Group

- Carbonated water in 2 clear plastic cups
- Hot water (about 50 °C)
- Cold water (about 5 °C)
- 2 deli containers (that cups easily fit in)

Procedure

1. Get the two cups of carbonated water that you moved aside earlier.
2. Fill one empty deli container about $\frac{1}{3}$ of the way with ice cold water and another about $\frac{1}{3}$ of the way with hot tap water.
3. Place each of the cups of carbonated water into the cold and hot water, as shown.
4. Watch the surface of the soda in each cup of carbonated water.



7. Does carbon dioxide stay dissolved better in hot water or in cold water?

How do you know?

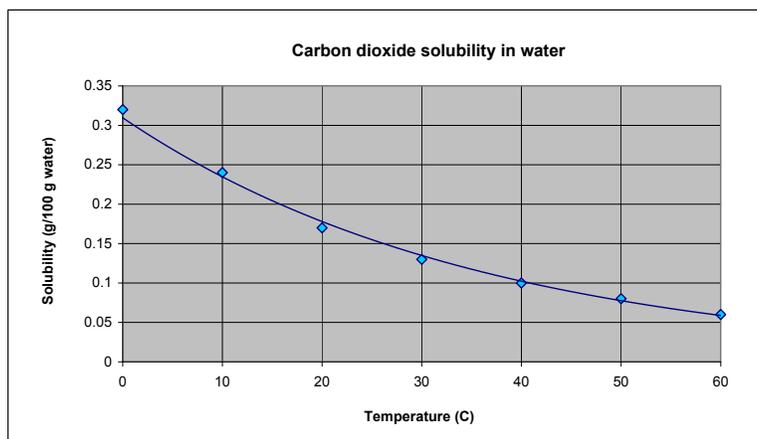
8. Based on what you observed in this experiment, why do you think people store soda pop in the refrigerator after the bottle has been opened?

EXPLAIN IT WITH ATOMS & MOLECULES

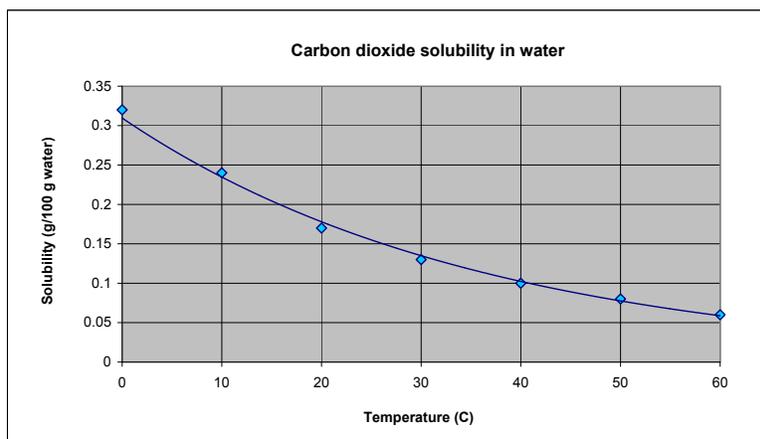
9. Why does warming carbonated water make it easier for carbon dioxide to come out of solution?
10. Look at the graph showing the solubility of carbon dioxide in water to answer the following questions.

As the temperature increases, is carbon dioxide more soluble in water or less soluble in water?

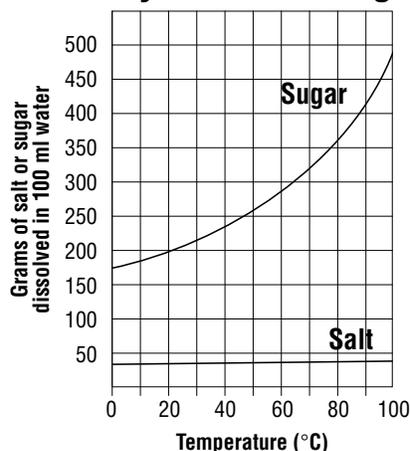
Does this graph match or not match your observations? Explain.



11. What do the graphs below tell you about the solubility of carbon dioxide compared to sucrose, as temperature increases?



Solubility of salt and sugar



TAKE IT FURTHER

12. During a long hot summer, you may notice fish gulping air at the surface of a pond. Why do you think the fish come to the surface like this, instead of breathing dissolved oxygen in the water the way that they normally do?
13. Coal-burning power plants heat water to turn turbines to make electricity. After using the water, it is cooled and then returned to the river or lake it came from. Why is it important to cool the water before returning it to the river?

14. What causes the fantastic “fountain” when a roll of Mentos mints is dropped in a bottle of Diet Coke?

Chapter 5, Lesson 9—Temperature Changes in Dissolving

Key Concepts

- The process of dissolving can be *endothermic* (temperature goes down) or *exothermic* (temperature goes up).
- When water dissolves a substance, the water molecules attract and “bond” to the particles (molecules or ions) of the substance causing the particles to separate from each other.
- The “bond” that a water molecule makes is not a covalent or ionic bond. It is a strong attraction caused by water’s polarity.
- It takes energy to break the bonds between the molecules or ions of the solute.
- Energy is released when water molecules bond the solute molecules or ions.
- If it takes more energy to separate the particles of the solute than is released when the water molecules bond to the particles, then the temperature goes down (endothermic).
- If it takes less energy to separate the particles of the solute than is released when the water molecules bond to the particles, then the temperature goes up (exothermic).

Summary

Students will feel the temperature change that occurs when a cold pack and a hot pack are activated. They will see that these temperature changes are due to a solid substance dissolving in water. Students will then compare the temperature changes that occur as four different solutes dissolve in water and classify these as either endothermic or exothermic. Students will be introduced to the concept that it takes energy to break bonds and energy is released when bonds are formed during the process of dissolving.

Objective

Students will be able to identify variables in an experiment to find out how much the temperature increases or decreases as each of four solutes dissolves in water. Students will be able to correctly classify the process of dissolving as either exothermic or endothermic for each solute. Students will be able to explain that the temperature changes in dissolving are a result of the amount of energy released compared to the amount of energy used as “bonds” are formed and broken.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Excess dry material from the hot and cold packs can be placed in the trash. Sodium carbonate may be a skin irritant. Have students wash hands after the activity.

Materials for the Demonstrations

- 2 disposable cold packs
- 2 disposable hot packs
- Graduated cylinder
- Water (room temperature)
- 2 clear plastic cups
- 2 thermometers

Materials for Each Group

- Potassium chloride
- Calcium chloride
- Sodium carbonate
- Sodium bicarbonate
- Water
- 5 small cups
- Graduated cylinder
- Thermometer
- Gram balance

Notes about the Materials

- The **cold and hot packs** used in this lesson contain urea (cold pack) and magnesium sulfate (hot pack) sealed in a bag with a water-filled bag inside. The hot packs (Rapid Aid Instant Warm Pack) can be purchased by the case (24) or singly from Lab Safety and Supply (product #144707). The urea-based cold pack (Morrison Medical Insta-Cold Compress) is available by the case (24) from Quick Medical Equipment and Supplies (product #6601). If you only need 2 urea-based cold packs, please visit www.middleschoolchemistry.com/contactus and drop us a line. We'll send you two.
- The **hand warmer** shown in the video in the Extend contains a super-saturated solution of sodium acetate with a small metal disk inside a clear plastic bag. This warmer is activated by bending the metal disk slightly. These hand warmers, called The Heat Solution, are available from Flinn Scientific, product number AP1933, and a variety of other vendors.
- **Potassium chloride** may be purchased at a grocery store under the brand name Nu-Salt Salt Substitute or from Flinn Scientific, product number P0042.
- **Calcium chloride** may be purchased at a hardware store under the brand name Damp-Rid or order calcium chloride, anhydrous product number C0016 from Flinn Scientific.
- **Sodium carbonate, anhydrous** may be purchased from Flinn Scientific, product number S0052.
- **Sodium bicarbonate** is baking soda. It may also be purchased from Flinn Scientific, product number S0043.

Note: This activity deals with a concept that is not often addressed in middle school—that a temperature change occurs during the process of dissolving. Most middle school textbooks and curricula associate a change in temperature only with chemical change. Dissolving is usually considered a physical change but also can result in a change in temperature. This change in temperature is based on the energy changes involved in breaking and making “bonds” in the process of dissolving.

Read more about the energy changes in making and breaking bonds in the context of dissolving in the additional teacher background section at the end of the lesson.

ENGAGE

1. Allow students to feel the temperature change in an activated cold pack and an activated hot pack.

Tell students that they will explore how some hot and cold packs work. Give them a hint that it has to do with dissolving, which they have been studying in this chapter.

Materials for the Demonstration

- 1 disposable cold pack
- 1 disposable hot pack



Procedure

1. Select two student volunteers—one to activate one cold pack and another to activate one hot pack.
2. Have each student feel each bag and guess what is inside each. They should notice a dry pellet-like solid and a fluid-filled bag. Point out that the bags do not feel cold or hot yet.
3. Direct the students to activate their packs by following the instructions on the package. Have them shake the packs to get the fluid to spread throughout the bag.
4. Pass the cold and hot packs around the room.



Expected Results

The cold pack quickly becomes cold while the hot pack quickly becomes hot.

2. Do a demonstration to show how cold and hot packs work.

Note: In this demonstration, 1 teaspoon of each substance is dissolved in 10 mL of water. Since the purpose of this demonstration is to show whether the temperature simply goes up or down, this type of volume measure is fine. But to compare which substance is more or less exothermic or endothermic than another, as students will do in the activity, the solute will be measured in grams.

Materials for the Demonstration

- 1 disposable cold pack
- 1 disposable hot pack
- Graduated cylinder, 50 mL or smaller
- Water (room temperature)
- 2 clear plastic cups
- 2 thermometers
- 1 teaspoon



Procedure

1. Carefully cut open one cold pack and one hot pack. Show students the contents, but do not handle or allow students to handle the solid substance inside the packs.
2. Tell students that the liquid inside the fluid-filled bags in both the cold and hot packs is water. Pour about 10 mL of room-temperature water in two separate clear plastic cups.
3. Place a thermometer in each cup and select two student volunteers to tell the class the starting temperature of the water in each cup.
4. With the thermometer still in the cup, add about 1 teaspoon of the solid substance from the cold pack to the water in one cup. Gently swirl the cup to help the substance dissolve. Have the class watch the thermometer and then ask a student to tell the class the lowest temperature of the solution.
5. With the thermometer still in the cup, place about 1 teaspoon of the substance from the hot pack in the water in the other cup. Gently swirl the cup to help the substance dissolve. Have the class watch the thermometer and then ask a student to tell the class the highest temperature of the solution.



Expected Results

Dissolving the substance from the cold pack will cause the temperature to decrease to less than 10 °C (endothermic). Dissolving the substance from the hot pack will cause the temperature to increase to over 40 °C (exothermic). Results may vary.

3. Introduce the terms endothermic and exothermic.

Tell students that scientists describe temperature changes that occur when substances interact as either *endothermic* or *exothermic*. When the temperature decreases, as it does in the cold pack, the process is endothermic. When the temperature increases, as it does in the hot pack, the process is exothermic.

EXPLORE

4. Introduce the dissolving activity students will do and help students identify the variables.

Tell students that they will compare how much the temperature changes when four household substances dissolve in water. Introduce the crystals students will dissolve:

- Potassium chloride is a common salt substitute.
- Calcium chloride is used to absorb moisture from the air. It is also included in some ice-melt mixtures to treat icy sidewalks during winter.
- Sodium carbonate is a common ingredient in detergents for dishwashing machines.
- Sodium bicarbonate, also known as baking soda, is used in baking, in toothpaste, and numerous other applications.

Ask students:

- **How can we set up a fair comparison to find out which solute is the most endothermic and which is most exothermic when dissolved in water?**

Be sure students know to use the same amount of each solute. Based on what they learned in Chapter 5, lesson 5, students should realize that each crystal should be weighed in grams. Students should also remember that they must use the same amount of water at the same temperature. Tell students that in this activity, they will need to swirl the solutes in the water to help them dissolve faster. They should be careful to swirl each cup in the same way for each test.

Note: Comparing the amount of temperature change for different substances by dissolving the same mass of each substance in the same amount of water is fine at the middle school level. However, a more rigorous approach is to dissolve the same number of particles (molecules or ionic units) of each substance in the same amount of water.

Read more about counting molecules in the additional teacher background section at the end of the lesson.

5. Have students monitor changes in temperature as they dissolve four different household solutes in water.



Question to Investigate

Which solute dissolves the most endothermically and which dissolves the most exothermically in water?

Materials for Each Group

- Potassium chloride
- Calcium chloride
- Sodium carbonate
- Sodium bicarbonate
- Water
- 5 small cups
- Permanent marker or masking tape and pen
- Graduated cylinder
- Thermometer
- Gram balance

Procedure

1. Label the small plastic cups Potassium Chloride, Calcium chloride, Sodium carbonate, and Sodium bicarbonate.
2. Weigh 2 g of each solute and place them in their labeled cups.
3. Add 10 mL of water to the small unlabeled cup and place a thermometer in the water. Record this initial temperature in the chart on the activity sheet.
4. Pour the potassium chloride into the water and swirl the cup. Watch the thermometer.
5. When the temperature stops changing, record the final temperature.
6. Repeat steps 3–5 for each solute.



Expected Results

Potassium chloride dissolved the most endothermically, and calcium chloride dissolved the most exothermically. Student temperature readings will vary, but will likely be similar to the following:

Solute dissolved in room-temperature water (22 °C)	Maximum or minimum temperature achieved during dissolving	Endothermic or exothermic?
potassium chloride	14 °C	endothermic
calcium chloride	45 °C	exothermic
sodium chloride	32 °C	exothermic
sodium bicarbonate	18 °C	endothermic

6. Discuss student observations.

Ask students:

- Which solutes dissolved endothermically?
Potassium chloride and sodium bicarbonate
- Which solute dissolved the most endothermically?
Potassium chloride
- Which solutes dissolved exothermically?
Calcium chloride and sodium carbonate
- Which solute dissolved the most exothermically?
Calcium chloride

EXPLAIN

7. Show students an animation of dissolving and explain that the energy of making and breaking “bonds” during dissolving causes a change in temperature.



Project the animation *Breaking and Making Bonds*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson9#breaking_and_making_bonds

Tell students that there is an important rule in chemistry: Energy is required to pull apart atoms, ions, or molecules that are attracted to each other. But when atoms, ions, or molecules come together, energy is released. One way to say it is, “It takes energy to break bonds, and energy is released when bonds are formed.”

Project the animation *Energy and Dissolving*.

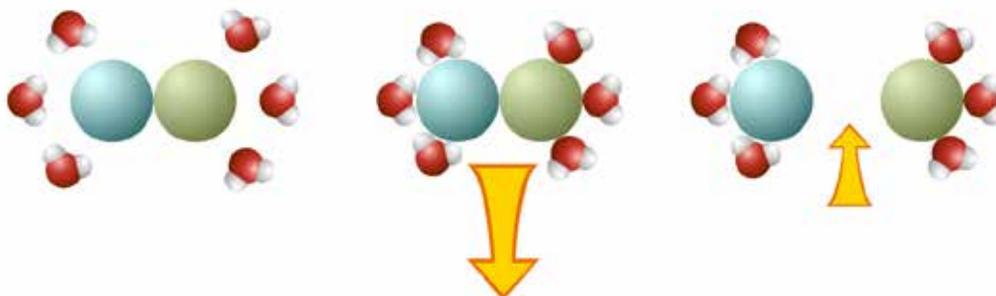
www.middleschoolchemistry.com/multimedia/chapter5/lesson9#energy_and_dissolving

Press the “next” button and explain that this happens in dissolving. When water molecules are attracted to and bond to the molecules or ions of a substance, some energy is released as shown by the arrow going out. Then the water molecules pull ions or molecules of the substance apart, which takes energy, as shown by the arrow going in.

Project the image *Exothermic Dissolving.*

www.middleschoolchemistry.com/multimedia/chapter5/lesson9#exothermic_dissolving

The process of dissolving is exothermic when more energy is released when water molecules “bond” to the solute than is used to pull the solute apart. Because more energy is released than is used, the molecules of the solution move faster, making the temperature increase.



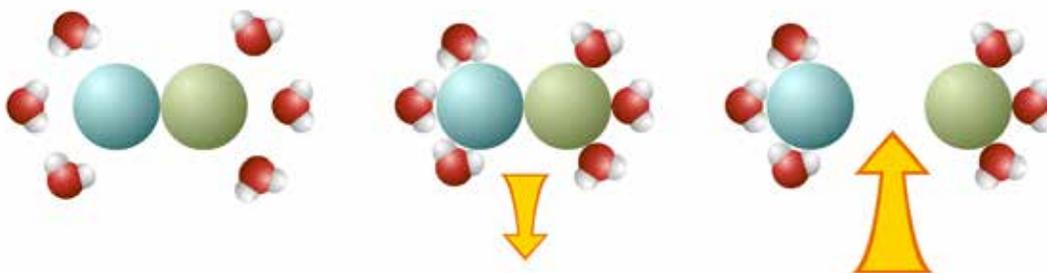
Energy released as water molecules “bond” to the solute.

Energy used when the solute is pulled apart.

Project the image *Endothermic Dissolving.*

www.middleschoolchemistry.com/multimedia/chapter5/lesson9#endothermic_dissolving

The process of dissolving is endothermic when less energy is released when water molecules “bond” to the solute than is used to pull the solute apart. Because less energy is released than is used, the molecules of the solution move more slowly, making the temperature decrease.



Energy released as water molecules “bond” to the solute.

Energy used when the solute is pulled apart.

EXTEND

8. Explore whether the process of crystallization can cause a change in temperature.



Project the video *Hand Warmer*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson9#hand_warmer

In the hand warmer, the water molecules and the ions of the solute come together to form a crystal. Bending the metal disk creates tiny scratches, which act as nucleation points where the sodium acetate crystal forms. As the water molecules and ions bond together in the growing crystal, energy is released. This results in an increase in temperature.

9. Liquids dissolving in a liquid can also cause a temperature change.

Project the video *Temperature Change Alcohol in Water*.

www.middleschoolchemistry.com/multimedia/chapter5/lesson9#temperature_change_alcohol

Explain that the “bonding” of water molecules to alcohol molecules releases more energy than it takes to separate the alcohol molecules from each other. This results in an increase in temperature.

DEMONSTRATION

1. The cold and hot packs you saw each contain a solid substance and water. What is the process that happens inside a cold or hot pack when it is activated?



2. Your teacher opened the cold and hot packs and showed you what was inside each. Then your teacher mixed a small amount of the solid substance from each pack with water.

What happened as each solid substance dissolved in water?		
Substance from the ...	Did the temperature of the solution <i>increase</i> or <i>decrease</i> ?	Is this process <i>endothermic</i> or <i>exothermic</i> ?
Cold pack		
Hot pack		

3. In this activity, you will place a thermometer in water and then add potassium chloride, calcium chloride, sodium carbonate, and sodium bicarbonate to find out which is most endothermic and which is most exothermic as it dissolves.

List three variables and how you might control them.

What is the only variable that should be changed?

ACTIVITY



Question to Investigate

Which solute dissolves the most endothermically, and which dissolves the most exothermically in water?

Materials for Each Group

- Potassium chloride
- Calcium chloride
- Sodium carbonate
- Sodium bicarbonate
- Water
- 5 small cups
- Permanent marker or masking tape and pen
- Graduated cylinder
- Thermometer
- Gram balance

Procedure

1. Label the small plastic cups Potassium chloride, Calcium chloride, Sodium carbonate, and Sodium bicarbonate.
2. Weigh 2 g of each solute and place them in their labeled cups.
3. Add 10 mL of water to the small unlabeled cup and place a thermometer in the water. Record this initial temperature in the chart on the activity sheet.
4. Pour the potassium chloride into the water and swirl the cup. Watch the thermometer.
5. When the temperature stops changing, record the final temperature.
6. Repeat steps 3–5 for each solute.

Substance dissolved in water	Initial temp °C	Final temp °C	Change in temp °C	Endothermic or exothermic?
potassium chloride				
calcium chloride				
sodium carbonate				
sodium bicarbonate				

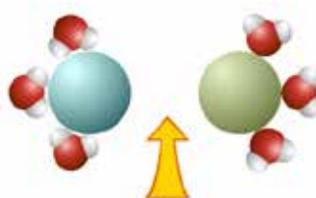
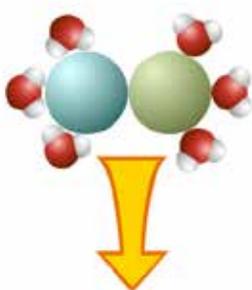
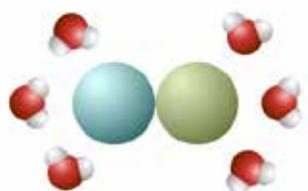
4. Which solute dissolves the most endothermically in water?

5. Which solute dissolves the most exothermically in water?

EXPLAIN IT WITH ATOM & MOLECULES

6. The two sets of illustrations below and on the next page show the energy changes that may occur during the process of dissolving. Title each either endothermic or exothermic and answer the question beneath each.

Title: _____

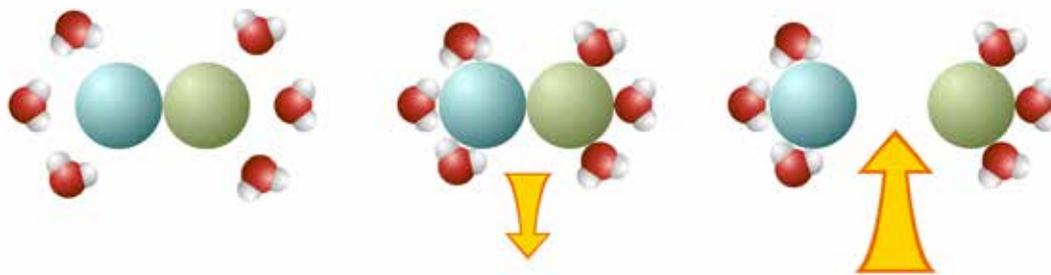


Energy released as water molecules “bond” to the solute.

Energy used when the solute is pulled apart.

How does the size of the arrows relate to the change in temperature of the solution?

Title: _____



Energy released as water molecules “bond” to the solute.

Energy used when the solute is pulled apart.

How does the size of the arrows relate to the change in temperature of the solution?

TAKE IT FURTHER

7. The hand warmer shown in the video heats up as molecules and ions come together to form a crystal. Does the process of making bonds to form a crystal use or release energy?
8. If you think about the energy of making and breaking “bonds,” why do you think there is an increase in temperature when isopropyl alcohol dissolves in water?

Additional Teacher Background

Chapter 5, Lesson 9, p. 503

The additional teacher background in Chapter 2 on evaporation and condensation discussed the concept that it takes energy to break bonds and that energy is released when bonds are formed. In the context of evaporation, “breaking bonds” refers to overcoming the attractions between water molecules as they change from a liquid to a gas. In the context of condensation, “making bonds” refers to the attractions of water vapor molecules bringing them together as they change from a gas to a liquid.

It was also explained that the energy “used” in breaking bonds is actually an *energy conversion* from kinetic to potential energy. The energy “released” in making bonds is an energy conversion from potential to kinetic energy.

Since evaporation is only a bond-breaking process, it only uses energy, resulting in a temperature decrease. Therefore, evaporation is endothermic. Since condensation is only a bond-making process, it only releases energy, resulting in a temperature increase. Therefore, condensation is exothermic.

But the process of dissolving involves both bond-making and bond-breaking. In the context of dissolving, “making bonds” refers to water molecules attracting and bonding to the ions or molecules of the solute. “Breaking bonds” refers to the action of the water molecules in separating the ions or molecules of the solute from one another as they go into solution. The energy conversions in bond-making and bond-breaking in dissolving are the same as in evaporation and condensation. Energy is released as water molecules bond to the solute and energy is used as the ions or molecules of the solute are separated.

The dissolving process is exothermic if more energy is released when water molecules bond to the ions or molecules of the solute than is used to break the bonds holding the solute together.

The dissolving process is endothermic if less energy is released when water molecules bond to the ions or molecules of the solute than is used to break the bonds holding the solute together.

Additional Teacher Background

Chapter 5, Lesson 9, p. 484

The idea of counting molecules was discussed in the Additional Teacher Background section in lesson 1. There, the question was whether it was fair to do an evaporation test between water and alcohol using one drop of each when these drops contain a different number of molecules. It was explained that there is a way to “count” molecules using the atomic mass and the mole concept. In this way, if needed, an evaporation test could be conducted using the same number of molecules of alcohol and water.

Here, the question is whether it is fair to compare the temperature change when equal masses of two different substances are dissolved. Again, since substances are composed of different atoms, they have different masses. So comparing the same mass of two different substances means that each sample contains a different number of particles.

For example, an equal mass of calcium chloride and potassium chloride must have a different number of units of each substance. Calcium chloride is CaCl_2 with a mass of $40 + 2(35.5) = 111$. This means that 111 grams of calcium chloride contain one mole or 6.02×10^{23} calcium chloride units.

Potassium chloride is KCl with a mass of $39 + 35.5 = 74.5$. This means that 74.5 grams of potassium chloride contains one mole or 6.02×10^{23} potassium chloride units. So it takes fewer grams of potassium chloride to have the same number of units as calcium chloride. This means that if you dissolve equal masses of these substances, you are actually using more units of potassium chloride than calcium chloride.

If you wanted to use the same number of units of each, you would need to use masses that are in the same ratio as 111 grams calcium chloride / 74.5 grams potassium chloride which is very close to 1.5/1.

Chapter 5—Student Reading

THE POLARITY OF THE WATER MOLECULE

Wonderful water

Water is an amazing substance. We drink it, cook and wash with it, swim and play in it, and use it for lots of other purposes. The water in our bodies helps keep us cool and the water in the oceans helps regulate the temperature and weather on Earth.



PHOTOS.COM



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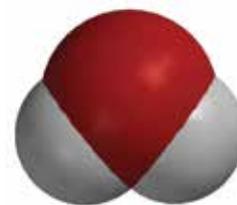


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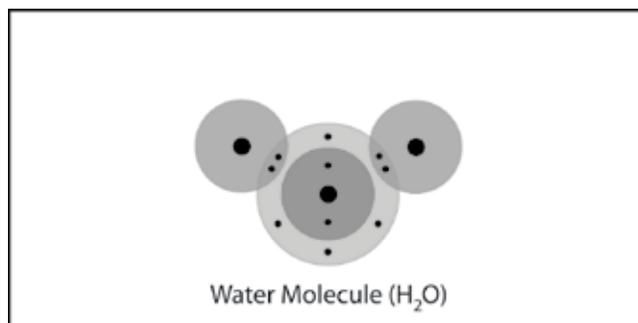
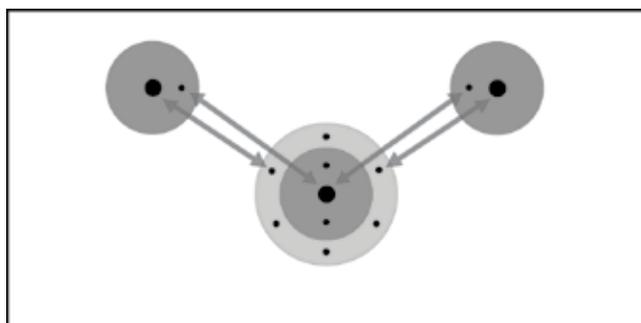
To understand what makes water so special, you need to look at the atoms that a water molecule is made from and how these atoms are bonded together. After you understand the characteristics of a single water molecule, it is easier to see why many water molecules together makes water act the way it does.

A molecular view of water

You already know that a water molecule is made up of two hydrogen atoms bonded to one oxygen atom. That's why the chemical formula for water is H_2O .

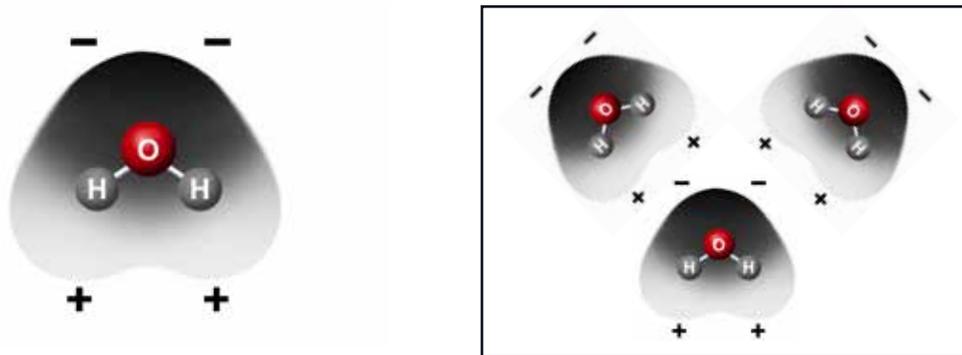


The hydrogen atoms and oxygen atom are bonded by covalent bonds. In a covalent bond, an electron from each atom is attracted to the protons of the other atom. If the attraction is strong enough in both directions and there is room for the electron in the outer energy level, the electrons end up being attracted and shared by both atoms.



Water is a polar molecule

Even though electrons are shared in the covalent bonds in a water molecule, the electrons are not shared equally. The oxygen atom has a stronger attraction for electrons than the hydrogen atoms. Therefore, the electrons tend to spend a little more time at the oxygen end of the molecule than at the hydrogen end. Since electrons are negative, this makes the oxygen end a bit negative. Since the electrons are not at the hydrogen end of the molecule as much, the hydrogen end is a bit positive.



The water molecule hasn't gained or lost any electrons, but the electrons are just spending more time in one area of the molecule than another. When electrons do this in a molecule, it is called a *polar* molecule. The positive and negative polar areas of the water molecule attract each other. This gives water many of its characteristic properties.

Polar molecules and evaporation

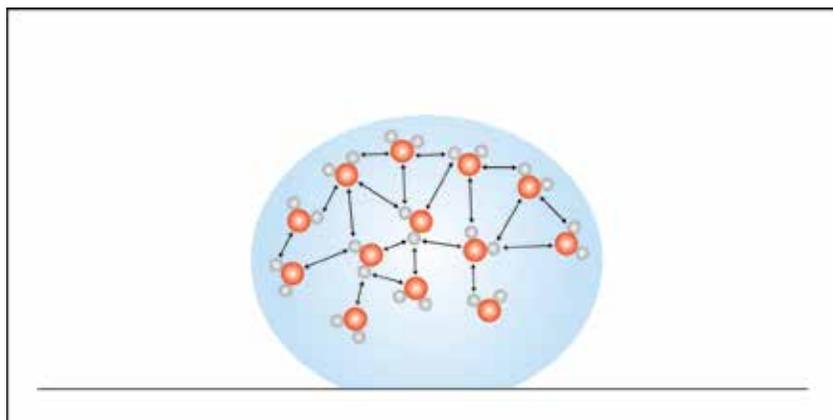
The polar characteristics of molecules of a liquid can affect how quickly the liquid evaporates. For instance, alcohol molecules evaporate at a faster rate than water molecules. One reason for this is that the attraction between alcohol molecules is not as strong as the attraction between water molecules.



Look at the models of the water and alcohol molecules. Water has two oxygen-hydrogen (O–H) bonds which are pretty polar. Alcohol only has one O–H bond and a few carbon-hydrogen (C–H) bonds which are not very polar. Because of this difference in polarity, the attraction between alcohol molecules is not as strong as the attraction between water molecules. This allows alcohol to evaporate more quickly than water.

Water's surface tension

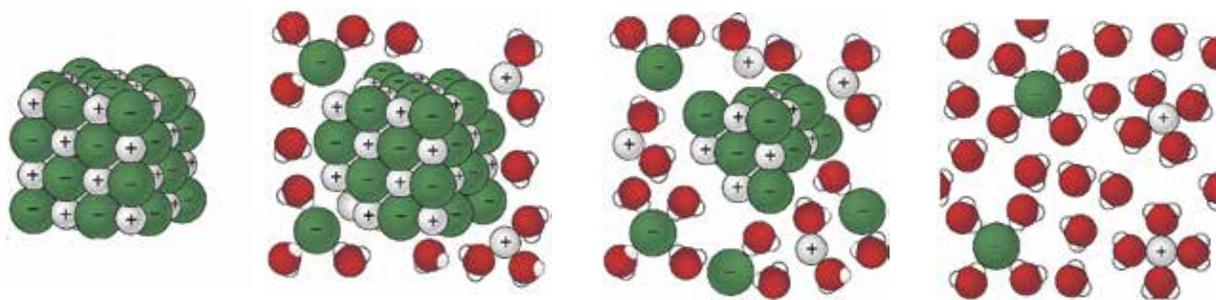
You have probably seen drops of water “bead up” on a piece of wax paper or on a newly-waxed car or other clean hard surface. The strong attraction that water molecules have for each other also helps explain why water beads up the way it does. Inside a drop of water, the water molecules are being attracted in every direction.



But the molecules at the surface only feel attractions from the molecules next to them and beneath them. These surface molecules are pulled together and inward by these attractions. This inward pull has results in a tight arrangement of molecules over the water's surface. This tight arrangement at the surface is called *surface tension*. The inward pull from the attractions of the molecules results in the smallest possible surface for a volume of water, which is a sphere.

Why water dissolves salt

A sodium chloride crystal is made up of positively charged sodium ions and negatively charged chloride ions. The positive and negative ends of the polar water molecules can arrange themselves near an ion and help remove it from the crystal. The positive area of a water molecule is attracted to the negative chloride ion. The negative area of a water molecule is attracted to the positive sodium ion.



Dissolving happens when the attractions between the water molecules and the sodium and chloride ions overcome the attractions of the ions to each other. This causes the ions to separate from one another and become thoroughly mixed into the water.

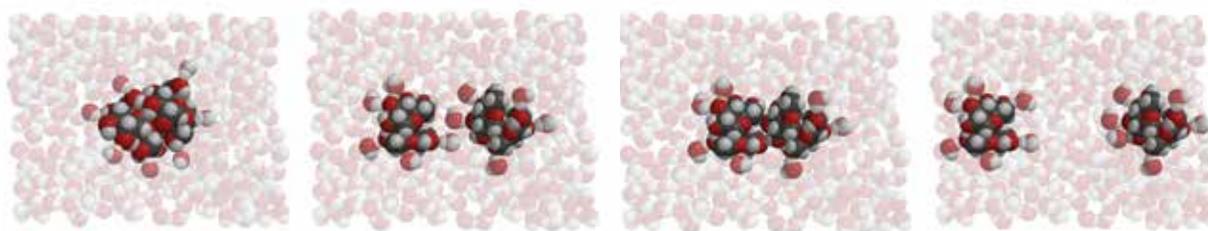
Mixtures and solutions

When talking about dissolving, the substance being dissolved is called the *solute*. The substance doing the dissolving is called the *solvent*. A solute is dissolved in a solvent when the particles of the solute are so thoroughly intermingled with the particles of the solvent that they will not settle out. When a solute is dissolved in a solvent, the combination is called a *solution*. A solution is a type of mixture. Some mixtures are not solutions. An example of a mixture that is not a solution is a *suspension* like a teaspoon of flour mixed in a cup of water. In a suspension, the particles are not as completely associated with the molecules of the solvent as they are in a solution. In a suspension, the solute particles will eventually settle to the bottom.

Why water dissolves sugar

Sucrose is a pretty big molecule. It is $C_{12}H_{22}O_{11}$. Sucrose has lots of oxygen-hydrogen (O–H) covalent bonds like the ones in a water molecule. The area near the hydrogen is positive and the area near the oxygen is negative.

When sucrose is placed in water, the positive area of a water molecule is attracted to the negative area of a sucrose molecule. And the negative area of a water molecule is attracted to the positive area of the sucrose molecule.



When the attraction between water molecules and sucrose molecules overcomes the attraction the sucrose molecules have to other sucrose molecules, they will separate from one another and *dissolve*. Notice how one whole sucrose molecule breaks away from another whole sucrose molecule. The molecule itself does not come apart into individual atoms.

Another solvent like mineral oil which has only carbon-hydrogen (C–H) bonds is not polar and does not dissolve salt or sugar as well as water does.

Solubility is a characteristic property of a substance

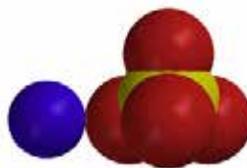
Each substance dissolves in water to a different extent. Another way of saying this is that each substance has its own characteristic solubility. Solubility is usually measured by the number of grams of a substance that dissolves in 100 mL of water at a particular temperature.

It should make sense that different substances would have different solubilities. Since substances are made up of different atoms and ions bonded together differently, they interact with water differently giving them each their own characteristic solubility.

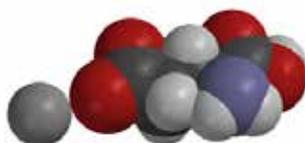
Common substances like sodium chloride (table salt), magnesium sulfate (Epsom salt), monosodium glutamate (MSG), and sucrose (sugar) all are made up of different ions or atoms bonded together differently. Since water interacts with each of these substances differently, they each have their own characteristic solubility.



Sodium chloride



Magnesium sulfate



Monosodium glutamate



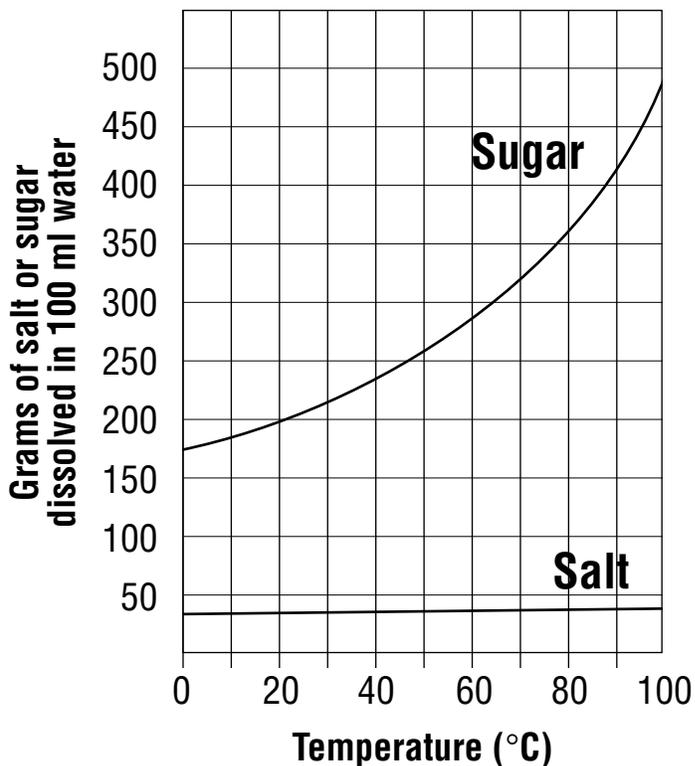
Sucrose

Temperature affects dissolving

Temperature affects dissolving but it doesn't have the same effect on all substances. Look at the table and graph below. They show that as the temperature increases, more sodium chloride and more sucrose dissolve in water. They also show that the dissolving of sugar increases with temperature much more than the dissolving of salt.

Comparing the solubility of sodium chloride and sucrose Measured in grams of solute dissolved in 100 mL of water						
Temperature °C	0	20	40	60	80	100
Sodium chloride	35.5	36	36.5	37.5	38	39
Sucrose	179	204	241	288	363	487

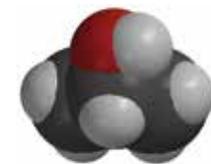
Solubility of salt and sugar



Liquids can dissolve in liquids

Solids aren't the only substances that can dissolve. Liquids can also dissolve in other liquids. For example, alcohol and corn syrup dissolve in water but mineral oil does not. The polarity of molecules can help explain this.

Even though alcohol has one polar area (O–H bond) and a larger nonpolar area (C–H bonds), polar water molecules and the polar area on alcohol molecules are attracted to each other, causing alcohol to dissolve in water.



Glucose has many areas where oxygen is bonded to hydrogen. These O–H bonds are polar. Polar water molecules and the polar areas of glucose molecules are attracted to each other, causing the corn syrup to dissolve.



The mineral oil molecule is made of carbon atoms bonded to hydrogen atoms. The bond between these atoms creates very little polarity. Water is not very attracted to the oil and so does not dissolve it.



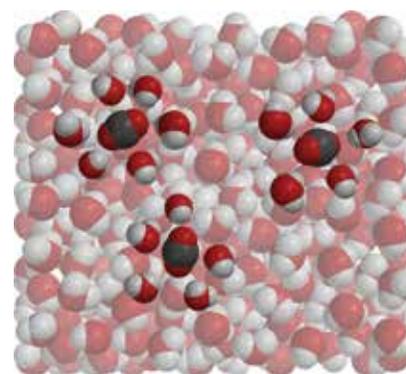
Gases can dissolve in liquids

Gases can also dissolve in liquids. A good example is the gas in soda pop. This gas is carbon dioxide (CO_2). In soda pop, molecules of carbon dioxide are thoroughly mixed and dissolved in water. When dissolved, the molecules of CO_2 are not like tiny little bubbles of gas mixed in the water. Instead single molecules of CO_2 are surrounded by water molecules.

A molecule of carbon dioxide has a slight negative charge near the oxygen and a slight positive charge near the carbon. CO_2 is soluble because water molecules are attracted to these polar areas. The bond between carbon and oxygen is not as polar as the bond between hydrogen and oxygen, but it is polar enough that carbon dioxide can dissolve in water.



Although the CO_2 dissolves, the molecules are not attracted as strongly by the water molecules as substances like salt or sugar. Due to these weaker attractions, the molecules of CO_2 come out of solution relatively easily. This is why soda becomes flat if it is left uncapped for too long.



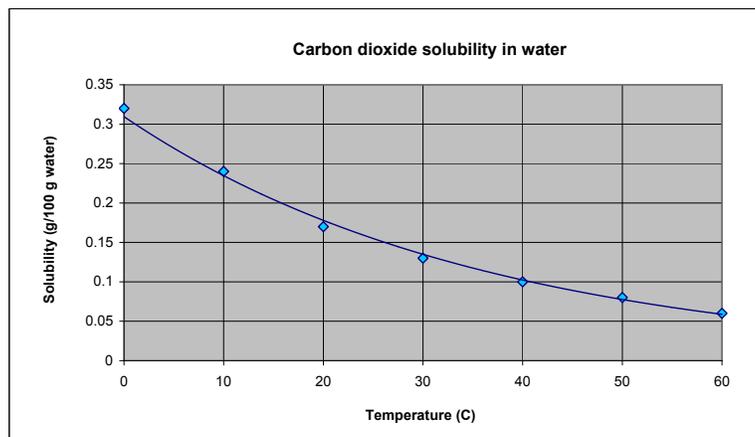
Heating and cooling affect how much gas stays in a liquid

As you have read, carbon dioxide molecules and water molecules are somewhat attracted to each other, but the attraction is not very strong. Carbon dioxide readily leaves an open container of soda pop. Increasing the temperature of the soda increases this effect. Warming the soda increases the motion of the water and carbon dioxide molecules making their attachments even looser and allowing the gas to escape even faster.

Look at the graph to see how the concentration of carbon dioxide in water changes with temperature.

You can see that at the lowest temperature, the concentration of CO_2 is the highest. At the highest temperature, the concentration of CO_2 is the lowest.

If you compare the solubility graph for carbon dioxide with the solubility graph for sucrose, you can see that the line moves in the opposite direction. The curve showing the solubility of carbon dioxide goes down as the temperature of the water increases, while the curve showing the solubility of sucrose goes up as the temperature of the water increases. More sucrose can dissolve in hot water than in cold. But for carbon dioxide, more can dissolve in cold water than in hot.



Dissolving can cause a change in temperature

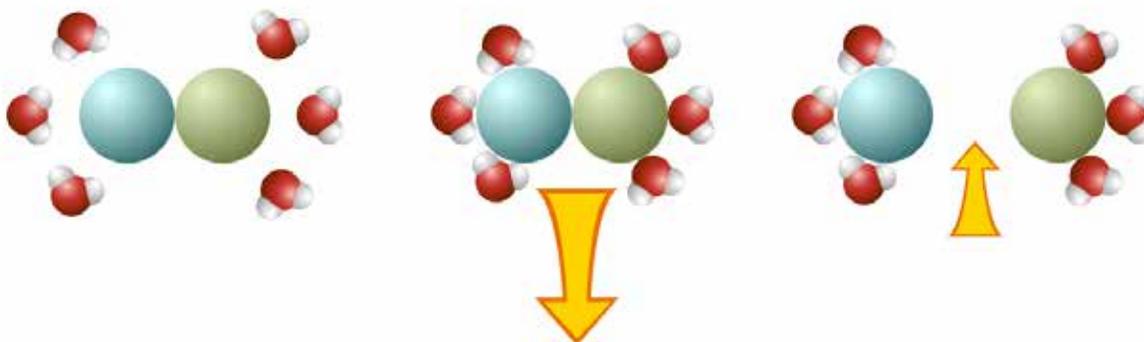
You have seen that the temperature of the solvent affects how much solute dissolves. But it is also true that the actual process of dissolving can cause a change in temperature. There is a principle in chemistry that states:

It takes energy to break bonds, and energy is released when bonds are formed

Normally these concepts are used when talking about chemical reactions which will be discussed in Chapter 6. But the same principles apply in a related way to the process of dissolving. When water molecules are attracted to and attach to the solute, energy is released. And when the movement of the molecules causes the solute to come apart, energy is absorbed. The absorbing and releasing of energy in dissolving can help explain why the temperature goes up when some solutes are dissolved and goes down when others are dissolved.

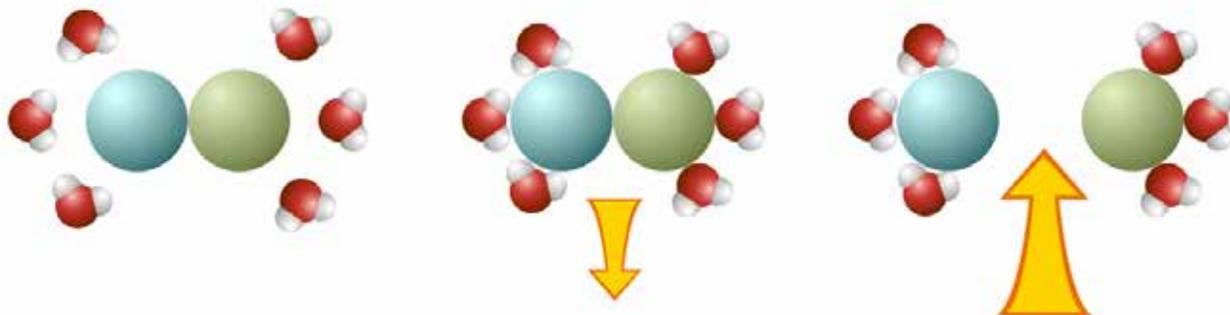
Exothermic Dissolving

If dissolving is exothermic, that means that it takes less energy for water molecules to break apart the solute than is released when the water molecules attract and attach to the solute. Overall, the temperature increases.



Endothermic Dissolving

If dissolving is endothermic, it takes more energy for water molecules to break apart the solute than is released when the water molecules attract and attach to the solute. Overall, the temperature decreases.



Chapter 6, Lesson 1: What is a Chemical Reaction?

Key Concepts:

- A physical change, such as a state change or dissolving, does not create a new substance, but a chemical change does.
- In a chemical reaction, the atoms and molecules that interact with each other are called *reactants*.
- In a chemical reaction, the atoms and molecules produced by the reaction are called *products*.
- In a chemical reaction, only the atoms present in the reactants can end up in the products. No new atoms are created, and no atoms are destroyed.
- In a chemical reaction, reactants contact each other, bonds between atoms in the reactants are broken, and atoms rearrange and form new bonds to make the products.

Summary

The teacher will use a small candle flame to demonstrate a chemical reaction between the candle wax and oxygen in the air. Students will see a molecular animation of the combustion of methane and oxygen as a model of a similar reaction. Students will use atom model cut-outs to model the reaction and see that all the atoms in the reactants show up in the products.

Objective

Students will be able to explain that for a chemical reaction to take place, the bonds between atoms in the reactants are broken, the atoms rearrange, and new bonds between the atoms are formed to make the products. Students will also be able to explain that in a chemical reaction, no atoms are created or destroyed.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Be careful when lighting the candle. Be sure that the match and candle are completely extinguished when you are finished with the demonstration.

Materials for the Demonstration

- Tea light candle or other small stable candle
- Matches
- Glass jar, large enough to be placed over the candle

Materials for Each Student

- Atom cut-outs from the activity sheet
- Sheet of colored paper or construction paper
- Colored pencils
- Scissors
- Glue or tape

ENGAGE

1. Review what happens during a physical change and introduce the idea of chemical change.

Tell students that in previous chapters they have studied different aspects of physical change. When atoms and molecules speed up or slow down, that is a physical change. When they change state from liquid to solid or from gas to liquid, that is a physical change. When a substance is dissolved by water or some other solvent, a new substance has not really been formed. The ions or molecules can still come back together to form the original substance.

Let students know that in this chapter they will explore what happens during a *chemical* change. In a chemical change, the atoms in the reactants rearrange themselves and bond together differently to form one or more new products with different characteristics than the reactants. When a new substance is formed, the change is called a chemical change.

2. As a demonstration, light a candle and explain what is happening using the terms *reactants*, *products*, and *chemical reaction*.

Explain that in most chemical reactions, two or more substances, called *reactants*, interact to create different substances called *products*. Tell students that burning a candle is an example of a chemical reaction.

Materials for the Demonstration

- Tea light candle or other small stable candle
- Matches
- Glass jar, large enough to be placed over the candle



Procedure

1. Carefully light a tea light candle or other small candle.
2. Keep the candle burning as you ask students the questions below. You will put the candle out in the second part of the demonstration.

Expected Results

The wick will catch on fire and the flame will be sustained by the chemical reaction.

The following question is not easy and students are not expected to know the answer at this point. However, thinking about a candle burning in terms of a chemical reaction is a good place to start developing what it means when substances react chemically.

Ask students:

- **What do you think are the reactants in this chemical reaction?**

Wax and oxygen from the air are the reactants.

Students often say that the string or wick is burning. It is true that the string of the wick does burn but it's the wax on the string and not so much the string itself that burns and keeps the candle burning. Explain that the molecules that make up the wax combine with oxygen from the air to make the products carbon dioxide and water vapor.

Point out to students that this is one of the major characteristics of a chemical reaction: **In a chemical reaction, atoms in the reactants combine in new and different ways to form the molecules of the products.**

Students may be surprised that water can be produced from combustion. Since we use water to extinguish a fire, it may seem strange that water is actually produced by combustion. You may want to let students know that when they “burn” food in their bodies, they also produce carbon dioxide and water.

3. Place a jar over the candle to help students realize that oxygen is a reactant in the burning of a candle.

Remind students that air is a mixture of gases. Explain that when something burns, it reacts with the oxygen in the air.

Ask students to make a prediction:

- **Will the candle still burn if one of the reactants (wax or oxygen) is no longer available?**

Students may guess that the candle will not burn because both reactants are required for the chemical reaction to continue.



Procedure

1. Carefully place a glass jar over the lit candle.

Expected Results

The flame goes out.

Ask students:

- **Why do you think the flame goes out when we put a jar over the candle?**

Placing a jar over the candle limits the amount of oxygen in the air around the candle. Without enough oxygen to react with the wax, the chemical reaction cannot take place and the candle cannot burn.



When a candle burns for a while, it eventually gets smaller and smaller. Where does the candle wax go?

When a candle burns, the candle wax seems to “disappear.” It doesn’t really disappear, though: It reacts chemically, and the new products go into the air.

Note: Some curious students may ask what the flame is made of. This is a great question and not trivial to answer. The flame is burning wax vapor. The light of the flame is caused by a process called chemiluminescence. Energy released in the chemical reaction makes electrons from different molecules move to a higher energy state. When the electrons come back down, energy is released in the form of light.

EXPLAIN

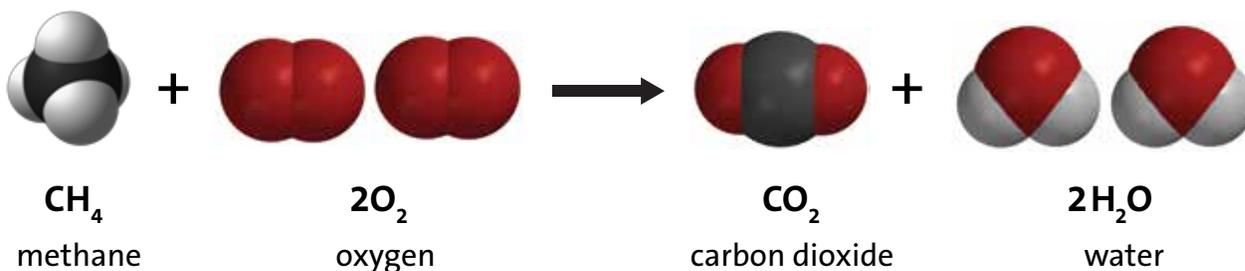
4. Introduce the chemical equation for the combustion of methane and explain that atoms rearrange to become different molecules.

Explain to students that wax is made of long molecules called *paraffin* and that paraffin is made up of only carbon atoms and hydrogen atoms bonded together. Molecules made of only carbon and hydrogen are called *hydrocarbons*. Tell students that you will use the simplest hydrocarbon (methane) as a model to show how the wax, or any other hydrocarbon, burns.

Project the image *Methane and Oxygen React*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson1#chemical_reaction_methane

Show students that there is methane and oxygen on the left side of the chemical equation and carbon dioxide and water on the right side. Explain that the molecules on the left side are the *reactants* and the ones on the right side are the *products*. When the candle was burning, the paraffin reacted with oxygen in the air to produce carbon dioxide and water, similar to the chemical reaction between methane and oxygen.



Explain to students that the chemical formula for methane is CH_4 . This means that methane is made up of one carbon atom and four hydrogen atoms.

Show students that the other reactant is two molecules of oxygen gas. Point out that each molecule of oxygen gas is made up of two oxygen atoms bonded together. It can be confusing for students that oxygen the atom, and oxygen the molecule, are both called *oxygen*. Let students know that when we talk about the oxygen in the air, it is always the molecule of oxygen, which is two oxygen atoms bonded together, or O_2 .

Ask students:

- **Where do the atoms come from that make the carbon dioxide and the water on the right side of the equation?**

The atoms in the products come from the atoms in the reactants. In a chemical reaction, bonds between atoms in the reactants are broken and the atoms rearrange and form new bonds to make the products.

Note: Leave this equation projected throughout the activity in the Explore section of this lesson. Students will need to refer to it as they model the chemical reaction.

Give Each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms and Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

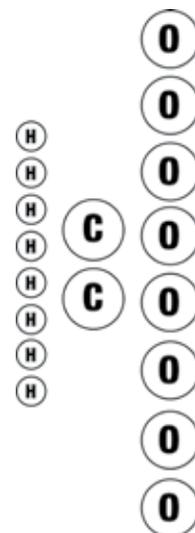
5. **Have students make a model to show that in a chemical reaction the atoms of the reactants rearrange to form the products.**

Question to Investigate

Where do the atoms in the products of a chemical reaction come from?

Materials for Each Student

- Atom model cut-outs (carbon, oxygen, and hydrogen)
- Sheet of colored paper or construction paper
- Colored pencils
- Scissors
- Glue or tape



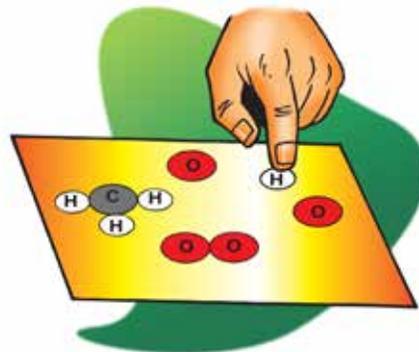
Procedure

Prepare the Atoms

1. Color the carbon atoms black, the oxygen atoms red, and leave the hydrogen atoms white.
2. Use scissors to carefully cut out the atoms.

Build the Reactants

3. On a sheet of paper, place the atoms together to make the molecules of the reactants on the left side of the chemical equation for the combustion of methane.
4. Write the chemical formula under each molecule of the reactants. Also draw a + sign between the reactants.



After you are sure that students have made and written the formula for the reactant molecules, tell students that they will rearrange the atoms in the reactants to form the products.

Build the Products

5. Draw an arrow after the second oxygen molecule to show that a chemical reaction is taking place.
6. Rearrange the atoms in the reactants to make the molecules in the products on the right side of the arrow.
7. Write the chemical formula under each molecule of the products. Also draw a + sign between the products.

Tell students that in a chemical reaction, the atoms in the reactants come apart, rearrange, and make new bonds to form the products.

Represent the Chemical Equation

8. Have students use their remaining atoms to make the reactants again to represent the chemical reaction as a complete chemical equation.
9. Glue or tape the atoms to the paper to make a more permanent chemical equation of the combustion of methane.

EXPLAIN

6. Help students count up the number of atoms on each side of the equation.



Project the animation *Combustion of Methane*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson1#combustion_of_methane

Show students that the atoms in methane and oxygen need to come apart just like in their models. Also point out that the atoms arrange themselves differently and bond again to form new products. This is also like their model. Be sure that students realize that the atoms in the products only come from the reactants. There are no other atoms available. No new atoms are created and no atoms are destroyed.

Explain to students that chemical reactions are more complicated than the simplified model shown in the animation. The animation shows that bonds between atoms in the reactants are broken, and that atoms rearrange and form new bonds to make the products. In reality, the reactants need to collide and interact with each other in order for their bonds to break and rearrange. Also, the animation shows all of the atoms in the reactants coming apart and rearranging to form the products. But in many chemical reactions, only some bonds are broken, and groups of atoms stay together as the reactants form the products.

Guide students as you answer the following question together:

- **How many carbon, hydrogen, and oxygen atoms are in the reactants compared to the number of carbon, hydrogen, and oxygen atoms in the products?**

Show students how to use the big number (coefficient) in front of the molecule and the little number after an atom of the molecule (subscript) to count the atoms on both sides of the equation. Explain to students that the subscript tells how many of a certain type of atom are in a molecule. The coefficient tells how many of a particular type of molecule there are. So if there is a coefficient in front of the molecule and a subscript after an atom, you need to multiply the coefficient times the subscript to get the number of atoms.

For example, in the products of the chemical reaction there are $2\text{H}_2\text{O}$. The coefficient means that there are two molecules of water. The subscript means that each water molecule has two hydrogen atoms. Since each water molecule has two hydrogen atoms and there are two water molecules, there must be 4 (2×2) hydrogen atoms.

Read more about the combustion of methane in the additional teacher background section at the end of the lesson.

Atoms	Reactant side	Product side
Carbon	1	1
Hydrogen	4	4
Oxygen	4	4

Note: The coefficients actually indicate the ratios of the numbers of molecules in a chemical reaction. It is not the actual number as in two molecules of oxygen and one molecule of methane since there are usually billions of trillions of molecules reacting. The coefficient shows that there are twice as many oxygen molecules as methane molecules reacting. It would be correct to say that in this reaction there are two oxygen molecules for every methane molecule.

7. Explain that mass is conserved in a chemical reaction.

Ask students:

- **Are atoms created or destroyed in a chemical reaction?**

No.

- **How do you know?**

There are the same number of each type of atom on both the reactant side and the product side of the chemical equation we explored.

- **In a physical change, like changing state from a solid to a liquid, the substance itself doesn't really change. How is a chemical change different from a physical change?**

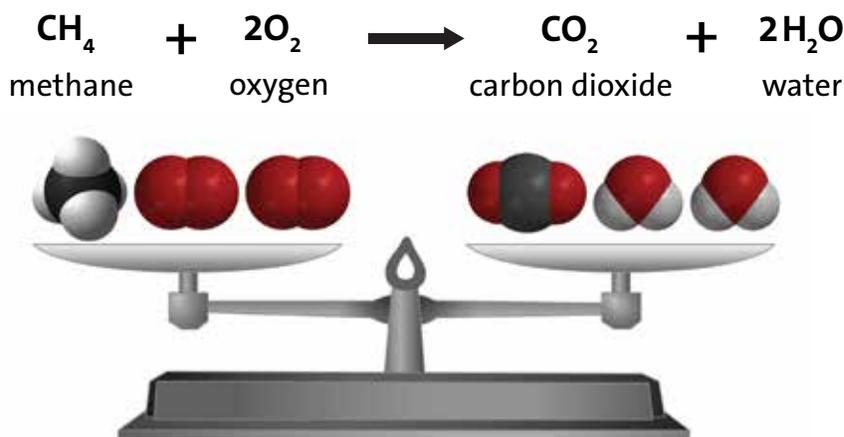
In a chemical change, the molecules in the reactants interact to form new substances. In a physical change, like a state change or dissolving, no new substance is formed.

Explain that another way to say that no atoms are created or destroyed in a chemical reaction is to say, "Mass is conserved."

Project the image *Balanced Equation*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson1#balanced_equation

Explain that the balance shows the mass of methane and oxygen on one side exactly equals the mass of carbon dioxide and water on the other. When an equation of a chemical reaction is written, it is "balanced" and shows that the atoms in the reactants end up in the products and that no new atoms are created and no atoms are destroyed.



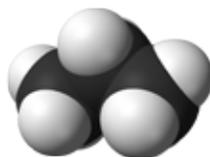
EXTEND

8. Introduce two other combustion reactions and have students check to see whether or not they are balanced.

Tell students that, in addition to the wax and methane, some other common hydrocarbons are propane (the fuel in outdoor gas grills), and butane (the fuel in disposable lighters). Have students count the number of carbon, hydrogen, and oxygen atoms in the reactants and products of each equation to see if the equation is balanced. They should record the number of each type of atom in the chart on their activity sheet.



CH_4
methane

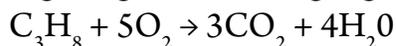


C_3H_8
propane

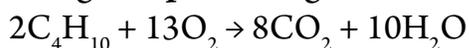


C_4H_{10}
butane

Lighting an outdoor gas grill—Combustion of propane



Using a disposable lighter—Combustion of butane



After students have counted up each type of atom, review their answers to make sure they know how to interpret subscripts and coefficients.

Activity Sheet
Chapter 6, Lesson 1
What is a Chemical Reaction?

Name _____

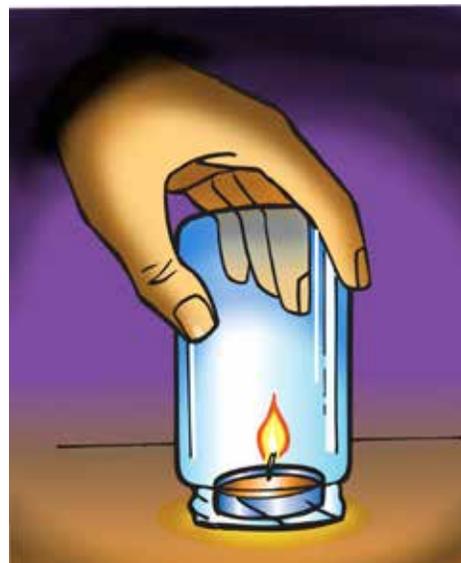
Date _____

DEMONSTRATION

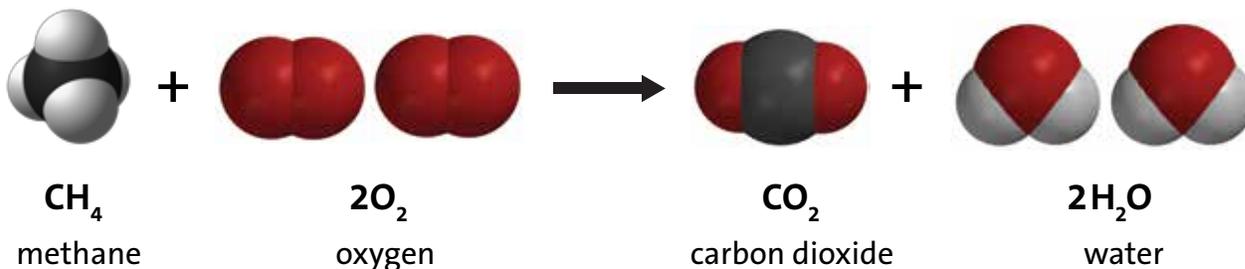
1. Your teacher lit a candle and told you that this was a chemical reaction. What are the *reactants* in this chemical reaction?

2. What are the *products* in this chemical reaction?

3. Why did the flame go out when your teacher put a jar over the candle?



4. Where do the atoms come from that make the carbon dioxide and the water on the right side of the equation?



ACTIVITY

Question to Investigate

Where do the atoms in the products of a chemical reaction come from?

Materials for Each Student

- Atom model cut-outs (carbon, oxygen, and hydrogen)
- Sheet of colored paper or construction paper
- Colored pencils
- Scissors
- Glue or tape

Procedure

Prepare the Atoms

1. Color the carbon atoms black, the oxygen atoms red, and leave the hydrogen atoms white.
2. Use scissors to carefully cut out the atoms.

Build the Reactants

3. On a sheet of paper, place the atoms together to make the molecules of the reactants on the left side of the chemical equation for the combustion of methane.
4. Write the chemical formula under each molecule of the reactants. Also draw a + sign between the reactants.

Build the products

5. Draw an arrow after the second oxygen molecule to show that a chemical reaction is taking place.
6. Rearrange the atoms in the reactants to make the molecules in the products on the right side of the arrow.
7. Write the chemical formula under each molecule of the products. Also draw a + sign between the products.

Tell students that in a chemical reaction, the atoms in the reactants come apart, rearrange, and make new bonds to form the products.

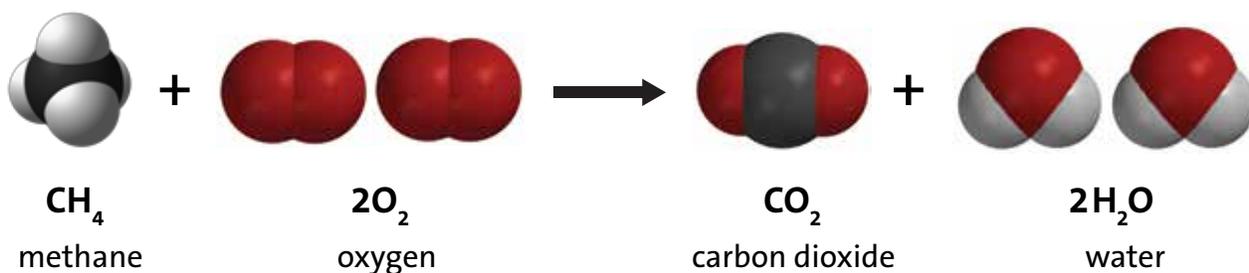
Represent the chemical equation

8. Use your remaining atoms to make the reactants again to represent the chemical reaction as a complete chemical equation.
9. Glue or tape the atoms to the paper to make a more permanent chemical equation of the combustion of methane.

EXPLAIN IT WITH ATOMS & MOLECULES

In a chemical equation, like the one below, you will notice that there are regular-sized numbers in front of some of the molecules and small numbers after certain atoms within a molecule. The little number is called the *subscript* and tells how many of a certain type of *atom* are in a molecule. The bigger number is called the *coefficient* and tells how many of a particular type of *molecule* there are.

If there is a coefficient in front of the molecule and a subscript after an atom, multiply the coefficient and the subscript to get the number of atoms. For example, in the products of the chemical reaction there are two water molecules, or $2\text{H}_2\text{O}$. The coefficient means that there are two molecules of water. The subscript means that each water molecule has two hydrogen atoms. Since each water molecule has 2 hydrogen atoms and there are two water molecules, there must be 4 (2×2) hydrogen atoms.

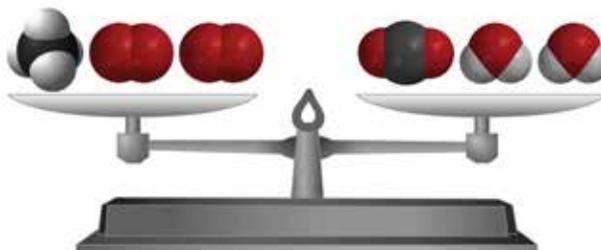


5. Count up the number of atoms on each side of the equation below and write this in the chart.

$\text{CH}_4 + 2\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$		
Atom	Reactant side	Product side
Carbon		
Hydrogen		
Oxygen		

6. Are atoms created or destroyed in a chemical reaction?

How do you know?



7. In a physical change, like changing state from a solid to a liquid, the substance itself doesn't really change. How is a chemical change different from a physical change?

TAKE IT FURTHER

Molecules made up of only carbon and hydrogen are called *hydrocarbons*. The candle and the hydrocarbons listed below react with oxygen in a chemical reaction called *combustion*.



CH₄
methane

Fuel in gas stoves
in many home
kitchens



C₃H₈
propane

Fuel in outdoor
gas grills



C₄H₁₀
butane

Fuel in disposable
lighters

8. Count the number of carbon, hydrogen, and oxygen atoms in the reactants and products of each equation to see if the equation is balanced. Record the number of each type of atom in each chart.

Combustion of Propane

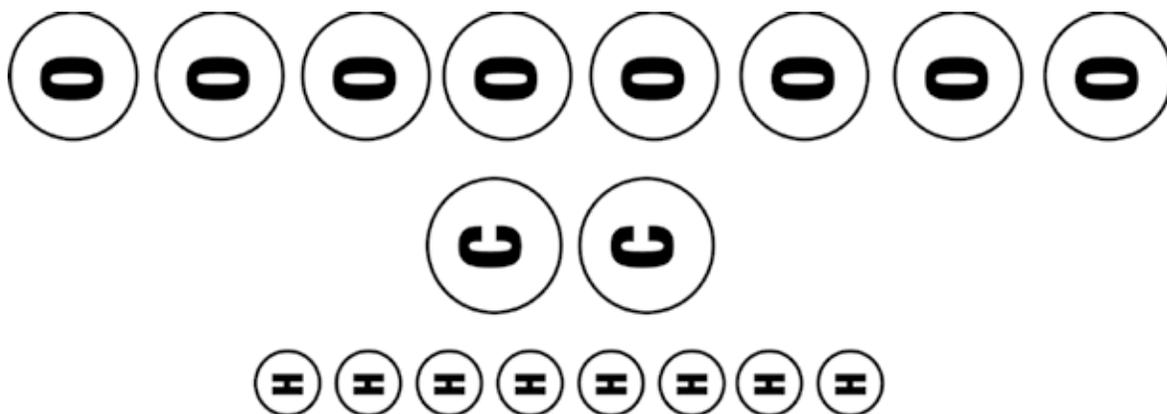


C₃H₈ + 5O₂ → 3CO₂ + 4H₂O		
Atom	Reactant side	Product side
Carbon		
Hydrogen		
Oxygen		

Combustion of Butane



$2\text{C}_4\text{H}_{10} + 13\text{O}_2 \longrightarrow 8\text{CO}_2 + 10\text{H}_2\text{O}$		
Atom	Reactant side	Product side
Carbon		
Hydrogen		
Oxygen		



Additional Teacher Background

Chapter 6, Lesson 1, p. 529

In the reaction between methane and oxygen, why do the atoms in methane and oxygen switch around and bond to form carbon dioxide and water?

This is a very good question. One of the main reasons has to do with the attractions that protons and electrons in one atom have for the electrons and protons in other atoms. If the conditions are right and atoms get a chance to bond with different atoms so that electrons are attracted to and closer to more protons, the atoms will switch around and rebond.

In methane, the 4 hydrogen atoms are each bonded to the 1 carbon atom by a covalent bond. Remember that this bond happens because the electron from the hydrogen atom is attracted to the protons in the carbon atom and an electron from the Carbon atom is attracted to the proton in the hydrogen atom. These attractions bring the atoms together and they end up sharing electrons to form a covalent bond.

The same is true about the other reactant – the oxygen molecules. The electrons from each oxygen atom are attracted to the protons in the other oxygen atom. This is true for both atoms so they come together and share electrons to form a covalent bond.

But the big question is why do the atoms in the methane and oxygen switch around and bond to different atoms in a chemical reaction to form carbon dioxide and water?

The answer is that the electrons and protons feel stronger attractions for each other and can get closer together if they switch around and bond to different atoms. The hydrogen atoms that are bonded to the carbon atom in methane switch around and end up being bonded to oxygen atoms to make water. The electrons in the hydrogen atoms were near 6 protons in the carbon atom but are now near 8 protons in oxygen. They are more attracted to 8 protons than to 6 so the trade is good to satisfy the attractions of electrons and protons.

The carbon atom that was attracted to the hydrogen atoms in the methane switched around and is now bonded to 2 oxygen atoms in the carbon dioxide. This is a really good trade. The electron from the carbon atom was near 1 proton from a hydrogen atom is now near 8 protons from an oxygen atom. This also satisfies the attractions of electrons and protons.

So one of the main reasons why atoms rearrange themselves in a chemical reaction is that by bonding to the other atoms, electrons and protons feel more attraction and get closer together.

Note:

This explanation works well for combustion reactions and other exothermic reactions but cannot fully explain endothermic reactions. In endothermic reactions, the electrons and protons in the products are actually in a less favorable situation for mutual attraction than they were in the reactants. There is a concept called entropy which helps explain why endothermic reactions occur but concepts related to entropy are beyond the scope of middle school chemistry and will not be introduced in this material.

Chapter 6, Lesson 2: Controlling the Amount of Products in a Chemical Reaction

Key Concepts

- Changing the amount of reactants affects the amount of products produced in a chemical reaction.
- In a chemical reaction, only the atoms present in the reactants can end up in the products.
- Mass is conserved in a chemical reaction.

Summary

Students will analyze the chemical equation for the reaction between vinegar (acetic acid solution) and baking soda (sodium bicarbonate). They will make the connection between the written chemical equation, the molecular model, and the real substances in the reaction. Students will see that the gas produced in the actual reaction is also written in the products of the equation. Students will also change the amount of one or more reactants and see how the change affects the amount of products.

Objective

Students will be able to explain that for a chemical reaction to take place, the bonds between atoms in the reactants are broken, the atoms rearrange, and new bonds between the atoms are formed to make the products. Students will be able to count the number of atoms on the reactant side and on the product side of a chemical equation. They will also be able to explain that the equal number of atoms on each side of the equation shows that mass is conserved during a chemical reaction. Students will also be able to explain, on the molecular level, why changing the amount of one or more reactants changes the amount of products. They will also be able to explain why simply adding more and more of one reactant will eventually not produce additional products.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Use vinegar in a well-ventilated room. Have students wash hands after the activity.

Materials for the Demonstrations

- Vinegar
- Baking soda
- Water
- Alka-Seltzer
- Detergent solution
- Graduated cylinder (50 mL)
- Graduated cylinder (100 mL)
- Measuring spoon (½ teaspoon)
- 1 clear plastic cup
- Small cup
- Dropper
- Plastic waste container

Materials for Each Group

- Vinegar in a cup
- Baking soda in a cup
- Detergent solution in a cup
- Dropper
- Graduated cylinder (50 mL)
- Measuring spoons (⅛, ¼, and ½ teaspoon)
- Plastic waste container

ENGAGE

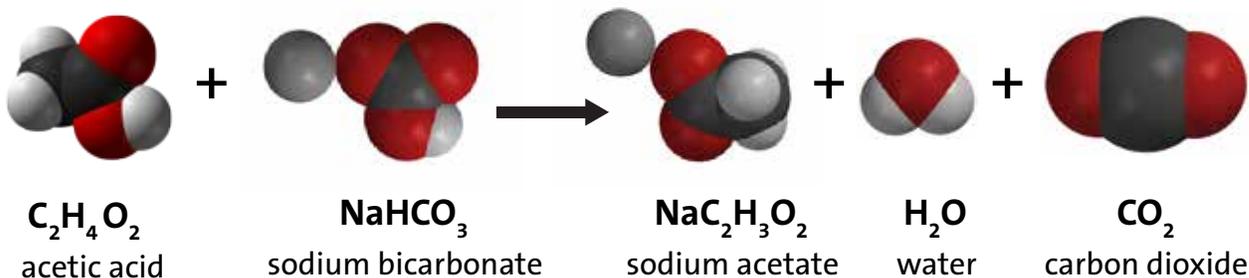
1. Have students look at the chemical equation for the vinegar and baking soda reaction as you discuss the reactants.

Remind students that in the last lesson, they learned that in a chemical reaction, certain atoms in the reactant molecules unbond from one another and then rearrange and rebond in different ways to form the products. Students saw that the same type and number of atoms were in the reactants as were in the products. Let students know that although the reaction in this lesson looks more complicated, these same principles still apply.

Project the image *Reactants*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson2#reactants

Show students the chemical equation for the reaction between vinegar and baking soda.



Ask students about vinegar:

- **Acetic acid mixed with water is vinegar. Usually vinegar is a solution of about 5% acetic acid and 95% water. When a reactant is in solution, the water is usually not listed as a reactant. Which atoms make up a molecule of acetic acid (vinegar)?**

Carbon, hydrogen, and oxygen (C, H, and O).

- **What do the little numbers below and to the right of each letter mean?**
These are the number of that particular atom in the acetic acid molecule. There are two carbon atoms, four hydrogen atoms, and two oxygen atoms in an acetic acid molecule.
- **Do you think every acetic acid molecule has this formula?**
Yes. The chemical formula for a substance is unique to that substance and defines what it is.

Ask students about baking soda:

- **Sodium bicarbonate is baking soda. What atoms is sodium bicarbonate made of?**
Sodium, hydrogen, carbon, and oxygen (Na, H, C, and O).
- **How many of each type of atom are there in the compound sodium bicarbonate?**
There are one sodium atom, one hydrogen atom, one carbon atom, and three oxygen atoms in every unit of sodium bicarbonate.



2. As a demonstration, combine vinegar and baking soda to show students the chemical reaction described in the equation.

Materials for the Demonstration

- Vinegar
- Graduated cylinder (50 mL)
- Baking soda
- Clear plastic cup

Procedure

- Use a graduated cylinder to measure 10 mL of vinegar.
- Place about $\frac{1}{2}$ teaspoon of baking soda in a clear plastic cup.
- While students watch, pour the vinegar into the baking soda.



Expected Results

Bubbles will form and rise up in the cup.

Ask students:

- **I combined a liquid and a solid, and you saw bubbling, which is made from gas. Do you think a chemical reaction occurred? Why?**

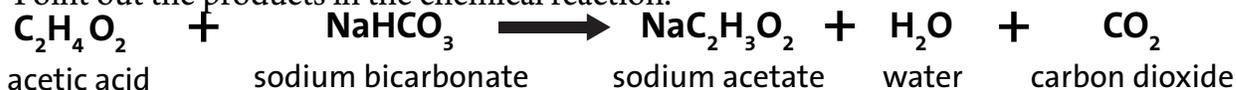
A chemical reaction occurred because a different substance was produced when the reactants combined.



Project the image *Products*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson2#products

Point out the products in the chemical reaction.



Ask students:

- **Look at the chemical equation. What is the gas produced in the chemical reaction between vinegar and baking soda?**

Carbon dioxide

- **What else is produced in this chemical reaction?**

When vinegar and baking soda react, atoms rearrange to form sodium acetate (the salty and sour flavor in salt-and-vinegar-flavored potato chips), water, and carbon dioxide.

Continue to project the chemical equation as you and students count the number of atoms on both the reactant side and product side of the equation.

3. Review the concept that mass is conserved in a chemical reaction.

Help students count the atoms in the reactants and in the products of the vinegar-baking soda reaction. Make sure students see that every type of atom on the left side of the equation is also on the right. Also be sure that they see that there is an equal number of each type on both sides of the equation.

Guide students as you answer the following questions together:

- **Is every type of atom on the left side of the equation also on the right side of the equation? Yes. Why?**

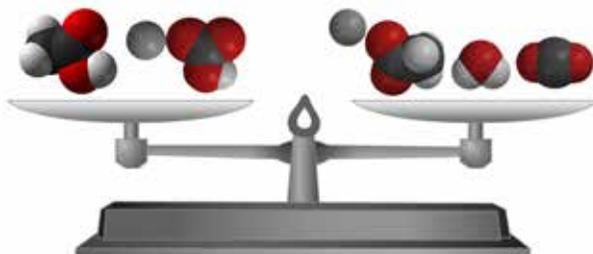
Atoms from the reactants rearrange to form the products. Atoms are not created or destroyed in a chemical reaction.

- **How many of each type of atom is on the reactant side of the equation?**

3 carbon atoms, 5 hydrogen atoms, 5 oxygen atoms, and 1 sodium atom.

- **How many of each type of atom is on the product side of the equation?**

3 carbon atoms, 5 hydrogen atoms, and 1 sodium atom.



5 carbon atoms, 5 hydrogen atoms, and 1 sodium atom.

Project the image *Mass is Conserved*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson2#mass_is_conserved



Point out that the type and number of atoms in the reactants and in the products are exactly the same. This is an important concept in chemistry: In a chemical reaction, all the atoms in the reactants end up in the products. When an equation of a chemical reaction is written, it is “balanced” to show this. A balanced chemical equation shows that no atoms are destroyed and no new atoms are created in the chemical reaction. Explain to students that another way of saying that no atoms are created or destroyed in a chemical reaction is that *mass is conserved*.

EXPLORE

4. As a demonstration, combine vinegar, detergent, and baking soda in a graduated cylinder so that foam rises and spills over the top.

Teacher Preparation for the Demonstration and for Each Group

- Make a detergent solution by adding 1 teaspoon of liquid dish detergent to 2 tablespoons of water. Divide this detergent solution equally into one small cup for each group.
- Place about 1 tablespoon of vinegar in a small cup for each group.
- Place about 2 teaspoons of baking soda in a small cup for each group.

Materials for the Demonstration

- Vinegar
- Baking soda
- Detergent solution
- Dropper
- Graduated cylinder (50 mL)
- Measuring spoon ($\frac{1}{2}$ teaspoon)

- Plastic waste container
- Small cup

Procedure

1. Use a graduated cylinder to measure 10 mL of vinegar.
2. Pour the vinegar in a small cup and add 1 drop of detergent. Swirl gently to mix.
3. Add $\frac{1}{2}$ teaspoon of baking soda to the empty graduated cylinder.
4. Place the graduated cylinder in a plastic waste container.
5. Pour the vinegar and detergent from the cup into the graduated cylinder. Have students observe the level of foam in the graduated cylinder.
6. Rinse the graduated cylinder over the waste container.



Expected Results

White foam will rise up in the graduated cylinder and overflow.

5. Discuss how to change the amount of foam produced so that it rises to the top of the cylinder without overflowing.

Ask students:

- What could you change to create a foam that rises as close as possible to the top of the cylinder without overflowing?

Students might mention variables such as:

- The amount of vinegar, detergent, or baking soda.
- The order in which the substances are added to the graduated cylinder.

Explain that the amount of detergent should not be varied in this activity because it is used as an indicator to help measure the amount of gas produced in the reaction. Also, the baking soda should be added to the cylinder first. The vinegar is poured in afterwards to cause better mixing of reactants.

Remind students that 10 mL of vinegar and $\frac{1}{2}$ teaspoon of baking soda caused the foam to overflow. Students should consider these amounts as they plan how much of each reactant they will use as they start their trials.

Ask students:

- **Can you add the baking soda first and then the vinegar on one trial and then switch it for the other trials? No. Why not?**

Every test should be conducted the same way. For example, in the demonstration baking soda was placed in the graduated cylinder before the vinegar and detergent were added. This method mixes the baking soda and vinegar well. All new trials should be conducted this same way.

- **Should you rinse the graduated cylinder each time? Yes. Why?**

Any products or leftover reactants that remain in the graduated cylinder may affect the next reaction. It is best to rinse the cylinder after each trial.

- **How will you remember the amounts you used in each trial?**

Students should realize the necessity of making and recording accurate measurements in the chart provided.

Give each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



6. Have each group experiment with different amounts of vinegar and baking soda in order to get the foam to rise to the top of the graduated cylinder without overflowing.

Tell students that they should try to get the foam to stop as close as possible to the top of the cylinder without overflowing. You may choose to limit students to a maximum of three tries or let them experiment further if time and supplies allow.



Question to Investigate

How can you make just the right amount of foam that rises to the top of the graduated cylinder without overflowing?

Materials for Each Group

- Vinegar in a cup
- Baking soda in a cup
- Detergent solution in a cup
- Dropper
- Graduated cylinder (50 mL)

- Measuring spoons ($\frac{1}{8}$, $\frac{1}{4}$, and $\frac{1}{2}$ teaspoon)
- Plastic waste container

Procedure

1. Decide on how much vinegar and baking soda you will use and write these amounts in the chart on the activity sheet.
2. Use a graduated cylinder to measure the amount of vinegar your group agreed on.
3. Pour the vinegar in a small cup and add 1 drop of detergent. Swirl gently to mix.
4. Add the amount of baking soda your group agreed on to the empty graduated cylinder.
5. Place the graduated cylinder in a plastic waste container.
6. Pour the vinegar and detergent from the cup into the graduated cylinder. Observe the level of foam in the graduated cylinder.
7. Rinse the graduated cylinder over the waste container.



Expected Results

Using $\frac{1}{8}$ teaspoon of baking soda, 5 mL of vinegar, and 1 drop of detergent will probably cause the foam to rise to the top of the cylinder without overflowing. Results may vary.

Have groups share their findings about the amounts of baking soda and vinegar that came closest to reaching the top of the cylinder. Did each group use similar amounts of baking soda and vinegar?

EXPLAIN

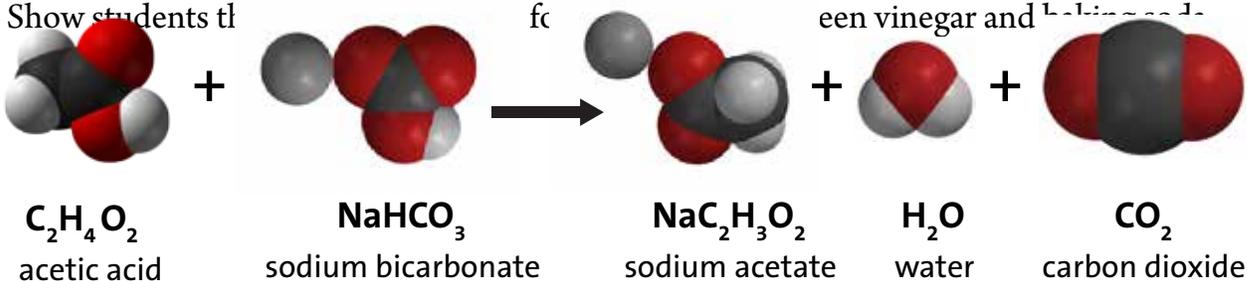
7. Discuss why adjusting the amount of reactants affects the amount of products.



Project the image *Controlling Amount of Products Formed*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson2#controlling

Show students th



Ask students:

- **Why, on the molecular level, does changing the amount of baking soda or vinegar affect the amount of carbon dioxide gas produced?**

Products are made from the reactants, so adding more reactants will produce more of the products.

The important point for students to realize is that atoms from *both* reactants are necessary to produce the products. Using less baking soda, for instance, produces less carbon dioxide gas because there are fewer atoms from the baking soda to produce the carbon dioxide. In general, using more of one or more reactants will result in more of one or more products. Using less of one or more reactants will result in less of one or more products. Let students know that this principle has limits.

Note: It is not necessary for students at the middle school level to know which particular atom in the reactants ended up in which product. It might seem strange, but sometimes a product can be made up of atoms from only one reactant. In the vinegar and baking soda reaction, the atoms in the CO_2 only come from the sodium bicarbonate.

Ask students:

- **What would you do if you wanted to make more carbon dioxide?**
Add more vinegar and more baking soda.
- **Could you just keep adding more and more baking soda to the same amount of vinegar to get more carbon dioxide?**
No. This might work for a while, as long as there is extra vinegar, but eventually there would be no atoms left of vinegar to react with the extra baking soda, so no more carbon dioxide would be produced.

EXTEND

8. **Do a demonstration using Alka-Seltzer or a similar effervescent tablet in water to show that citric acid reacts with sodium bicarbonate to produce carbon dioxide gas.**



Tell students that an Alka-Seltzer tablet contains aspirin, sodium bicarbonate, and citric acid. Citric acid interacts with the sodium bicarbonate similar to the way the acetic acid in vinegar interacts with sodium bicarbonate.

Ask students to make a prediction:

- **What will happen when an Alka-Seltzer tablet is placed in water with a drop of detergent solution?**

Materials for the Demonstration

- Alka-Seltzer
- Water
- Graduated cylinder (100 mL)
- Detergent solution
- Dropper

Procedure

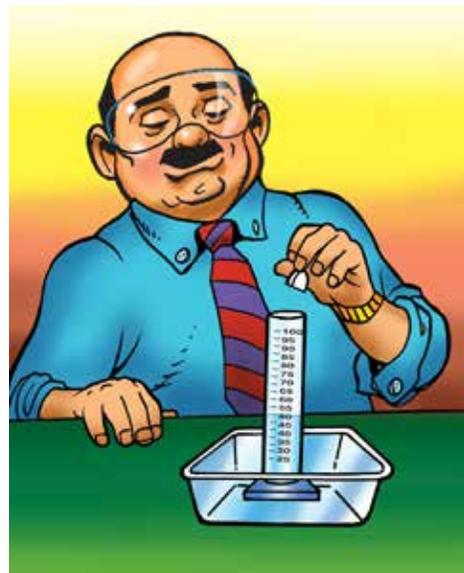
1. Place 50 mL of water in a 100 mL graduated cylinder.
2. Add 1 drop of detergent solution and swirl gently to mix.
3. Drop half of an Alka-Seltzer tablet in the graduated cylinder.
4. Place the graduated cylinder in a waste container.

Expected Results

White foam will rise up in the graduated cylinder and overflow as the tablet becomes smaller and smaller.

Ask students:

- **Do you think this is a chemical reaction? Yes. Why?**
Because a gas was produced. This gas was not in one of the reactants, so it must have been produced during the chemical reaction.
- **Why do you think this reaction is similar to the reaction of vinegar and baking soda?**
Citric acid and vinegar are both acids and so interact with sodium bicarbonate in a similar way to produce carbon dioxide gas.



Activity Sheet
Chapter 6, Lesson 2
Controlling the Amount of
Products in a Chemical Reaction

Name _____

Date _____

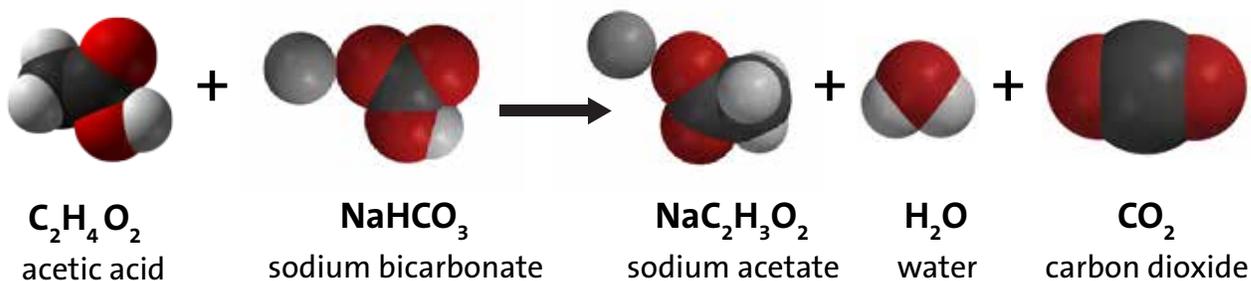
DEMONSTRATION

1. Your teacher combined a liquid (vinegar) and a solid (baking soda). You observed bubbling, which is made from gas. Do you think a chemical reaction occurred?

Why?



2. Look at the chemical equation for the reaction between vinegar and baking soda to answer the following questions.



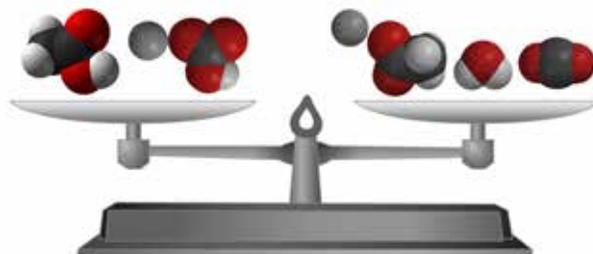
What are the *reactants* in this chemical reaction?

What are the *products* in this chemical reaction?

How many of each type of atom appears on each side of the chemical equation?

$C_2H_4O_2 + NaHCO_3 \longrightarrow NaC_2H_3O_2 + H_2O + CO_2$		
Atom	Reactant side	Product side
Carbon		
Hydrogen		
Oxygen		
Sodium		

3. What does the statement “*Mass is conserved during a chemical reaction*” mean?



ACTIVITY

Question to investigate

How can you make just the right amount of foam that rises to the top of the graduated cylinder without overflowing?



Materials for each group

- Vinegar in a cup
- Baking soda in a cup
- Detergent solution in a cup
- Dropper
- Graduated cylinder (50 mL)
- Measuring spoons ($\frac{1}{8}$, $\frac{1}{4}$, and $\frac{1}{2}$ teaspoon)
- Plastic waste container

Procedure

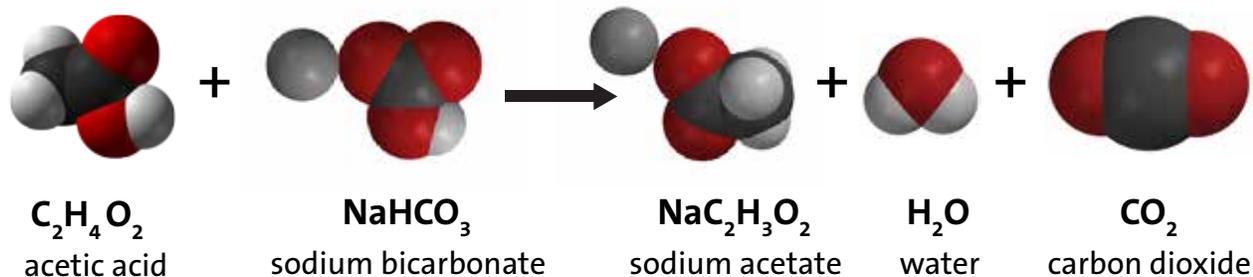
1. Decide on how much vinegar and baking soda you will use and write these amounts in the chart on the activity sheet.
2. Use a graduated cylinder to measure the amount of vinegar your group agreed on.
3. Pour the vinegar in a small cup and add 1 drop of detergent. Swirl gently to mix.
4. Add the amount of baking soda your group agreed on to the empty graduated cylinder.
5. Place the graduated cylinder in a plastic waste container.
6. Pour the vinegar and detergent from the cup into the graduated cylinder. Observe the level of foam in the graduated cylinder.
7. Rinse the graduated cylinder over the waste container.



Adjust the amounts of baking soda and vinegar to create just enough foam to rise to the top of the graduated cylinder without overflowing.

	Demonstration	First try	Second try	Third try
Vinegar	10 mL			
Baking soda	$\frac{1}{2}$ teaspoon			
Detergent	1 drop	1 drop	1 drop	1 drop
How close did the foam get to the top of the cylinder?	Overflowed			

EXPLAIN IT WITH ATOMS & MOLECULES



4. Why, on the molecular level, does changing the amount of baking soda or vinegar affect the amount of carbon dioxide gas produced?
5. What would you do if you wanted to make more carbon dioxide?
6. Could you just keep adding more and more baking soda to the same amount of vinegar to get more carbon dioxide?

Why or why not?

TAKE IT FURTHER

7. An Alka-Seltzer tablet contains aspirin, sodium bicarbonate, and citric acid. Your teacher placed an Alka-Seltzer tablet in water with a drop of detergent. Do you think placing an Alka-Seltzer in water causes a chemical reaction?

Why?



Chapter 6, Lesson 3: Forming a Precipitate

Key Concepts

- The ions or molecules in two solutions can react to form a solid.
- A solid formed from two solutions is called a *precipitate*.

Summary

Students will combine two clear colorless solutions (baking soda solution and calcium chloride solution) and see the formation of a solid and a gas. Students will analyze the chemical equation for the reaction and see that all atoms in the reactants end up in the products. They will make the connection between the chemical equation and the real substances and see that the solid and gas produced in the actual reaction are also in the products of the equation.

Objective

Students will be able to explain that for a chemical reaction to take place, the reactants interact, bonds between certain atoms in the reactants are broken, the atoms rearrange, and new bonds between the atoms are formed to make the products. Students will also be able to explain that this definition applies to the production of a solid called a *precipitate*.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. If you do the demonstration with copper II sulfate solution, ammonia, and hydrogen peroxide at the end of the lesson, pour the resulting solution and precipitate in a cup or beaker and allow it to evaporate. Put the small amount of solid in a paper towel and dispose in the trash or use a disposal method required by local regulations. Sodium carbonate may irritate skin. Wash hands after the activity. Magnesium sulfate dust can irritate respiratory tract.

Materials for the Demonstrations

- Sodium carbonate
- Epsom salt (magnesium sulfate)
- 2 clear plastic cups
- Test tube
- Water
- Copper II sulfate
- Household ammonia
- Hydrogen peroxide (3%)
- Graduated cylinder
- 2 droppers

Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Measuring spoon ($\frac{1}{2}$ teaspoon) or balance
- 2 clear plastic cups
- Masking tape
- Pen

About the Materials

Copper II sulfate is available from various chemical suppliers, including Sargent Welch, Product #WLC94770-06 or Flinn Scientific, Product #C0110. Follow all safety precautions regarding use, storage, and disposal of copper II sulfate and sodium carbonate. Sodium carbonate is Product #WLC94291-06 or #S0052.

ENGAGE

1. Do a demonstration by combining two clear colorless solutions that produce a white solid and introduce the term *precipitate*.



Materials for the Demonstration

- Magnesium sulfate (Epsom salt)
- Sodium carbonate
- Water
- 2 clear plastic cups
- 1 tablespoon
- 1 teaspoon

Teacher Preparation

- Pour 100 mL of water in one clear plastic cup and add 10 g (about 1 tablespoon) of magnesium sulfate. Stir until the solution is clear.
- Pour 50 mL of water in another clear plastic cup and add 5 g (about 1 teaspoon) of sodium carbonate. Stir until the solution is clear.

Procedure

1. Hold up the two clear colorless solutions and slowly pour the smaller amount into the larger.



Expected Results

Particles of a white solid will form.

Ask students:

- **Would you consider adding a sodium carbonate solution to a magnesium sulfate solution a chemical reaction?**

Yes.

Why or why not?

Combining the two clear colorless liquids is a chemical change because a different solid substance is formed.

Tell students that a precipitate is an insoluble solid that forms when two solutions are combined and react chemically. Insoluble means that the solid will not dissolve.

Give Each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. Have students combine two liquids to observe another precipitate.

Question to Investigate

How do you know when a precipitate is formed in a chemical reaction?

Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Measuring spoon ($\frac{1}{2}$ teaspoon) or balance
- 2 clear plastic cups
- Masking tape
- Pen



Note: If you would like students to practice using a balance to weigh grams, have them weigh two grams each of baking soda and calcium chloride.

Procedure

1. Use masking tape and a pen to label 2 plastic cups *baking soda solution* and *calcium chloride solution*.
2. Use a graduated cylinder to add 20 mL of water to each cup.
3. Add 2 g (about $\frac{1}{2}$ teaspoon) of calcium chloride to the water in its labeled cup. Swirl until as much of the calcium chloride dissolves as possible.
4. Add 2 g (about $\frac{1}{2}$ teaspoon) of baking soda to the water in its labeled cup. Swirl until as much



of the baking soda dissolves as possible. There may be some undissolved baking soda remaining in the bottom of the cup.

- Carefully pour the baking soda solution into the calcium chloride solution. Try not to pour in any undissolved baking soda. Observe.

Expected Results

Bubbling and a white precipitate appear.

3. Discuss student observations.

Ask students:

- What did you observe when you mixed the baking soda solution and the calcium chloride solution?**
The solutions bubbled and little white particles of solid formed.
- Did you observe a precipitate?**
Yes. The white particles appeared after the two solutions were combined.
- Do you think this was a chemical reaction? Yes. Why?**
The two substances that were combined were liquids and the substances that were produced were a solid and a gas. These products seem to be different from the reactants.

EXPLAIN

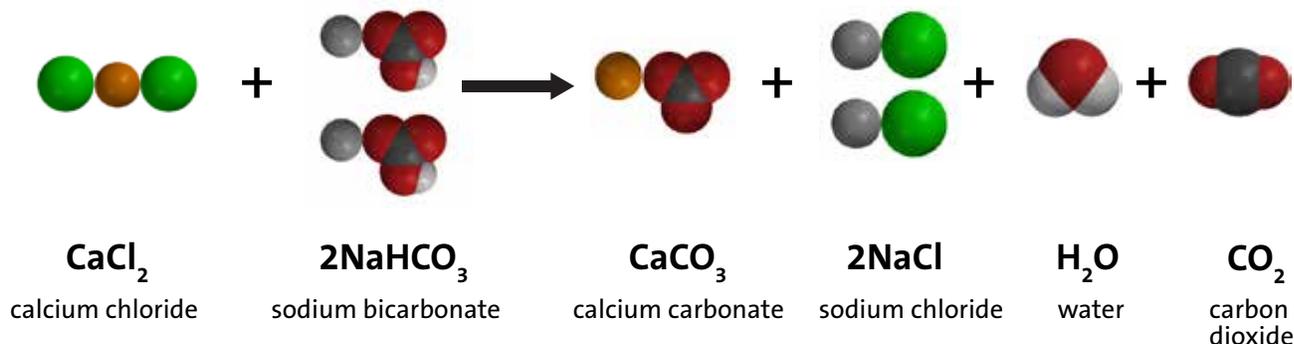
4. Discuss the products produced in this chemical reaction.

Remind students that in the chemical reactions they have seen so far, certain atoms in the reactant molecules unbond from one another and then rearrange and rebond in different ways to form the products. They saw that the same type and number of atoms were in the reactants as were in the products.



Project the image *Calcium Chloride and Sodium Bicarbonate*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson4#calcium_chloride



Ask students:

- **What products of the reaction do you recognize?**
Students should recognize sodium chloride (NaCl), water (H₂O), and carbon dioxide (CO₂).
- **Look at the product side of the chemical equation. What gas is produced in the chemical reaction?**
Carbon dioxide gas.
- **What do you think is the precipitate?**
The salt and water are clear and colorless as a solution, so the precipitate must be CaCO₃, which is calcium carbonate. Tell students that calcium carbonate is ordinary chalk.
- **How many of each type of atom is on the reactant side of the equation?**
1 calcium atom, 2 chlorine atoms, 2 sodium atoms, 2 hydrogen atoms, 2 carbon atoms, and 6 oxygen atoms.
- **How many of each type of atom is on the product side of the chemical equation?**
1 calcium atom, 2 chlorine atoms, 2 sodium atoms, 2 hydrogen atoms, 2 carbon atoms, and 6 oxygen atoms.
- **Is this a balanced chemical equation? Yes. Why?**
The same type and number of atoms are in the reactants and products.

Make sure students see that every type of atom on the left side of the equation is also on the right. Also be sure that they see that there is an equal number of each type on both sides of the equation.

EXPLORE

5. Separate the products to show that the precipitate is a solid.

Ask students:

- **How do you think we could separate the precipitate from the other products?**

Question to Investigate

Can you separate the calcium carbonate from the rest of the products?



Materials for Each Group

- Coffee filter or paper towel
- Tall clear plastic cup

Procedure

1. Use a large enough coffee filter (or paper towel) so that you can push it about $\frac{1}{3}$ of the way into the cup and still have enough left to hold it around the outside of the cup.
2. While holding the coffee filter in place, pour the products into the center of the coffee filter.
3. Allow the liquid to drip through the filter. This may take a while.
4. Set the precipitate aside and allow the water to evaporate.



Expected results

A white solid will remain in the coffee filter. After the water evaporates, the calcium carbonate will be a white powder.

Note: If you'd like to separate the sodium chloride from the water that flowed through the filter, pour the liquid into a clean empty cup and allow the water to evaporate for a few days. As the water evaporates, students will begin to see cubic-shaped salt crystals forming in the solution. Eventually only salt crystals will remain in the cup.

Ask students:

- **What is the solid white substance on the paper?**
Calcium carbonate (chalk).
- **Is filtering out the calcium carbonate and allowing the water to evaporate a chemical change or a physical change? Why?** Physical change. Why? These substances were already present in the water, so no new chemicals are made.
- **How do you know that a chemical reaction occurred when you combined baking soda solution and calcium chloride solution?**
Because different products were formed.

EXTEND

6. **Do a demonstration to show students another example of a precipitate and a color change.**

Tell students that you will show them another reaction that forms a precipitate and a little something extra.

Materials for the Demonstration

- Copper II sulfate
- Household ammonia

- Hydrogen peroxide (3%)
- Water
- Graduated cylinder
- Test tube
- 2 droppers
- 1 clear plastic cup (empty)

Note: The copper compound is called “copper II” because copper can make different types of ions. It can lose one electron and be just Cu^+ or it can lose two electrons and be Cu^{2+} . This type of copper ion is called copper II. The “sulfate” in copper II sulfate is also an ion. This ion is made up of more than one atom. It is one of the polyatomic ions discussed in Chapter 4, Lesson 3. The sulfate ion is made up of a sulfur atom bonded to four oxygen atoms and is treated as one ion (SO_4^{2-}).

Teacher Preparation

Make a copper II sulfate solution by adding 5 g of copper II sulfate to 50 mL of water.

Procedure

1. Pour 15–20 mL of copper II sulfate solution into a test tube so it is about $\frac{1}{2}$ full.
2. Add about 10–20 drops of ammonia.
3. Add about 10–20 drops of hydrogen peroxide.

Expected Results

After adding the ammonia, a whitish precipitate will form at the top of the copper II sulfate solution. As more ammonia is added, the color on top of the liquid will change to a deeper darker blue. As the hydrogen peroxide is added, the dark blue area at the top of the solution will turn dark green and a dark precipitate will form.

Note: The details of the chemical reactions that produce the different precipitates and different color changes are fairly complicated. The main idea for students is that atoms or groups of atoms in the reactants rearranged and bonded in different ways to form different substances in the products.

Let students know that when they see the production of a gas, a precipitate, or a color change, that this is evidence that a chemical reaction has taken place.

Ask students:

- **How can you tell that something new was made when the copper II sulfate and ammonia reacted?**

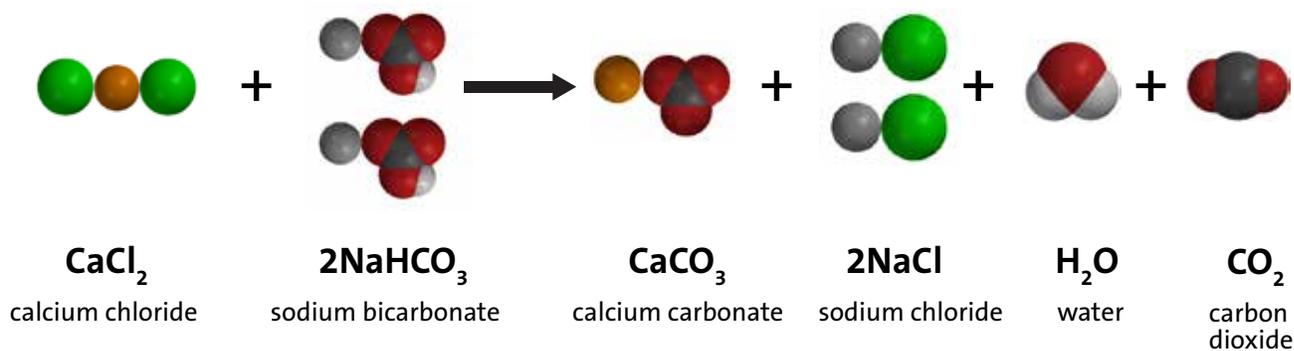
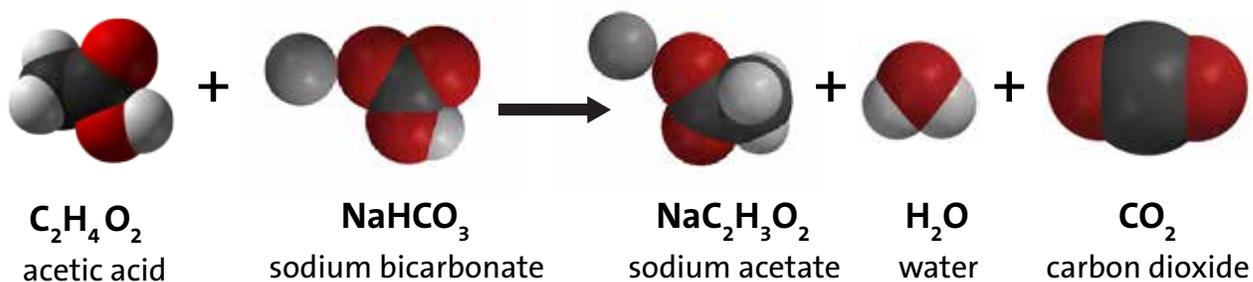
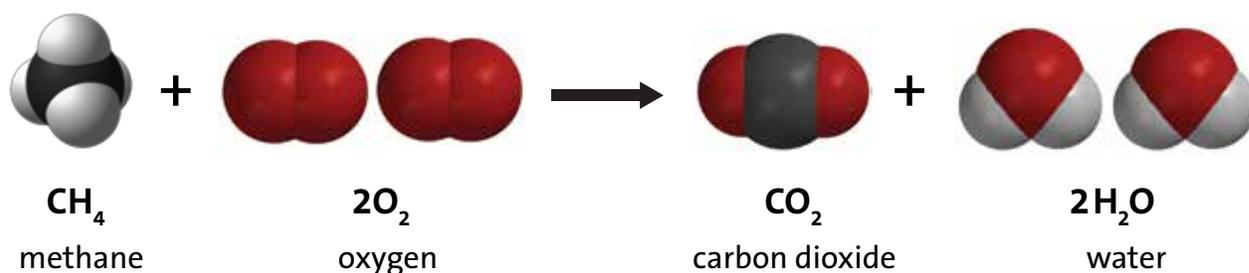
A precipitate was produced.

- How can you tell that something new was made when these substances reacted with hydrogen peroxide?

The color change and other precipitate are evidence of another chemical reaction.

In-Class or At-Home Project.

Have students use objects such as gum drops, beads, M&Ms, Legos, or other small objects to represent the atoms in two of the three reactions they have explored so far. Students can tape or glue the objects to poster board and write down the chemical formula for the reactants and products.



Activity Sheet
Chapter 6, Lesson 3
Forming a Precipitate

Name _____

Date _____

DEMONSTRATION

1. Your teacher combined two clear colorless solutions. One was a sodium carbonate solution and the other was a magnesium sulfate solution. Do you think a chemical reaction occurred when these two substances were combined?

Why or why not?



2. What is a precipitate?

ACTIVITY

Question to Investigate

How do you know when a precipitate is formed in a chemical reaction?



Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Measuring spoon ($\frac{1}{2}$ teaspoon) or balance
- 2 clear plastic cups
- Masking tape
- Pen

Procedure

1. Use masking tape and a pen to label 2 plastic cups *Baking Soda Solution* and *Calcium Chloride Solution*.
2. Use a graduated cylinder to add 20 mL of water to each cup.
3. Add 2 g (about $\frac{1}{2}$ teaspoon) of calcium chloride to the water in its labeled cup. Swirl until as much of the calcium chloride dissolves as possible.
4. Add 2 g (about $\frac{1}{2}$ teaspoon) of baking soda to the water in its labeled cup. Swirl until as much of the baking soda dissolves as possible. Some undissolved baking soda may remain in the bottom of the cup.
5. Carefully pour the baking soda solution into the calcium chloride solution. Try not to pour in any undissolved baking soda. Observe.

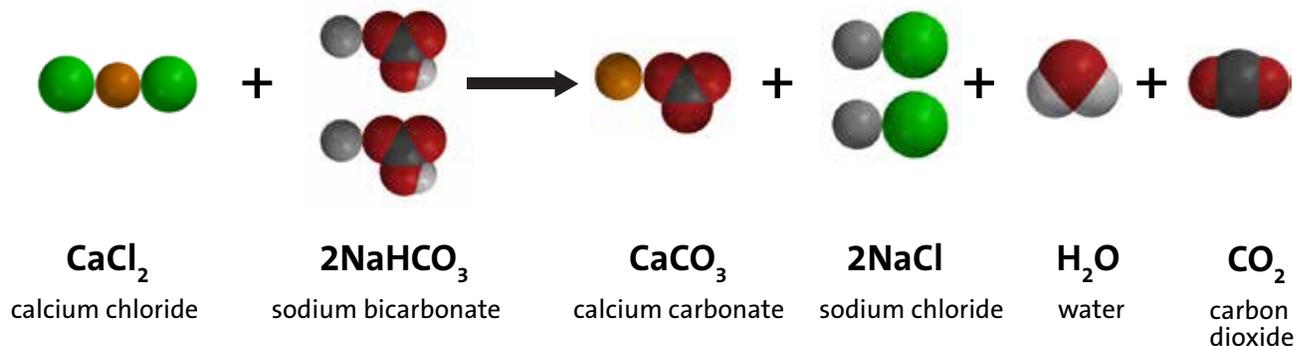


3. What do you observe when you combine baking soda solution and calcium chloride solution?

4. How do you know that a chemical reaction occurs when you combine baking soda solution and calcium chloride solution?

EXPLAIN IT WITH ATOMS & MOLECULES

5. Take a look at the chemical equation for the reaction between calcium chloride and sodium bicarbonate and answer the following questions.



What gas is produced in the chemical reaction?

What do you think is the precipitate?

How many of each type of atom appears on each side of the chemical equation?		
Atom	Reactant side	Product side

ACTIVITY

Question to Investigate

Can you separate the calcium carbonate from the rest of the products?



Materials for Each Group

- Coffee filter or paper towel
- Tall clear plastic cup

Procedure

1. Use a large enough coffee filter (or paper towel) so that you can push it about $\frac{1}{3}$ of the way into the cup and still have enough left to hold it around the outside of the cup.
 2. While holding the coffee filter in place, pour the products into the center of the coffee filter.
 3. Allow the liquid to drip through the filter. This may take a while.
 4. Set the precipitate aside and allow the water to evaporate.
6. Is filtering the calcium carbonate and allowing the water to evaporate a chemical change or a physical change?



Why?

TAKE IT FURTHER

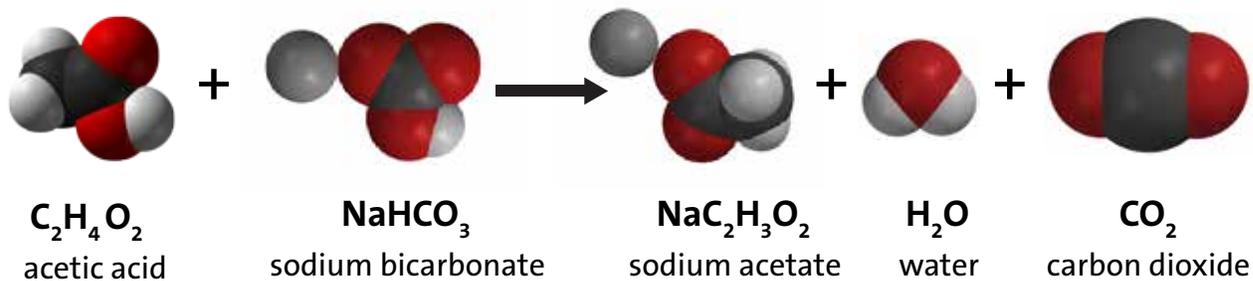
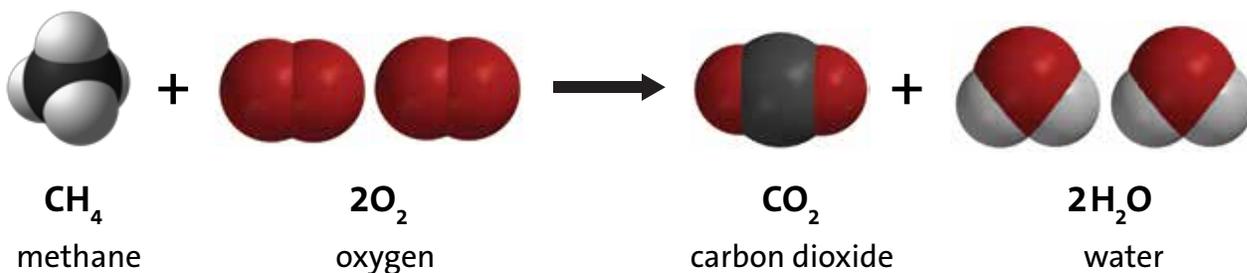
Your teacher added drops of ammonia to copper II sulfate solution.

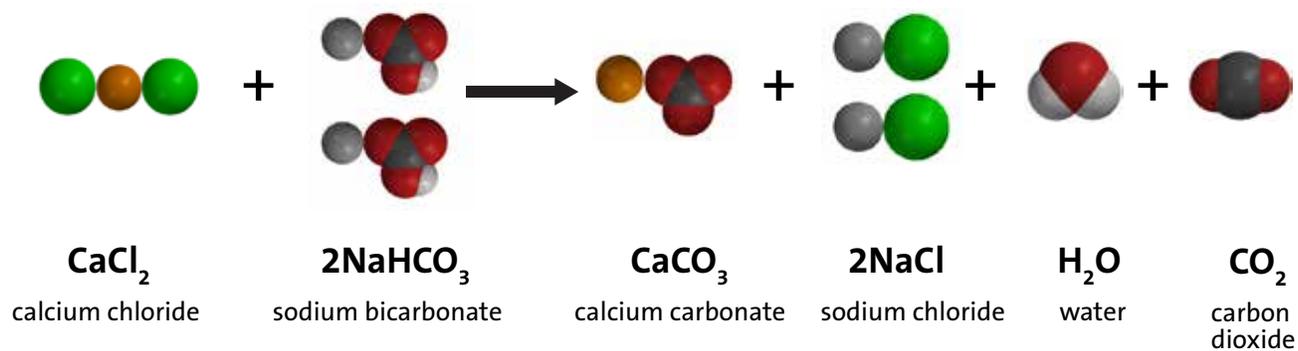
7. How can you tell that something new was made when the copper II sulfate and ammonia reacted?

8. How can you tell that something new was made when these substances reacted with hydrogen peroxide?



9. Use objects such as gum drops, beads, M&Ms, Legos, or other small objects to represent the atoms in two of the three chemical reactions you have covered in chapter 6. The three chemical equations are written below. Tape or glue the objects to poster board and write down the chemical formula for the reactants and products.





Chapter 6, Lesson 4: Temperature and the Rate of a Chemical Reaction

Key Concepts

- Reactants must be moving fast enough and hit each other hard enough for a chemical reaction to take place.
- Increasing the temperature increases the average speed of the reactant molecules.
- As more molecules move faster, the number of molecules moving fast enough to react increases, which results in faster formation of products.

Summary

Students will make the same two clear colorless solutions (baking soda solution and calcium chloride solution) from Lesson 3. They will help design an experiment to see if the temperature of the solutions affects how fast they react. Students will then try to explain, on the molecular level, why the temperature affects the rate of the reaction.

Objective

Students will be able to identify and control variables to design an experiment to see if temperature affects the rate of a chemical reaction. Students will be able to explain, on the molecular level, why the temperature of the reactants affects the speed of the reaction.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. Use caution when handling hot water.

Materials for the Demonstration

- Hot water in an insulated cup
- Ice water in an insulated cup
- 2 glow sticks

Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Balance or measuring spoon (teaspoon)
- 2 wide (9 oz) clear plastic cups
- 4 small clear plastic cups
- 2 plastic deli-style containers
- Hot water (40–50 °C)
- Cold water (0–5 °C)
- Masking tape
- Pen

ENGAGE

1. Do a demonstration with glow sticks to introduce the idea that temperature can affect the rate of a chemical reaction.

Question to Investigate

How does warming or cooling a glow stick affect its chemical reaction?



Materials for the Demonstration

- Hot water in an insulated cup
- Ice water in an insulated cup
- 2 glow sticks

Teacher preparation

Be sure not to start the glow sticks as you prepare for the demonstration. Place one glow stick in hot water and another in ice water before students arrive. The glow sticks will need to be in the water for at least a couple of minutes before the demonstration.

Tell students that you have heated one glow stick and cooled another.

Ask students:

- **How do you start a glow stick?**
Bend the stick until you hear a popping sound.
- **What should you do if you want your glow stick to last longer?**
Place the glow stick in the freezer when you are not using it.

Explain that when students bend the stick to start it, they are breaking a small container filled with a chemical inside the light stick. Once broken, the chemicals, which were separate, combine and react with each other. If putting a glow stick in the freezer makes it last longer, temperature may have something to do with the rate of the chemical reaction.

Procedure

1. Remove the glow sticks from both the hot and cold water.
2. Have two students bend and start the glow sticks.
3. Show students both glow sticks and ask them what they observe. You may pass the sticks around the class so that they can feel the difference in temperature.



Expected Results

The warm glow stick will be brighter than the cold one.

Ask students:

- **How can you tell whether the chemical reaction is happening faster or slower in each glow stick?**

The warm glow stick is brighter, so the chemical reaction may be happening faster. The cool glow stick is not as bright, so the chemical reaction may be happening slower.

- **Some people place glow sticks in the freezer to make them last longer. Why do you think this works?**

The chemical reaction that happens in a light stick is slower when cold.

- **Do you think that starting with warmer reactants increases the rate of other chemical reactions? Why?**

It is reasonable to think that temperature will affect the rate of other chemical reactions because temperature affected this reaction.

EXPLORE

2. **Ask students how they could set up an experiment to find out if the temperature of the reactants affects the speed of the reaction.**

Review with students the chemical reactions they did in the last lesson. They combined a calcium chloride solution with a baking soda solution. They saw that when the solutions were combined, a solid and a gas were produced. Tell students that they will warm and cool a calcium chloride solution and a baking soda solution to find out whether temperature affects the rate of the chemical reaction.

Ask students:

- **How many sets of solutions should we use?**

Students should use two sets—one that is heated and one which is cooled. Tell students that they will use hot and cold water baths, like in the demonstration, to warm and cool the solutions.

- **Should the warmed samples of baking soda solution and calcium chloride solution be the same as the samples that are cooled?**

Yes. Samples of the same solution should be used and the same amount of cold solution as warm solution should be used.

- **In the glow stick demonstration, we could tell that the reaction was happening faster if the light was brighter. How can we tell if the reaction is happening faster in this chemical reaction?**

The chemical reaction is happening faster, if more products are produced. We should look for more bubbling (carbon dioxide) and more white precipitate (calcium carbonate).

3. Have students warm a pair of reactants and cool another and compare the amount of products in each reaction.

Question to Investigate

Does the temperature of the reactants affect the rate of the chemical reaction?

Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Balance or measuring spoon ($\frac{1}{2}$ teaspoon)
- 4 small plastic cups
- 2 plastic deli-style containers
- Hot water (about $50\text{ }^{\circ}\text{C}$)
- Cold water ($0\text{--}5\text{ }^{\circ}\text{C}$)
- Masking tape
- Pen

Procedure

Make the Baking Soda Solution

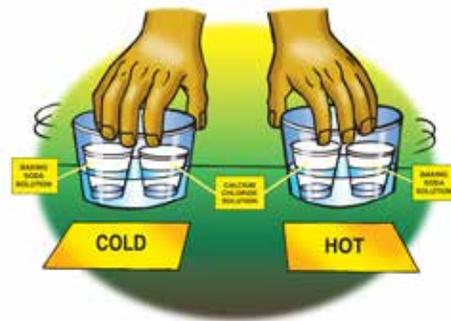
1. Use masking tape and a pen to label 2 small plastic cups baking soda solution, and 2 small plastic cups calcium chloride solution.
2. Use a graduated cylinder to add 20 mL of water to one of the baking soda solution cups.
3. Add 2 g (about $\frac{1}{2}$ teaspoon) of baking soda to the water in its labeled cup. Swirl until as much of the baking soda dissolves as possible. (There may be some undissolved baking soda in the bottom of the cup.)
4. Pour half of your baking soda solution into the other baking soda solution cup.

Make the Calcium Chloride Solution

5. Use a graduated cylinder to add 20 mL of water to one of the calcium chloride solution cups.
6. Add 2 g (about $\frac{1}{2}$ teaspoon) of calcium chloride to the water in its labeled cup. Swirl until the calcium chloride dissolves.
7. Pour half of your calcium chloride solution into the other calcium chloride solution cup.

Heat and Cool the Solutions

8. Pour hot water into one plastic container and cold water into the other until each is about $\frac{1}{4}$ filled. The water should not be very deep. These are your hot and cold water baths.
9. Place and hold one cup of baking soda solution and one cup of calcium chloride solution in the hot water. Gently swirl the cups in the water for about 30 seconds to heat up the solutions.
10. Your partner should place and hold one cup of baking soda solution and one cup of calcium chloride solution in the cold water. Gently swirl the cups in the water for about 30 seconds to cool the solutions.



Combine the Solutions

11. At the same time, you and your partner should combine the two warm solutions with each other, and the two cold solutions with each other.

Expected Results

The warm solutions will react immediately and much faster than the cold solutions. Bubbling and particles of white solid will quickly appear in the combined warm solutions. The cold solutions will turn a cloudy grayish and stay that way for a while. Eventually the combined solutions will gradually turn white and bubble, and particles of white solid will appear.

4. Discuss student observations.

Ask students:

- **Does the temperature of the reactants affect the rate of the chemical reaction?**
Yes. The warm solutions react much faster than the cold solutions.

EXPLAIN

5. **Show students that the faster moving molecules in the warm reactants hit each other with more energy and so are more likely to react.**

Ask students:

- **On the molecular level, why do you think the warm solutions react faster than the cold solutions?**

Explain to students that for reactant molecules to react, they need to contact other reactant molecules with enough energy for certain atoms or groups of atoms to come apart and recombine to make the products. When the reactants are heated, the average kinetic

energy of the molecules increases. This means that more molecules are moving faster and hitting each other with more energy. If more molecules hit each other with enough energy to react, then the rate of the reaction increases.

Project the animation *Molecules collide and react.*

www.middleschoolchemistry.com/multimedia/chapter6/lesson4#molecules_collide

Point out that the slower-moving molecules hit and bounce off without reacting. But the faster-moving molecules hit each other with enough energy to break bonds and react.



EXTEND

6. Introduce the idea that energy must be added for some chemical reactions to occur.

Tell students that the reaction between baking soda solution and calcium chloride solution happens at room temperature. Students saw that the rate of the reaction can be sped up if the reactants are warmed and slowed down if the reactants are cooled.

Explain that there are many reactions that will not occur at all at room temperature. For these reactions to occur, the reactants need to be heated. When they are heated, the reactants have enough energy to react. Often, once a reaction has started, the energy produced by the reaction itself is enough to keep it going.

Project the video *Volcano.*

www.middleschoolchemistry.com/multimedia/chapter6/lesson4#volcano

Tell students that this reaction requires heat to get started but produces enough heat to keep reacting. You could also mention to students that a common burning candle works the same way. The wax of the candle and oxygen do not react until the heat of a match is added. Then, the heat from the burning wax supplies the heat to keep the reaction going.



DEMONSTRATION

1. Your teacher warmed one glow stick and cooled another. Once the light sticks were started, there was a noticeable difference in their brightness.

How can you tell whether the chemical reaction is happening faster or slower in each glow stick?



Some people place glow sticks in the freezer to make them last longer. Why do you think this works?

Do you think that starting with warmer reactants increases the rate of other chemical reactions?

Why or why not?

ACTIVITY

Question to Investigate

Does the temperature of the reactants affect the rate of the chemical reaction?

Materials for Each Group

- Baking soda
- Calcium chloride
- Water
- Graduated cylinder
- Balance or measuring spoon ($\frac{1}{2}$ teaspoon)
- 4 small plastic cups
- 2 plastic deli-style containers
- Hot water (about $50\text{ }^{\circ}\text{C}$)
- Cold water ($0\text{--}5\text{ }^{\circ}\text{C}$)
- Masking tape
- Pen



Procedure

Make the Baking Soda Solution

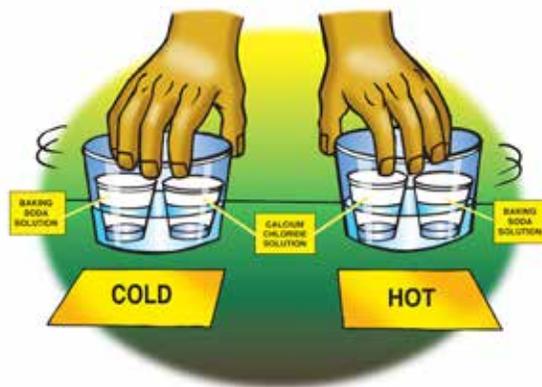
1. Use masking tape and a pen to label 2 small plastic cups baking soda solution, and 2 small plastic cups calcium chloride solution.
2. Use a graduated cylinder to add 20 mL of water to one of the baking soda solution cups.
3. Add 2 g (about $\frac{1}{2}$ teaspoon) of baking soda to the water in its labeled cup. Swirl until as much of the baking soda dissolves as possible. (Some undissolved baking soda may remain in the bottom of the cup.)
4. Pour half of your baking soda solution into the other baking soda solution cup.

Make the calcium chloride solution

5. Use a graduated cylinder to add 20 mL of water to one of the calcium chloride solution cups.
6. Add 2 g (about $\frac{1}{2}$ teaspoon) of calcium chloride to the water in its labeled cup. Swirl until the calcium chloride dissolves.
7. Pour half of your calcium chloride solution into the other calcium chloride solution cup.

Heat and Cool the Solutions

8. Pour hot water into one plastic container and cold water into the other until each is about $\frac{1}{4}$ filled. The water should not be very deep. These are your hot and cold water baths.
9. Place and hold one cup of baking soda solution and one cup of calcium chloride solution in the hot water. Gently swirl the cups in the water for about 30 seconds to heat up the solutions.
10. Your partner should place and hold one cup of baking soda solution and one cup of calcium chloride solution in the cold water. Gently swirl the cups in the water for about 30 seconds to cool the solutions.



Combine the solutions

11. At the same time, you and your partner should combine the two warm solutions with each other, and the two cold solutions with each other.

EXPLAIN IT WITH ATOMS & MOLECULES

2. Does the temperature of the reactants affect the rate of the chemical reaction?

How do you know?

3. On the molecular level, why do you think the warm solutions react faster than the cold solutions?

TAKE IT FURTHER

4. You saw a video showing the ammonium dichromate volcano. How is heat involved in this chemical reaction?

Chapter 6, Lesson 5: A Catalyst and the Rate of Reaction

Key Concepts

- A catalyst is a substance that can help the reactants in a chemical reaction react with each other faster.
- A catalyst does not actually become part of the products of the reaction.

Summary

Students watch a video and do a quick activity to see that a catalyst can increase the rate of the breakdown (decomposition) of hydrogen peroxide. Students will then use salt as a catalyst in a reaction between aluminum foil and a solution of copper II sulfate. Students will be introduced to the concept that a catalyst increases the rate of a chemical reaction but is not incorporated into the products of the reaction.

Objective

Students will be able to define a catalyst as a substance that increases the rate of a chemical reaction but is not incorporated into the products of the reaction.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. When using hydrogen peroxide, follow all warnings on the label. After students have conducted the activity with the copper II sulfate solution and aluminum foil, allow the contents of the cup to evaporate. Put the small amount of solid in a paper towel and dispose in the trash or use a disposal method required by local regulations.

Materials for Each Group

- Graduated cylinder (50 mL or 100 mL)
- Hydrogen peroxide (3%)
- Yeast
- 2 Popsicle sticks
- Detergent solution
- Dropper
- Small cup
- Clear plastic cup
- Copper II sulfate solution (in cup)
- Salt
- Aluminum foil (5 cm × 5 cm)
- Thermometer

Notes about the Materials

Copper II sulfate is available from various chemical suppliers including Sargent Welch, Product #WLC94770-06 or Flinn Scientific, Product #C0110.

ENGAGE

1. Show students two demonstrations and have them look for evidence that a gas is produced in the chemical reactions.

Tell students that you will show them video of two demonstrations where water vapor and oxygen gas are produced in the exact same chemical reaction. Because gases are invisible, ask students to watch closely for evidence that a gas is produced.



Project the video *Elephant's Toothpaste*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson5#elephant_toothpaste

The foaming shows that gases (oxygen and water vapor) are being produced very quickly. The amount of foam produced in a period of time is a way of measuring the rate of the reaction.

Project the video *Genie in a Bottle*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson5#genie_bottle

The steam coming out of the bottle is water vapor that is condensing as it leaves the bottle. Oxygen is also leaving the bottle but it is invisible.

Ask students:

- **How could you tell that a gas is produced in the chemical reaction?**
The foaming in the elephant toothpaste demonstration means that a gas is produced. Production of a gas is a clue that a chemical reaction has occurred. The water vapor in the genie-in-a-bottle demonstration also shows the production of a gas.

Tell students that this lesson is about speeding up chemical reactions. Some reactions occur very slowly, but chemicals called catalysts can be added in order to make them happen faster. Both of these demonstrations relied on a catalyst.

EXPLAIN

2. Describe how the decomposition of hydrogen peroxide produced oxygen gas in both of the videos.

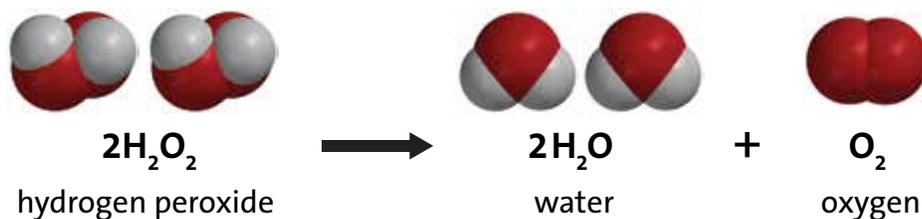
Tell students that both of the demonstrations use a 30% hydrogen peroxide solution. Typically the hydrogen peroxide you can buy at the store is only 3% hydrogen peroxide.

Explain to students that the chemical formula for hydrogen peroxide is H_2O_2 . Point out that hydrogen peroxide is not very stable and breaks down into water and oxygen on its own. This kind of change is a chemical reaction called *decomposition*. The decomposition of hydrogen peroxide is slow and is not usually noticeable.

Project the image *Decomposition of Hydrogen Peroxide*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson5#hydrogen_peroxide

Explain that hydrogen peroxide decomposes to form water and oxygen according to this chemical equation:



Tell students that this chemical reaction happens on its own, and that even the energy from the light in a room can cause hydrogen peroxide to decompose faster. This is why hydrogen peroxide is sold in opaque containers.

Tell students that in the video, a substance (potassium permanganate or manganese dioxide) was used to make the decomposition of hydrogen peroxide happen a lot faster. Even though it made the reaction go faster, the substance itself didn't change during the reaction. A substance that increases the rate of a reaction but does not become part of the products of the reaction is called a *catalyst*.

Ask students:

- **Your teacher showed you a demonstration where a catalyst is added to hydrogen peroxide and a great deal of oxygen gas is produced. If the catalyst is involved in the chemical reaction, why isn't it included as a product in the chemical equation?**
A catalyst does not end up in the products so is not included in the chemical reaction.
- **What does a catalyst do in a chemical reaction?**
Catalysts help a reaction happen faster but do not change themselves during the reaction.

Give Each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

3. Have students use yeast to catalyze the decomposition of hydrogen peroxide.



Question to Investigate

Can another substance catalyze the decomposition of hydrogen peroxide?

Materials for Each Group

- Graduated cylinder
- Hydrogen peroxide (3%)
- Yeast
- Popsicle stick
- Detergent solution
- Dropper

Teacher Preparation

Make a detergent solution by adding 1 teaspoon of liquid dish detergent to 2 tablespoons of water. Divide this detergent solution equally into one small cup for each group.

Procedure

1. Add 10 mL of hydrogen peroxide to a graduated cylinder. Add 1 drop of detergent solution. Swirl gently and watch the solution for any bubbling.

Explain to students that the detergent is added only to make bubbles if any gas is produced. Since the breakdown of hydrogen peroxide produces oxygen gas, bubbling shows that the hydrogen peroxide is breaking down or decomposing. The lack of bubbling shows that not much oxygen gas is being produced.

2. Use the end of a popsicle stick to add a small amount of yeast to the hydrogen peroxide in the graduated cylinder and swirl.
3. Place the graduated cylinder on the table and watch for any bubbling.
4. Hold the graduated cylinder to see if there seems to be any change in temperature.



Expected Results

Before the yeast is added, there is no observable bubbling. After the yeast is added, bubbling will cause foam to move up the graduated cylinder. Also, the graduated cylinder should feel a little warmer because the decomposition of hydrogen peroxide releases energy. Energy changes in chemical reactions will be investigated in more detail in Chapter 6, Lesson 7.

4. Discuss student observations.

Ask students:

- **What clues did you have that a chemical reaction occurred in this activity?**
Bubbling. Tell students that a change in temperature is also a sign that a chemical reaction may be occurring. Endothermic and exothermic chemical reactions will be addressed in Chapter 6, Lesson 7.
- **What is the catalyst in this activity?**
A substance in yeast.
- **What evidence do you have that hydrogen peroxide decomposed faster when you added yeast?**
Bubbles of oxygen gas were produced after the yeast was added.
- **When you write the chemical equation for this reaction, should yeast be included on the product side of the chemical equation?**
Explain to students that the catalyst in the yeast does not end up in the products but is a substance that helps the decomposition happen faster. Sometimes a catalyst is written above or below the arrow in a chemical equation, but it is never included with the reactants or products.

In general, catalysts work by providing a place where reactants can come together to react. Explain to students that cells in yeast and other organisms contain a catalyst called *catalase*. Through normal cell processes, living things produce hydrogen peroxide in their cells. But hydrogen peroxide is a poison so the cells need a way to break it down very quickly. Cells contain catalase, which breaks down hydrogen peroxide at a very fast rate. A single molecule of catalase can catalyze the breakdown of millions of hydrogen peroxide molecules every second.

Students may continue to explore the effect of catalase on hydrogen peroxide by adding a piece of raw fresh potato to a small amount of hydrogen peroxide.

EXTEND

5. Have students identify the changes that occur when copper II sulfate reacts with a piece of aluminum foil.

Note: This is a reaction between copper II sulfate and aluminum. The copper is called “copper II” because copper can make different types of ions. It can lose one electron and be just Cu^+ or it can lose two electrons and be Cu^{2+} . This type of copper ion is called copper II. Also the “sulfate” in copper II sulfate is also an ion. This ion is made up of more than one atom. It is one of the polyatomic ions discussed in Chapter 4, Lesson 3. The sulfate ion is made up of a sulfur atom bonded to four oxygen atoms and is treated as one ion (SO_4^{2-}).

There are several interesting aspects of the reaction between copper II sulfate and aluminum, but it is different from the other reactions students have seen so far. In this reaction, the movement of electrons, rather than entire atoms, ions, or molecules, causes the reaction to occur. This type of reaction is called an oxidation/reduction reaction. This particular reaction is fun to do because it is exothermic, generates a gas, and copper metal appears as aluminum metal disappears.

Salt can be considered a catalyst in the reaction but has a different role than most catalysts. Copper II sulfate and aluminum react very slowly because aluminum is coated with a very thin layer of tarnish (aluminum oxide). This reaction can be sped up if the layer of aluminum oxide is removed or compromised. Adding salt does this and allows electrons from the aluminum to react with the copper ions in the solution, causing them to become copper metal.

Question to Investigate

What is the catalyst in the following activity?

Materials for Each Group

- Copper II sulfate solution (in cup)
- Clear plastic cup (empty)
- Salt
- Piece of aluminum foil
- Thermometer
- Popsicle stick

Teacher preparation

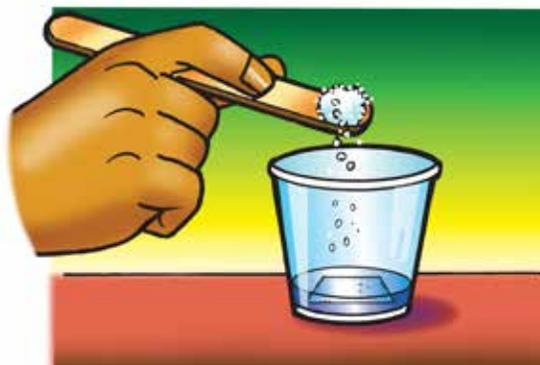
Make a copper II sulfate solution by adding 20 g of copper II sulfate to 200 mL of water.

Pour about 25 mL of copper II sulfate solution into a cup for each group.

Cut aluminum foil into pieces big enough to cover the bottom of a cup (about 5 cm long × 5 cm wide).

Procedure

1. Place the piece of aluminum foil in an empty cup. Use your fingers or a Popsicle stick to push the foil firmly down so that it lays flat and covers the bottom of the cup.
2. Add all of the copper II sulfate solution to the cup with the aluminum foil.
3. Gently swirl the solution for a few seconds and let it stand still. Watch the aluminum for any bubbling or color change.
4. Use your Popsicle stick to place a small amount of salt in the copper II sulfate solution. Gently swirl the solution for a few seconds and let it stand still. Watch for any bubbling or color change.
5. Carefully place a thermometer in the cup and see if the temperature changes.



Expected Results

Before the salt is added, there is no bubbling or color change. After the salt is added, the color turns greenish and bubbles begin to form on the aluminum. Soon, brownish material (copper) begins to form on the aluminum. The bubbling becomes more vigorous and the solution loses its blue color as the aluminum disappears and more copper is produced. The solution also gets warmer.

6. Discuss student observations.

Ask students:

- **How do you know that a chemical reaction occurs when a piece of aluminum foil and sodium chloride is placed in copper II sulfate solution?**
There was bubbling, a color change, an increase in temperature, and a different solid was formed.
- **What is the catalyst in this activity?**
Salt.
- **How is adding salt to the aluminum similar to adding yeast to the hydrogen peroxide?**
Both can be seen as catalysts. Adding yeast helps the hydrogen peroxide decompose faster and adding salt helps the aluminum react with the copper II sulfate.

Tell students that the blue solution contains copper ions (Cu^{2+}). Adding salt to the solution helps remove a layer of tarnish from the piece of aluminum that was in the solution. This exposes some aluminum and allows electrons from the aluminum to react with the copper ions. These negative electrons are attracted to the positive copper ions. When the electrons join with the copper ions, the ions become neutral copper atoms and look like copper metal in the solution. As the aluminum loses its electrons, it becomes aluminum ions and goes into the solution and seems to disappear.

DEMONSTRATION

1. Your teacher showed you videos of two chemistry demonstrations: Elephant's toothpaste and genie in a bottle. Are both of these chemical changes?

How do you know?

EXPLAIN IT WITH ATOMS & MOLECULES

2. Even though the two demonstrations seem different, the chemical reaction behind both is exactly the same—the decomposition of hydrogen peroxide. Refer to the following equation as you answer the questions below.



Which new substances are created when hydrogen peroxide decomposes?

Each demonstration used a substance called a catalyst. What does a catalyst do in a chemical reaction?

If the catalyst is involved in the chemical reaction, why isn't it included as a product in the chemical equation?

ACTIVITY

Question to Investigate

Can another substance catalyze the decomposition of hydrogen peroxide?



Materials for Each Group

- Graduated cylinder
- Hydrogen peroxide (3%)
- Yeast
- Popsicle stick
- Detergent solution
- Dropper

Procedure

1. Add 10 mL of hydrogen peroxide to a graduated cylinder. Add 1 drop of detergent solution. Swirl gently and watch the solution for any bubbling.
2. Use the end of a Popsicle stick to add a small amount of yeast to the hydrogen peroxide in the graduated cylinder and swirl.
3. Place the graduated cylinder on the table and watch for any bubbling.
4. Hold the graduated cylinder to see if there seems to be any change in temperature.



3. What clues did you have that a chemical reaction occurred in this activity?

4. What is the catalyst in this activity?

5. What evidence do you have that hydrogen peroxide decomposed faster when you added yeast?
6. When writing the chemical equation for this reaction, should yeast be included on the product side of the chemical equation?

TAKE IT FURTHER

Question to Investigate

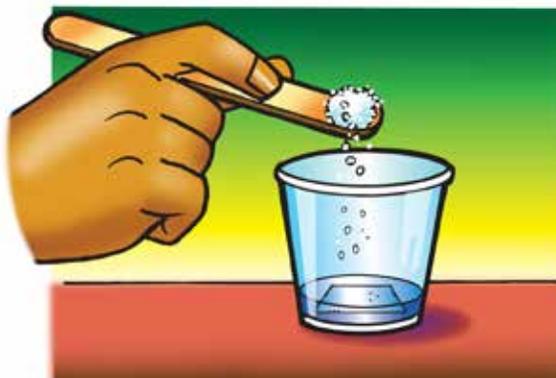
What is the catalyst in the following activity?

Materials for Each Group

- Copper II sulfate solution (in cup)
- Clear plastic cup (empty)
- Salt
- Piece of aluminum foil
- Thermometer
- Popsicle stick

Procedure

1. Place the piece of aluminum foil in an empty cup. Use your fingers or a popsicle stick to push the foil firmly down so that it lays flat and covers the bottom of the cup.
2. Add all of the copper II sulfate solution to the cup with the aluminum foil.
3. Gently swirl the solution for a few seconds and let it stand still. Watch the aluminum for any bubbling or color change.
4. Use your Popsicle stick to place a small amount of salt in the copper II sulfate solution. Gently swirl the solution for a few seconds and let it stand still. Watch for any bubbling or color change.
5. Carefully place a thermometer in the cup and see if the temperature changes.



7. How do you know that a chemical reaction occurs when a piece of aluminum foil and sodium chloride is placed in copper II sulfate solution?
8. What is the catalyst in this activity?
9. How is adding salt to the aluminum similar to adding yeast to the hydrogen peroxide?

Chapter 6, Lesson 6: Using Chemical Change to Identify an Unknown

Key Concepts

- Substances react chemically in characteristic ways
- A set of reactions can be used to identify an unknown substance

Summary

Students will use test liquids on different known powders and observe their reactions. Then students will use these characteristic chemical changes to help them identify an unknown powder.

Objective

Students will be able to identify and control variables to develop a test to identify an unknown powder. Students will be able to explain that a substance reacts chemically in characteristic ways and that these characteristics can be used to identify an unknown substance.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles. When using tincture of iodine, follow all warnings on the label. Have students wash hands after the activity.

Materials for the Demonstrations

- Baking soda
- Cornstarch
- Cream of tartar
- Tincture of iodine
- Vinegar
- Water
- Universal indicator
- Graduated cylinder or beaker
- 2 droppers
- ¼ teaspoon
- 5 clear plastic cups
- 3 Popsicle sticks

Materials for Each Group

- Baking soda
- Baking powder
- Cream of tartar
- Cornstarch
- Water
- Vinegar
- Tincture of iodine
- Universal indicator
- 10 small plastic cups
- 4 droppers
- 8 Popsicle sticks
- Testing chart (laminated or covered with wax paper)

ENGAGE

1. Add iodine solution to baking soda and cornstarch to introduce the idea that different substances react chemically in characteristic ways.

Materials for the Demonstration

- Tincture of iodine
- Baking soda
- Cornstarch
- Water
- Graduated cylinder or beaker
- 2 Popsicle sticks
- Dropper
- $\frac{1}{4}$ teaspoon
- 3 clear plastic cups



Teacher Preparation

- Make a dilute tincture of iodine solution by adding about 10 drops of tincture of iodine to 100 mL of water. Pour 50 mL of this solution into a clear plastic cup for this demonstration.
- Set the other 50 mL aside for the student activity. Instructions for preparing the rest of the materials for the student activity are in the *Explore* section of this lesson.
- Place $\frac{1}{4}$ teaspoon corn starch in a clear plastic cup and $\frac{1}{4}$ teaspoon baking soda in another cup. Do not tell students which powder is in each cup.

Procedure

1. Tell students that you have a different powder in each cup. Both are white and look alike, but they are chemically different.
2. Pour about 25 mL of the iodine solution in each cup and swirl.



Expected Results

The iodine solution stays light brown when added to the baking soda. The iodine solution and corn starch turns a very dark purple.

Ask students:

- **Both powders looked similar at first. How do you know that they are different?**
The iodine changed color in one powder, but not in the other.

- **Which do you think is probably a chemical change?**

The iodine and the cornstarch are probably the chemical change because the dramatic color change seems like something new may have been produced. The iodine does not change color when it combines with the baking soda.

- **What other tests could you conduct with baking soda and cornstarch to compare their characteristic properties?**

Let students know that there can be no tasting or smelling of the powders.

Students might suggest adding water to see if they dissolve differently or maybe adding another substance to see if a different chemical reaction takes place.

Give Each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



Give Each Group a Testing Chart.

Make one copy of the testing chart, found at the end of the downloaded lesson, for each group. Either laminate this testing chart or have students lay a piece of wax paper over it.

EXPLORE

2. **Introduce the activity and ask students how they might test and compare the four different powders with four different test solutions.**

Tell students that in this activity they will test four different similar-looking powders with four different test solutions. The four powders are baking soda, baking powder, cream of tartar, and cornstarch. The four test solutions are water, vinegar, iodine solution, and universal indicator. Explain that each powder will react in a certain way with each solution used to test it. Each powder and solution pair is one set of reactants. Let students know that in some cases, no chemical reaction will occur.

Students will need to observe and record the reactions the liquids have with each powder.

Have students look at the testing chart. Point out that the names of the four test solutions are on the left and the names of the different powders are on the top.

Test solutions	Baking Soda	Baking Powder	Cream of tartar	Cornstarch	Unknown
Water					
Vinegar					
Iodine					
Indicator					

There is also one column for an unknown powder. Explain that after testing all four known powders and recording their observations, you will give students an unknown powder to identify.

Ask students questions like the following to help them plan how they will organize and conduct their tests:

- **Do we need more than one pile of each powder placed on the chart?**
Yes. Each powder will be tested with each of the four solutions so there needs to be four piles of each powder in the squares under its name.
- **Should all the squares on the entire chart have samples of powder on them before you start testing?**
It is best if students place the four samples of one particular powder in its column on the testing sheet. Then students should test that powder with each of the four solutions. Students should test a single powder with each of the test liquids before moving on to the next powder.
- **Do the piles have to be about the same size?**
The size of the piles is not particularly important as long as enough powder is used to see a reaction, if there is one. When testing the unknown, try to make the piles of unknown about the same size as the piles of the other powders.
- **Should the number of drops placed on each pile be the same?**
The precise number of drops is not particularly important, although enough liquid should be added to see if there is a reaction. When testing the unknown, be sure that the number of drops used on the unknown is the same as the number used on the other powders.
- **How will you remember your observations for each reaction?**
Students should record their observations immediately after a single test solution is added to a powder. These will be recorded in a chart in the corresponding box on the activity sheet.

3. As an example, guide the class as they test baking soda with water, vinegar, iodine solution, and universal indicator.

Question to Investigate

Can you use the characteristic ways substances react to tell similar-looking substances apart?



Materials for Each Group

- Baking soda in cup
- Baking powder in cup
- Cream of tartar in cup
- Cornstarch in cup

- Water in cup
- Vinegar in cup
- Tincture of iodine solution in cup
- Universal indicator solution in cup
- 4 Popsicle sticks
- Testing chart, either laminated or with a piece of wax paper over it
- 4 droppers

Teacher Preparation

Each group will need five labeled cups each containing one of the powders and four labeled cups each containing one of the four test solutions to complete all three of the activities in this lesson.

Prepare the powders

1. Label five small cups Baking Soda, Baking Powder, Cream of Tartar, Cornstarch and Unknown.
2. Place about $\frac{1}{2}$ teaspoon of each powder into its labeled cup. These powders will be tested in this activity.
3. Place about $\frac{1}{2}$ teaspoon of baking powder in the cup labeled unknown. Place the unknown cups aside. Students will test this unknown powder after they have tested each of the “known” powders and recorded their observations.

Prepare the Test Solutions

4. Label four cups Water, Vinegar, Iodine, and Indicator.
5. Use the iodine solution left over from the demonstration or make a new solution by adding 5 drops of tincture of iodine to 50 mL of water. Place about 5 mL iodine solution in a small labeled cup for each group.
6. Make a universal indicator solution by adding 5 mL of universal indicator to 100 mL of water. Place about 5 mL (or 1 teaspoon) indicator solution in a small labeled cup for each group. Some indicator solution will be left over for the demonstration at the end of the lesson.
7. Place about 5 mL of water and vinegar into their small labeled cups.

Procedure

1. Use the end of a popsicle stick to place four equal piles of baking soda on the testing chart in the baking soda column. Let students know that they should not use all of the powder at this time. The remaining powder will be used in the *Extend* portion of this lesson.



2. Add 5 drops of water to the first pile of baking soda. Record your observations in the chart on the activity sheet.
3. Continue testing each pile of baking soda with a different test solution and recording your observations.

Expected Results

There will be no change with water, bubbling with vinegar, and little to no change with the iodine or indicator solutions.

Ask students:

- **What did you observe when each test solution was added to a sample of baking soda?**
Baking soda bubbles with vinegar. There is no change with water or iodine solution. Students may see a slight color change with the indicator solution. Explain to students that their results show them the characteristic set of reactions that baking soda has with these four test solutions.
- **Would you expect each test solution to react with baking powder the same way as it did with baking soda?**
No. Each powder should have a different set of reactions.

Have students conduct the tests on the remaining powders and record their observations.

Procedure

4. Test each of the powders with the test solutions the way you tested baking soda and record your observations.



Expected Results

Test solution	Baking soda	Baking powder	Cream of tartar	Cornstarch
Water	No change	Bubbling	No change	No change
Vinegar	Lots of bubbling that ends quickly	Bubbling	No change	No change
Iodine	No change	Bubbling, purple	Orange	Purple
Indicator	Greenish-blue	Bubbling, orange changes to yellow with some green	Dark orange or pink	Brighter green, may have some orange

4. Give students the unknown powder and have them use their test solutions and observation chart to identify it.

Give each group the unknown powder. Explain to students that the unknown is one of the four powders they have tested and their job is to find out which one.

Ask students:

- **How could you test the unknown powder so that you could identify it?**
Students should realize that they will need to test the unknown powder the same way they tested all of the other known powders and compare the results. If the unknown powder reacts with each test solution the same way one of the known powders did, then these two powders must be the same.

Question to Investigate

Can you use the characteristic ways substances react to identify an unknown powder?



Materials for each group

- Unknown in cup
- 1 Popsicle stick
- Testing sheet
- 4 test solutions
- 4 droppers

Note: The unknown is baking powder.

Procedure

1. Place four samples of your group's unknown powder in the "Unknown" column on the testing chart.
2. Test the unknown with each test solution in the same way you tested each of the other powders.
3. Compare the set of reactions for the unknown with those of the other powders.

Expected Results

The unknown will react with each test solution the same way that baking powder does because the unknown is baking powder.

EXPLAIN

5. Have students report the identity of the unknown and discuss what evidence led them to their conclusion.

Tell students that they were able to use their observations to identify the unknown because each powder had its own set of characteristic chemical reactions with the test solutions.

Ask students:

- **What is the identity of the unknown?**
Baking powder is the unknown.
- **Which observations led you to your conclusion?**
The unknown reacted with each test solution the same way baking powder did.

Explain that each substance is made up of certain molecules which interact with the molecules in each test liquid in a characteristic way. Some of these interactions result in a chemical reaction and others do not. However, each observation students made is based on the way the molecules of each powder interact with the molecules of each test solution.

EXTEND

6. Have students identify the two substances in baking powder that make it bubble when water is added.

Remind students that baking powder was the only substance that bubbled when water was added to it. Bubbling in a chemical reaction is a sign that a gas is one of the products. Tell students that baking powder is a combination of different powders—baking soda, cream of tartar, and cornstarch. Two of these three react with one another and produce a gas when water is added. Students will need to test all of the possible combinations of two powders with water. This way they can figure out which two powders cause baking powder to bubble with water.

Question to Investigate

Which two substances in baking powder react with one another and produce a gas when water is added?



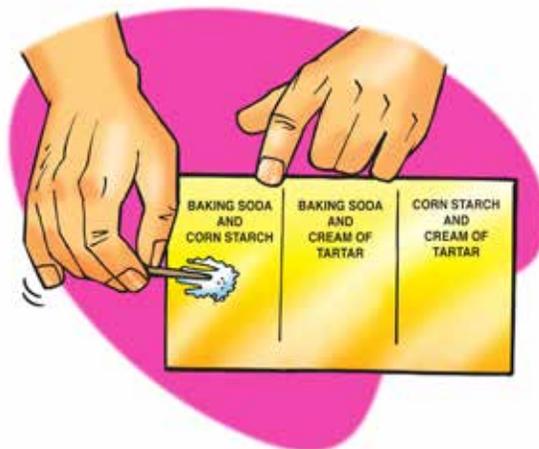
Materials for Each Group

- Baking soda in a cup
- Cornstarch in a cup
- Cream of tartar in a cup
- 3 Popsicle sticks

- Toothpicks
- Wax paper
- Water
- Dropper

Procedure

1. Use separate popsicle sticks to place a small amount of two powders on a piece of wax paper.
2. Use a toothpick to mix the powders.
3. Use a dropper to add about 5 drops of water to the combined powders and record your observations.
4. Repeat steps 1 and 2 until you have tested all three combinations.



Expected results

The two combined powders that bubble with water are baking soda and cream of tartar.

7. Demonstrate that vinegar and cream of tartar are both acids.

Point out that the mixture of baking soda and cream of tartar reacts with water to produce a gas. Baking soda and vinegar also react to produce a gas. Explain that carbon dioxide gas is produced in both reactions. Tell students that cream of tartar and vinegar have something else in common that they will investigate in the next demonstration.

Question to Investigate

Will vinegar turn universal indicator solution pink the way that cream of tartar does?

Materials for the Demonstration

- Universal indicator
- Cream of tartar
- Vinegar
- 2 clear plastic cups
- Popsicle stick
- Dropper

Teacher Preparation

- Use the indicator left over from the lesson.
- Use vinegar and cream of tartar left over from one of the student groups.

Procedure

1. Pour about 25 mL of universal indicator solution in two separate clear plastic cups.
2. Add 2 or 3 drops of vinegar to one cup.
3. Scoop up a small amount of cream of tartar with the tip of a popsicle stick and add it to the other cup.
4. Swirl both cups.



Expected Results

The indicator in both cups will turn pink.

Tell students that the color changes of indicator solution can tell you whether a substance is an acid or not. Because universal indicator turns pink when acids are added to it, we can say that both vinegar and cream of tartar are acids. When an acid reacts with baking soda, carbon dioxide gas is produced.

Ask students

- **Why do you think the baking soda and cream of tartar reaction is similar to the baking soda and vinegar reaction?**
Cream of tartar and vinegar are both acids and interact with sodium bicarbonate in a similar way to produce carbon dioxide gas.

DEMONSTRATION

1. Your teacher poured iodine solution on top of two white powders. How do you know that these two similar-looking powders are really different?



2. Adding iodine solution to one powder caused a physical change, while adding the iodine solution to the other powder caused a chemical change. Which powder probably reacted chemically with the iodine solution?

How do you know?

ACTIVITY

Question to Investigate

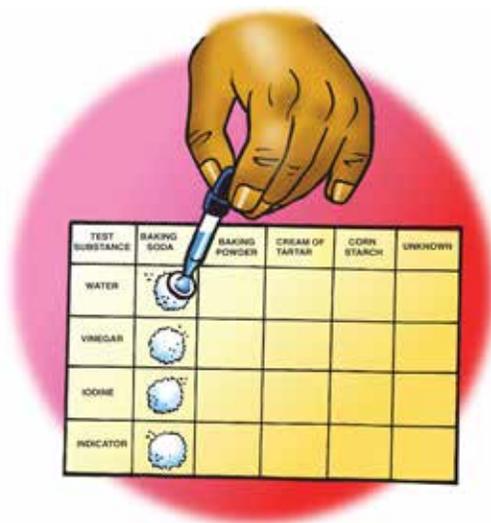
Can you use the characteristic ways substances react to tell similar-looking substances apart?

Materials for Each Group

- Baking soda in cup
- Baking powder in cup
- Cream of tartar in cup
- Cornstarch in cup
- Water in cup
- Vinegar in cup
- Tincture of iodine solution in cup
- Universal indicator solution in cup
- 4 Popsicle sticks
- Testing chart, either laminated or with a piece of wax paper over it
- 4 droppers

Procedure

1. Use the end of a Popsicle stick to place four equal piles of baking soda on the testing chart in the baking soda column. Let students know that they should not use all of the powder at this time. The remaining powder will be used in the *Extend* portion of this lesson.
2. Add 5 drops of water to the first pile of baking soda. Record your observations in the chart on the activity sheet.
3. Continue testing each pile of baking soda with a different test solution and recording your observations.
4. Test each of the powders with the test solutions the way you tested baking soda and record your observations.



Question to Investigate

Can you use the characteristic ways substances react to identify an unknown powder?



Materials for Each Group

- Unknown in cup
- 1 Popsicle stick
- Testing sheet
- 4 test solutions
- 4 droppers

Procedure

1. Place four samples of your group's unknown powder in the "Unknown" column on the testing chart.
2. Test the unknown with each test solution in the same way you tested each of the other powders.
3. Compare the set of reactions for the unknown with those of the other powders.

Test solutions	Unknown
Water	
Vinegar	
Iodine solution	
Indicator solution	

3. What is the identity of the unknown?

Which observations led you to your conclusion?

EXPLAIN IT WITH ATOMS & MOLECULES

4. On the molecular level, why did the different substances react in a characteristic way with the test solutions?

TAKE IT FURTHER

Baking powder is a combination of different powders—baking soda, cream of tartar, and cornstarch. Two of these three powders react with one another and produce carbon dioxide gas when water is added.

Question to Investigate

Which two substances in baking powder react with one another and produce a gas when water is added?

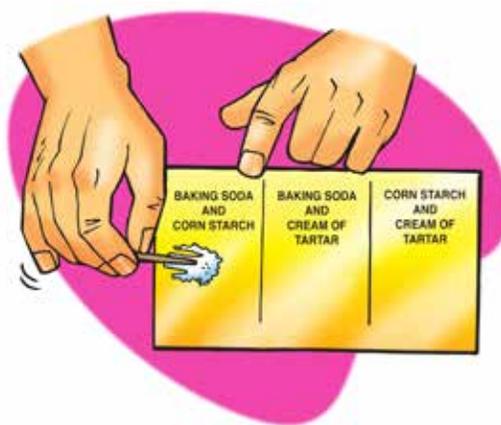


Materials for Each Group

- Baking soda in a cup
- Cornstarch in a cup
- Cream of tartar in a cup
- 3 Popsicle sticks
- Toothpicks
- Wax paper
- Water
- Dropper

Procedure

1. Use separate popsicle sticks to place a small amount of two powders on a piece of wax paper.
2. Use a toothpick to mix the powders.
3. Use a dropper to add about 5 drops of water to the combined powders and record your observations.
4. Repeat Steps 1 and 2 until you have tested all three combinations.



Baking soda + cornstarch	Baking soda + cream of tartar	Cornstarch + cream of tartar

5. Your teacher did a demonstration comparing the way vinegar and cream of tartar react with indicator solution. Based on your observations, why do you think the baking soda and cream of tartar reaction is similar to the baking soda and vinegar reaction?



Testing Chart

Test solutions	Baking Soda	Baking Powder	Cream of tartar	Cornstarch	Unknown
Water					
Vinegar					
Iodine					
Indicator					

Results Table

Test solutions	Baking Soda	Baking Powder	Cream of tartar	Cornstarch	Unknown
Water					
Vinegar					
Iodine					
Indicator					

Chapter 6, Lesson 7: Energy Changes in Chemical Reactions

Key Concepts

- If two substances react and the temperature of the mixture decreases, the reaction is endothermic.
- If two substances react and the temperature of the mixture increases, the reaction is exothermic.
- A chemical reaction involves the breaking of bonds in the reactants and the forming of bonds in the products.
- It takes energy to break bonds.
- Energy is released when bonds are formed.
- If a reaction is endothermic, it takes more energy to break the bonds of the reactants than is released when the bonds of the products are formed.
- If a reaction is exothermic, more energy is released when the bonds of the products are formed than it takes to break the bonds of the reactants.

Summary

Students will conduct two chemical reactions. In the first, the temperature will go down (endothermic) and in the second, the temperature will go up (exothermic). Students will see an animation to review a concept that was introduced in Chapter 5— that it takes energy to break bonds and that energy is released when new bonds are formed. Students will use this idea to explain why a reaction is either endothermic or exothermic.

Objective

Students will be able to define an endothermic and exothermic reaction. Students will be able to use the concept of energy in bond breaking and bond making to explain why one reaction can be endothermic and another reaction can be exothermic.

Evaluation

The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan. The activity sheets are formative assessments of student progress and understanding. A more formal summative assessment is included at the end of each chapter.

Safety

Be sure you and the students wear properly fitting goggles.

Materials for the Each Group

- Vinegar
- Baking soda
- Calcium chloride

- Water
- Thermometer
- 4 small cups
- Disposable self-heating hand warmer
- Self-inflating balloon

Additional Materials if You Choose to do the Extra Extend

- Magnesium sulfate
- Sodium carbonate
- Citric acid
- Universal indicator

About the Materials

Calcium chloride is available from Sargent Welch, Product #WLC94075-06, or Flinn Scientific, Product #C0016, or other suppliers. Calcium chloride is also available in hardware stores for absorbing moisture and for melting ice in the winter.

Hand warmers may be purchased from Flinn Scientific, catalog #AP1931 or from camping, sporting goods, or discount stores. Look for hand warmers that are disposable and sealed in a package and will only warm up when the package is opened. Self-inflating mylar balloons may be purchased from Joissu, product number #43-712 or Educational Innovations, product #AS-800.

ENGAGE

1. Discuss the temperature changes in chemical reactions students have conducted so far.

Remind students that the decomposition reaction of hydrogen peroxide and the reaction with copper II sulfate and aluminum both caused the temperature of the solution to increase. Tell students that you will show them three chemical reactions where the temperature increases dramatically.

Project the video *Thermite Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#thermite

After adding one or more catalysts, iron oxide (rust) and aluminum react to produce elemental iron and aluminum oxide. So much heat is produced in this reaction that the iron becomes a liquid. The heat is so intense that the molten iron can be used to weld railroad tracks together.



Project the video *Nitrogen Triiodide Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#nitrogen_triiodide

This is a decomposition reaction where nitrogen triiodide decomposes into nitrogen gas and purple iodine vapor. Nitrogen triiodide crystals are so unstable that just a light touch will cause them to rapidly decompose generating a great deal of heat.

Project the video *White Phosphorous Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#white_phosphorous

White phosphorous is dissolved in a solvent and spread on a piece of paper. When the solvent evaporates, the phosphorous reacts with oxygen in the air in a combustion reaction.

Ask students to make a prediction:

- **Do you think substances can react and cause the temperature of the mixture to decrease?**

Tell students that this lesson is going to explore temperature changes in chemical reactions.

Give Each Student an Activity Sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. Look at the teacher version of the activity sheet to find the questions and answers.



EXPLORE

2. **Have students measure the change in temperature of the reaction between baking soda and vinegar.**



Question to Investigate

Does the temperature increase, decrease, or stay the same in the reaction between baking soda and vinegar?

Materials

- Vinegar in a cup
- Baking soda in a cup
- Thermometer

Materials Note: The amount of the solutions must be enough to cover the bulb of the thermometer. If they aren't, use a smaller cup or clip the end of a plastic-backed thermometer so that the backing is flush with the bottom of the bulb.

Teacher Preparation

- Place about 10 mL of vinegar in a small plastic cup for each group.
- Place about $\frac{1}{2}$ teaspoon of baking soda in a small cup for each group.

Procedure

1. Place a thermometer in the vinegar. Read the thermometer and record the temperature on the activity sheet.
2. While the thermometer is in the cup, add all of the baking soda from your cup.
3. Watch the thermometer to observe any change in temperature. Record the temperature after it has stopped changing.



Expected Results

If you begin with room-temperature vinegar, the temperature will decrease by about 7°C . The amount of temperature decrease will vary. Carbon dioxide gas is also produced.

3. Discuss student observations.

Ask students:

- **Did the temperature increase, decrease, or stay the same when you combined baking soda and vinegar?**
The temperature decreased.
- **What is the lowest temperature reached during your group's reaction?**
There will likely be some variation.

Tell students that when the temperature of a chemical reaction decreases, the reaction is called an *endothermic* reaction. The first part of the word, *endo*, means in or into and *thermic* has to do with heat or energy. So an endothermic reaction means that more energy goes into making the reaction happen than is released by the reaction. This leaves the reaction mixture at a lower temperature.

4. Have students measure the change in temperature of the reaction between baking soda solution and calcium chloride.

Question to Investigate

Does the temperature increase, decrease, or stay the same in the reaction between baking soda solution and calcium chloride?

Materials

- Baking soda solution in a cup
- Calcium chloride in a cup
- Thermometer

Teacher Preparation

- Make a baking soda solution by dissolving about 2 tablespoons of baking soda in 1 cup of water. Stir until no more baking soda will dissolve.
- Place about 10 mL of baking soda solution in a small plastic cup for each group.
- Place about $\frac{1}{2}$ teaspoon of calcium chloride in a small cup for each group.

Procedure

1. Place a thermometer in the baking soda solution. Read the thermometer and record the temperature on the activity sheet.
2. While the thermometer is in the cup, add all of the calcium chloride from the cup.
3. Watch the thermometer to observe any change in temperature. Record the temperature when it stops changing.

Expected Results

The temperature of the solution should increase by about 15–20 °C. The temperature increase will vary. Carbon dioxide gas is produced, and a white cloudy precipitate, calcium carbonate, is formed.

Note: Some of the temperature increase in this reaction may be due to the chemical reaction between baking soda and calcium chloride, but some is also due to the exothermic way calcium chloride dissolves in water. Chapter 5 Lesson 9 addresses temperature changes as bonds between a solute are broken and the bonds between the solute and water are formed during the physical change of dissolving.

Read more about exothermic and endothermic chemical reactions in the additional teacher background section at the end of the lesson.

5. Discuss student observations.

Ask students:

- **Did the temperature increase, decrease, or stay the same when you combined baking soda solution and calcium chloride?**
The temperature increased.
- **What is the highest temperature reached during your group's reaction?**
There will likely be some variation.

Tell students that when the temperature of a chemical reaction increases, the reaction is called an *exothermic* reaction. The first part of the word, *exo*, means out or out of, and *thermic* has to do with heat or energy. So an exothermic reaction means that more energy goes out or is released by the reaction than goes into it. This leaves the reaction mixture at a higher temperature.

EXPLAIN

6. Explain how differences in the energy required to break bonds and make bonds cause temperature changes during chemical reactions.

Tell students that an example of a very exothermic reaction is the combustion or burning of fuel like the gas in a kitchen stove. Even if students have seen the animation of the combustion of methane from Chapter 6, Lesson 1, remind them that methane (CH_4) reacts with oxygen (O_2) from the air to produce carbon dioxide gas (CO_2) and water vapor (H_2O) and a lot of energy.

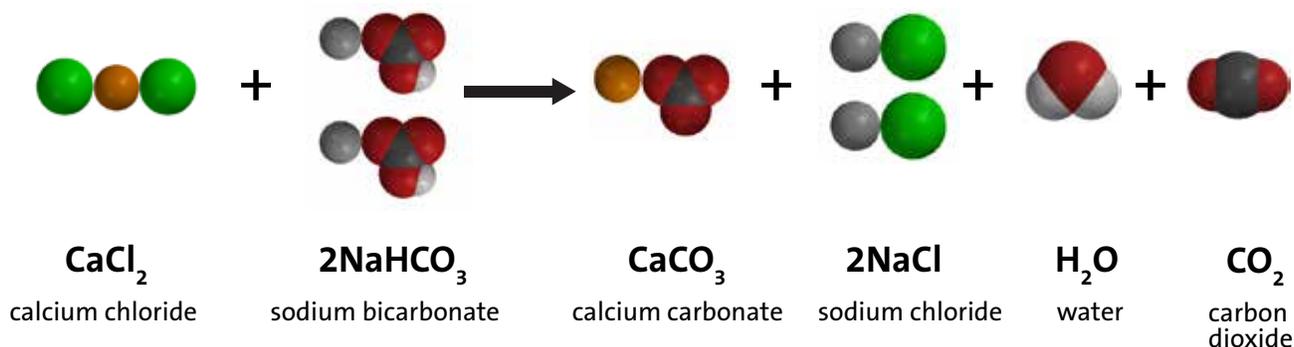
Project the animation *Methane Combustion Energy*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#methane_combustion_energy

Click on the methane and the oxygen to show that it takes energy to break the bonds of the reactants. This is shown by “energy arrows” going *into* the molecules of the reactants. Then click on the carbon dioxide and the water to show that energy is released when the atoms bond to make the products. This is shown by the energy arrows coming *out* of the molecules in the products. Show students that more energy was released when the bonds in the products were formed than was used to break the bonds of the reactants. This is shown by larger energy arrows coming out of the products and smaller energy arrows going into the reactants. Since more energy was released than was used, this reaction gets warmer and is exothermic.

Project the image *Baking Soda and Calcium Chloride Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#baking_soda_calcium_chloride

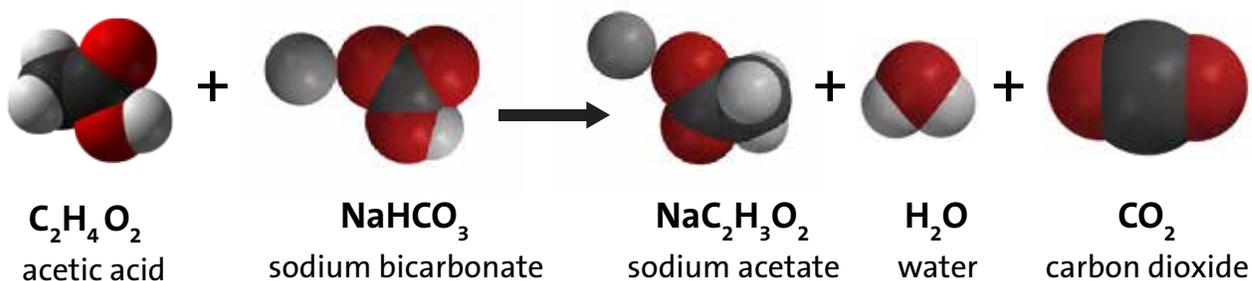


Ask students:

- **Is this an endothermic or exothermic reaction?**
Exothermic
- **What do you know about the amount of energy required to break the bonds of the reactants compared to the amount of energy released when bonds in the products are formed?**
More energy was released when bonds in the products were formed than was required to break the bonds in the reactants.
- **If we were using energy arrows, where would the bigger and smaller arrow go?**
A smaller arrow going in would be on the reactant side and a bigger arrow coming out would be on the product side.

Project the image *Baking Soda and Vinegar Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#baking_soda_and_vinegar



Ask students:

- **Is this an endothermic or exothermic reaction?**
Endothermic.
- **What do you know about the amount of energy required to break the bonds of the reactants compared to the amount of energy released when the products are formed?**
It took more energy to break the bonds of the reactants than was released when the bonds in the products were formed.

- If we were using energy arrows, where would the bigger and smaller arrows go? A bigger arrow going in would be on the reactant side and a smaller arrow coming out would be on the product side.

Project the animation *Endothermic Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#endothermic_reaction

Remind students that a chemical reaction involves the breaking of bonds in the reactants and the making of bonds in the products. Also remind them that it takes energy to break bonds and that energy is released when bonds are formed.

In an endothermic reaction, it takes more energy to break the bonds of the reactants than is released when the bonds in the products are formed. In an endothermic reaction, the temperature goes down.

Project the animation *Exothermic Reaction*.

www.middleschoolchemistry.com/multimedia/chapter6/lesson7#exothermic_reaction

Explain that in an exothermic reaction it takes less energy to break the bonds of the reactants than is released when the bonds in the products are formed. In an exothermic reaction, the temperature goes up.

EXTEND

7. Have students explain how changes in energy during chemical reactions cause them to be either endothermic or exothermic.

Tell students that they will use their knowledge of endothermic and exothermic reactions to describe the energy changes that occur when hand warmers and self-inflating balloons are activated. These two different products use chemical reactions to make them work.

Materials for Each Group

- Disposable self-heating hand warmer
- Self-inflating balloon



Hand Warmer

Tell students that the outer packaging on the hand warmer keeps air away from it and prevents the chemical reaction from happening, until the moment when a consumer wants it to start producing heat. Oxygen in the air is one of the reactants in the chemical reaction. So once the package is opened, the iron powder in the hand warmer reacts with the oxygen in the air.

Self-Inflating Balloon

Tell students that the chemical reaction that causes the self-inflating balloon to inflate is very similar to a chemical reaction students have explored already. Have students gently feel the self-inflating balloon to guess what the reactants are. They should notice a tablet and a sealed packet of liquid inside the balloon. Explain that the tablet is baking soda and the liquid in the packet is citric acid, which reacts with baking soda in a way similar to vinegar. The sealed packet prevents the citric acid from reacting with the baking soda.

Procedure

1. Open the package the hand warmer is in to begin the chemical reaction.
2. Shake the hand warmer and feel for any temperature change.
3. Activate the self-inflating balloon by either pressing down or stepping on the packet of citric acid to rupture it.
4. Shake the balloon and feel the area on the balloon where the liquid is.
5. Be sure everyone in your group has a chance to feel both the hand-warmer and the self-inflating balloon.

Expected Results

The hand warmer will become warmer, and the liquid in the self-inflating balloon will become colder. The balloon will inflate as carbon dioxide gas is produced.

Ask students:

- **Which is an example of an endothermic reaction? An exothermic reaction?**
The self-inflating balloon is an example of an endothermic reaction, and the hand warmer is an example of an exothermic reaction.
- **What can you say about the amount of energy required to break bonds in the reactants compared to the amount of energy that is released when bonds are formed in the products in the self-inflating balloon?**
In the self-inflating balloon, more energy is required to break the bonds in the reactants than is released when the new bonds in the products are formed. Therefore, the reaction is endothermic.
- **What can you say about the amount of energy required to break bonds in the reactants compared to the amount of energy that is released when bonds are formed in the products in the hand warmer?**
In the hand warmer, more energy is released when the new bonds in the products are formed than is used to break the bonds in the reactants. Therefore, the reaction is exothermic.
- **What do you think is the gas inside the self-inflating balloon?**
Carbon dioxide gas is produced when citric acid and baking soda react.

If you do not have access to a self-inflating balloon, you may choose to have students make their own.

Materials for Each Group

- Alka-Seltzer tablet
- Water
- Graduated cylinder
- Snack size zip-closing plastic bag

Procedure

1. Place 10 mL of water in a zip-closing plastic bag.
2. Tilt the open bag at an angle so that the water flows into one corner. Hold the bag while a partner places the Alka-Seltzer tablet in the opposite corner. Remove as much air as possible and seal the bag.
3. Shake the bag to help water and tablet react. Place the bag in a bowl or other container in case it pops.
4. Feel the liquid near the tablet to see if there is any temperature change.

Expected Results

The liquid will get colder and the bag will inflate and may pop.

EXTRA EXTEND

8. Have students identify clues of chemical change in the following reactions.

Remind students that in this chapter, they have seen different clues of chemical change. Ask students to identify all of the clues they observe in this pair of chemical reactions.

Question to Investigate

What clues do you observe that a chemical reaction is taking place?

Materials for Each Group

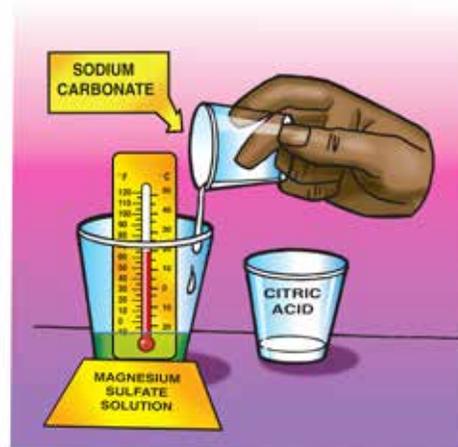
- Magnesium sulfate solution in cup
- Sodium carbonate solution in cup
- Citric acid solution in cup
- Universal indicator
- Thermometer
- Dropper

Teacher Preparation

- Label 3 small cups Magnesium Sulfate, Sodium Carbonate, and Citric Acid for each group.
- Make each solution by adding:
 - 1 tablespoon of magnesium sulfate to 250 mL of water.
 - 1 teaspoon of sodium carbonate to 125 mL of water.
 - 1 teaspoon of citric acid to 125 mL water.
- Pour 30 mL of magnesium sulfate, 10 mL of sodium carbonate, and 10 mL of citric acid solution into their labeled cups for each group.

Procedure

1. Add 5 drops of universal indicator to the magnesium sulfate solution.
2. Place a thermometer in the cup and record the temperature of the solution.
3. Add 10 mL of sodium carbonate solution.
4. Add 10 mL of citric acid.



Expected Results

The magnesium sulfate, universal indicator, and sodium carbonate will turn purple and form a precipitate. With the addition of citric acid, the mixture will turn yellow or pink and bubble as the precipitate disappears. There is no noticeable temperature change.

Ask Students:

- **What clues do you observe that let you know that a chemical reaction is taking place?**
There is a color change, formation of a precipitate, another color change, and bubbling.
- **How can it be that in this chemical reaction, you did not notice a temperature change?**
Maybe the amount of energy required to break bonds was about the same as the amount of energy released when bonds are formed. Or the temperature change was so small that it was not noticeable with student thermometers.

Activity Sheet
Chapter 6, Lesson 7
Energy Changes in Chemical Reactions

Name _____

Date _____

ACTIVITY

Question to Investigate

Does the temperature increase, decrease, or stay the same in the reaction between baking soda and vinegar?

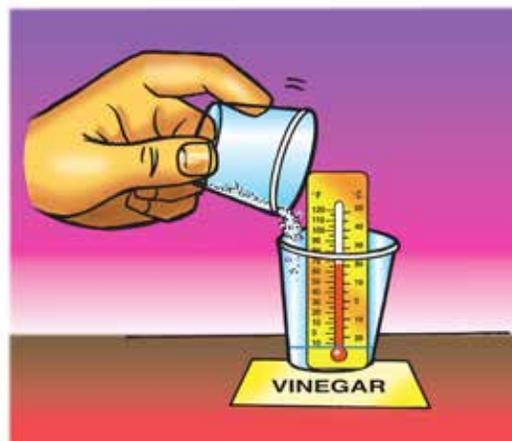


Materials

- Vinegar in a cup
- Baking soda in a cup
- Thermometer

Procedure

1. Place a thermometer in the vinegar. Read the thermometer and record the temperature on the activity sheet.
2. While the thermometer is in the cup, add all of the baking soda from your cup.
3. Watch the thermometer to observe any change in temperature. Record the temperature after it has stopped changing.



1. Did the temperature increase, decrease, or stay the same when you combined baking soda and vinegar?

2. What is the lowest temperature reached during your group's reaction?

Question to Investigate

Does the temperature increase, decrease, or stay the same in the reaction between baking soda solution and calcium chloride?

Materials

- Baking soda solution in a cup
- Calcium chloride in a cup
- Thermometer

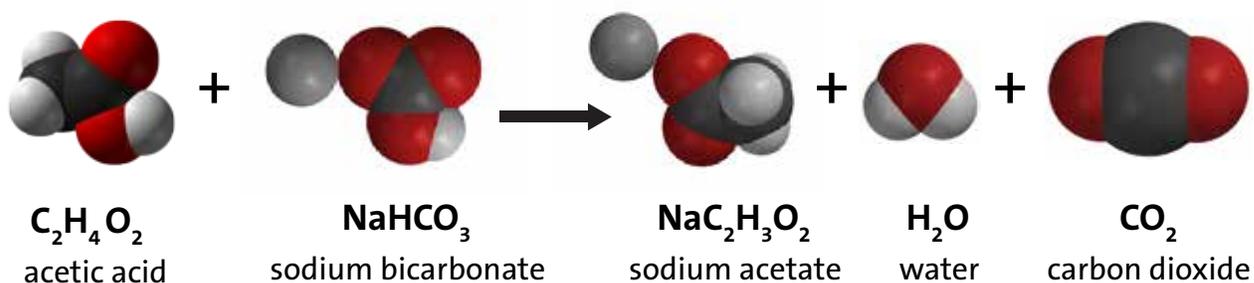
Procedure

1. Place a thermometer in the baking soda solution. Read the thermometer and record the temperature on the activity sheet.
 2. While the thermometer is in the cup, add all of the calcium chloride from the cup.
 3. Watch the thermometer to observe any change in temperature. Record the temperature when it stops changing.
3. Did the temperature increase, decrease, or stay the same when you combined baking soda solution and calcium chloride?
4. What is the highest temperature reached during your group's reaction?

EXPLAIN IT WITH ATOMS & MOLECULES

When the temperature of a chemical reaction decreases, the reaction is called an *endothermic* reaction. When the temperature of a chemical reaction increases, the reaction is called an *exothermic* reaction.

Vinegar and baking soda reaction



5. Is this an endothermic or exothermic reaction?
6. Draw an energy arrow on the reactant side and another on the product side to compare the amount of energy used and released during the reaction.
7. What do the arrows show about the amount of energy required to break the bonds of the reactants compared to the amount of energy released when the products are formed?

TAKE IT FURTHER

Disposable hand warmers and self-inflating balloons use different chemical reactions to make them work. Both are packaged so that the reactants are kept separate. Once the consumer causes the reactants to combine, the chemical reactions begin.

Question to Investigate

How can endothermic and exothermic chemical reactions be useful?

Materials for Each Group

- Disposable self-heating hand warmer
- Self-inflating balloon

Procedure

1. Open the package the hand warmer is in to begin the chemical reaction.
2. Shake the hand warmer and feel for any temperature change.
3. Activate the self-inflating balloon by either pressing down or stepping on the packet of citric acid to rupture it.
4. Shake the balloon and feel the area on the balloon where the liquid is.
5. Be sure everyone in your group has a chance to feel both the hand-warmer and the self-inflating balloon.

11. Which is an example of an endothermic reaction?

Which is an example of an exothermic reaction?

12. For the hand warmer, what can you say about the amount of energy required to break bonds in the reactants compared to the amount of energy that is released when bonds are formed in the products?

13. For the self-inflating balloon, what can you say about the amount of energy required to break bonds in the reactants compared to the amount of energy that is released when bonds are formed in the products?

EXTRA EXTEND

Question to investigate

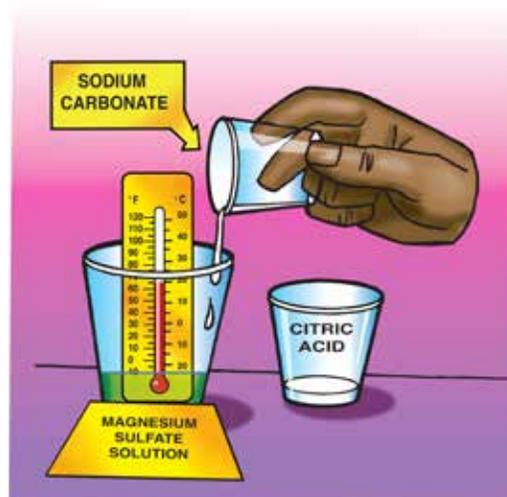
What clues do you observe that a chemical reaction is taking place?

Materials for each group

- Magnesium sulfate solution in cup
- Sodium carbonate solution in cup
- Citric acid solution in cup
- Universal indicator
- Thermometer
- Dropper

Procedure

1. Add 5 drops of universal indicator to the magnesium sulfate solution.
2. Place a thermometer in the cup and record the temperature of the solution.
3. Add 10 mL of sodium carbonate solution.
4. Add 10 mL of citric acid.



14. What clues do you observe that let you know that a chemical reaction is taking place?
15. In this chemical reaction, you may not have noticed a temperature change. Use what you know about energy and the breaking and making of bonds to explain how this can be.

Additional Teacher Background

Chapter 6, Lesson 7, p. 609

Chemical Reactions can be Exothermic or Endothermic

Previous Extra Teacher Background sections in Chapters 2 and 5 discuss the energy changes in the process of evaporation/condensation and dissolving. The processes in both contexts involve the breaking and making of bonds. In both cases the point was made that:

It takes energy to break bonds, and energy is released when bonds are formed

These same principles apply in the context of chemical reactions which can be either exothermic or endothermic. If it takes more energy to break the bonds of the reactants than is released when the bonds of the products are formed, then the reaction is endothermic and the temperature decreases.

If more energy is released when the bonds in the products form than it took to break the bonds of the reactants, then the reaction is exothermic and the temperature increases.

“Using” energy in bond-breaking and “releasing” energy in bond-making are really energy conversions between kinetic and potential energy. It takes a certain amount of kinetic energy to break the bonds holding atoms together in the molecules of the reactants. When the bonds are broken, this kinetic energy is converted to the potential energy of attraction between the atoms. When the atoms rebond to form the products, this potential energy is converted to kinetic energy. Depending on the combinations of bonds broken and made, a reaction is either endothermic or exothermic.

Chapter 6, Lesson 8: pH and Color Change

Key Concepts

- Whether a solution is acidic or basic can be measured on the pH scale.
- When universal indicator is added to a solution, the color change can indicate the approximate pH of the solution.
- Acids cause universal indicator solution to change from green toward red.
- Bases cause universal indicator to change from green toward purple.
- Water molecules (H_2O) can interact with one another to form H_3O^+ ions and OH^- ions.
- At a pH of 7, there are equal numbers of H_3O^+ ions and OH^- ions in water, and this is called a neutral solution.
- Acidic solutions have a pH below 7 on the pH scale.
- Basic solutions have a pH above 7 on the pH scale.

Summary

Students will see a demonstration of a color change using universal pH indicator. Students will change the concentrations of an acid and a base and use universal indicator to test the pH of the resulting solutions. Students will see an animation showing that water molecules interact and separate into the H_3O^+ ion and the OH^- ion. Students will see that the pH of a solution is related to the concentration of these ions in water.

Objective

Students will be able to explain, on the molecular level, that pH is a measure of the concentration of the H_3O^+ ions in water and that adding an acid or a base to water affects the concentration of these ions.

Evaluation

Download the student activity sheet, and distribute one per student when specified in the activity. The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan.

Safety

Be sure you and the students wear properly fitting goggles during the activity and wash hands afterwards. Sodium carbonate may irritate skin. Citric acid is an eye irritant. Universal indicator is alcohol-based and flammable. Read and follow all safety warnings on the label. At the end of the lesson, have students pour their used solutions in a waste container. Dispose of this waste down the drain or according to local regulations. The leftover citric acid and sodium carbonate powders can be disposed of with the classroom trash.

Materials for the Demonstration

- 3 clear plastic cups
- Citric acid
- Sodium carbonate
- Universal indicator solution
- Water

Materials for Each Group

- 3 clear plastic cups
- Masking tape and pen or permanent marker
- Universal indicator solution
- pH color chart
- Water
- Citric acid
- Sodium carbonate
- Graduated cylinder
- At least 12 flat toothpicks
- 2 6-well spot plates or 1 12-well spot plate
- 3 droppers

About the Materials

For this lesson, each group will need a Universal Indicator pH Color Chart. Print enough pages of these charts on a color printer so that each group can have its own chart, or purchase them from Flinn Scientific, Product #AP8765.

Each group will also need Universal Indicator Solution, Flinn Product #U0002, citric acid (anhydrous), Product #C0136 (500 grams) and sodium carbonate (anhydrous – Laboratory grade), Product #S0052. Each group will need either two 6-well spot plates or one 12-well spot plate. A porcelain 6-well spot plate is available from NASCO, Product #SB40727M. A polystyrene 12-well spot plate is available from Flinn Scientific, Product #AP6399.

About this Lesson

Because of their chemical properties, reactions involving acids and bases are different from the chemical reactions students have seen so far in Chapter 6. In the previous lessons, it was always the electrons that were being shared or transferred when atoms interacted. In the next three lessons about acids and bases, things are a little different. With acids and bases, it is a proton from a hydrogen atom that is transferred from one substance to another.

The main aspect of acids and bases that students will explore in the next three lessons deals with the influence of acids and bases on water. The reactions of acids and bases with water are measured using the pH scale. Understanding pH on the molecular level will give students a better appreciation for some of the environmental issues involving acids and bases. The meaning of pH and the way it is affected by acids and bases can be a little tricky, but by using animations, drawings, and some simplifications, students should be able to understand the main ideas.

ENGAGE

1. Add universal indicator solution to an acid and a base hidden in “empty” cups to demonstrate how an acid and a base can change the color of a pH indicator.



Materials for the Demonstration

- 3 clear plastic cups
- Citric acid
- Sodium carbonate
- Universal indicator solution
- Water

Note: Your local tap water is likely fine for the demonstration and activities in this lesson. If the indicator solution you make is not green, this means that your water is either acidic or basic. If this happens, use distilled water, which is available in supermarkets and pharmacies.

Teacher Preparation

Make indicator solution for student groups

- Make a dilute universal indicator solution for this demonstration and for each student group by combining 250 mL water with 10 mL universal indicator solution.
- Pour about 25 mL of this dilute universal indicator solution into a clean cup for each student group.

Note: In the activity, students will fill 12 wells with universal indicator solution. Check to make sure that 25 mL of solution is enough. You will need about 50 mL of indicator solution for your demonstration. If 250 mL of solution is not enough, make more using the same proportions.

Prepare for the Demonstration

- Pour about 50 mL indicator solution into a clear plastic cup for you to use in the demonstration.
- Using two empty clear plastic cups, add about $\frac{1}{8}$ teaspoon of citric acid to one cup and $\frac{1}{8}$ teaspoon of sodium carbonate to the other. Do not tell students that you have added anything to the cups.

Procedure

1. Pour about $\frac{1}{3}$ of the indicator solution into the citric acid cup and $\frac{1}{3}$ into the sodium carbonate cup. Leave $\frac{1}{3}$ in the indicator cup.



Expected Results

The citric acid turns the indicator from green to reddish. The sodium carbonate turns the indicator from green to purple.

Reveal to students that you put something in the cups beforehand.

Ask students:

- **Do you think this was a chemical reaction? Why or why not?**

A color change is often a clue that a chemical reaction has taken place. So the color change in each cup is likely the result of a chemical reaction. (This point is made in Chapter 6, Lesson 6.)

- **Would you say that the substances that were in the cups before the liquid was added were the same or different? Why?**

The liquid in each cup turned a different color during the reaction. Because substances react chemically in characteristic ways and the substances reacted differently, the substances in each cup must be different.

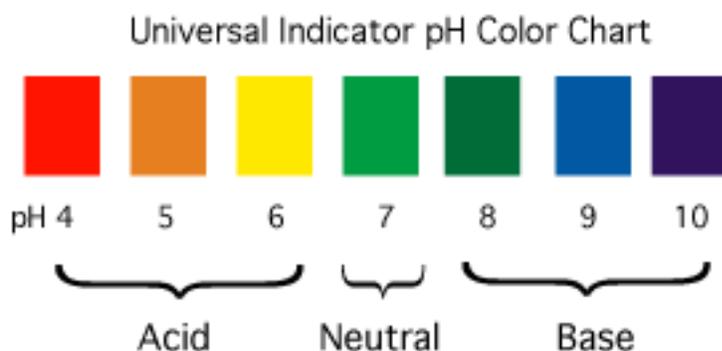
Tell students that the green solution was made by adding a substance called *universal indicator* to water. Explain that you put a small amount of a substance, one an acid and one a base, in each cup. Don't tell students which cup contained the acid or base.

Tell students that when you poured universal indicator solution into the cups, the acid and base each reacted with the indicator and changed its color. Usually, when two substances are mixed and a color change results, that is a clue that a chemical reaction has taken place. The cause of this color change will be discussed later in this lesson when students do their own activity.

Tell students that they will use an acid, a base, and universal indicator solution to learn about how acids and bases affect water. They will also learn how to measure the effect with colors and numbers on the pH scale.

2. Have students compare the color of the solutions made in the demonstration to the colors on the Universal Indicator pH Color Chart.

Distribute one Universal Indicator pH Color Chart to each group. The charts can be found right before the student activity sheets. Explain that the chart shows the range of color changes for universal indicator when acidic or basic solutions are added to the indicator. Point out that each color has a number associated with it and that students will learn more about these numbers later in the lesson. As the solution becomes more acidic, the color changes from green toward red. As the solution becomes more basic, the color changes from green toward purple.



Hold up the cups from the demonstration and ask the following questions:

- **What does the color of the liquid in each cup tell you about the substance that was already in the cup when the indicator was added?**

The cup that turned reddish initially contained an acid, and the cup that turned purple initially contained a base.

- **What does the green color of the indicator tell you about the water in that cup? Is it acidic, basic, or neither?**

The green indicator left in the cup is neither acidic nor basic, so it must be neutral.

3. Introduce the acid and base used in the demonstration and discuss how the color of universal indicator may change with other common acids and bases.

Explain that before class, you placed a small amount of citric acid in the cup that turned red and a small amount of sodium carbonate in the cup that turned purple. So citric acid is an acid and sodium carbonate is a base.

Acids and Universal Indicator Solution

Explain that citric acid is in citrus fruits such as lemons, limes, and oranges.

Ask students:

- **What are some other common examples of acids?**

Students might say that vinegar is an acid. You could point out that there are also stronger acids, like sulfuric acid used in car batteries.

- **What colors would you expect to see if you placed any of these substances in universal indicator?**

The color may change to yellow, orange, or red for these acids.

Bases and Universal Indicator Solution

Explain that sodium carbonate is one of the chemicals commonly used in detergents made for dishwashing machines.

Ask students:

- **What are some other common examples of bases?**

Students may not know any examples of bases but you can tell them that soaps, ammonia, and other cleaners are often bases.

- **What colors would you expect to see if you placed any of these substances in universal indicator?**

The color may change to dark green, blue, and purple for any of these bases. (For universal indicator, the changes in color for bases are not as different as they are for acids.)

Tell students that next they will explore the color changes of universal indicator with small amounts of citric acid and sodium carbonate.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. To find the answers to the activity sheet, go to the downloads area within the online version of this lesson.



EXPLORE

4. Have students prepare the solutions for the activity.

Explain to students that they will first make their solutions for the activity. Either go through each step with them or have them follow the procedure described on their activity sheet.



Teacher Preparation

Students will need small amounts of sodium carbonate and citric acid for the activity.

- Label two small plastic cups *citric acid* and *sodium carbonate* for each group.
- Place about $\frac{1}{4}$ teaspoon of **citric acid** and **sodium carbonate** in the labeled cups.
- Distribute the cups with universal indicator solution to each student group.

Materials for Each Group

- 2 clear plastic cups
- 3 droppers
- Masking tape and pen or permanent marker
- Universal indicator in cup
- Water

- Graduated cylinder
- Sodium carbonate
- Citric acid
- 2 flat toothpicks

Procedure

Label your equipment

1. Use masking tape and a pen to label one cup **citric acid solution** and another cup **sodium carbonate solution**.
2. Use a small piece of masking tape and a pen to label one dropper **citric acid solution** and the other dropper **sodium carbonate solution**.



Make a citric acid solution

3. Use your graduated cylinder to add 5 mL of water to the cup labeled citric acid.
4. Use a flat toothpick to pick up as much citric acid as you can on the end of the toothpick as shown.
5. Add this citric acid to the water in the citric acid cup. Gently swirl until the citric acid dissolves.



Make a sodium carbonate solution

6. Use your graduated cylinder to add 5 mL of water to the cup labeled sodium carbonate.
7. Use a flat toothpick to pick up as much sodium carbonate as you can on the end of a toothpick.
8. Add this sodium carbonate to the water in the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.



5. Explain what students will do in the next activity and discuss the purpose of having a control.

Explain to students that in this activity they will fill the wells in each spot plate with universal indicator solution. Then in the first spot plate, they will test how different concentrations of citric acid affect the color of universal indicator solution. In the other spot plate, they will test how different concentrations of sodium carbonate affect the color of universal indicator solution.

Tell students that in each spot plate, they will add nothing to the indicator solution in the first well. This is because the first well will serve as the control.

Ask students:

- **Why is it important to have a control?**

The control is left alone and not changed so that any color changes in the other wells can be compared to the original color in the control.

6. Have students test increasing concentrations of citric acid solution.

Question to Investigate

How does the concentration of citric acid affect the color of universal indicator solution?



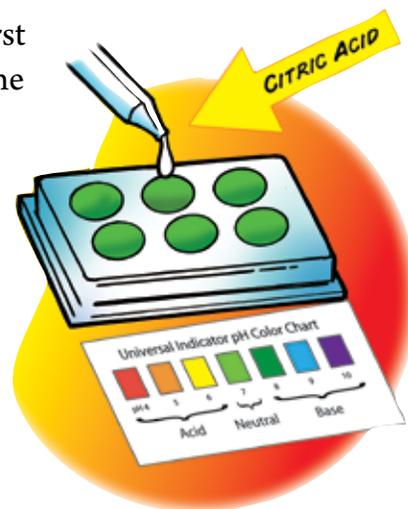
Materials for Each Group

- Universal indicator solution
- pH color chart
- Citric acid solution
- At least 6 toothpicks
- Spot plate
- 2 droppers

Procedure

Test your citric acid solution

1. Use one of your droppers to nearly fill 6 wells in your first spot plate with the universal indicator solution. Place the Universal Indicator pH Color Chart in front of the spot plate.
2. Use your dropper to add 1 drop of citric acid solution to the second well. Gently mix the liquid with a clean toothpick.
3. Compare the color of the liquid to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of citric acid, and the pH number in the chart on the activity sheet for well 2.



Expected Results

The color of the indicator should turn yellow-green or yellow. If there is no obvious color change after adding a toothpick of citric acid, have students add a little more citric acid to the solution. Tell them to be sure to pick up as much citric acid as they can on the end of a toothpick.

Record Observations

Help students fill out the chart on their activity sheet. Students may say that the color of the solution in well 2 is yellow or yellow-green. Then have students assign a number for pH. Tell students that if the color in the well seems to be between two colors on the chart, they should assign a pH value between the two.

Tell students that in the next part of the activity they will add a little more citric acid to the citric acid solution. This will make the citric acid solution more **concentrated**. Just as they did before, they will add one drop of citric acid solution, but this time the citric acid solution will be more concentrated.

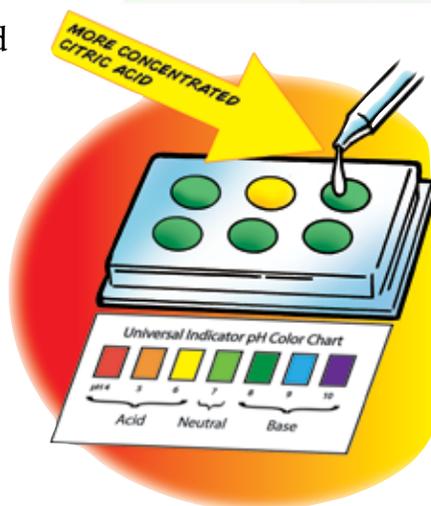
Ask students to make a prediction:

- How do you think the color will change if you add one drop of a more concentrated citric acid solution to the universal indicator in the next well?

Procedure

Test a more concentrated citric acid solution

4. Add another toothpick scoop of citric acid to the citric acid cup. Gently swirl until the citric acid dissolves.
5. Add 1 drop of this more concentrated citric acid solution to the *third* well. Gently mix the solution with a clean toothpick.
6. Compare the color of the solution to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpick scoops of citric acid added, and the pH number in the chart for well 3.
7. Continue adding toothpicks of citric acid and testing the solution in the last three wells to see how many different colors you can get.



Expected Results

As the citric acid solution becomes more concentrated, the color should change to variations of yellow-green, yellow, yellow-orange, orange, orange-red, and red. The colors obtained will vary from group to group because of the different amounts of citric acid students can pick up on the end of a toothpick. Students may be able to get 4 or 5 different colors. The answers and colors included in the chart below will vary.

The color and pH of different concentrations of citric acid			
Well number	Number of toothpicks of citric acid used in 5 mL of water	Color	pH
1	0	Green	7
2	1	Yellow-green	6.5
3	2	Yellow	6
4	3	Light orange	5.5
5	4	Peach	5
6	5	Pink	4

Ask students:

- **How does the color of the indicator solution change as the citric acid solution becomes more concentrated?**
As the citric acid solution becomes more concentrated, the color moves from green toward red on the pH color chart.
- **How does the number on the pH scale change as the concentration of citric acid solution increases?**
As the citric acid solution becomes more concentrated (more acidic), the number on the pH scale decreases.

7. Have students test increasing concentrations of sodium carbonate solution.

Note: The differences in color on the base side of the pH scale for universal indicator are not as obvious as those on the acid side. Students will have to look harder to see the difference between green-blue, blue, blue-purple, and purple.

Question to Investigate

How does the concentration of sodium carbonate affect the color of universal indicator solution?



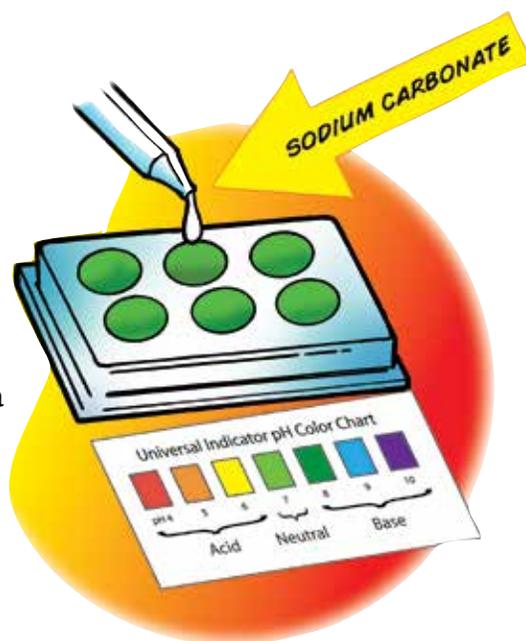
Materials for Each Group

- Universal indicator solution
- pH color chart
- Sodium carbonate solution
- At least 6 toothpicks
- Spot plate
- 2 droppers

Procedure

Test your sodium carbonate solution

1. Use a dropper to nearly fill the 6 wells in your other spot plate with universal indicator solution. You will not add anything else to the first well.
2. Add 1 drop of sodium carbonate solution to the second well. Gently mix the solution with a clean toothpick.
3. Compare the color of the solution to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of sodium carbonate used to make the solution, and the pH number in the chart for well 2.



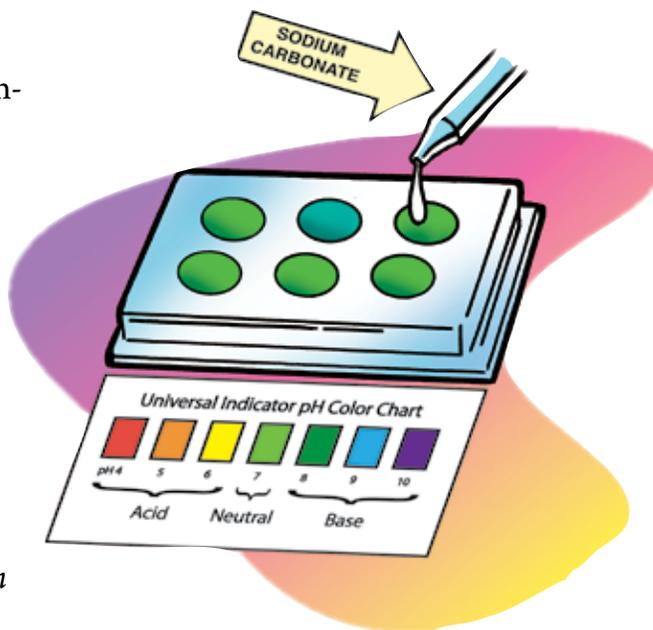
Expected Results

The color of the indicator should turn green-blue or blue.

Ask students to make a prediction:

- **How do you think the color will change if you add one drop of a more concentrated sodium carbonate solution to the universal indicator in the next well?**

Tell students that if you add more base to the same amount of water, the *concentration* of the base increases.



Procedure

Test a more concentrated sodium carbonate solution

4. Add another toothpick of sodium carbonate to the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.
5. Add 1 drop of sodium carbonate solution to the next well. Gently mix the liquid with a clean toothpick.
6. Compare the color of the liquid to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of sodium carbonate used, and the pH number in the chart for well 3.
7. Continue adding toothpicks of sodium carbonate and testing the solution in the last three wells to see how many different colors or shades you can make.



Expected Results

The more concentrated sodium carbonate solution should cause the color to change to a darker blue moving toward purple. Answers and colors in the chart below will vary.

The color and pH of different concentrations of sodium carbonate			
Well number	Number of toothpicks of sodium carbonate used in 5 mL of water	Color	pH
1	0	Green	7
2	1	Green-blue	8
3	2	Blue	8.5
4	3	Blue-purple	9
5	4	Purple	9.5
6	5	Purple	10

Ask students:

- **How does the color of the indicator solution change as the sodium carbonate solution becomes more concentrated?**

As the sodium carbonate solution becomes more concentrated, the color moves from green toward purple on the pH color chart

- **How does the number on the pH scale change as the concentration of base increases?**

As the sodium carbonate solution becomes more concentrated (more basic), the number on the pH scale increases.

EXPLAIN

8. Explain how water molecules interact with each other to form ions.

Tell students that pH has to do with the way acids and bases interact with water. Explain that first you will show students how water molecules interact with each other before an acid or a base is added.

Project the animation *Proton Transfer in Water*

Play the first part of the animation.

Remind students that each hydrogen atom in a water molecule has both a proton and an electron. The hydrogen atoms share their electrons with the oxygen atom.

Click “next” to show how the water molecules become ions.

Water molecules continuously move and bump into one another. Sometimes when two water molecules come together, a proton from one hydrogen atom leaves its water molecule and becomes part of another water molecule. Only the positively charged proton moves; the negatively charged electron stays behind. So, these two H_2O molecules become the ions H_3O^+ and OH^- .

Click “next” again to show how the ions become water molecules again

Explain that when these ions bump into each other, the proton from the H_3O^+ can move over to the OH^- ion, forming two regular water molecules again. Because protons go back and forth between the water molecules or between ions continuously, there is always the same amount of H_3O^+ and OH^- ions in water.

Project the illustration *Water Molecules Trade Protons.*

This illustration shows the chemical equations that explain how water molecules can become ions and how ions can become water molecules again.

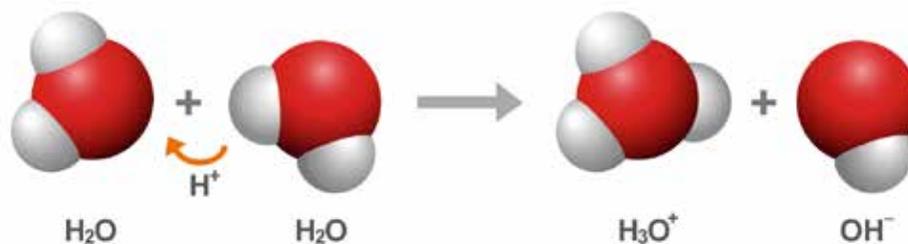
Explain to students that the first chemical equation shows two water molecules coming together. Point out the chemical formula for each water molecule, H_2O .

- *Explain the formation of the H_3O^+ ion.*

After the proton is transferred, the water molecule that now has the extra proton is called an H_3O^+ ion. The reason why the number of hydrogen atoms changed from two (the subscript in H_2) to three (the subscript in H_3) is because having an extra proton is like having an extra hydrogen atom, even though the electron did not come over with it. Because one proton was added, there is one more proton than electrons, making this a positive ion.

- *Explain the formation of the OH⁻ ion.*

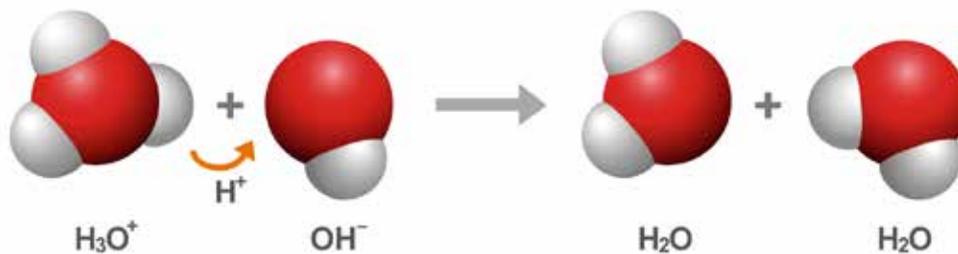
The water molecule that lost a proton now has an extra electron, so it is called the OH⁻ ion. The reason why the number of hydrogen atoms changed from two (the subscript in H₂) to one (no subscript after the H means 1 hydrogen) is because losing a proton is like losing a hydrogen atom. Because only the proton was transferred, there is one more electron than proton, making this a negative ion.



Tell students that the second chemical equation shows an H₃O⁺ ion and OH⁻ ion coming together to become water molecules again.

- *Explain the reformation of two H₂O molecules.*

Explain to students that water molecules and ions are always colliding. When an H₃O⁺ ion and an OH⁻ ion bump into each other, the proton can be transferred from the H₃O⁺ ion over to the OH⁻ ion so that each ion becomes an H₂O molecule again.



At any given time in an ordinary sample of water, a small percentage of water molecules are transferring protons and becoming ions. Also, the H₃O⁺ and OH⁻ ions are transferring protons and becoming water molecules again.

9. Explain how acids and bases cause the indicator to change color.

Project the animation *Acids Donate Protons.*

Tell students that when an acid is added to an indicator solution, the acid donates protons to the water molecules. This increases the concentration of H₃O⁺ ions in the solution.

The H₃O⁺ ions donate protons to the indicator molecules causing the indicator to change color toward red.

Project the animation *Bases Accept Protons*.

When a base is added to an indicator solution, it accepts protons from the water molecules, creating OH^- ions. The H_3O^+ ions and indicator molecules donate protons to the OH^- ions, causing the indicator to change color toward purple.

Read more about strength vs. concentration in acids and bases in the additional teacher background section at the end of the lesson.

EXTEND

10. Have students slowly pour their remaining acidic and basic solutions into the indicator solution to introduce the idea that acids and bases can neutralize each other.

Ask students to make a prediction:

- How do you think the color will change if you pour a small amount of each leftover solution into your universal indicator solution?

Materials for Each Group

- Universal indicator solution
- pH color chart
- Citric acid solution
- Sodium carbonate solution

Procedure

1. Pour a small amount of either your citric acid solution or sodium carbonate solution into your indicator solution. Swirl and compare the color to your Universal Indicator pH Color Chart.
2. Pour a small amount of the other solution into your indicator solution. Swirl and compare the color to your color chart.
3. Continue pouring small amounts of the acid and base solutions into your indicator until the solutions are used up.



Expected Results

The colors of the indicator solution will vary, but students should see that acids and bases mixed together cause the color of the indicator to change toward neutral.

Have students describe what they did and their observations. Then explain that in Chapter 6, Lesson 9, they will combine acids and bases in an indicator solution with the goal of making the pH of the final solution neutral.

Activity Sheet
Chapter 6, Lesson 8
pH and Color Change

Name _____

Date _____

DEMONSTRATION

1. Your teacher poured green universal indicator into each of two cups. What does the change in color of the indicator solution tell you about the substance your teacher placed in each cup?



PREPARE FOR THE ACTIVITY

Materials for Each Group

- 2 clear plastic cups
- 3 droppers
- Masking tape and pen or permanent marker
- Universal indicator in cup
- Water
- Graduated cylinder
- Sodium carbonate
- Citric acid
- 2 flat toothpicks



Procedure

Label your equipment

1. Use masking tape and a pen to label one cup **citric acid solution** and another cup **sodium carbonate solution**.
2. Use a small piece of masking tape and a pen to label one dropper **citric acid solution** and the other dropper **sodium carbonate solution**.



Make a citric acid solution

3. Use your graduated cylinder to add 5 mL of water to the cup labeled **citric acid**.
4. Use a flat toothpick to pick up as much citric acid as you can on the end of the toothpick as shown.
5. Add this citric acid to the water in the citric acid cup. Gently swirl until the citric acid dissolves.



Make a sodium carbonate solution

6. Use your graduated cylinder to add 5 mL of water to the cup labeled **sodium carbonate**.
7. Use a flat toothpick to pick up as much sodium carbonate as you can on the end of a toothpick.
8. Add this sodium carbonate to the water in the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.



ACTIVITY

Question to Investigate

How does the concentration of citric acid affect the color of universal indicator solution?



Materials for Each Group

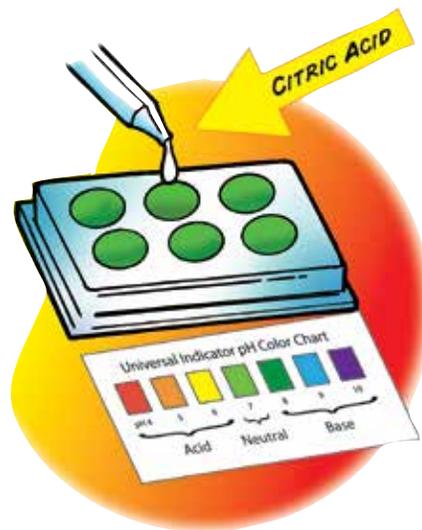
- Universal indicator solution
- pH color chart
- Citric acid solution
- At least 6 toothpicks
- Spot plate
- 2 droppers

Procedure

Test your citric acid solution

1. Use one of your droppers to nearly fill 6 small wells in your first spot plate with the universal indicator solution. Place the Universal Indicator pH Color Chart in front of the spot plate.

- Use your dropper to add 1 drop of citric acid solution to the second well. Gently mix the liquid with a clean toothpick.
- Compare the color of the liquid to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of citric acid, and the pH number in the chart for well 2.



Test a more concentrated citric acid solution

- Add another toothpick scoop of citric acid to the citric acid cup. Gently swirl until the citric acid dissolves.
- Add 1 drop of this more concentrated citric acid solution to the *third* well. Gently mix the solution with a clean toothpick.
- Compare the color of the solution to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpick scoops of citric acid added, and the pH number in the chart for well 3.
- Continue adding toothpicks of citric acid and testing the solution in the last three wells to see how many different colors you can get.



The color and pH of different concentrations of citric acid			
Well number	Number of tiny scoops of citric acid used in 5 mL of water	Color	pH
1	0		7
2	1		
3	2		
4	3		
5	4		
6	5		

2. How does the color of the indicator solution change as the citric acid solution becomes more concentrated?
3. How does the number on the pH scale change as the concentration of citric acid solution increases?

Question to Investigate

How does the concentration of sodium carbonate affect the color of universal indicator solution?



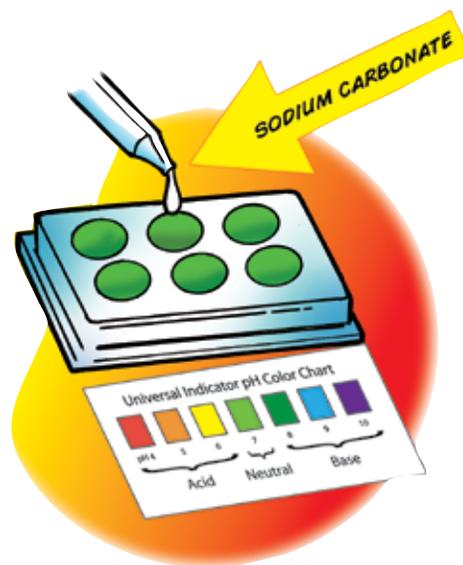
Materials for Each Group

- Universal indicator solution
- pH color chart
- Sodium carbonate solution
- At least 6 toothpicks
- Spot plate
- 2 droppers

Procedure

Test your sodium carbonate solution

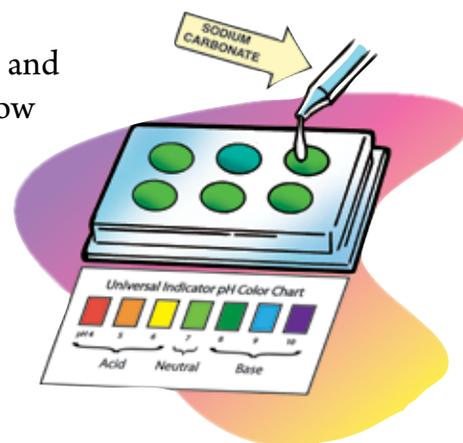
1. Use a dropper to nearly fill the 6 wells in your other spot plate with universal indicator solution. You will not add anything else to the first well.
2. Add 1 drop of sodium carbonate solution to the second well. Gently mix the solution with a clean toothpick.



3. Compare the color of the solution to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of sodium carbonate used to make the solution, and the pH number in the chart for well 2.

Test a more concentrated sodium carbonate solution

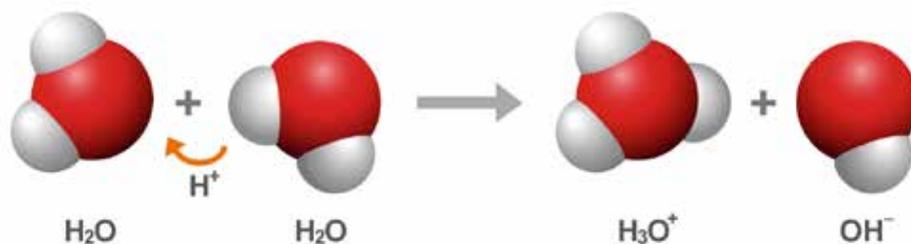
4. Add another toothpick of sodium carbonate to the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.
5. Add 1 drop of sodium carbonate solution to the next well. Gently mix the liquid with a clean toothpick.
6. Compare the color of the liquid to the control and to the Universal Indicator pH Color Chart. Record the color of the indicator, the number of toothpicks of sodium carbonate used, and the pH number in the chart for well 3.
7. Continue adding toothpicks of sodium carbonate and testing the solution in the last three wells to see how many different colors or shades you can make.



The color and pH of different concentrations of sodium carbonate			
Well number	Number of tiny scoops of sodium carbonate used in 5 mL of water	Color	pH
1	0		7
2	1		
3	2		
4	3		
5	4		
6	5		

EXPLAIN IT WITH ATOMS & MOLECULES

7. The chemical formula for water is H_2O . Sometimes two water molecules can bump into each other and form the ions H_3O^+ and OH^- .



What is happening in the chemical equation above?

Why is one ion positive and the other ion negative?

8. The pH scale is a measure of the concentration of H_3O^+ ions in a solution. Use the words *increases*, *decreases*, or *stays the same* to describe how the concentration of H_3O^+ ions changes as different substances are added to water.

How does the concentration of H_3O^+ ions change as each substance is added to water?	
Type of substance	Concentration of H_3O^+ ions
Acid	
Base	
Neutral	

TAKE IT FURTHER

Question to Investigate

How will the color change as you slowly pour your acid and base solutions into the indicator?



Materials for Each Group

- Universal indicator solution
- pH color chart
- Citric acid solution
- Sodium carbonate solution

Procedure

1. Pour a small amount of either your citric acid solution or sodium carbonate solution into your indicator solution. Swirl and compare the color to your Universal Indicator pH Color Chart.
2. Pour a small amount of the other solution into your indicator solution. Swirl and compare the color to your color chart.
3. Continue to pour small amounts of the acid and base solutions into your indicator until the solutions are used up.



9. What did you observe as you slowly poured your acid and base solutions into the indicator solution?

Chapter 6, Lesson 9: Neutralizing Acids and Bases

Key Concepts

- pH is a measure of the concentration of H_3O^+ ions in a solution.
- Adding an acid increases the concentration of H_3O^+ ions in the solution.
- Adding a base decreases the concentration of H_3O^+ ions in the solution.
- An acid and a base are like chemical opposites.
- If a base is added to an acidic solution, the solution becomes less acidic and moves toward the middle of the pH scale. This is called *neutralizing* the acid.
- If an acid is added to a basic solution, the solution becomes less basic and moves toward the middle of the pH scale. This is called *neutralizing* the base.

Summary

Students will use citric acid and sodium carbonate solutions to see that adding a base to an acidic solution makes the solution less acidic. Students will then use a base to help them identify which of two acidic solutions is more concentrated.

Objective

Students will be able to explain, on the molecular level, that pH is affected by the concentration of the H_3O^+ ions in water. They will also be able to explain why adding a base to an acid or an acid to a base can make the pH of the solution closer to 7.

Evaluation

Download the student activity sheet, and distribute one per student when specified in the activity. The activity sheet will serve as the “Evaluate” component of each S-E lesson plan.

Safety

Be sure you and the students wear properly fitting goggles during the activity and wash hands afterwards. Sodium carbonate may irritate skin. Citric acid is an eye irritant. Universal indicator is alcohol-based and flammable. Read and follow all safety warnings on the label. At the end of the lesson, have students pour their used solutions in a waste container. Dispose of this waste down the drain or according to local regulations. The leftover citric acid and sodium carbonate powders can be disposed of with the classroom trash.

Materials for the Demonstration

- 4 clear plastic cups
- Graduated cylinder
- Universal indicator
- Water
- Sodium carbonate
- Citric acid
- Flat toothpicks
- 2 droppers
- Masking tape and pen or permanent marker

Materials for Each Group

- Universal indicator solution in cup
- Citric acid in cup
- Sodium carbonate in cup
- Water
- Solution A, sodium carbonate solution
- Solution B, more concentrated sodium carbonate solution
- At least 8 flat toothpicks
- Graduated cylinder
- Spot plate
- 4 droppers
- 3 clear plastic cups
- Masking tape and pen or permanent marker

About the Materials

Each group will need Universal Indicator Solution, Flinn Product #U0002, citric acid (anhydrous), product # C0136 (500 grams) and sodium carbonate (anhydrous—Laboratory grade), Product #S0052. Each group will also need a spot plate. A porcelain 6-well spot plate is available from NASCO, Product #SB40727M. A polystyrene 12-well spot plate is available from Flinn Scientific, Product #AP6399.

ENGAGE

1. Do a demonstration to show students that an acidic solution becomes less acidic when drops of a base are added.

Materials for the Demonstration

- 4 clear plastic cups
- Graduated cylinder
- Universal indicator
- Water
- Sodium carbonate
- Citric acid
- Flat toothpicks
- 2 droppers
- Masking tape and pen or permanent marker



Teacher Preparation

Make indicator solution for student groups

- Make a dilute universal indicator solution for this demonstration and for each student group by combining 125 mL water with 5 mL universal indicator solution.
- Pour about 15 mL of this dilute universal indicator solution into a clean cup for each student group.

Note: Your local tap water is likely fine for the demonstration and activities in this lesson. If the indicator solution you make is not green, this means that your water is either acidic or basic. If this happens, use distilled water, which is available in supermarkets and pharmacies.

Note: In the engage and extend activities, students will fill 6 wells with universal indicator solution. Check to make sure that 15 mL of solution is enough. You will need about 25 mL of indicator solution for your demonstration. If 125 mL of solution is not enough, make more using the same proportions.

Prepare for the Demonstration

- Divide the remaining indicator solution into two clear plastic cups for you to use in the demonstration.
- Use masking tape and a pen to label two empty cups **citric acid** and **sodium carbonate**.
- Use your graduated cylinder to add 5 mL of water to each labeled cup.
- Use a flat toothpick to pick up as much citric acid as you can on the end of the toothpick as shown. Add this citric acid to the water in the citric acid cup. Gently swirl until the citric acid dissolves.
- Use a flat toothpick to pick up as much sodium carbonate as you can on the end of a toothpick. Add this sodium carbonate to the water in the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.

Procedure

1. Hold up the two cups of universal indicator solution, which are both green.
2. Also show students that you have a citric acid solution and a sodium carbonate solution.

Ask students:

- **What color will the green indicator solution turn if I add a few drops of citric acid solution?**
The indicator solution will change color toward red.

Procedure

3. Add 3–5 drops of citric acid solution to one of the cups.

Expected Results

The color of the solution should change from green to reddish.

Ask students:

- **What do you think you could add to the reddish indicator to make it less acidic and go back toward green?**

Students should suggest adding sodium carbonate (a base) to the acidic (red) solution.

Procedure

4. While holding up the cup of reddish indicator solution, add 1 drop of sodium carbonate solution, swirl, and compare the color of the solution to the color of the control.
5. Add another drop if necessary to get closer to the green color of the control. Continue adding drops until the color gets close to green. If you add a drop and the color goes past green to blue, ask students what the blue color tells you about the solution. The blue indicates that the solution has gone from being acidic to basic.



Explain that acids and bases are like chemical opposites. Tell students that they will experiment to figure out how many drops of a basic solution it takes to cause an acidic solution to move to the middle of the pH scale. This is called *neutralizing* the acid.

Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. To find the answers to the activity sheet, go to the downloads area within the online version of this lesson.



EXPLORE

2. Have students prepare the solutions for the activity.



Teacher Preparation

Students will need small amounts of sodium carbonate and citric acid for the activity.

- Label two small plastic cups **citric acid solution** and **sodium carbonate solution** for each group.
- Place about $\frac{1}{4}$ teaspoon of citric acid and sodium carbonate in the labeled cups.
- Distribute the cups with universal indicator solution to each student group.

Materials for Each Group

- Sodium carbonate in cup
- Citric acid in cup
- Universal indicator in cup
- Water
- 3 clear plastic cups
- Graduated cylinder
- Flat toothpicks
- 2 droppers
- Spot plate
- Masking tape and pen or permanent marker



Procedure

Label your equipment

1. Use masking tape and a pen to label one cup **citric acid solution** and another cup **sodium carbonate solution**.
2. Use a small piece of masking tape and a pen to label one dropper **citric acid solution** and the other dropper **sodium carbonate solution**.



Make a citric acid solution

3. Use your graduated cylinder to add 5 mL of water to the cup labeled citric acid.
4. Use a flat toothpick to pick up as much citric acid as you can on the end of the toothpick as shown.
5. Add this citric acid to the water in the citric acid cup. Gently swirl until the citric acid dissolves.

Make a sodium carbonate solution

6. Use your graduated cylinder to add 5 mL of water to the cup labeled sodium carbonate.
7. Use a flat toothpick to pick up as much sodium carbonate as you can on the end of a toothpick.
8. Add this sodium carbonate to the water in the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.



3. Have students neutralize an acidic solution.

Question to Investigate

How many drops of sodium carbonate solution will it take to neutralize your citric acid solution?

Materials for Each Group

- Universal indicator solution
- Citric acid solution
- Sodium carbonate solution
- At least 6 flat toothpicks
- Spot plate
- 3 droppers

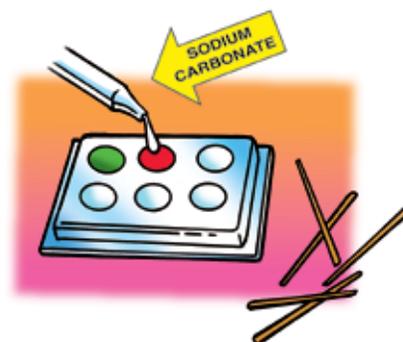
Procedure

1. Use a dropper to nearly fill two small wells in your spot plate with universal indicator solution. Do not add anything else to the first well. This will be your control.
2. Add 3 drops of citric acid solution to the indicator in one of the wells. Use a clean toothpick to mix the solution. If it is not reddish, add more drops, but be sure to count the total number of drops added.

Ask students:

- **What could you then add in order to make the indicator solution less acidic?**
Adding a base, like the sodium carbonate solution, will make the solution less acidic.
- **Should you add one drop of sodium carbonate solution at a time or a few drops at once?**
You should add one drop at a time to better monitor how many more drops of the sodium carbonate solution should be added.
- **How will you know when the solution is neutralized?**
The color of the solution will be similar to the color of the control. Tell students that if the solution turns blue, it has gone from an acid, past neutral, and is now a base. If this happens, try adding one or more drops of citric acid until the color is close to neutral. Be sure to keep track of the total number of drops of acid and base you add.

3. Add single drops of sodium carbonate to the same well in which you added the acid. Be sure to count the drops you use and stir with a toothpick after adding each drop.



Expected Results

With each drop of sodium carbonate, the citric acid solution will move toward neutral, eventually becoming green.

Note: The solution may get close to the green color of the control, but will probably not be exact. This is because the citric acid and sodium carbonate solutions are not exactly equal in the way they act as acid and base. Also, to be very exact, students would need to be able to use half-drops or even quarter-drops, which is not possible with the droppers the students are using. As long as students see a trend toward the green control color, that is good enough.

How many drops of sodium carbonate does it take to neutralize your citric acid solution?		
Acidic solution	Number of drops of citric acid solution added to the indicator	Number of drops of sodium carbonate solution needed to neutralize the citric acid solution
First citric acid solution	3 drops	
More concentrated citric acid solution		

4. Discuss student observations.

- **How many drops of sodium carbonate did it take to bring the color back to the color of the control?**

Results will vary but it should take fewer drops of sodium carbonate than drops of citric acid to neutralize the solution.

- **Does the solution become more acidic or less acidic as each drop of sodium carbonate is added to the indicator?**

The solution becomes less acidic.

- **How do you use the color of the control to help you neutralize an acid?**

When the color of the universal indicator solution becomes near green, the acidic solution has been neutralized.

EXPLAIN

5. Explain how adding a base to an acidic solution affects the concentration of H_3O^+ ions.

Project the animation *Neutralizing an Acidic Solution*.

Explain to students that adding drops of citric acid to the indicator solution increased the concentration H_3O^+ ions. When you add a base to this acidic solution, the base accepts protons from the water molecules creating OH^- ions. The H_3O^+ ions and indicator molecule transfer protons to the OH^- ions. When enough base is added so that the concentration of H_3O^+ and OH^- ions becomes equal, the solution is neutralized.

EXPLORE

6. Have students compare how many more drops of a base it takes to neutralize a more concentrated acidic solution.

Question to Investigate

How many more drops of sodium carbonate solution will it take to neutralize a more concentrated citric acid solution?

Materials for Each Group

- Citric acid
- Citric acid solution
- Sodium carbonate solution
- Universal indicator solution
- 2 flat toothpicks
- 3 droppers
- Spot plate

Procedure

Neutralize a citric acid solution

1. Use a flat toothpick to add two scoops of citric acid to your citric acid solution to make it even more acidic. Gently swirl until the citric acid dissolves.
2. Add universal indicator solution to a clean well in the spot plate.
3. Add 3 drops of the more concentrated citric acid solution to the indicator and stir with a clean toothpick.



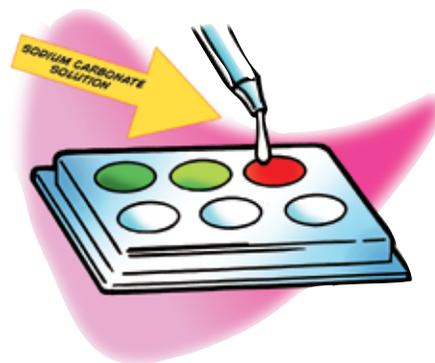
Ask students:

- **Do you think it will take more, less, or the same amount of sodium carbonate solution to neutralize this more concentrated citric acid solution?**
It will take more drops of the base to neutralize the more concentrated citric acid solution.
- **Thinking about the animation, why will you need more drops of sodium carbonate solution?**
Since the solution is more acidic, there are more H_3O^+ ions. So it takes more molecules of the base to accept the extra protons and move more toward neutral.

Procedure

Neutralize a more concentrated citric acid solution

4. Add single drops of sodium carbonate solution to the same well in which you added the acid. Be sure to count the drops you use and stir with a toothpick after adding each drop. Record this number in the chart.



EXTEND

7. Have students neutralize two basic solutions to determine which is most concentrated.

Materials for each group

- Universal indicator solution
- Citric acid solution
- Solution A
- Solution B
- At least 6 toothpicks
- Spot plate
- 3 droppers

Teacher Preparation

Make two mystery solutions using different amounts of sodium carbonate.

- Label two cups *Solution A* and *Solution B* for each group.
- Make a class set of solutions A and B.
 - Solution A: 50 mL of water and 5 toothpicks of sodium carbonate
 - Solution B: 50 mL of water and 10 toothpicks of sodium carbonate
- Place about 5 mL of each solution in their labeled cups.

Ask students:

- **Solutions A and B are both basic solutions made with sodium carbonate and water. One of these solutions has more sodium carbonate than the other. How can you figure out which solution is more concentrated?**

Students should describe a procedure very similar to the one they used to neutralize the two citric acid solutions. They should suggest that they neutralize each sodium carbonate solution with drops of citric acid and count how many drops it takes to neutralize each solution. When the color of the solution is close to the color of the control, the solution is neutralized.

- **How will you know which solution is the most concentrated?**

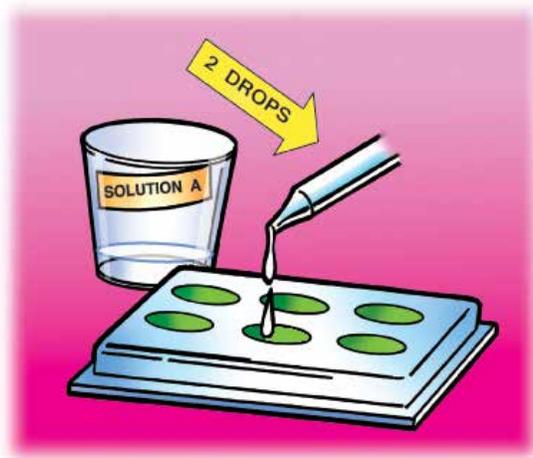
The solution that requires the greatest number of drops of citric acid to become neutral is the most basic.

Discuss what students will do:

- For best results, have students place 2 drops of Solution A in one well filled with indicator solution and 2 drops of Solution B in another well filled with indicator solution.
- Then they should add single drops of citric acid solution, stir, and compare the color to the color of the control.
- Students should keep track of the number of drops of citric acid it takes to neutralize each solution.

Procedure

1. Add universal indicator solution to three wells in a clean spot plate.
2. Leave the first well alone so that it can be used as a control. Add 2 drops of Solution A to the second well.
3. Add 2 drops of Solution B to the third well.
4. Neutralize Solution A. Record the number of drops used in the chart.
5. Neutralize Solution B. Record the number of drops used in the chart.



Which solution is the most concentrated?		
Solution	Number of drops of solution added to the indicator	Number of drops of citric acid solution needed to neutralize the sodium carbonate solution
Solution A	2 drops	
Solution B	2 drops	

Ask students:

- **Which solution is the most concentrated? How do you know?**
Students should discover that it takes more drops of citric acid to neutralize Solution B. Therefore, Solution B must be more concentrated than Solution A.
- **Antacids are medicines people take when the acid in their stomach is causing them discomfort. One advertisement says that the medicine provides relief for acid indigestion and sour stomach. What type of chemical do you think is in the medicine?**
Bases neutralize acids, so the chemical is probably a base.

DEMONSTRATION

1. Your teacher added drops of an acid to a universal indicator solution and then neutralized the solution by adding drops of a base. How did you know when the solution became close to neutral?



PREPARE FOR THE ACTIVITY

Materials for Each Group

- Sodium carbonate in cup
- Citric acid in cup
- Universal indicator in cup
- Water
- 3 clear plastic cups
- Graduated cylinder
- Flat toothpicks
- 2 droppers
- Masking tape and pen or permanent marker



Procedure

Label your equipment

1. Use masking tape and a pen to label one cup **citric acid solution** and another cup **sodium carbonate solution**.
2. Use a small piece of masking tape and a pen to label one dropper **citric acid solution** and the other dropper **sodium carbonate solution**.



Make a citric acid solution

3. Use your graduated cylinder to add 5 mL of water to the cup labeled citric acid.
4. Use a flat toothpick to pick up as much citric acid as you can on the end of the toothpick as shown.
5. Add this citric acid to the water in the citric acid cup. Gently swirl until the citric acid dissolves.



Make a sodium carbonate solution

6. Use your graduated cylinder to add 5 mL of water to the cup labeled sodium carbonate.
7. Use a flat toothpick to pick up as much sodium carbonate as you can on the end of a toothpick.
8. Add this sodium carbonate to the water in the sodium carbonate cup. Gently swirl until the sodium carbonate dissolves.



ACTIVITY

Question to Investigate

How many drops of sodium carbonate solution will it take to neutralize your citric acid solution?



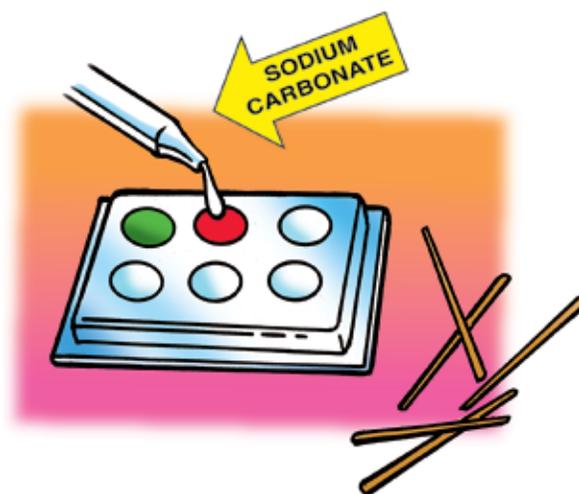
Materials for Each Group

- Universal indicator solution
- Citric acid solution
- Sodium carbonate solution
- At least 6 flat toothpicks
- Spot plate
- 3 droppers

Procedure

1. Use a dropper to nearly fill two small wells in your spot plate with universal indicator solution. Do not add anything else to the first well. This will be your control.
2. Add 3 drops of citric acid solution to the indicator in one of the wells. Use a clean toothpick to mix the solution. If it is not reddish, add more drops, but be sure to count the total number of drops added.

3. Add single drops of sodium carbonate to the same well in which you added the acid. Be sure to count the drops you use and stir with a toothpick after adding each drop.



How many drops of sodium carbonate does it take to neutralize your citric acid solution?		
Acidic solution	Number of drops of citric acid solution added to the indicator	Number of drops of sodium carbonate solution needed to neutralize the citric acid solution
First citric acid solution	3 drops	
More concentrated citric acid solution		

2. Does the solution become more acidic or less acidic as each drop of sodium carbonate is added to the indicator?

EXPLAIN IT WITH ATOMS & MOLECULES

3. What happens to the protons from the H_3O^+ ions when a base is used to neutralize an acid?
4. What do you know about the concentration of H_3O^+ ions and OH^- ions when a solution is neutralized?

ACTIVITY

Question to Investigate

How many more drops of sodium carbonate solution will it take to neutralize a more concentrated citric acid solution?

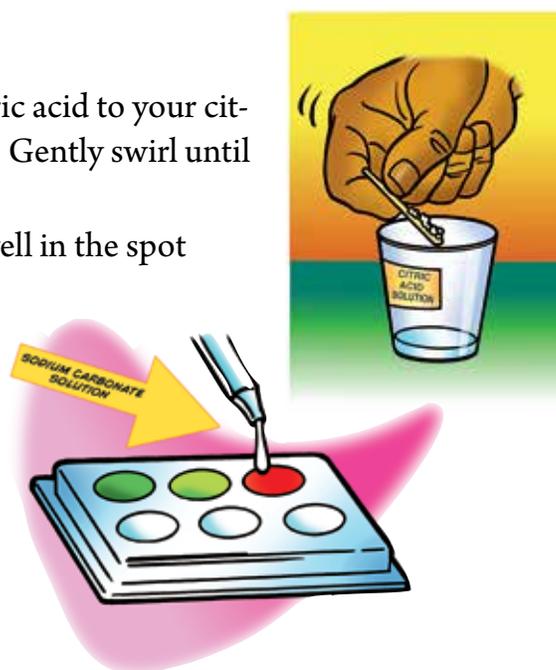


Materials for Each Group

- Citric acid
- Citric acid solution
- Sodium carbonate solution
- Universal indicator solution
- 2 flat toothpicks
- 3 droppers
- Spot plate

Procedure

1. Use a flat toothpick to add two scoops of citric acid to your citric acid solution to make it even more acidic. Gently swirl until the citric acid dissolves.
2. Add universal indicator solution to a clean well in the spot plate.
3. Add 3 drops of the more concentrated citric acid solution to the indicator and stir with a clean toothpick.
4. Add single drops of sodium carbonate to the same well in which you added the acid. Be sure to count the drops you use and stir with a toothpick after adding each drop. Record this number in the chart.



5. Did it take *more, less, or the same* amount of sodium carbonate solution to neutralize this more concentrated citric acid solution?
6. Thinking about the animation, why will you need more drops of sodium carbonate solution?

TAKE IT FURTHER

Question to Investigate

Is Solution A or Solution B a more concentrated basic solution?

Materials for Each Group

- Universal indicator solution
- Citric acid solution
- Solution A
- Solution B
- At least 6 toothpicks
- Spot plate
- 3 droppers

Procedure

1. Add universal indicator solution to three wells in a clean spot plate.
2. Leave the first well alone so that it can be used as a control. Add 2 drops of Solution A to the second well.
3. Add 2 drops of Solution B to the third well.
4. Neutralize Solution A. Record the number of drops used in the chart.
5. Neutralize Solution B. Record the number of drops used in the chart.



Which solution is the most concentrated?		
Solution	Number of drops of solution added to the indicator	Number of drops of citric acid solution needed to neutralize the sodium carbonate solution
Solution A	2 drops	
Solution B	2 drops	

7. Which solution is the most concentrated? How do you know?

8. Antacids are medicines people take when the acid in their stomach is causing them discomfort. One advertisement says that the medicine provides relief for acid indigestion and sour stomach. What type of chemical do you think is in the medicine?

Chapter 6, Lesson 10: Carbon Dioxide Can Make a Solution Acidic

Key Concepts

- Carbon dioxide (CO_2) gas dissolved in water can cause water to become acidic.
- The acidity of water from dissolved CO_2 can be reduced by a base such as baking soda (sodium bicarbonate).

Summary

The teacher blows into a universal indicator solution until it changes color. Students interpret this color change and explain that the solution becomes acidic. Students explore whether carbon dioxide from other sources, namely carbonated water and a chemical reaction between baking soda and vinegar, can also make a solution acidic and whether a base can neutralize it. Students then apply their learning about acids, bases, and neutralization from lessons 8, 9, and 10 as they place an Alka-Seltzer[®] tablet in indicator solution.

Objective

Students will be able to explain that carbon dioxide from any source reacts chemically with water to form carbonic acid. They will also be able to use the color changes of universal indicator to monitor the changing pH of a solution during a chemical reaction.

Evaluation

Download the student activity sheet, and distribute one per student when specified in the activity. The activity sheet will serve as the “Evaluate” component of each 5-E lesson plan.

Safety

Be sure you and the students wear properly fitting goggles during the activity and wash hands afterwards. Universal indicator is alcohol-based and flammable. Use vinegar in a well-ventilated room. Read and follow all safety warnings on the label. Because Alka-Seltzer contains aspirin, check with school administrators before doing the extend portion of this lesson. Students do not touch the tablet because it is sealed in a zip-closing plastic bag.

At the end of the lesson, collect the sealed inflated plastic bags. Have students pour their other used solutions in a common waste container. Later, dispose of all liquid waste down the drain or according to local regulations. The leftover baking soda and plastic bags can be disposed of with the classroom trash.

Materials for the Demonstration

- Universal indicator solution
- Water
- 2 clear plastic cups
- Straw

Materials for Each Group

- Water
- Universal indicator solution in cup
- Universal indicator pH color chart
- Carbonated water (club soda or seltzer water) in wide, clear, plastic cup
- Baking soda in wide, clear, plastic cup
- Vinegar
- Alka-Seltzer tablet
- 2 small clear plastic cups
- 4 wide clear plastic cups
- 4 taller, clear, plastic cups
- Graduated cylinder
- Snack-sized zip-closing plastic bag

About the Materials

For this lesson, each group will need a Universal Indicator pH color chart. Print enough pages of these charts on a color printer so that each group can have their own chart. You may also choose to purchase them from Flinn Scientific, Product number AP8765.

Each group will also need universal indicator solution, Flinn Product #U0002, Flinn Product #C0136 (500 grams), and sodium bicarbonate (baking soda).

ENGAGE

- 1. Do a demonstration to show that adding CO₂ gas to water can make the water become acidic.**

Materials for the Demonstration

- Universal indicator solution
- Water
- 2 clear plastic cups
- Straw

Teacher Preparation

Make indicator solution for student groups

- Make a dilute universal indicator solution for this demonstration and for each student group by combining 625 mL water with 25 mL universal indicator solution.
- Pour at least 80 mL of this dilute universal indicator solution into a clean plastic cup for each student group.

Note: *Your local tap water is likely fine for the demonstration and activities in this lesson. If the indicator solution you make is not green, this means that your water is either acidic or basic. If this happens, use distilled water, which is available in supermarkets and pharmacies.*

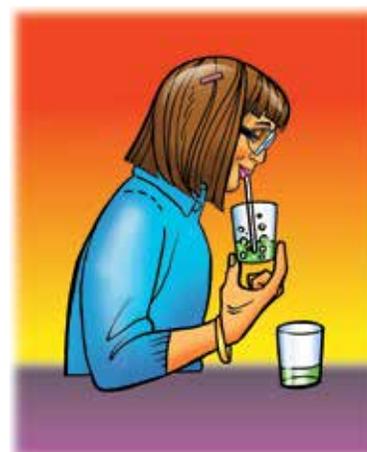
Note: In the activities, each group will need 80 mL of indicator solution. Check to make sure that you prepare enough solution. You will need about 50–60 mL of indicator solution for your demonstration. If 650 mL of solution is not enough, make more using the same proportions.

Prepare for the Demonstration

Pour about 25–30 mL of indicator solution into each of two clear plastic cups for you to use in the demonstration.

Procedure

1. Show students both samples of universal indicator solution. Place a straw in one of the samples so that the straw goes all the way to the bottom of the cup.
2. Hold the cup so that students can clearly see the liquid. Blow into the straw until the indicator solution changes from green to yellow.



Ask students:

- **Does blowing into the indicator solution change its pH?**
Yes, the color changes, so there must be a change in pH, too.
- **Does the solution become a little more acidic or a little more basic?**
The color change shows that the solution is a little more acidic.

Tell students that a chemical reaction occurs between the molecules of CO_2 and the molecules of H_2O to create a very small amount of an acid called *carbonic acid* (H_2CO_3).



Give each student an activity sheet.

Students will record their observations and answer questions about the activity on the activity sheet. The *Explain It with Atoms & Molecules* and *Take It Further* sections of the activity sheet will either be completed as a class, in groups, or individually, depending on your instructions. To find the answers to the activity sheet, go to the downloads area within the online version of this lesson.

EXPLORE

2. Have students use club soda as a source of CO_2 to see if the gas will change the pH of an indicator solution.



Question to Investigate

Will carbon dioxide from carbonated water change the pH of an indicator solution?

Materials for Each Group

- Universal indicator solution in a plastic cup
- Water
- Carbonated water (club soda or seltzer water) in a wide clear plastic cup
- 1 wide, clear, plastic cup
- 2 taller, clear, plastic cups
- Graduated cylinder
- Universal Indicator pH Color Chart

Teacher Preparation

Pour 25 mL of carbonated water into a wide, clear, plastic cup for each group.

Procedure

1. Measure 30 mL of universal indicator solution and divide it evenly into two small, clear, plastic cups.
2. Add 25 mL of water to a wide plastic cup and 25 mL of carbonated water to another wide cup.
3. Stand the small cups with indicator solution in the liquid in the wider cups as shown.
4. Turn the two tall cups upside down and place them over the two wider cups.
5. While holding the top and bottom cups to keep them together, gently swirl both sets of cups. Watch the color of the indicator in both cups to see if there is any change.
6. Compare the color of the indicator to the pH Color Chart to find out whether the solution is acidic, neutral, or basic.



Expected Results

The indicator inside the cups with water remained green, while the indicator with the carbonated water turned yellow.

3. Discuss student observations and what will happen in the following activity.

Ask students:

- **Did either indicator change color?**
Only the indicator with the carbonated water changed color.
- **What does the color change tell you about the pH of the indicator solution? Is it acidic or basic?**
The indicator solution is now acidic.
- **The carbonated water should not have splashed into the indicator. Why did the indicator solution change color in one set of cups?**
The carbon dioxide from the carbonated water dissolved in the indicator solution. The molecules of carbon dioxide reacted with the water, forming carbonic acid, and changed the color of the indicator.

Tell students that they have seen carbon dioxide gas from your breath and carbon dioxide gas from carbonated water turn an indicator solution acidic.

Ask students:

- **Do you think carbon dioxide gas produced during a chemical reaction will also turn an indicator solution acidic?**
Carbon dioxide from any source should cause the indicator solution to become acidic. The amount of carbon dioxide gas produced and dissolved in the indicator solution may cause the color of the indicator to vary, but on the acidic side.
- **What chemical reaction do you know of that can produce carbon dioxide gas?**
Students should remember that vinegar and baking soda react, producing carbon dioxide gas. Tell students that they will combine baking soda and vinegar in the next activity.

4. Use a chemical reaction to produce CO₂ to see if it changes the pH of an indicator solution.

Question to Investigate

Will carbon dioxide gas produced in the baking soda and vinegar reaction change the pH of an indicator solution?

Materials for Each Group

- Universal indicator solution in cup
- Universal indicator pH color chart
- Water
- Baking soda in wide clear plastic cup
- Vinegar in cup
- 2 small clear, plastic, cups
- 1 wide clear, plastic, cups
- 2 taller clear, plastic, cups
- Graduated cylinder

Teacher Preparation

- Pour about 25 mL of vinegar in a small plastic cup for each group.
- Place about $\frac{1}{2}$ teaspoon of baking soda into a wide, clear, plastic cup for each group.

Procedure

1. Measure 30 mL of universal indicator solution and pour it into a clean small plastic cup.
2. Pour half the indicator solution into another small cup so that you have two equal samples.
3. Pour the vinegar on top of the baking soda in one of the wide clear plastic cups. The other wide, clear, plastic cup should be empty.
4. Stand the small cups with indicator solution in both of the wider cups as shown.
5. Turn the two tall cups upside down and place them over the two wider cups.
6. While holding the top and bottom cups to keep them together, gently swirl both sets of cups. Watch the color of the indicator in both cups to see if there is any change.
7. Compare the color of the indicator to the pH Color Chart to find out whether the solution is acidic, neutral, or basic.



Expected Results

The indicator inside the cups without liquid in the outside remained green while the indicator inside the cups with the chemical reaction turned yellow.

5. Discuss student observations.

Ask students:

- **Did either indicator change color?**
Only the indicator with the chemical reaction changed color.
- **Why did one set of cups have nothing in the bottom?**
It is possible that swirling the indicator solution causes it to change color. Since this indicator did not change color, it must be the carbon dioxide gas produced in the other set of cups that is responsible for the color change. The indicator solution in the set of cups with nothing in the bottom serves as a control.
- **What does the color of the indicator solution tell you about the pH of each solution? Is it acidic, neutral, or basic?**
The color change shows that the indicator solution is slightly acidic.
- **What could you add to the acidic indicator solution to neutralize it?**
Because the indicator solution is acidic, students should suggest adding a base. Tell students that baking soda is a base.

6. Neutralize the acidic indicator solution.

Procedure

8. Take the tall cup off and remove the small cup from the wide cup.
9. Add a very small amount of baking soda to the indicator solution. Gently swirl to mix. Compare the color of this indicator solution to the color of the control.

Expected Results

The indicator solution should become a shade of green.

Ask students:

- **What did you observe?**
Students should see that a small amount of sodium bicarbonate can neutralize a carbonic acid solution.

EXPLAIN

7. Explain that carbon dioxide from any source can make water acidic.

Ask students:

- **What did carbon dioxide from breath, carbonated water, and the baking soda and vinegar reaction all do to water?**
The CO_2 from each source reacted with the water and made it acidic.

Project the illustration CO_2 Reacting with Water.

Tell students that carbon dioxide reacts with water to produce carbonic acid. Students may count up the number of atoms on each side of the equation to show that it is balanced. Point out that the double arrow in this equation means that carbonic acid breaks down readily to form carbon dioxide and water again.

Explain to students that the water in raindrops can react with certain pollutants in the air, making the rain very acidic. We call this “acid rain”, which is bad for fish, other aquatic creatures, marble statues, buildings, and natural rock formations. However, even when the rain falls through unpolluted atmosphere, rainwater is slightly acidic. This is because carbon dioxide is one of the gases in air. As rainwater falls through the air, it reacts with carbon dioxide, forming carbonic acid.

EXTEND

8. Place an Alka-Seltzer[®] tablet in indicator solution and have students interpret what the color changes say about the pH of the solution.

Explain that Alka-Seltzer[®] contains powdered acids and a base. The acids are citric acid, which tastes a little sour, and acetylsalicylic acid, which is aspirin. The base is baking soda, which is also known by its chemical name sodium bicarbonate.

Tell students that they will observe an Alka-Seltzer tablet in a universal indicator solution. Then they will use what they know about universal indicator and its color changes to describe whether the solution is acidic or basic as the substances in the tablet react.

Question to Investigate

How does the pH of the solution change during a chemical reaction between the ingredients in an Alka-Seltzer tablet in water?

Materials for Each Group

- Universal indicator solution in cup
- Water
- Alka-Seltzer tablet
- Graduated cylinder
- Snack-sized zip-closing plastic bag

Procedure

1. Add 20 mL of universal indicator solution to a snack-sized zip-closing plastic bag.
2. Seal the bag.

Note: So that students do not handle the Alka-Seltzer, which is a medicine, you should place an Alka-Seltzer tablet in each group's bag.

Procedure for the Teacher

1. Add an Alka-Seltzer tablet to each group's bag by opening the corner of the bag just enough so that the tablet can fit through.
2. Remove as much air as possible and drop the Alka-Seltzer tablet through the small opening.
3. Seal the bag and hand it to one of the students. Instruct this student to shake the bag and pass it around so that each group member has an opportunity to hold the bag.



Expected Results

As soon as the Alka-Seltzer tablet is placed in the bag, the color of the indicator solution changes to red. Bubbles appear in the solution and the bag inflates. The solution also becomes cold. Over time the solution becomes orange, yellow, and finally returns to green.

9. Discuss student observations.

As the colors are changing and the bags are inflating, ask students:

- **What do the color changes tell you about the pH of the solution at the beginning, middle, and end of the chemical reaction?**
Beginning: The solution is acidic.
Middle: The solution is becoming less acidic.
End: The solution is neutralized.

Students should conclude that the acid and base ingredients in the tablet neutralized one another.

EXTRA EXTEND

10. Have students research the effects of too much carbon dioxide in ocean water.

Too much CO_2 in the atmosphere can cause something called the greenhouse effect, which makes Earth and its atmosphere warmer. But CO_2 can do something else, which you have seen in your experiments. Too much carbon dioxide can cause water to become too acidic. This can be a big problem for fish, aquatic plants, and especially coral.

Have students research questions such as the following:

- What is coral and why would too much acid be bad for them?
- How could we reduce the amount of CO_2 that gets into the atmosphere?
- Even if a large amount of CO_2 gets in the atmosphere, what could we do to trap some of it so that it doesn't get into the oceans?

You may choose to do related activities found on the National Oceanographic and Atmospheric Administration (NOAA) website at <http://coralreef.noaa.gov/education/educators/resourcecd/activities/>. Look for the “Climate Change & Coral Activities”.

The site also offers a wealth of resources on ocean acidification and climate change.

DEMONSTRATION

1. Your teacher blew through a straw into a universal indicator solution until it changed color. Did the indicator solution become acidic or basic?
2. What chemical from your teacher's breath caused the indicator to change color?



ACTIVITY

Question to Investigate

Will carbon dioxide from carbonated water change the pH of an indicator solution?



Materials for Each Group

- Universal indicator solution in a plastic cup
- Water
- Carbonated water in a wide clear plastic cup
- 1 wide, clear, plastic cup
- 2 taller, clear, plastic cups
- Graduated cylinder
- Universal Indicator pH Color Chart

Procedure

1. Measure 30 mL of universal indicator solution and divide it evenly into two small clear plastic cups.
2. Add 25 mL of water to a wide plastic cup and 25 mL of carbonated water to another wide cup.
3. Stand the small cups with indicator solution in the liquid in the wider cups as shown.
4. Turn the two tall cups upside down and place them over the two wider cups.
5. While holding the top and bottom cups to keep them together, gently swirl both sets of cups. Watch the color of the indicator in both cups to see if there is any change.
6. Compare the color of the indicator to the pH Color Chart to find out whether the solution is acidic, neutral, or basic.



Describe the color of the indicator solution in each set of cups	
Carbonated Water	Water

3. What does the color of the indicator solution tell you about the pH of each solution? Is it acidic, neutral, or basic?

4. The carbonated water and water should not have splashed into the indicator solutions. Why did the indicator solution change color in one set of cups?

ACTIVITY

Question to Investigate

Will carbon dioxide gas produced in the baking soda and vinegar reaction change the pH of an indicator solution?

Materials for Each Group

- Universal indicator solution in cup
- pH color chart
- Water
- Baking soda in wide clear plastic cup
- Vinegar in cup
- 2 small, clear, plastic cups
- 1 wide, clear, plastic cups
- 2 taller, clear, plastic cups
- Graduated cylinder

Procedure

1. Measure 30 mL of universal indicator solution and pour it into a clean, small, plastic cup.
2. Pour half the indicator solution into another small cup so that you have two equal samples.
3. Pour the vinegar on top of the baking soda in one of the wide, clear, plastic cups. The other wide, clear, plastic cup should be empty.
4. Stand the small cups with indicator solution in both of the wider cups as shown.



- Turn the two tall cups upside down and place them over the two wider cups.
- While holding the top and bottom cups to keep them together, gently swirl both sets of cups. Watch the color of the indicator in both cups to see if there is any change.
- Compare the color of the indicator to the pH Color Chart to find out whether the solution is acidic, neutral, or basic.



Describe the color of the indicator solution in each set of cups	
Reacting baking soda and vinegar	Nothing extra, just the cups

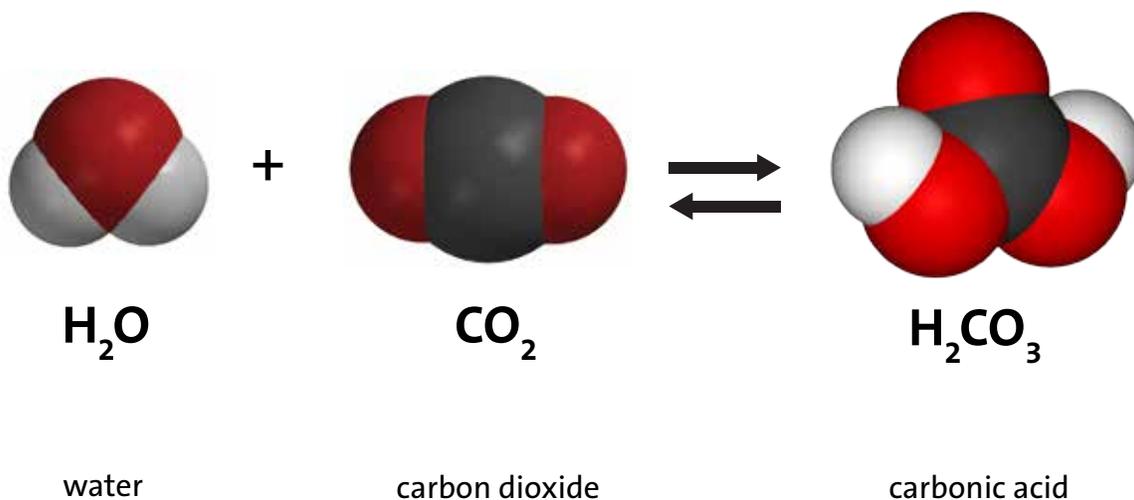
Neutralize the acidic indicator solution

- Take the tall cup off and remove the small cup of indicator solution from the wide cup.
 - Add a very small amount of baking soda to the indicator solution. Gently swirl to mix. Compare the color of this indicator solution to the color of the control.
5. What does the color of the indicator solution tell you about the pH of each solution? Is it acidic, neutral, or basic?
6. Why did one set of cups have nothing in the bottom?

7. The baking soda and vinegar should not have splashed into their indicator solutions. Why did the indicator solution change color in one set of cups?

EXPLAIN IT WITH ATOMS & MOLECULES

8. When raindrops react with certain pollutants in the air, rain can become very acidic. We call this “acid rain”. But even when rain falls through clean, non-polluted air, the rain becomes slightly acidic. Use the equation below to explain why this happens.



TAKE IT FURTHER

Question to Investigate

How does the pH of the solution change during a chemical reaction between the ingredients in an Alka-Seltzer tablet in water?



Materials for Each Group

- Universal indicator solution in cup
- Water
- Alka-Seltzer tablet
- Graduated cylinder
- Snack-sized zip-closing plastic bag

Procedure

1. Add 20 mL of universal indicator solution to a snack-sized zip-closing plastic bag.
2. Seal the bag.

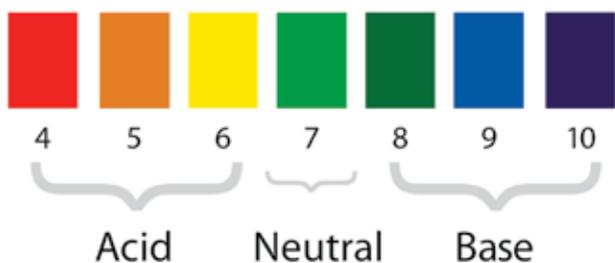


9. What do you think caused the bag to inflate?

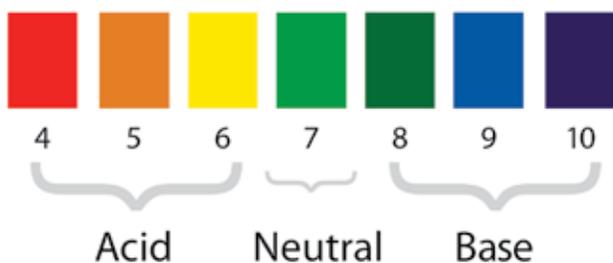
What do the color changes tell you about the pH of the indicator solution during the chemical reaction?	
Beginning	
Middle	
End	

10. Why do you think the indicator solution turned green at the end of the reaction?

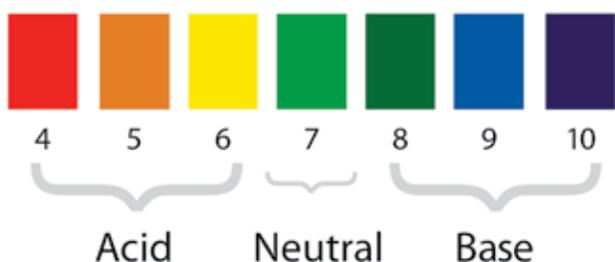
Universal Indicator pH Color Chart



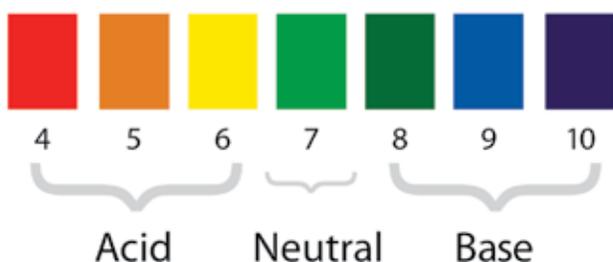
Universal Indicator pH Color Chart



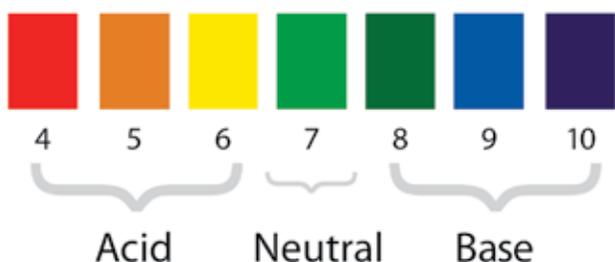
Universal Indicator pH Color Chart



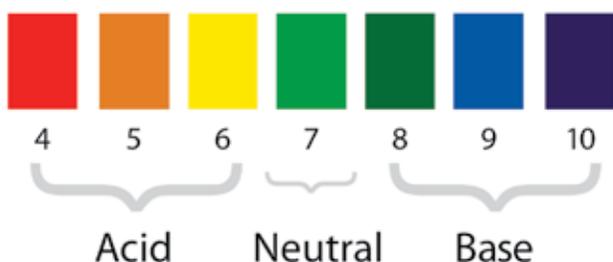
Universal Indicator pH Color Chart



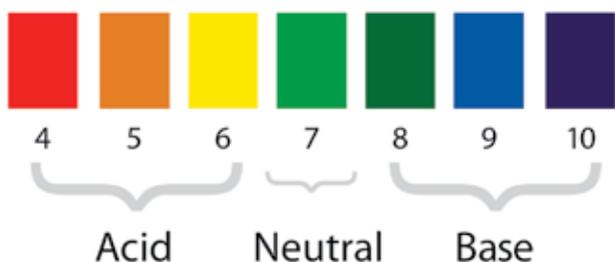
Universal Indicator pH Color Chart



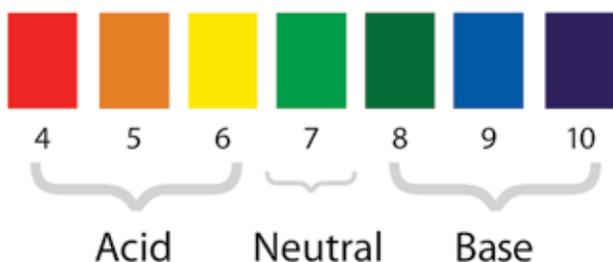
Universal Indicator pH Color Chart



Universal Indicator pH Color Chart



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Chapter 6—Student Reading

What is a chemical reaction?

There are many common examples of chemical reactions. For instance, chemical reactions happen when baking cookies and in your digestive system when you eat the cookies. Rusting iron and burning gasoline in a car engine are chemical reactions. Adding baking soda to vinegar also causes a chemical reaction. In a chemical reaction, the molecules in the reactants interact to form new substances. A chemical reaction causes a *chemical* change. Other processes, like dissolving or a change of state, cause a *physical* change in which no new substance is formed.



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Another chemical reaction that you have seen many times is a burning candle.

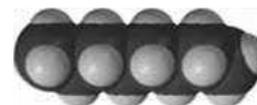


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When a candle burns, molecules in the wax react with oxygen in the air. This reaction, called combustion, releases energy in the form of the heat and light of the flame. The reaction also produces something else which is not as obvious – carbon dioxide and water vapor.

A closer look at a burning candle

The wax in the candle is made of long molecules called *paraffin*. These paraffin molecules are made up of only carbon atoms and hydrogen atoms bonded together.

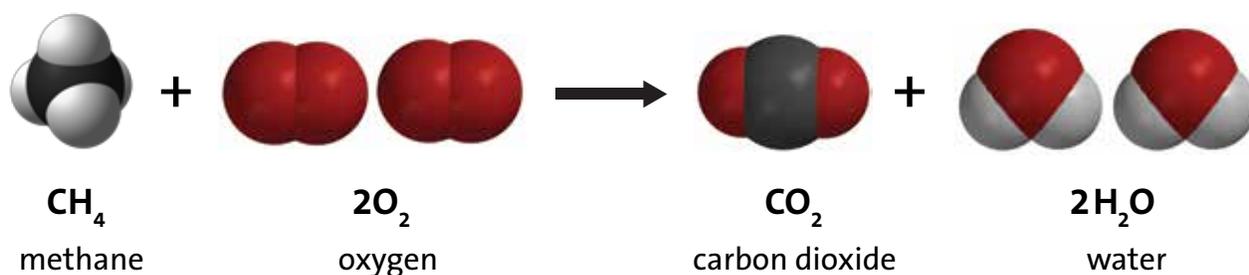


Molecules made of only carbon and hydrogen are called *hydrocarbons*. The simplest hydrocarbon (methane) can be used as a model to show how the wax or any other hydrocarbon burns.



The chemical formula for methane is CH_4 . This means that methane is made up of one carbon atom and 4 hydrogen atoms.

This is the chemical equation for the reaction of methane and oxygen.



The reactants

The methane and oxygen on the left side of the equation are called the *reactants*. Each molecule of oxygen gas is made up of two oxygen atoms bonded together.



It can be confusing that oxygen the atom, and oxygen the molecule, are both called “oxygen”. When we talk about the oxygen in the air, it is always the molecule of oxygen which is two oxygen atoms bonded together, or O_2 . The reason why there is a “2” in front of the O_2 shows that there are two molecules of O_2 .

The products

The carbon dioxide and water on the right side of the equation are called the products. The chemical formula for carbon dioxide is CO_2 . This means that carbon dioxide is made up of one carbon atom and 2 oxygen atoms.



The other product is two molecules of water. Each molecule of water is made up of two hydrogen atoms bonded to one oxygen atom or H₂O.



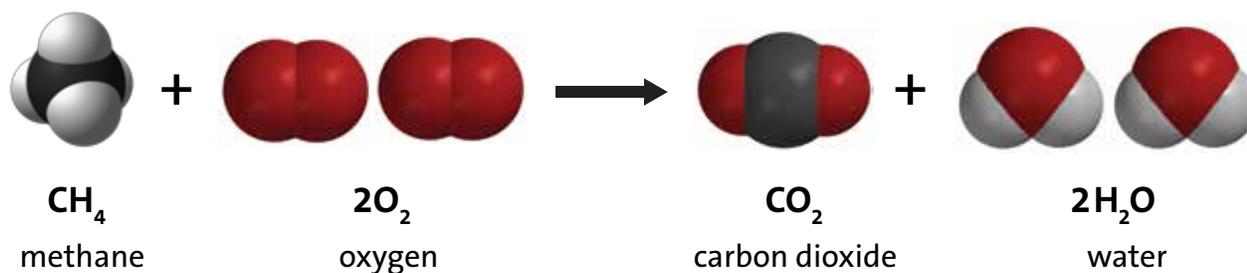
The reason why there is a “2” in front of the H₂O shows that there are two molecules of H₂O.

Where do the products come from?

The atoms in the products come from the atoms in the reactants. In a chemical reaction, the reactants interact with each other, bonds between atoms in the reactants are broken, and the atoms rearrange and form new bonds to make the products.

Counting the atoms in the reactants and products

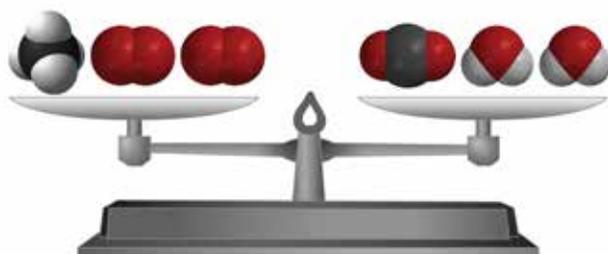
To understand a chemical reaction, you need to check that that the equation for the reaction is *balanced*. This means that the same type and number of atoms are in the reactants as are in the products. To do this, you need to be able to count the atoms on both sides.



Look again at the equation for methane reacting with oxygen. You see a big number (coefficient) in front of some of the molecules and a little number (subscript) after an atom in some of the molecules. The coefficient tells how many of a particular type of *molecule* there are. The subscript tells how many of a certain type of *atom* are in a molecule. So if there is a coefficient in front of the molecule and a subscript after an atom, you need to multiply the coefficient times the subscript to get the number of atoms.

Example: In the products of the reaction there are 2H₂O. The coefficient means that there are two molecules of water. The subscript means that each water molecule has 2 hydrogen atoms. Since each water molecule has 2 hydrogen atoms and there are two water molecules, there must be 4 (2×2) hydrogen atoms.

If you look closely at the equation, you can see that there is 1 carbon atom in the reactants and 1 carbon atom in the products. There are 4 hydrogen atoms in the reactants and 4 hydrogen atoms in the products. There are 4 oxygen atoms in the reactants and 4 oxygen atoms in the products. This equation is balanced.

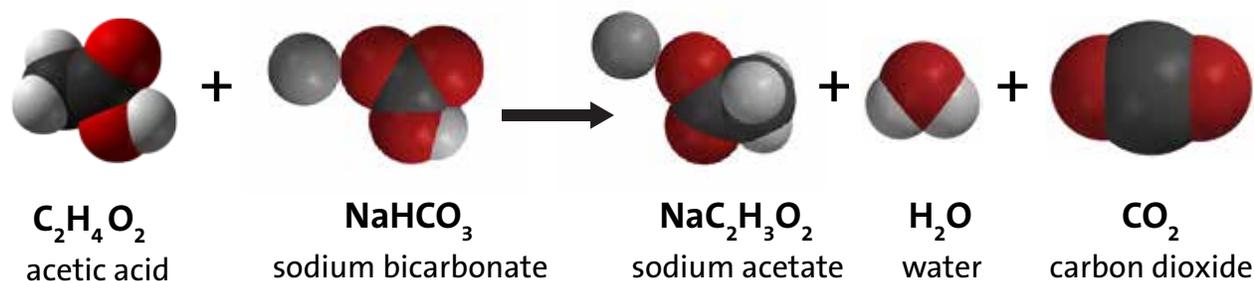


Another way of saying that an equation is balanced is that “mass is conserved”. This means that the atoms in the reactants end up in the products and that no new atoms are created and no atoms are destroyed.

Changing the amount of products

If you want to change the amount of products formed in a chemical reaction, you need to change the amount of reactants. This makes sense because atoms from the reactants need to interact to form the products.

An example is the popular reaction between vinegar (acetic acid) and baking soda (sodium bicarbonate).



When you do this reaction, one of the most noticeable products, which you see on the right side of the equation, is carbon dioxide gas (CO_2). If you wanted to produce more CO_2 , you could use more baking soda because there would be more baking soda to react with the vinegar to produce more carbon dioxide. In general, using more of one or more reactants will result in more of one or more products. Using less of one or more reactants will result in less of one or more products. But this principle has limits. If you wanted to make a lot of carbon dioxide, you couldn't just keep adding more and more baking soda to the same amount of vinegar. This might work for a while, as long as there was enough vinegar, but eventually there would be no atoms left of vinegar to react with the extra baking soda so no more carbon dioxide would be produced.

EVIDENCE OF A CHEMICAL REACTION

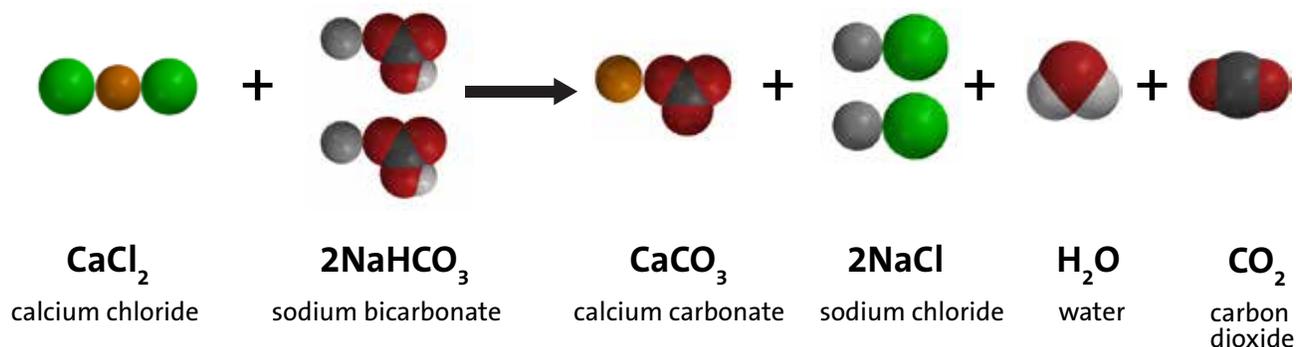
Production of a gas

The gas produced from mixing vinegar with baking soda is evidence that a chemical reaction has taken place. Since the gas was produced from mixing a solid (baking soda) and a liquid (vinegar), the gas is a new substance formed by the reaction.

Formation of a precipitate

Another clue that a chemical reaction has taken place is a solid is formed when two solutions are mixed. When this happens, the solid is called a precipitate. The precipitate does not dissolve in the solutions. One example of solutions that form a precipitate are calcium chloride solution and sodium bicarbonate solution.

When these solutions are combined, a precipitate called calcium carbonate is produced. Calcium carbonate is the main ingredient in chalk and sea shells, and does not easily dissolve.



Color change

When two substances are mixed and a color change results, this color change can also be evidence that a chemical reaction has taken place. The atoms that make up a molecule and the structure of the molecules determines how light interacts with them to give them their color. A color change can mean that new molecules have been formed in a chemical reaction with different structures that produce different colors.

Temperature change

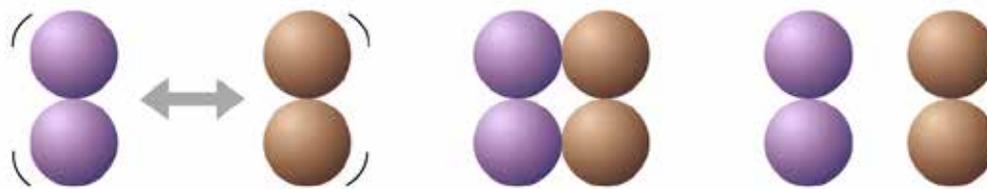
Another clue that a chemical reaction has occurred is a change in temperature of the reaction mixture. You can read more about the change in temperature in a chemical reaction under Chemical reactions and energy below.

RATE OF A CHEMICAL REACTION

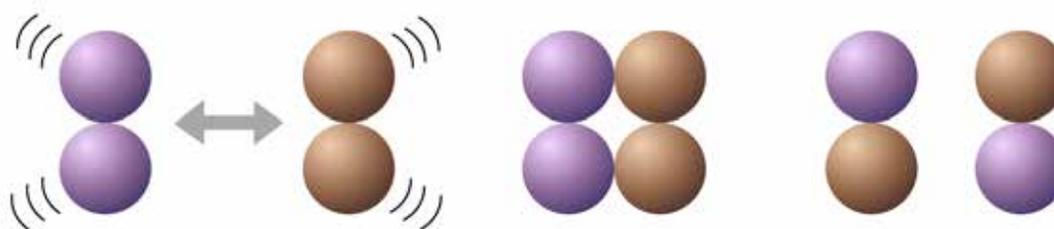
Increasing the temperature increases the rate of the reaction

The rate of a chemical reaction is a measure of how fast the reactants are changed into products. This can be increased by increasing the temperature of the reactants.

For reactant molecules to react, they need to contact other reactant molecules with enough energy for atoms or groups of atoms to come apart and recombine to make the products. If they do not have enough energy, most reactant molecules just bounce off and do not react.



But if the reactants are heated, the average kinetic energy of the molecules increases. This means that more molecules are moving faster and hitting each other with more energy. If more molecules hit each other with enough energy to react, then the rate of the reaction increases.



A catalyst can increase the rate of the reaction

Another way to increase the rate of the reaction is by adding a substance that helps bring the reactants together so they can react. A substance which helps speed up a chemical reaction in this way but does not become a product of the reaction is called a *catalyst*.

A common catalyst in the cells of living organisms is called *catalase*. During normal cell processes, living things produce hydrogen peroxide in their cells. But hydrogen peroxide is a poison so the cells need a way to break it down very quickly. Catalase helps break down hydrogen peroxide at a very fast rate. Catalase and many other catalysts in living things, are large complex molecules. Reactants attach to specific parts of the catalyst which helps the reactants to come apart or bond together. A single molecule of catalase can catalyze the breakdown of millions of hydrogen peroxide molecules every second.

Substances react chemically in characteristic ways

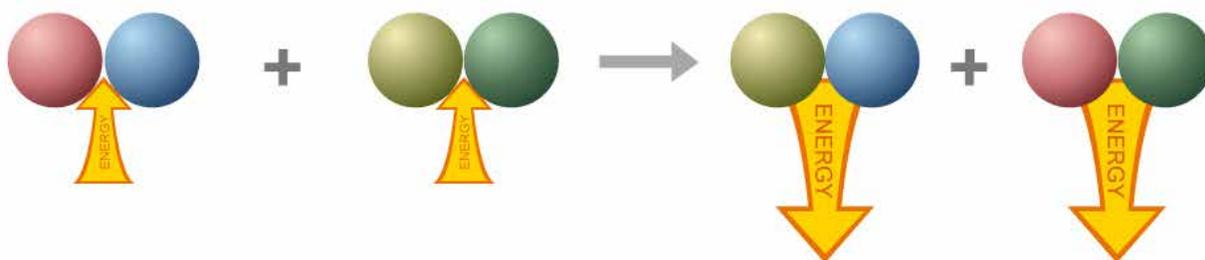
If you tested different substances with a particular liquid to see how the substances react, each would react in its own characteristic way. And each substance that reacted would react the same way each time it was tested with the same liquid. Substances react in characteristic ways because every substance is different. Each one is made up of certain atoms bonded in a particular way that makes it different from any other substance. When it reacts with another substance, certain atoms or groups of atoms unbond, rearrange, and rebond in their own way.

Chemical reactions and energy

Chemical reactions involve breaking bonds in the reactants and making new bonds in the products. It takes energy to break bonds in the reactants. Energy is released when new bonds are formed in the products. The using and releasing of energy in a chemical reaction can help explain why the temperature of some reactions goes up (exothermic) and the temperature of other reactions goes down (endothermic).

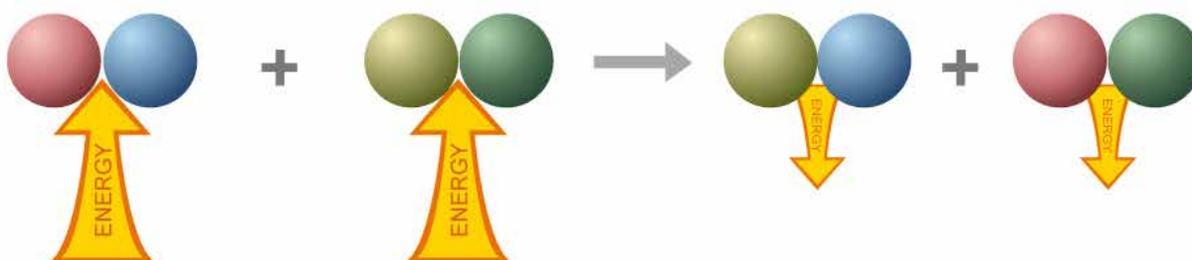
Exothermic

If a reaction is exothermic, that means that it takes less energy to break the bonds of the reactants than is released when the bonds in the products are formed. Overall, the temperature increases.



Endothermic

If a reaction is endothermic, it takes more energy to break the bonds in the reactants than is released when the products are formed. Overall, the temperature decreases.

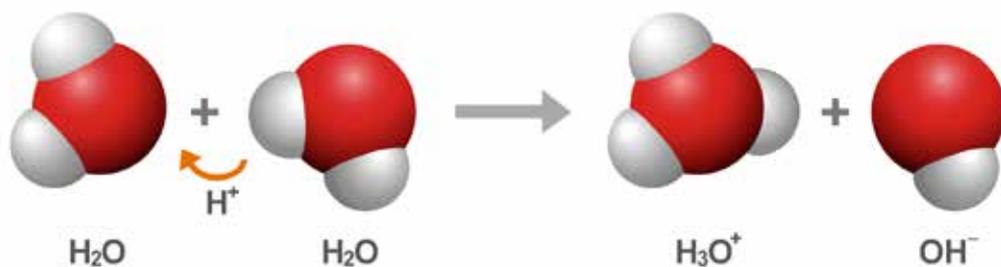


Acids, bases, and pH

You may have heard of the term “pH” when talking about the water in a pool or fish tank. You may have seen people take a sample of water and compare it to colors on a chart to test the pH of the water. The pH scale is a way to measure whether the water is acidic or basic.



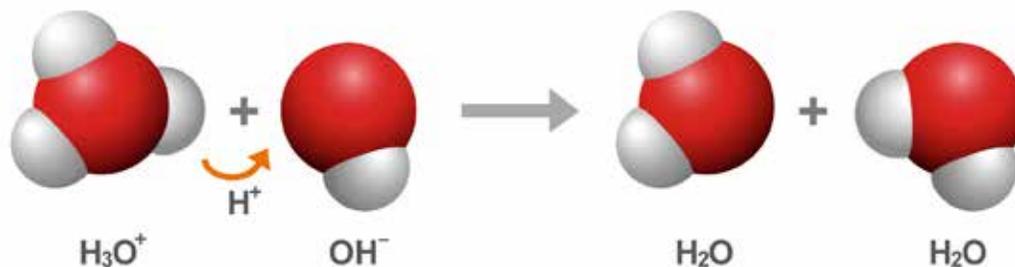
Normally we think of water as good ol' H_2O , but in fact some water molecules react with each other and become something different. When two water molecules bump into each other and react, a proton from a hydrogen atom in one of the water molecules gets transferred to the other water molecule. This proton leaves its electron behind in the water molecule it came from.



When a proton is transferred from one water molecule to another, it's as if the molecule gaining the proton is actually gaining another hydrogen atom (but without the electron). So in the reaction between the two water molecules, the one that gained the extra proton has one more proton than electron and changes from H_2O to become the ion H_3O^+ .

It works the other way around for the water molecule that lost the proton. It's as if the water molecule lost a hydrogen atom (but held on to the electron). So the water molecule that lost the proton has one more electron than proton and changes from H_2O to become the ion OH^- .

But the H_3O^+ ions and the OH^- ions also react with one another. In this reaction, the extra proton on the H_3O^+ can be transferred back to the OH^- to form two water molecules again.

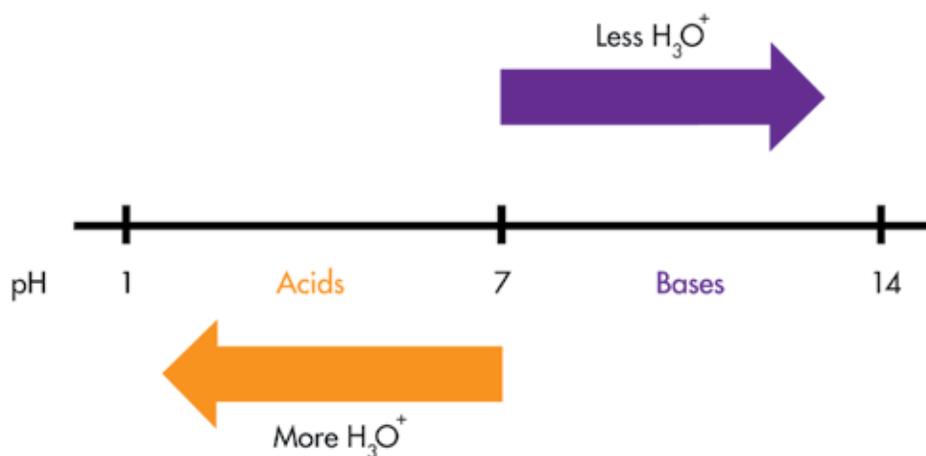


In pure water, these reactions balance one another and result in a small but equal concentration of H_3O^+ and OH^- .

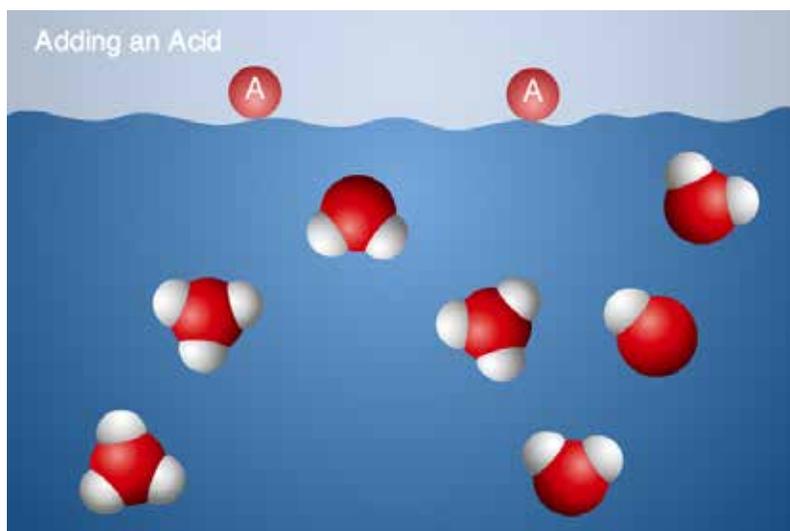
The concentration of H_3O^+ in water determines how acidic or basic a solution is. The pH scale is a measure of the concentration of H_3O^+ in water. Pure water is neutral and measures 7 on the pH scale.

How do acids and bases make water acidic or basic?

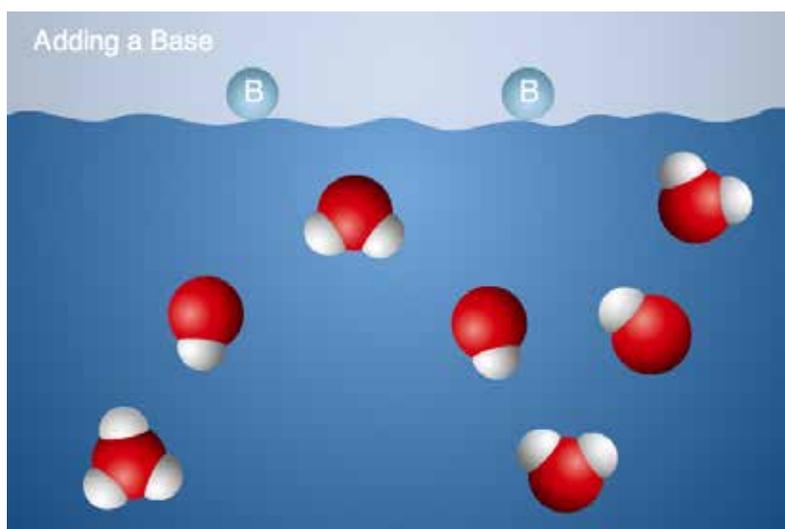
If a solution has a higher concentration of H_3O^+ than OH^- , it is considered an acid. An acid measures less than 7 on the pH scale. If the solution has a lower concentration of H_3O^+ than OH^- , it is considered a base. A base measures greater than 7 on the pH scale.



Acids are sometimes called “proton donors”. This means that when an acid is added to water, the acid molecule transfers a proton to water molecules forming more H_3O^+ . Since the solution has a higher concentration of H_3O^+ than OH^- , it is an acid.



Talking about bases is a little trickier because you have to look at two steps to see how they affect the pH. Bases are sometimes called “proton acceptors”. This means that when a base is added to water, the base molecule accepts a proton from water forming more OH^- . When there is extra OH^- in the water, the H_3O^+ ions transfer protons to the OH^- ions causing the concentration of H_3O^+ to go down. Since the solution has a lower concentration of H_3O^+ than OH^- , it is a base.



Acids and bases are like chemical opposites

An acid can neutralize a base and base can neutralize an acid. This makes sense because if an acid is a proton donor and a base is a proton acceptor, they have the opposite effect on water and can cancel each other.

The acid donates protons and increases the concentration of H_3O^+ . The base accepts protons from water molecules making more OH^- . The H_3O^+ transfers a proton to the OH^- and causes the concentration of H_3O^+ to decrease and become closer to neutral again.

STRENGTH AND CONCENTRATION IN ACIDS AND BASES

The effect that an acid or base has in a chemical reaction is determined by its strength and concentration. It is easy to confuse these two terms.

Strength

There are different kinds of acids. There are strong acids, weak acids, and acids in-between. Some acids are so strong that they can make a hole in a piece of metal. Other acids, like citric acid or ascorbic acid (Vitamin C), are weaker and are even safe to eat.

The factor that determines the strength of an acid is its ability to donate a proton, increasing the amount of H_3O^+ in water. A strong acid produces a lot of H_3O^+ in water, while the same amount of a weak acid produces a smaller amount of H_3O^+ .

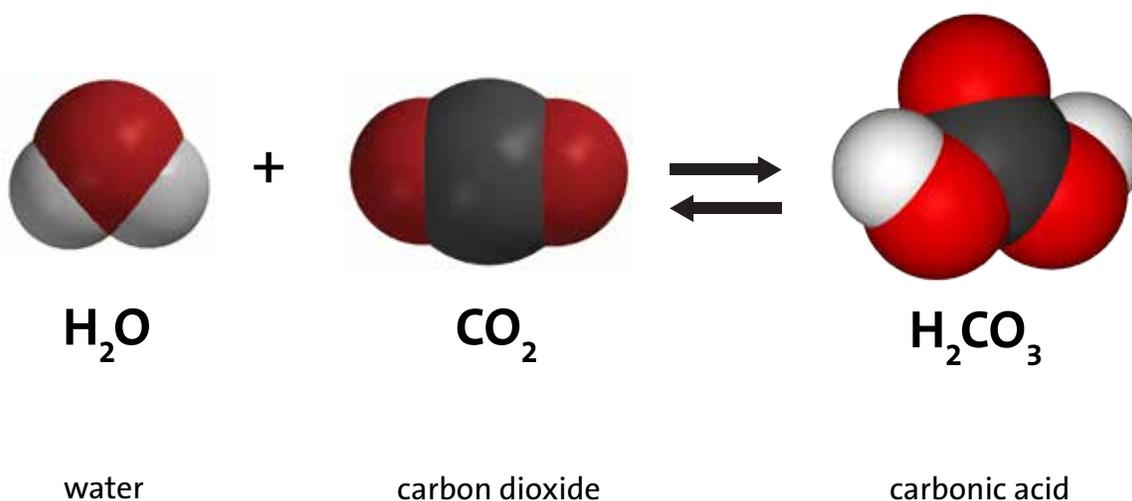
Concentration

Concentration is different from strength. Concentration has to do with the amount of acid added to a certain amount of water.

It is the combination of the concentration and the strength of an acid that determines the amount of H_3O^+ in the solution. And the amount of H_3O^+ is a measure of the acidity of the solution.

Acids and the environment

There's been a lot of news lately about too much carbon dioxide (CO_2) gas going into the atmosphere and contributing to global warming. This is a big problem but CO_2 also does something else which is not in the news as much. Carbon dioxide gas goes into the ocean and reacts with water to form a weak acid called carbonic acid.



This extra carbonic acid affects the pH of the ocean. The ocean is actually slightly basic. The extra acid makes the ocean less basic or more acidic than it would normally be. The change in ocean pH has an effect on organisms in the ocean, particularly ones that build shells like corals.



These organisms need calcium ions and carbonate ions to make the material for their shell which is calcium carbonate. The extra H_3O^+ from the acid interacts with the carbonate ion and changes it so that it can't be used for making shells. Reducing the amount of CO_2 that gets into the ocean is the first step to helping to solve this problem.