Classic Science

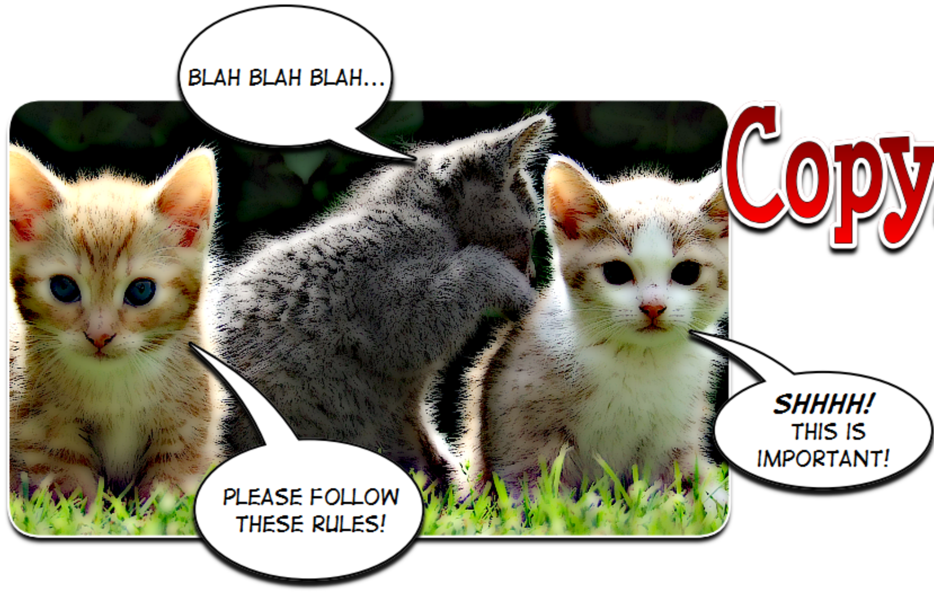
For the Family

STUDENT COPY

ADVANCED CHEMISTRY

The lab of MR.Q

Scott McQuerry



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First of all thank you very much for choosing to use this book with your family. You will not be disappointed! I have been asked by several families the same question, "**Who** are you and **why** are you doing this?" Without going into great detail, E=McQ is owned, operated and stressed over by me. Yep… little o' me. I am an educator by profession and began working with homeschool families several years ago while offering free programs to area families to explore various concepts in science. I guess I can't stop doing what I love!

This product is the fruit of my 13-year labor in science education. Having worked with homeschool families over these years I have gained an appreciation for your needs, struggles and wants. I could not make this curriculum any simpler for your child to master the concepts of science. It is completely reusable, relatively cheap (I tried to keep it under the cost of a tank of gas), adaptable to various needs at home and as fun as humanly possible.

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Conversion Chart: U.S. to Metric

7 drops = 1 mL

1 teaspoon = 5 mL

1 tablespoon = 15 mL

1/4 cup = 60 mL

1 cup = 240 mL

2 cups (1 pint) = 480 mL

4 cups (1 quart) = 0.95 liter

4 quarts (1 gal.) = 3.8 liters

1 fluid oz. = 30 mL

1 oz. = 28 grams

1 pound = 454 grams

1 inch = 2.54 cm

Chapter I

You don't have to look too deeply to see how chemistry affects your life.

Every breath you take, every meal you eat, everything has something to do with the interaction of chemicals! And I know for a fact that you know how chemicals affect your life every time you finish a large meal (or go too long without a meal!)

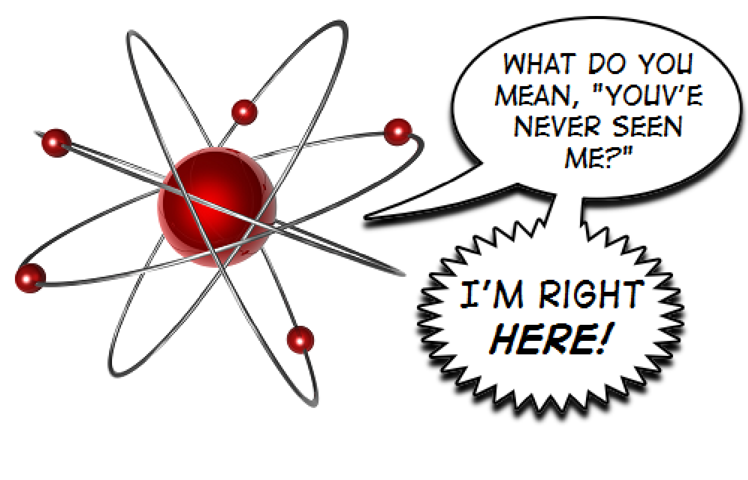
But you need to know something that many people do not understand – even though we will be spending a lot of time studying the tiniest pieces of matter within this book...

Nobody on the planet has ever actually seen an atom before!

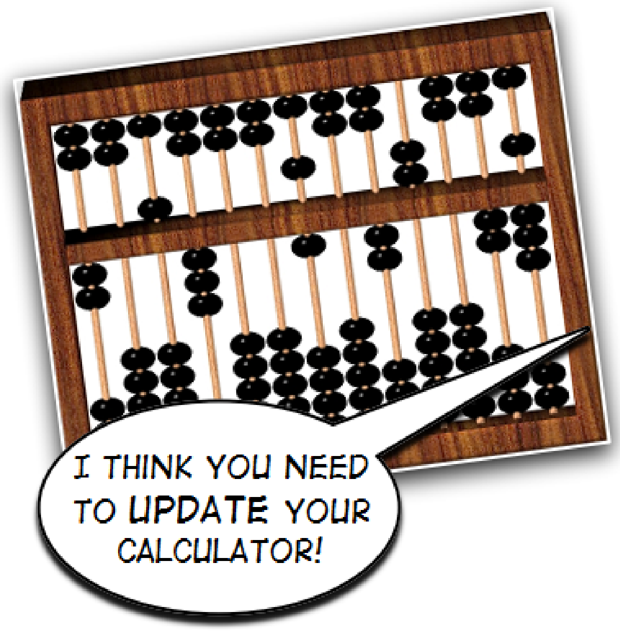
This does not mean that scientists have been making all of this stuff up over the years. Thousands of experiments have been run in order to create the best models for our understanding of chemistry. Remember - the study of chemistry is the study of very small particles! Much like the scientists who have created these chemistry models, you too will be successful in measuring very tiny pieces of matter to incredibly precise details! The nature of how we measure or quantify these particles is the topic of study this first week.

Since the study of chemistry has to do with measuring many different things about incredibly tiny pieces of matter, a new way of writing down massively large or small numbers had to be developed. Let's get real here, who really wants to write out that there are 3,011,500,000,000,000 molecules of sugar* in every cup?

* I'm really not joking here. This is pretty close to the actual number!



So a short cut was needed to make all of these large numbers a little easier to manage. This short cut is known as **scientific notation**. The rules of this method are very easy to follow as long as you can multiply or divide any number by 10.



Instead of writing out the number 100, you would write 10^2. The "2" that is hovering on top of the 10 represents the multiplication of the number 10 by itself. The number 1000 would be written as 10^3; 10,000 would be 10^4, and so on. The following chart can be used to follow how these exponents work:

100 = 10 x 10 = $\mathbf{10^2}$ = **one hundred**
1,000 = 10 x 10 x 10 = $\mathbf{10^3}$ = **one thousand**
10,000 = 10 x 10 x 10 x 10 = $\mathbf{10^4}$ = **ten thousand**
1,000,000 = 10 x 10 x 10 x 10 x 10 x 10 = $\mathbf{10^6}$ = **one million**
1,000,000,000 = 10 x 10 x 10 x 10 x 10 x 10 x 10 x 10 x 10 = $\mathbf{10^9}$ = **one billion**

This chart only tells us how to make our number larger. We can use this same procedure <u>in reverse</u> to determine how small an object can be as well. For example:

1/10 = $\mathbf{10^{-1}}$ = **one tenth**
1/100 = 1/10 x 10 = $\mathbf{10^{-2}}$ = **one hundredth**
1/1,000 = 1/10 x 10 x 10 = $\mathbf{10^{-3}}$ = **one thousandth**
1/10,000 = 1/10 x 10 x 10 x 10 = $\mathbf{10^{-4}}$ = **one ten thousandth**
1/1,000,000 = 1/10 x 10 x 10 = $\mathbf{10^{-6}}$ = **one millionth**
1/1,000,000,000 = 1/10 x 10 x 10 = $\mathbf{10^{-9}}$ = **one billionth**

So how do all of these short cuts help us? Well, instead of writing out the number 2,500 when you are doing your calculations, you can write it as
$2.5 \times 10 \times 10 \times 10 = 2{,}500$

Or, you can make it REALLY easy on yourself and use scientific notation to write it as 2.5×10^3.

In scientific notation, the first number you wrote down (the 2.5) is known as the **coefficient** and the 10^3 is called the **exponent**. There is one little rule with coefficients:

The coefficient must always be a number between 1.0 and 9.9

This may seem to be more work that it is really worth. And you may be right when we are using small numbers like 2,500. However, it is very possible that you will have to work with much larger numbers like the one you read about a little while ago - 3,011,500,000,000,000. This is where scientific notation comes in handy. Let's use this short cut once again:

Question #1:

$$3{,}011{,}500{,}000{,}000{,}000 = 3.0115 \times 10^{15}$$

How can you figure out how many times you have to divide this large number by 10 to get 3.0115×10^{15}?

After a few practice problems, you will probably become very good at solving this problem. Until then, you can use a little device to help you count the number of 10's you'll need for your answer.

3.011500000000000.

Simply count the number of loops you draw from the beginning decimal to where it will create a coefficient between 1.0 and 9.9.

The same is true if you have a very small number and you need to put it into scientific notation. For example, take a look at the following example:

$$0.0112556 = 1.12556 \times 10^{-2}$$

In order to convert 0.0112556 into scientific notation, you need to multiply this number by 100 to get a coefficient between 1.0 and 9.9. This can be represented with the following pic:

$$0.01.12556$$

Another way to determine how to convert a large or small number into scientific notation is to remember this simple rule:

If the coefficient needs to get smaller during a conversion, the exponent will need to get bigger, and vice versa!

For example, let's assume we need to convert the following number into scientific notation:

$$3{,}011{,}500{,}000{,}000{,}000$$

This number will need to get smaller, right? So if this number gets smaller, its exponent will need to get larger. Therefore...

$$3{,}011{,}500{,}000{,}000{,}000 = 3.0115 \times 10^{15}$$

If you have a small number like this one... 0.016115

...and you need to convert it into scientific notation, it will have to become LARGER in order to be a proper coefficient. Therefore...

$$0.016115 = 1.6115 \times 10^{-2}$$

Question #2:
So how do you know where you should place the final decimal?
That is easy! Your coefficient should always be between 1 and 9.999...

Question #3:
Why are we doing this again?
Okay, I will be honest with you. Crunching numbers and practicing scientific notation problems may not be the most exciting thing to do. However, this skill is very important not only to chemists, but to mathematicians, and engineers too!

Chemists need the best data possible when running experiments on such tiny pieces of matter. The ability to accurately AND precisely identify very large or small numbers is extremely important. You may be thinking that accuracy and precision is the same thing. But they're not!

Precision is how close a series of measurements are to each other.

Accuracy is how close a measured value is to the real value of the object.

You can see how these two terms work in the kitchen:

Imagine baking three batches of cookies. Everything is going fine until you get distracted with phone calls, TV, etc. and end up burning all three sheets of cookies.

This is an example of being precise but not accurate (they all look like charcoal, but that's not what you were hoping for!)

So you cook three more sheets of cookies and make certain to set a timer this time. The bell goes off and you pull out the cookies only to find that the sheet on the bottom of the oven burned all of the cookies but the sheets on top of the oven were still gooey. Your oven does not heat evenly!

This is an example of being NEITHER precise nor accurate.

You decide to give it one more try. The timer is set and you have the brilliant idea to rotate your cookie sheets within the oven as they are cooking. When the time is up, you pull out your cookies and find all of them to be cooked perfectly!

This is an example of being BOTH accurate and precise. All of the cookies have been cooked precisely the same as each other AND the accuracy of how long they were to be cooked was perfect!

This week, you need to practice how to place large and small numbers into scientific notation. I know there are a lot of problems to do. But practice makes perfect. Actually, **perfect practice makes perfect.** Good luck and get ready for next week because we are going to see how to take these huge numbers and apply a little mathematics to them. Stay tuned....

Scientific notation practice (Part I)

Convert the following to scientific notation:

1) 45,700 ______________________
2) 0.009 ______________________
3) 23 ______________________
4) 0.9 ______________________
5) 24,212,000 ______________________
6) 0.000665 ______________________
7) 21.9 ______________________
8) 0.00332 ______________________
9) 321 ______________________
10) 0.119 ______________________
11) 1492 ______________________
12) 0.2713 ______________________
13) 314159 ______________________
14) 6022 ______________________
15) 0.12011 ______________________

Convert the following numbers in scientific notation to expanded form:

16) 3.825×10^{3} ______________________________

17) 6.3×10^{4} ______________________________

18) 2.3×10^{-2} ______________________________

19) 4.44×10^{-6} ______________________________

20) 7.121×10^{9} ______________________________

21) 1.2×10^{-1} ______________________________

22) 1.8×10^{2} ______________________________

23) 8.1×10^{-4} ______________________________

24) 6.7×10^{5} ______________________________

25) 3.4×10^{7} ______________________________

If you did okay with these first problems, let's see how you do with a few more...

Scientific notation practice (Part II)

Put these numbers into scientific notation.

26) 0.000034 ____________________
27) 65000 ____________________
28) 36000×10^{10} ____________________
29) 549 ____________________
30) 0.0000403×10^{12} ____________________
31) 0.00000000082 ____________________
32) 0.000000000205 ____________________
33) 21.8×10^{-4} ____________________
34) 0.00973×10^{8} ____________________
35) 0.0000070 ____________________
36) 3,621.471 ____________________
37) 3,752.6 ____________________
38) 456.83 ____________________
39) 215 ____________________
40) 0.0428 ____________________
41) 0.00005673 ____________________
42) 0.00000000900 ____________________
43) 0.000039256 ____________________
44) 0.000000010 ____________________

45) 0.0037004 ______________________________

46) 0.002 ______________________________

47) 0.0080×10^{-3} ______________________________

48) 36000×10^{-10} ______________________________

49) 0.156 ______________________________

50) 0.045×10^{-3} ______________________________

Chapter 2

Last week, you were introduced to a method of shortening huge numbers down into smaller pieces. This is really going to be helpful in the coming weeks!

In order to get started this week, let's imagine running an experiment and collecting all kinds of data. Let's say our experiment was on the time it takes a cup of saltwater to boil on the stove. We can collect data on the temperature of the water when it boils, the amount of salt added, the amount of heat added to the solution, etc. Let's assume that we ran our experiment three times and documented that it took 12.2 minutes, 11.9 minutes, and 12.4 minutes for the saltwater to boil each time. An average of these three numbers would give us:

$$12.2 \text{ minutes} + 11.9 \text{ minutes} + 12.4 \text{ minutes} = 36.5 \text{ minutes}$$

$$36.5 \text{ minutes} / 3 = 12.1666666 \text{ minutes}$$

The questions that we need to be asking now are:

How precise is our data?

and

How accurate is our data?

You learned the difference between accuracy and precision last week:

Precision is how close a series of measurements are to each other.

Accuracy is how close a measured value is close to the real value of the object.

We are going to tackle the "accuracy" question later in this chapter. Let's focus on how precise our data is first!

In order to see how precise our data is, we have to look at the number of decimal places that exist within our data.

In our example, the measurements of 12.2, 11.9, and 12.4 minutes are all assumed to be precise to the nearest 0.1 minute. Why can we make that assumption? Because why else would you go to the trouble of writing out each measurement to the tenths category?

The .2, .9, and .4 in our three measurements give us a very precise amount of information about the time it took our saltwater to boil. Any digits that give us some amount of precision about a set of data are known as **significant figures (numbers)**.

This does not mean that ALL numbers which exist within our data are significant. Some of them are not significant at all!

So how do you identify the significant from the non-significant numbers?

Well, there are set of easy rules you can follow to determine which numbers are significant?

Rule #1: All non-zero numbers in measurements are always significant.

For example...

12.2 minutes = three significant figures

11.9 minutes = three significant figures

12.4 minutes = three significant figures

Each of these numbers is precise to the nearest tenth (0.1)

Rule #2: Zeros between non-zero numbers are always significant (these are the zeros that are "sandwiched" between two significant figures!)

For example...

1801 grams = four significant figures

2001 grams = four significant figures

Both of these numbers are assumed to be precise to the nearest 1 gram.

Rule #3: Zeros to the left of all non-zero numbers are never significant.

For example...

0.01 grams = one significant figure
(both of the zeros are to the left of the significant figure)

0.12 grams = two significant figures
(the zero is not to the right of a non-zero number)

Both numbers are assumed to be precise to the nearest hundredth (0.01)

Rule #4: Zeros to the right of all non-zero numbers are only significant if there's a decimal within the number.

For example...

0.1220 = four significant figures
(the first zero is not significant because it is on the left of the non-zero numbers; however, the second zero is on the right of a non-zero number so it is significant.)

45.0000 = six significant figures

Both of these numbers are assumed to be precise to the nearest 0.0001

Rule #5: All other zeros are insignificant.

For example...

0.56 = two significant figures

7840000 = three significant figures

Rule #6: Within scientific notation, only the coefficient may have significant figures, not the exponent.

For example...

2.348×10^7 = four significant figures
(pay no attention to the 10^7)

So why do we need to learn this?

Because knowing how precise our data is makes all the difference in the world in science!

Let's look back at our saltwater experiment – when we start applying math and calculate averages with our data, we need to be certain that our averages are a reflection of how precise our data was to begin with! Here's the data we are working with:

12.2 minutes + 11.9 minutes + 12.4 minutes = 36.5 minutes

36.5 minutes / 3 = 12.1666666 minutes

How precise is our data with 12.2, 11.9, and 12.4? All of them are assumed to be precise to the nearest tenth (0.1) of a minute. Now, how precise was our average – 12.166666 minutes? This answer is assumed to be precise to the nearest millionth (0.000001) of a minute. That's very precise! In fact, it is much more precise than the data we collected! This answer is way too precise and something needs to be done about it.

This is why we have to follow a simple rule when multiplying and dividing significant figures:

When multiplying and/or dividing significant figures, the answer should have the same number of significant figures as the number with the **least** number of significant figures. You can't have an answer that is more precise than all of the numbers that were used in the problem, right?

An analogy to this problem can be summed up like this:

A team can only be as fast as its slowest runner; similarly, a set of data can only be as precise as its most imprecise numbers.

With this in mind, let's fix our problem:

12.2 minutes + 11.9 minutes + 12.4 minutes = 36.5 minutes

36.5 minutes / 3 = 12.2 minutes

* You should notice that we rounded the original 12.166666 minutes to the nearest tenth of a minute.

Since significant figures give you important information about the precision of your data, it is important to always write your answers with the correct amounts of significant figures.

Now it's time to look at the second question we asked at the beginning of this chapter:

How accurate is our data?

Since accuracy is how close a measured value is close to the real value of the object, let's use our saltwater experiment again to determine the accuracy of our data.

Let's assume that before we ran our experiment, scientists had already determined the length of time it would take to boil our saltwater. For the sake of argument, let's say the accepted time for our saltwater to boil (according to these scientists) would be 12.35 minutes.

This amount is different from the 12.2 minute average we discovered. So how do we determine the accuracy of our data? By using...

Percent Error

The accuracy of a measurement is determined by the **percent error** of the measurement and can be calculated with the following equation:

$$\text{Percent error} = \frac{(\text{Accepted measurement} - \text{Experimental measurement})}{\text{Accepted measurement}} \times 100\%$$

So what is the percent error of our saltwater experiment?

$$\frac{12.35 \text{ minutes} - 12.2 \text{ minutes}}{12.35 \text{ minutes}} \times 100\% = 1.2\% \text{ Percent Error}$$

A high percent error tells us that the experimental data was not very accurate and should probably be investigated (or thrown out!)

So what does all this mean?

Nothing! That is, if you can't use it.

That is why you will be practicing the rules of significant figures and percent error a lot this week! You have a lot of number crunching to do before we dig into…

The Metric System!

Significant figures and percent error practice

How many significant figures do the following numbers have?

1) 1234 ______
2) 0.023 ______
3) 890 ______
4) 91010 ______
5) 9010.0 ______
6) 1090.0010 ______
7) 0.00120 ______
8) 3.4×10^4 ______
9) 9.0×10^{-3} ______
10) 9.010×10^{-2} ______
11) 0.00030 ______
12) 1020010 ______
13) 780. ______
14) 1000 ______
15) 918.010 ______
16) 0.01 ______
17) 0.00390 ______
18) 8120 ______
19) 7.991×10^{-10} ______
20) 72 ______

Solve the following mathematical problems such that the answers have the correct number of significant figures:

21) $34.1 \div 1.1$ = ____________________
22) $2.11 \times 10^3 \div 34$ = ____________________
23) $450 \div 114$ = ____________________
24) 84×31.221 = ____________________

25) $1.267 \times 42 \times 0.9963$ = ____________________

26) $4.993 \times 10100 \times 7$ = ____________________

27) $63 \times 49 \div 6.664$ = ____________________

28) 23.4×14 = ____________________

29) 0.0945×1.47 = ____________________

30) $7.895 \div 34$ = ____________________

Solve the following percent error problems:

31) For years, many sightings of the legendary monster – Godzilla have been reported along the Asian continent and, most recently, in downtown New York City. Little is known about his physiology as he creatively disguises himself by altering his body and facial structure from citing to citing. Most individuals would agree however, that he must be reaching the twilight of his years and should be at least 50 years old. If Godzilla's actual age is 56 years old, what is the percent error in this estimation?

32) The mass of a sample of a compound was experimentally found to be 45.0 grams. If the actual mass of the compound was 55.0 grams, what is the percent error of this calculation?

Chapter 3

Throughout the remainder of this book we will be using the metric system as it is used throughout the world as a standard method of measurement. Within the metric system, we can measure nearly everything in the known world. For our needs, we will be looking at only a few items to measure during our studies:

Time, Distance, Mass/Weight, Pressure, and Temperature

In addition, we will be using some basic metric units (called **base units**) for each of these measurements. Below you will find the base units for these types of measurements:

Base Unit	Measurements
Second	Time
Meter (m)	Distance
Kilogram (kg)	Mass/Weight
Liter (L)	Volume
Pascal (Pa)	Pressure
Kelvin (K)	Temperature

Why are they called "**base units**"? Because every metric measurement we make is based on these units!

Sometimes these units are not large or small enough to give us an accurate measurement. For example, you wouldn't measure the distance between two letters in this sentence with a meter stick! No way! You would probably want to use a measuring device that is much smaller.

Therefore, in order to make these base units as accurate as possible, we place prefixes in front of them to identify rather large or small numbers. Here's what I mean:

Prefix	What the prefix means is...
Kilo- (k)	1000 (10^3)
Hecto- (h)	100 (10^2)
Deka- (da)	10
Deci- (d)	0.1 (10^{-1})
Centi- (c)	0.01 (10^{-2})
Milli- (m)	0.001 (10^{-3})
Micro- (μ)	0.000001 (10^{-6})
Nano- (n)	0.000000001 (10^{-9})

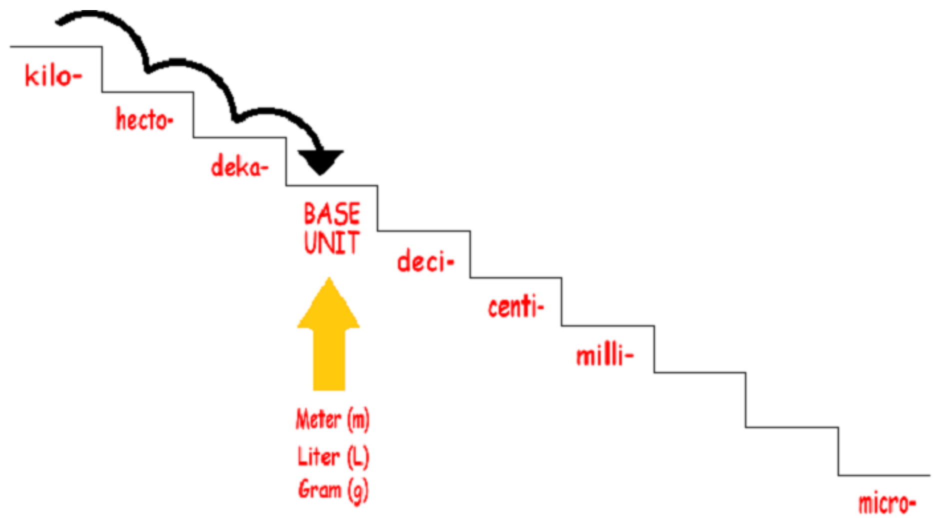
Does this look a little familiar? It should! We started using these types of exponents in Chapter 1. The exponents within the chart (i.e. 10^3, 10^{-1}, 10^{-9}) simply tell us **how many steps away from the base units they are located.** For example, 10^3 is **three** steps away from the base unit and 10^{-9} is **nine** steps away from the base unit.

Before we get started looking at how to use the math, I think you need to be certain how all these prefixes work.

When you use these prefixes in front of a number you always add the prefix symbol (i.e. k, h, da, d, c, m, μ, n) to the base unit symbol of your measurement (i.e. m, L, or g).

For example, if you have measurement of 1,200 deciliters, you would write it as 1,200 dL. A measurement of 230 micrometers would be 230 μm. And a measurement of 23 kilograms would be 23 kg. Simple, right?

Now let's look at the previous chart in a different way:



The purpose of drawing out the metric system like stairs is so that you can easily convert one metric measurement to another by counting how many times you have to move the decimal within the conversion. For example, let's say you have to travel a distance of 100 kilometers (km) and would like to know how many meters (m) it contains. By looking at our stairs you will find kilometers on the top and meters (a base unit) three steps down.

So to make your conversion from kilometers to meters you can place your 100 km on the top of the stairs in decimal form (which is 100.0 km) and then you would move the decimal three times to the right. This gives you a total of 100,000 m. Another way to look at these types of problem is…

Every step you move <u>down</u> the stairs, you multiply your number by 10.

If you start out with 100 km and you want to convert it to meters, you simply multiply 100 by 10 a total of three times.

$$100 \times 10 \times 10 \times 10 = 100{,}000 \text{ m}$$

*Don't forget! You can always use exponents with these numbers as well! "10 x 10 x 10" can also be written as 10^3

What if you wanted to know how many centimeters (cm) there are in 100 kilometers?

Instead of stopping at the base unit (meters), you would go down two more steps and keep multiplying by 10's.

100 km x 10 x 10 x 10 x 10 x 10 or 100 km x 10^5 cm = 10,000,000 cm

Can this work in reverse? Is it possible to use the same method for finding smaller numbers?

Oh yeah!

Instead of multiplying when moving up the stairs, you simply **divide by 10's** on each step. For example, let's say you have 700 centigrams (cg) of candy and want to know how many hectograms (hg) that equals.

First count how many steps exist between centigrams and hectograms:

= 4 steps

Next divide your 700cg of candy four times (one for each step) by the number 10:

$$700\text{cg} \div 10 \div 10 \div 10 \div 10 \quad \text{or} \quad 700\text{cg} \times 10^{-4} = 0.07\text{hg}$$

*Don't be confused by the **multiplication** within 700cg x 10^{-4} in this example. Remember that 10^{-4} really means 0.0001

Remember, you can either count up and down the metric stairs or you can multiply by an exponent. Either way, you will receive the correct conversion!

One final rule you can use to ALWAYS check your conversions is this:

During a conversion, if the units go up (like from cm to km), its number will go down in size. Likewise, if the units go down (like from km to m), then its number will be greater.

I know this all may seem a little simple, but believe me it is going to get much harder in the near future! The only way you are going to be able to tackle the harder problems is to completely understand the basics of metric conversions! You will thank me for practicing this week's problems. So let's get to work! Next week, you are going to learn how to take this metric conversion "stuff" to the next level.

Let's go!

Metric conversion practice

Convert the following numbers:

1) 1000 mg = _________ g
2) 1000 g = _________ kg
3) 1000 ml = _________ L
4) 47 mm = _________ cm
5) 130 cm = _________ m
6) 1200 m = _________ km
7) 3456 mm = _________ m
8) 55 mm = _________ cm
9) 4568 m = _________ km

State the name of the metric prefix and its symbol for each of the following numbers:

10) 10^{-2} _________
11) 10^{-9} _________
12) 1000 _________
13) 10^{3} _________
14) 10^{-1} _________
15) 0.000001 _________
16) 10^{-3} _________

Write the symbol for the following metric units:

17) microgram ________
18) micrometer ________
19) kilometer ________
20) millisecond ________
21) deciliter ________
22) kilocalorie ________
23) centiliter ________
24) nanogram ________

Write the name of the unit indicated by the following symbols:

25) m ________
26) μs ________
27) mL ________
28) dm ________
29) cL ________
30) kg ________

Perform the following metric conversions and write the answers in scientific notation.

31) 1.55 km to m ____________________

32) 0.486 g to cg ____________________

33) 1.885 L to dL ____________________

34) 0.388 m to mm ____________________

35) 0.00125 s to ms ____________________

36) 10.6 dg to g ____________________

37) 125 mL to L ____________________

38) 100 ns to s ____________________

39) 94.6 kcal to cal ____________________

40) 3.15×10^5 mg to g ____________________

Chapter 4

Last week, you learned how to convert a unit from one metric measurement to another. This is a very important skill to master! This week, you are going to use a new way of converting measured units which is known as:

Dimensional Analysis

Dimensional analysis is a method of converting one unit into another unit through the use of **conversion factors**.

A conversion factor is a shortcut that can be used to help with your conversion. For example, you probably know that every time a basketball player gets the ball into the hoop it counts as two points. The receiving of two points per basket is a conversion factor towards measuring how many baskets were made in the game.

If a player made a total of 42 points in one game, how many (2 point) baskets did he make?

You probably said "21" pretty quickly because 21 x 2 = 42. What you didn't realize is that you used a conversion factor to get your answer. Let's write it out and see what dimensional analysis can do:

One basket equals 2 points

You could interpret this by writing out the conversion factor like this:

2 points per basket

Or it can be written as a fraction like this:

$$\frac{1\text{ basket}}{2\text{ points}} \qquad \text{or} \qquad \frac{2\text{ points}}{1\text{ basket}}$$

It doesn't matter which is on top, because they both equal the same thing! Now let's look at what we are trying to convert:

42 points need to be converted to number of baskets

So, we can write out what we need to convert like this:

$$\begin{array}{c|c} \underline{42\text{ points}} & \underline{1\text{ basket}} \\ 1 & 2\text{ points} \end{array}$$

Why is there a "1" under the "42 points"? Well, any number that is divided by 1 is equal to itself. In future examples, you may not see the "1" present at all! This is not a problem as it does not affect any of your calculations. However, we put the "42 points" in the numerator position to make it easier for the next step in dimensional analysis...

The cool thing about using dimensional analysis is that it helps you cancel out unneeded units. Like with this basketball example, do we really need the unit of "points" in our answer? No!

We are trying to convert points into baskets. Therefore, dimensional analysis is being used to cancel out the unit of points. And this is done very easily:

$$\frac{42\ \cancel{\text{points}}}{1}\left|\frac{1\ \text{basket}}{2\ \cancel{\text{points}}}\right.$$

Since we no longer have the units of "points" in our problem, we can now multiply our numerators together.

42 x 1 basket = 42 baskets

We then multiply our denominators together. In this case it would be:

1 x 2 = 2

We then divide the numerator and the denominator to get our answer:

$$\frac{42\ \cancel{\text{points}}}{1}\left|\frac{1\ \text{basket}}{2\ \cancel{\text{points}}}\right. = \frac{42\ \text{baskets}}{2} = \textbf{21 baskets}$$

Alright! Let's review what we just did and start diving into how this can help us in the field of chemistry:

1. First we figured out what we are going to convert (points) and what we are going to convert it into (baskets.)
2. We then lined up our conversion factor (2 points per basket) so that our unneeded units (points) would be cancelled out.
3. To complete this problem we multiplied the top numbers together (42 x 1) and the bottom numbers together (1 x 2) and divided the numerator by the denominator (42 ÷ 2).

Now let's use this method to solve a more scientific problem!

How many milliliters (mL) are there in 5 liters (L)?

First, let's make certain you know we are converting liters (L) into milliliters (mL).

Second, let's line up our conversion factor so that we can cancel out our unneeded units (L).

5 ~~Liters~~	1000 milliliters
	1 ~~Liter~~

* Remember your metric conversions! There are 1000 milliliters in a liter because there are three steps between milliliters and liters! You multiply 10 by itself for each step you take between the units (10 x 10 x 10).

Third, multiply the top numbers (5 x 1000) together and divide it from the bottom number (1).

5 ~~Liters~~	1000 milliliters	= **5000 milliliters**
	1 ~~Liter~~	

Got it? I'm sure you did! Now let's look at some other variables that can exist while using dimensional analysis:

A pound of butter costs $1.39. What is the cost of a pound of butter in kilograms? (453.6 grams = 1 pound)

To solve this problem, you need to know the conversion factor of pounds to kilograms. However, you are only given the amount of grams per pound! No worries. Just follow through the steps like we did earlier and you'll get the answer.

We are converting dollars per pound to dollars per kilogram so we can set up our problem like this:

$1.39	1 pound	1000 grams
pound	453.6 grams	1 kilogram

We added the 1000 grams per kilogram conversion factor into this problem so that we could:

a) cancel out the units of "grams"; and,
b) place the units of kilograms into the bottom part of our equation

Now we need to start canceling our units:

$1.39	1 ~~pound~~	1000 ~~grams~~
~~pound~~	453.6 ~~grams~~	1 kilogram

Next, we need to start multiplying the top numbers together and the bottom numbers together. Once we have both of these numbers, we can divide them to get our answer.

$$\begin{array}{c|c|c} \$1.39 & 1\ \cancel{\text{pound}} & 1000\ \cancel{\text{grams}} \\ \hline \cancel{\text{pound}} & 453.6\ \cancel{\text{grams}} & 1\ \text{kilogram} \end{array} = \frac{\mathbf{\$1390}}{\mathbf{453.6\ kilograms}} = \mathbf{\$3.06\ per\ kilogram}$$

* Please look at how we carried the units of kilograms across when we multiplied 453.6 x 1 kilogram. Remember, the use of dimensional analysis is to cancel out **units**. During the conversion process, the units are multiplied with the numbers as well!

The use of dimensional analysis is going to be very important throughout this course as it is used to calculate the tiniest of particles within the universe. Of course, before we can measure these objects you need to understand what they are! This is the topic for our upcoming unit:

The Structure and Properties of Matter

Dimensional analysis practice

Perform the following conversions using dimensional analysis:

1) 894.2 cm = ____ km

2) 97 mL = _____ gallons

3) 133.5 mm = _____ inches

4) 14.5 feet = _____ meters

5) 337 milliseconds = _____ kiloseconds

6) How many centigrams are in 43 kilograms?

7) How many miles will a person run during a 10 km race (1 mi = 1.609 km)?

8) The moon is 250,000 miles away. How many inches is it from earth (1 mi = 63,360 inches)?

9) A family pool holds 10,000 gallons of water. How many liters is this (1 L = 0.264 gal)?

10) How many seconds are there in 3 months, assuming there are 30 days in a month?

11) If you have 3.37 oz of a fluid, how many milliliters is this? (1 mL = 0.034 oz)

12) A weather station measures 0.8 yards of rain. Express this amount in centimeters (1 cm = 0.01 yds).

13) How many meters are in 750 kilometers?

14) How many liters are in 52.3 milliliters?

15) Determine the number of years in 835,000,000 minutes.

16) Light travels through space at 186,000 miles per second or 300,000 kilometers per second. How far will light travel in one year? [This distance is called a light-year (ly)]

17) The sun is 93 million miles or 150 million kilometers away. If you could drive a car at 65 miles per hour to the sun, how long would it take?

18) How many blank CDs will Mr. Q need if he wants to digitize eight (8) of his old photo albums? (1 old album = 450 pictures, 1 picture = 2.2 Mb, 1 CD = 760 Mb)

19) If Mr.Q were to drink 13 cups of coffee during a day, how many gallons would this be?

Chapter 5

I'm certain that you have already learned before that nearly everything you see in the universe is made up of tiny particles called **atoms**. These atoms can be broken down into even smaller, **subatomic particles** called **protons**, **neutrons**, and **electrons**.

Each proton has a mass of 1.66×10^{-27} kg which is referred to as an **atomic mass unit (amu).** These units are used to determine the overall **atomic mass** of an element. Stay tuned for more on this topic… A proton carries with it a certain amount of force that is known as a charge. The unit we use to determine the charge of a proton is +1.

Much like the proton, the neutron has nearly the same mass. Therefore, both of these subatomic particles have a mass of 1 amu. However, unlike the proton, this particle has no charge whatsoever. It is said to have a neutral charge.

The electron is nothing like its other two counterparts. This particle is much smaller than both the neutron and proton. In fact, it is nearly 2000 times smaller (1/1836 amu).

However, if you would put an electron and a proton in a tug-of-war contest, neither of them would win!

Despite its tiny size, an electron carries an equal and opposite -1 charge as compared to a proton.

A total of around 120 different kinds of atoms have been discovered in the known world. Each of these atoms, known as **elements**, has their own unique properties that make them different from each other. The single most important difference between elements is the number of protons each element contains. Since the number of protons change with each unique element, the mass of each element will also change as well.

Since we are talking about mass, let's look at how these subatomic particles fit together.

The structure of an atom places the neutrons and protons in its center which is known as a **nucleus**. Since the protons and neutrons are the largest subatomic particles, they make up the majority of an atom's mass. The sum of the protons and neutrons within an atom is known as the element's **mass number**.

But what about the ratio of protons to electrons?

For now, it is safe to assume that an atom contains an equal number of protons and electrons. You'll see how this changes in future chapters.

(We are only assuming because electrons don't always stay put! Stay tuned for more details.)

The electrons are typically found moving around the nucleus and do not contribute much to the mass of the atom due to their small size. They do use their strong charges to balance out the overall charge of an atom!

The strong negative charge and their constant motion allow electrons to move between atoms at times, but we'll save that story for a later chapter. Let's spend our time learning how to calculate these subatomic particles for our future use.

Scientists use two very important measurements to determine the characteristics of elements:

Atomic Number and Mass Number

The **atomic number** of an element is simply its number of protons. This measurement is the most precise way of identifying an element.

The **mass number** of an element is the mass of the protons **AND** neutrons within an element's nucleus which is measured in amu's. Don't worry about calculating the mass of the electrons. These guys are so small, chemists tend to ignore them in most calculations.

Even though we are assuming that the number of protons equal the number of electrons for now, this is not the case for the neutrons.

In fact, most elements can contain a variety of neutrons, each of which is called an element's **isotope**. For example, the element carbon has three isotopes that occur naturally in the world:

Carbon-12 has six protons and six neutrons

Carbon-13 has six protons and seven neutrons

Carbon-14 has six protons and eight neutrons

Since there are so many different isotopes of elements, scientists needed a way to determine the average mass of an element with all of its isotopes.

How could this be calculated?

After much experimentation, a measured percentage of each element that exists in nature was determined. This is known as the **relative abundance** of an element. With this knowledge, a more precise average mass of an element and all of its isotopes could be calculated. This average mass is known as the **atomic mass**.

Here's how you calculate the average atomic mass of an element:

1. Make a list of each of its isotopes along with the mass numbers and their percent relative abundance.
2. Multiply the mass number of each isotope by its relative abundance.
3. When you add all of these products together, you have calculated the atomic mass.

Let's look at the two naturally occurring isotopes of another element (chlorine) as an example:

Chlorine-35 (^{35}Cl) has 17 protons and 18 neutrons and has a relative abundance of 75.8%

Chlorine-37 (^{37}Cl) has 17 protons and 20 neutrons and appears 24.2% of the time.

So how do we find the average atomic mass?

We have all of the data ready to calculate the atomic mass so let's do some math! First, we multiply the mass number of each isotope with its relative abundance. The percentages will be placed in their decimal form:

(Chlorine-35) 35 amu × 0.758 = **26.53 amu**
(Chlorine-37) 37 amu × 0.242 = **8.95 amu**

Next we add together both of these products to get the atomic mass of chlorine:

26.53 amu + 8.95 amu = **35.48 amu**

Okay, let's review:

- The number of protons = the number of electrons within an atom (for now!)
- The number of protons is known as the atomic number.
- The mass number is the sum of the atom's protons and neutrons.
- Since most elements have isotopes, an average atomic mass can be calculated once we know their relative abundance in nature.

Now that you understand the basics of subatomic particles, let's see what they can do when we set them in motion. See you next week!

Periodic Table of the Elements

Key

11 — Atomic number
Na — Element symbol
Sodium — Element name
22.99 — Average atomic mass

	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1	1 H Hydrogen 1.01																	2 He Helium 4.00
2	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
3	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
4	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
5	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
6	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)
7	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)									

Lanthanides	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
Actinides	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)

Isotopes and average atomic mass practice problems

Here are three isotopes of an element: ^{12}C ^{13}C ^{14}C

1) The name of the element is: ____________________
2) The numbers 12, 13, and 14 refer to the ____________________
3) Complete the table below:

atomic symbol	atomic #	mass #	# of protons	# of neutrons	# of electrons
^{12}C					
^{13}C					
^{14}C					

4) Naturally occurring europium (Eu) consists of two isotopes was a mass of 151 and 153. ^{151}Eu has an abundance of 48.03% and ^{153}Eu has a relative abundance of 51.97%. What is the atomic mass of europium?

5) Strontium consists of four isotopes with masses of 84 (abundance 0.50%), 86 (abundance of 9.9%), 87 (abundance of 7.0%), and 88 (abundance of 82.6%). Calculate the atomic mass of strontium.

6) Rubidium is a soft, silvery-white metal that has two common isotopes, ^{85}Rb and ^{87}Rb. If the abundance of ^{85}Rb is 72.2% and the abundance of ^{87}Rb is 27.8%, what is the average atomic mass of rubidium?

7) Uranium is used in nuclear reactors and is a rare element on earth. Uranium has three common isotopes. If the abundance of ^{234}U is 0.01%, the abundance of ^{235}U is 0.71%, and the abundance of ^{238}U is 99.28%, what is the average atomic mass of uranium?

8) Titanium has five common isotopes: ^{46}Ti (8.0%), ^{47}Ti (7.8%), ^{48}Ti (73.4%), ^{49}Ti (5.5%), ^{50}Ti (5.3%). What is the average atomic mass of titanium?

9) Why is the atomic mass in amu of a carbon atom reported as 12.011 in the periodic table of the elements?

Chapter 6

Last week, you explored the subatomic particles that make up an atom. But it is pretty rare to find only a single atom hanging around all by itself. Usually, atoms are found in extremely large numbers either bound to each other or mixed together. This brings us to two new categories of matter:

Pure Substances and Mixtures

Throughout this book, we will be studying two different types of pure substances: **Elements** and **Compounds**

If you recall from last week, **elements** occur when only one kind of atom is present within a material. **Compounds** are pure substances that are made up of two or more different atoms that are bonded together.

So what is the difference between a compound and a molecule?

Well, a **molecule** is a general term which describes a combination of any two or more atoms that are bonded together. However, a compound is a molecule that contains at least two different elements. Therefore...

All compounds are molecules, but not all molecules are compounds!

One major rule you should know about **pure substances** is they cannot be physically separated (like with your hands) because they tend to be bound together! This is much different than **mixtures** which are not bound together.

Mixtures exist in two different forms:

Homogeneous and Heterogenous

("hah-mah-jen-us") ("het-tur-oh-geen-us")

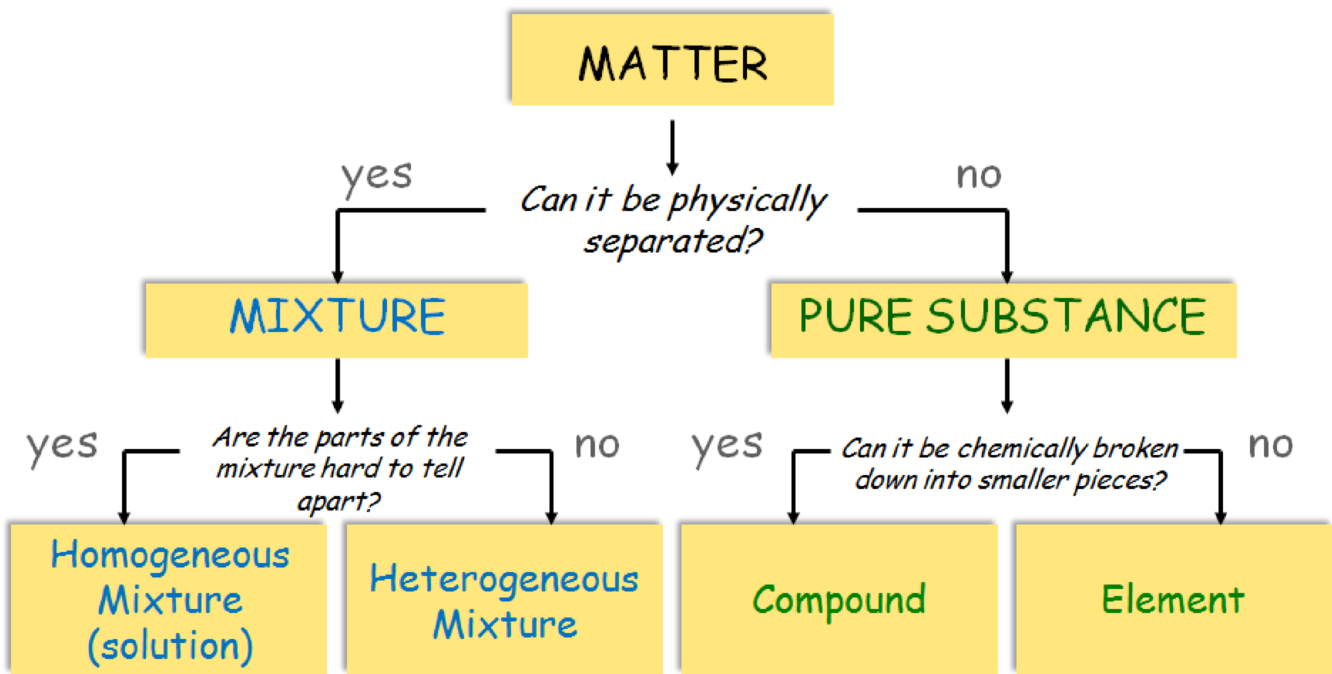
Homogenous mixtures (also known as **solutions**) are made up of two or more different kinds of atoms that are not easy to separate (but they are not bound together.) For example, iced tea would be a solution of water and the dissolved molecules from the tea leaves. It would not be easy to separate the tea-flavored molecules from its watery solution. (I'd like to see you try!) Your breath is another example of a solution of different kinds of gas molecules. Even a chocolate chip cookie can be considered a homogenous mixture – do you think it would be easy to separate the sugar, flour, baking soda, vanilla extract, etc. from a baked cookie? I don't think so!

Heterogeneous mixtures are made up of two or more different kinds of molecules that are not bound together too. However, this kind of mixture is made up of molecules that are so different in size and shape they are easy to separate.

For example, a bowl of M&M's would be a heterogeneous mixture (if they are all different colors.) The oil and vinegar mixture in Italian salad dressing naturally separates into two layers. And the reasons why many families purchase radon or carbon monoxide detectors in their home are examples of how these gases form heterogeneous mixtures with the air.

**We'll be talking more about these gases in future chapters!*

That's right! Solids, liquids, and gases can all be homogenous and heterogeneous mixtures!



Throughout this book we will investigate the properties of solids, liquids, and gases as the main states of matter studied in chemistry. A collection of atoms can change between these three states (also known as **phases**) when one or both of their following properties change:

Temperature and Pressure

At room temperature (around 25 degrees Celsius), most elements are either solids or liquids. However, as the temperature of an object increases matter tends to move from a solid phase to a liquid phase (**melting**) and then from a liquid phase to a gas phase (**boiling**.) This process can be reversed when energy is removed from a gas during cooling. The gas can turn into a liquid (**condensation**) and then back into a solid (**freezing**.)

You will be looking at the concept of **pressure** in a few chapters, for now let's just define pressure as "the force of gas molecules hitting the sides of the container in which they are stored."

Under most circumstances:

High pressures and low temperatures = Solids

Low pressures and high temperatures = Gases

Everything in the middle = Liquids

It is very easy to assume that everything either melts or boils at temperatures that are easy to create here on Earth. This is not always true!

Some of these temperatures are so incredibly high or low that a thermometer we typically use in the kitchen or outdoors would not work. Therefore, scientists created a new temperature scale called the **Kelvin scale** which sets its zero degree (0°) at the most extreme measurement known as **absolute zero**. At absolute zero, all elements are in the solid phase.

So what is the conversion factor for Celsius to Kelvin?

This is easy!

To convert any Celsius degree to Kelvin, simply add *273* to your Celsius degree.

Freezing point of water = 0° Celsius = 273° Kelvin

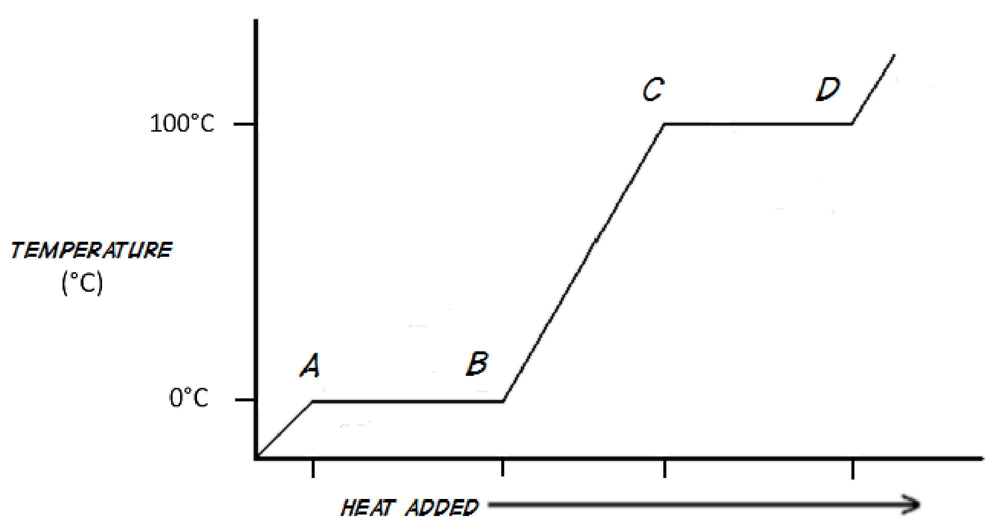
Boiling point of water = 100° Celsius = 373° Kelvin

Here's an important thing to remember:

Phase changes occur due to the transfer of energy into or out of atoms.

Transferring heat energy into a solid causes it to melt and removing heat from a liquid causes it to freeze. The cool thing is that every atom has the ability to absorb heat energy in different amounts. Therefore, every molecule has its own unique melting and boiling points. With the massive numbers of possible combinations of these atoms, scientists needed a way to organize this information in the easiest way possible; therefore, **phase change diagrams** were created!

This image displays a phase change diagram for water. It should be fairly obvious that as heat is added, the temperature of water increases as well.



You could probably also figure out that point A is the melting point of water and point C is the boiling point.

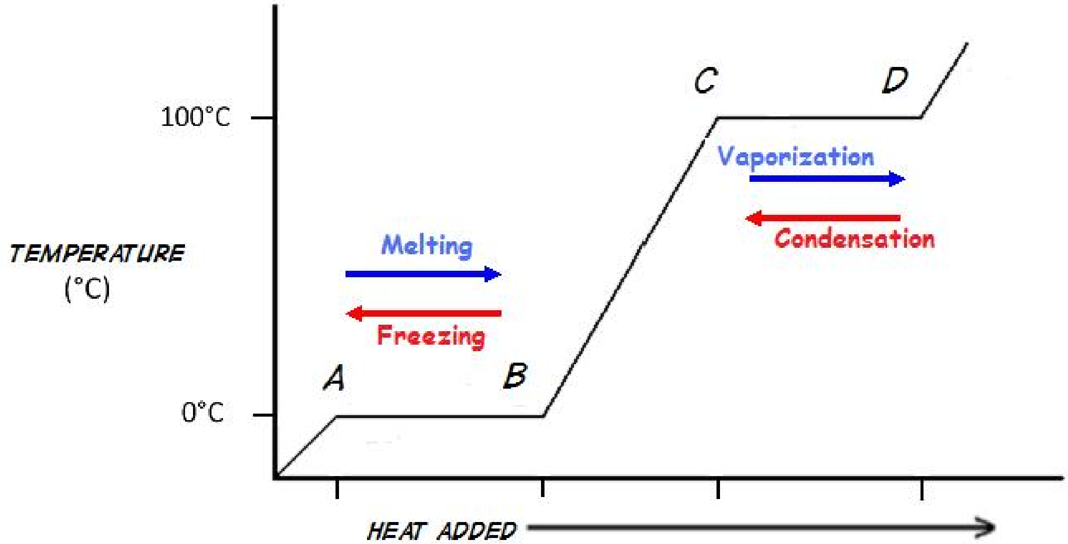
What may not be so easy to figure out is why the temperature does not rise between points A-B and C-D.

Any guesses?

Hopefully, you predicted that as solid ice begins to melt at point A, energy continues to be absorbed by the solid until all of the solid ice is turned into liquid. Point B describes the point in which all of the solid ice has melted and only liquid water remains. At this point, the liquid begins to absorb more energy and the molecules of water begin to move around more. This causes its temperature to rise.

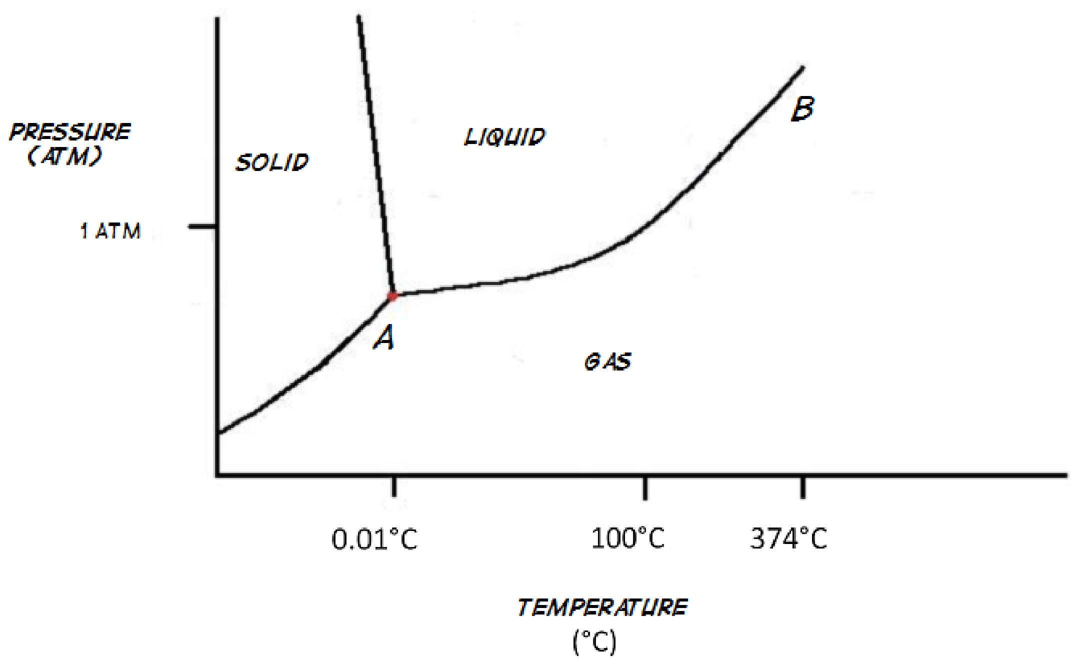
The same thing happens again at point C; however, the phase changes are moving from a liquid into a gas. The liquid molecules absorb enough energy at point C to move away from each other and escape as a gas. The temperature of the liquid will remain constant until all of the molecules in a liquid state have turned into a gas at point D. Here, the gas molecules will begin to heat up again!

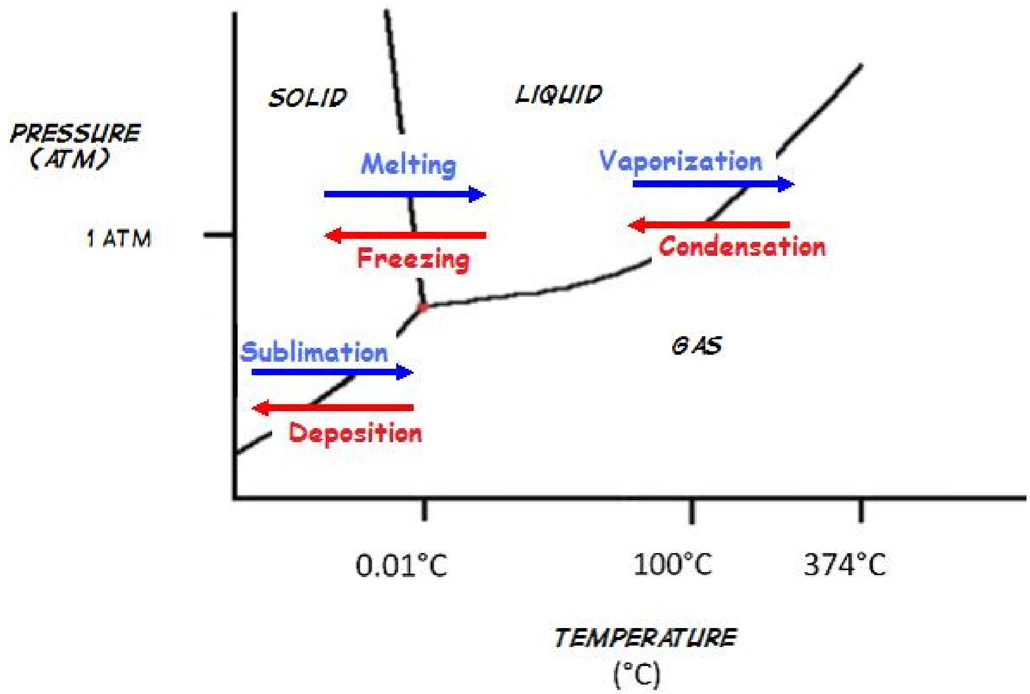
And don't forget that we can reverse this entire process as well! All you have to do is remove heat energy from the gas to create a liquid (condensation) or remove heat from the liquid to create a solid (freezing). Here' another way to look at this diagram:



Now that you know what happens between points A through D…

…let's throw another variable into the diagram and see what **pressure** does to our graph:

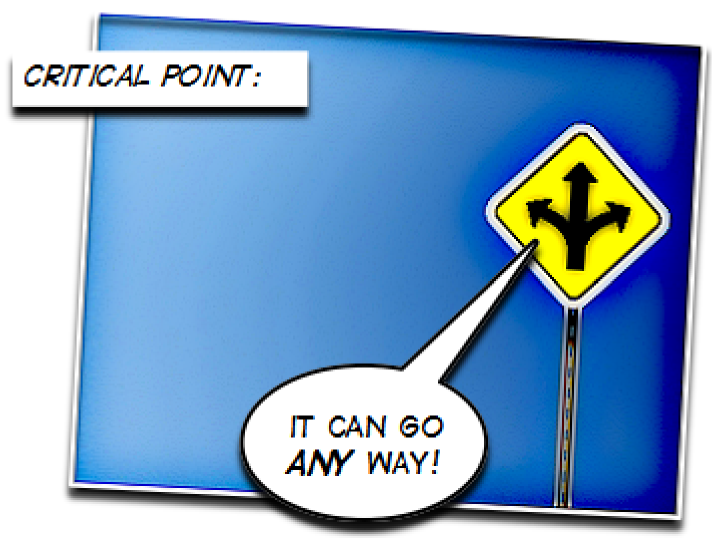




Air pressure is measured in units called atmospheres (atm). One (1) atm is considered to be the "normal" atmospheric pressure. So if you were to draw an imaginary horizontal line from the 1 atm mark on the graph across the graph until you reached room temperature (~25°C), your line would end within the liquid section of the graph. This means that at room temperature, water is a liquid! Here's a different look at this phase change diagram:

Point A within this temperature/pressure graph is called the **triple point** and is a place where water can be either a solid, liquid, or gas. Weird, huh?

Point B is just as weird! At 374°C, water begins to act like both a gas and a liquid. This point is known as the **critical point**. Try to imagine a gas that can dissolve solid material. Yep…pretty weird. Luckily we will not be running any experiments at that temperature in this book!



Every element and molecule has their own unique melting and boiling points.

You have only seen the phase change diagrams for water. You should remember one important thing…

Every element and molecule has its own phase change diagram!

I think you understand why we can't put all of them in this chapter!

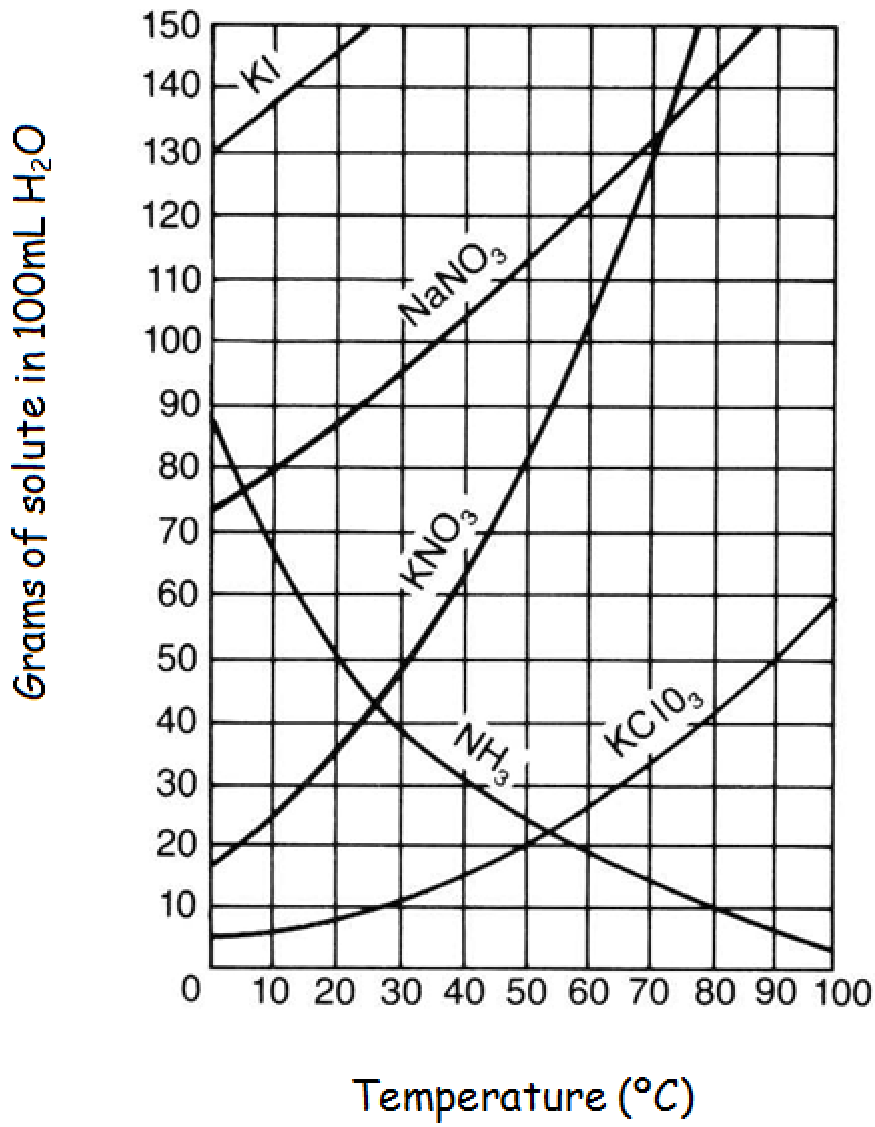
Okay, you probably have dozens of questions right now like:

- How to these atoms stick together to form molecules? And how do the molecules stick to each other?
- What causes molecules to absorb and release energy?
- Why do all atoms have unique melting and boiling points?

These questions will all be answered in later chapters. I promise you! We are only laying the foundation for a much stronger understanding of our submicroscopic world. Please be patient! Before we can tackle these questions, we need to look at how we organize the elements into a logical and easy to follow pattern. Stay tuned!

Solubility graph and classifying matter practice

A solubility graph can be used to solve a variety of questions concerning the temperature and pressure of mixtures and pure substances. The first type is simply to identify a substance when you are given the solubility in g/100 mL of water and the temperature. All you do is see which solubility curve the solubility and temperature intersect at. For example, what substance has a solubility of 90 g in 100 mL of water at a temperature of 25°C? The only substance whose solubility curve is located at the intersection of 90 g in 100 mL and 25°C is sodium nitrate.



1) How many grams of NH_3 will dissolve in 100 mL of water at 20°C AND at 80 °C?

2) How many grams of $NaNO_3$ will dissolve in 100 mL of water at 70°C?

3) How hot does 100 mL of water have to be in order to dissolve 150g of KI?

4) How hot does 100 mL of water have to be in order to dissolve 120g of KNO_3?

5) How hot does 100 mL of water have to be in order to dissolve 80 g of $NaNO_3$?

6) How many grams of $KClO_3$ will dissolve in 100 mL of water at 80°C?

Classify each of the materials below. In the center column, state whether the material is a **pure substance** or a **mixture**. If the material is a pure substance, further classify it as either an **element** or **compound** in the right column. Similarly, if the material is a mixture, further classify it as **homogeneous** or **heterogeneous** in the right column.

Material	Pure Substance or Mixture	Element, Compound, Homogeneous, Heterogeneous
concrete		
sugar + pure water ($C_{12}H_{22}O_{11}$ + H_2O)		
iron filings (Fe)		
limestone ($CaCO_3$)		
orange juice (w/pulp)		
Pacific Ocean		

Material	Pure Substance or Mixture	Element, Compound, Homogeneous, Heterogeneous
air inside a balloon		
aluminum (Al)		
magnesium (Mg)		
acetylene (C_2H_2)		
tap water in a glass		
soil		
pure water (H_2O)		
chromium (Cr)		
Chex mix		
salt + pure water ($NaCl + H_2O$)		
benzene (C_6H_6)		
muddy water		
Brass (Cu mixed with Zn)		
baking soda ($NaHCO_3$)		

Chapter 7

It's about time we start looking at the most important tool scientists use in the field of chemistry:

The Periodic Table

Chapter seven is going to set the stage for a deep understanding of this invaluable tool. I hope you don't think I'm going to ask you to start memorizing the periodic table right now? That would not be fun at all. You will be using the elements within the periodic table throughout this book. And most of them will be the first 18 elements! Through all of your practice problems, you will become very familiar with a good portion of the table. Now, let's get to work...

You already have a strong understanding of how the nucleus of the atom is arranged with protons and neutrons. And with this knowledge you should be able to see a couple of easy patterns within the table and how it is set up.

But there is so much more to know!

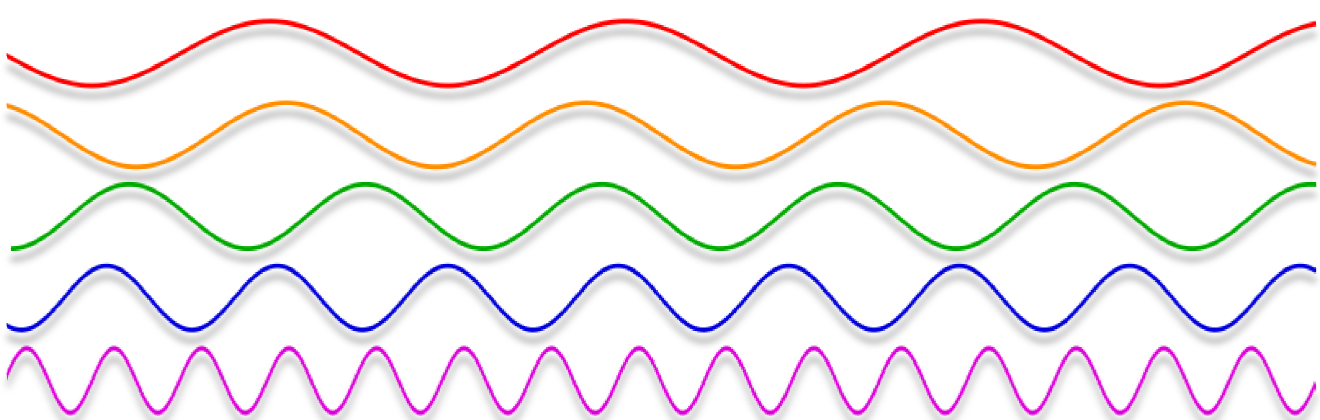
What we are going to do today is to move away from the nucleus and start exploring the tiny negatively-charged particles buzzing around the core of an atom. That's right! In order to understand the periodic table, you need to have a strong understanding about...

Electrons

Before we go any further, it is very important to know that the movement of electrons is the cause for all forms of electricity. Every time you connect a battery to a remote control, turn a key in your car, or flip a switch on the wall, pathways of electrons are allowed to move freely through circuits which have been designed to allow buttons to relay signals to your TV, the car battery to turn on your engine, or to simply pass through filaments in light bulbs to give us light.

With this knowledge, you should be able to predict why we use metal to construct massive radio towers and antennas on your cars. Metals are made up of atoms that allow at least some of their electrons to move around rather freely. But to understand **how** these structures work is vital towards our understanding of the atom.

Radio or TV broadcasters send their signals all over the world by moving electrons back and forth and side to side really fast. How? Because every time an electron is oscillated in this way, energy is released in the form of **electromagnetic waves**. By controlling the rate of these oscillations (known as **frequency**), different amounts of energy can be sent out.



In the drawing above, the electromagnetic waves on the bottom have higher frequency and therefore contain a higher amount of energy. The waves on top have a lower frequency and much lower energy. What's even more beneficial is what happens after these waves are created.

Not only can these waves be created by oscillating electrons, waves can also cause other electrons very far away to oscillate too!

Therefore, the electrons within the antenna in your cell phone or on your radio begin to oscillate at the same frequency of the wave that is slamming into its metal surface!

We need to look a little deeper at electromagnetic waves to see what all of this has to do with chemistry. Hang on…

First of all, the frequency of electromagnetic waves provides not only radio and cell phone signals; they also provide a very wide range of energies that give us the microwaves, x-rays, and radiation. Even the light we see is a range of electromagnetic wave frequencies!

Look around you right now. Every time light is generated from a light bulb, a heated metal surface like a cooking pan, or a heated non-metal surface like a block of wood, waves of electromagnetic energy are being released. Cool, huh?

Second, you need to imagine holding onto a string or rope, tying one end of it to a door, taking a few steps back, and swinging your arm back and forth (oscillating) to create the patterns of waves much like the picture on the previous page. If you have never done this before, you may not realize that it takes a certain amount of energy to create each of these wave patterns. Once the rate of your arm moving up and down (frequency) provides a specific amount energy into the rope, one of these waves are formed! And once you create one of these wave patterns, it is pretty easy to keep it going.

But if you try to change the length of the rope, you are going to have to change your frequency to keep the same wave pattern.

Think of it this way: The rope you are swinging around will only accept certain frequencies from your arm in order to create a stable wave pattern. By changing the length of rope, you are going to have to change the frequency in order to achieve the same wave pattern.

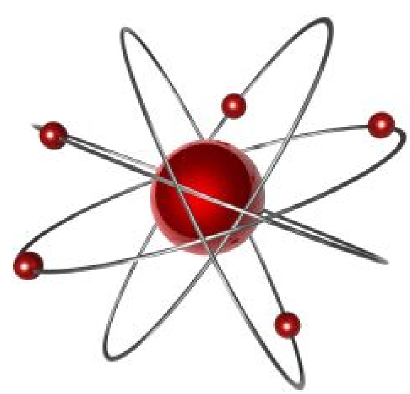
Now, how does all this tie into chemistry?

1. By oscillating electrons we create electromagnetic (EM) waves;
2. Stable waves will only be created by oscillating materials (such as electrons) that are moving with a precise frequency;
3. Changing the amount of oscillating material (electrons) will change the frequency of wave patterns;
4. We know that all elements contain unique amounts of electrons;
5. Therefore, each element can produce a unique EM wave pattern which provides its own unique amount of energy!

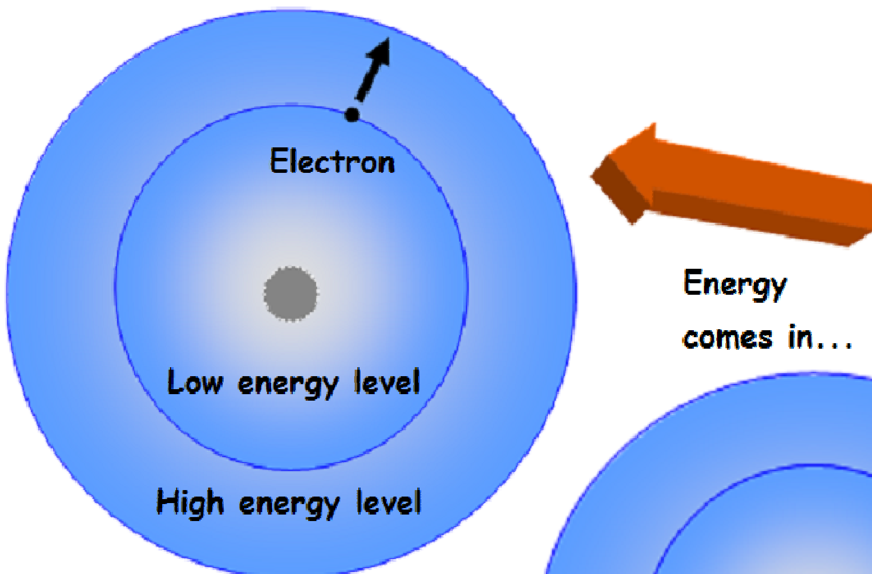
Now that you know that the electrons within each atom determine the amount of energy it can provide, think about all of the different colors that are created from **fireworks**. Each color you see in the night sky from a firework is a specific EM wave. That wave was caused through the rapid burning of a specific element. For example, yellow sparks are created from burning the element sulfur.

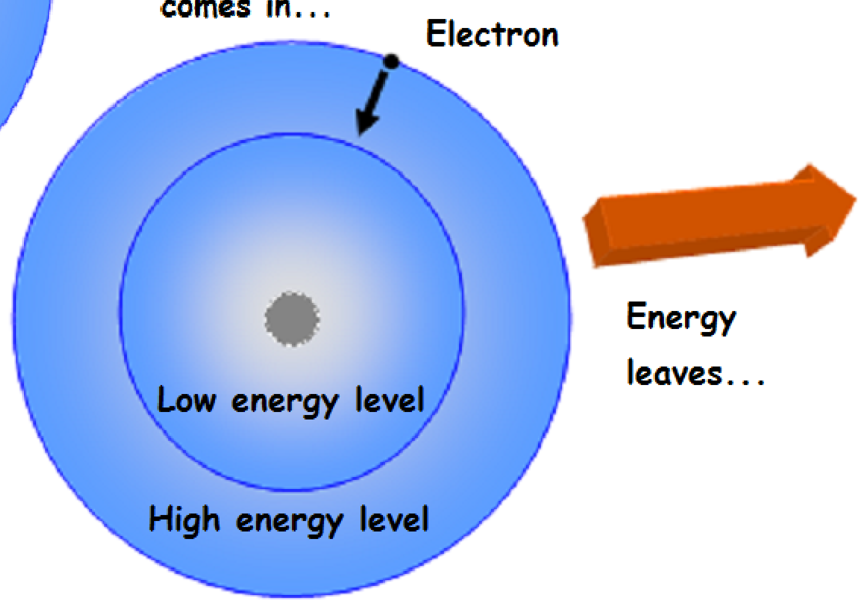
But how do atoms give off light?

In order to answer this question, you need to visualize a model of the atom. First of all, throw away anything that looks like this in your mind:



A better model for the atom places its electrons in different **energy levels** from the nucleus. The differences between these levels identify the frequencies of light energy that each atom can give off. For example, in order to get our yellow light from an atom of sulfur we first have to add energy to get the process working. This comes from the fuse or other source of ignition inside the firework. This energy is absorbed by the atoms of sulfur and causes electrons in their low energy level to jump into a high energy level. However, these electrons do not stay long within this high energy level and jump back to the lower levels. When this "jumping back" occurs, energy is released in the form of light. Here's what it looks like:





The light energy (yellow light) that is given off by sulfur is equal to the difference in energy between its energy levels!

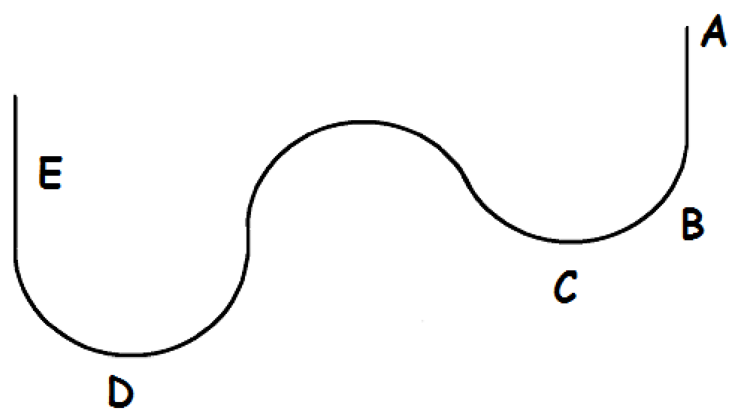
The next question you may be asking yourself is:

Why is the lowest energy near the nucleus of the atom?

Let me answer that question with another question. Would you rather have a ball dropped on your head from a distance of 1cm or 1km? Hopefully you said 1cm! The energy of the ball increases as you move it farther from your head. The same is true with electrons within an atom. The farther they are from the nucleus, the more energy they have.

Another question you may be asking is:

Why do electrons jump back to a lower energy level?



Imagine a skater at the top of a ramp like this:

We just learned that the lowest possible energy of a system would be found the area closest to the ground. <u>In this skate park, point D would have the lowest possible energy.</u>

Now if we set our skater in motion and begin at points A or E, she will probably end up at the lowest area of the ramp which is point D. However, if she begins at point B, it is unlikely she will make it all the way to point D. The lowest possible area she could hope for would be the trough at point C.

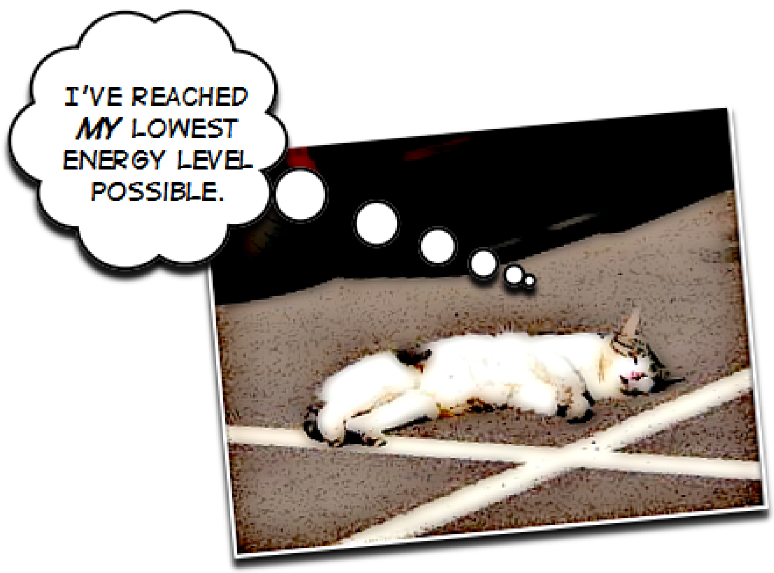
The reason why electrons move back from a high energy level to a low energy level can be found with the actions of our skater:

Everything in the universe tends to move towards the lowest energy possible.

*This is a very very important rule in science and will be seen many times throughout our studies!

Even if she couldn't always reach point D, the area of lowest energy, she reached the lowest **possible** energy she could achieve!

When electrons jump from a lower energy level to a higher energy level, they cannot stay in that higher energy level for long. Why? Because electrons, like our skater, tend to move towards the lowest energy level possible. That is why they jump back. And when they do, what happens to the energy that they absorbed? You guessed it! It gets released out of the atom in the form of light.



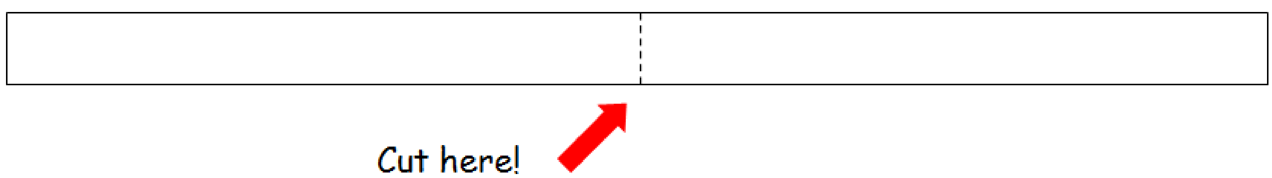
The movement of electrons to their lowest possible energy within and among atoms is one of the central ideas we will cover next week. You've absorbed a lot of new information for now. So relax, you deserve a small break! See you next week.

Electrons and light energy practice

Your first activity this week may seem a little simple; however, it is very important that you understand the size of the particles you are studying. You are going to need a piece of paper, a ruler and a pair of scissors to begin your practice work this week...

First, cut out a 28cm X 1cm strip of paper.

Second, cut the strip of paper in half as many times as possible. Be certain to count how many times you are able to cut the strip of paper.



Third, compare the number of cuts you were able to make with the table below. The following conversions may be helpful:

1 meter (m) = ~39 inches
1 centimeter (cm) = 0.01 m
1 millimeter (mm) = 0.001 m
1 micron (μm) = 0.000001 m

The following table that shows how many cuts it takes to get to the size of an electron

Cut #	Metric	English	Items of similar size
Cut 1	14.0 cm	5.5 in	Small book, hand
Cut 2	7.0 cm	2.75 in	Fingers, apple
Cut 3	3.5 cm	1.38 in	Watch, mushroom
Cut 4	1.75 cm	.69 in	Keyboard keys, rings, insects
Cut 6	.44 cm	.17 in	Peas, poppy seeds

Cut 8	1 mm	.04 in	Thread, sharp pencil width
Cut 10	.25mm	.01 in	If you can cut this small, you are superhuman!
Cut 12	.06	.002 in	Microscopic range, human hair
Cut 14	.015 mm	.006 in	Width of paper, microchip components
Cut 18	1 μm	.0004 in	Water purification openings, bacteria
Cut 19	.5 μm	.000018 in	Visible light waves
Cut 24	.015 μm	.0000006 in	Electron microscope range, DNA, membranes
Cut 31	1×10^{-10} m	4.5×10^{-9} in	The size of an atom!
Cut 41	~ 1×10^{-15} m	~ 4×10^{-14} in	The size of the nucleus of an atom (the largest nuclei would be this amount x10)
Cut 58	~ 1×10^{-19} m	~ 4×10^{-18} in	**The upper limit on the size of electrons!**

How many cuts could you make? Most people can make about 7-8 cuts. Even though the size of an electron is nearly impossible to understand, try to visualize cutting your strip of paper 58 times in order to obtain a piece of matter the size of an electron!

When energy is added to atoms, they give off light. For example, turning on an ordinary light bulb causes an electric current to flow through a metal filament that heats the filament and produces light. The electrical energy absorbed by the filament "excites" the electrons within the atoms, causing them to jump towards higher energy levels. When they jump back, the absorbed energy is released from the atoms in the form of light.

All of this movement by electrons provides different wavelengths of light to be emitted by each element. Each wavelength can be measured by the specific pattern of light that is created. These patterns are known as **line spectra.**

When normal white light, such as that from the sun, is passed through a prism, the light separates into a continuous spectrum of colors:

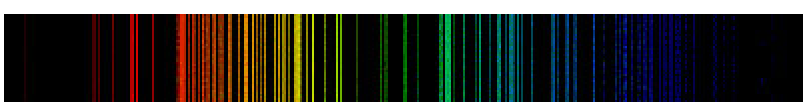
Continuous (white light) spectra

When light from an excited element is passed through a prism, only specific lines (or wavelengths) of light can be seen. For example, when hydrogen is heated and the light is passed through a prism, the following line spectra can be seen:

Hydrogen line spectra

As stated before, each element has its own distinct line spectra. For example:

Helium line spectra



Neon line spectra

Questions:

1) What is a continuous spectrum?

2) What does a line spectra phenomenon tell you?

3) What happens when an atom is excited?

4) What happens when electrons fall back to lower energy levels?

5) Why do electrons jump back to lower energy levels?

Chapter 8

We have a lot to cover today! Let's start off by briefly reviewing last week's main points:

- Each element can produce a unique EM wave pattern (frequency) which provides its own unique amount of energy.
- Electrons are found within energy levels surrounding the nucleus.
- Electrons can jump from low to high energy levels after absorbing some form of activation energy; however, they quickly jump back again releasing energy in the form of waves.
- Energy levels that are farther away from the nucleus have the highest amount of energy.
- And most importantly…

Everything in the universe tends to move towards the lowest energy possible.

This week we need to create an image in your head of how these energy levels work. Our focus will be primarily on where the electrons may or may not be found around the atom. This may seem like an easy task; however, nothing could be farther from the truth. In fact, it is very difficult to determine exactly where an electron is while it is buzzing around its energy level!

Electrons do not orbit the nucleus like the Earth orbits the sun. It is important that you understand this fact because you don't want to get confused with another term scientists use to describe an energy level – an **orbital**. If you can imagine an atom being a stadium, the rows of seats would be the orbitals. And just like any stadium, there are a limited number of seats that can be found within each row.

The following chart will show you the names and the maximum amount of electrons that can "fill up each seat" within the row:

Name of the orbital	Maximum numbers of electrons the orbital can hold
1s	2
2s	2
2p	6
3s	2
3p	6
4s	2
3d	10
4p	6
5s	2
4d	10
5p	6
6s	2
4f	14
5d	10
6p	6
7s	2
5f	14
6d	10
7p	6

If you are looking for a pattern in these numbers, you should easily see the following:

s orbitals = 2 electrons
p orbitals = 6 electrons
d orbitals = 10 electrons
f orbitals = 14 electrons

The reason for the s, p, d, and f titles is beyond the scope of this book as it deals with a branch of science called **quantum mechanics**. However, I will tell you that these letters signify different ways scientists have quantified (measured) the movements of atoms within these energy levels.

Each element has its own unique **electron configuration**, which is a method of describing these orbitals for us to study. For example, the element helium has two protons and two electrons. The electron configuration for helium would be:

$$1s^2$$

The superscript "2" is used to identify how many electrons are within this orbital. Another example would be the element beryllium which has four protons and four electrons. The electron configuration for beryllium would be:

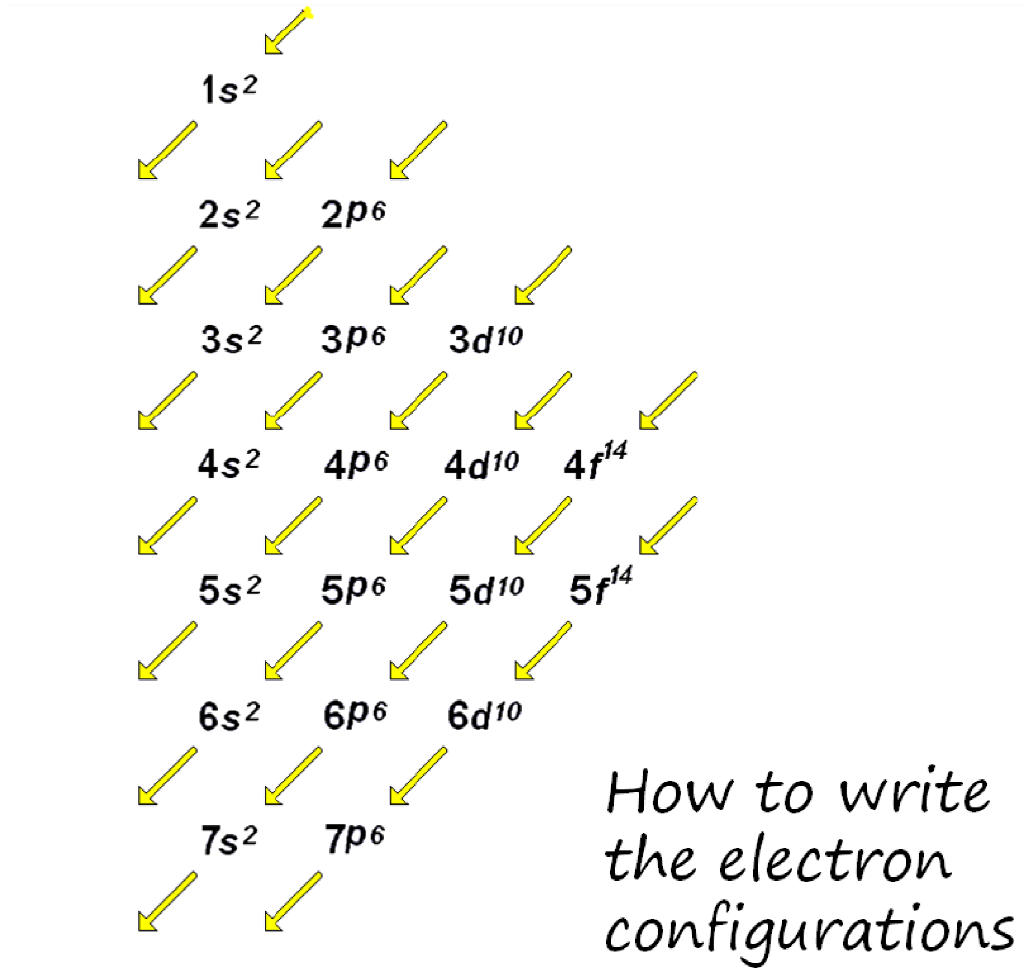
$$1s^2 2s^2$$

You might be asking yourself, what about elements with odd numbered electrons? Take, for example, the element lithium which has three protons and three electrons. Its electron configuration would be:

$$1s^2 2s^1$$

Remember – the superscript tells you how many electrons are found within the orbital. If the orbital is not filled up, you simply write out how many electrons can be found in that energy level. Easy, right?

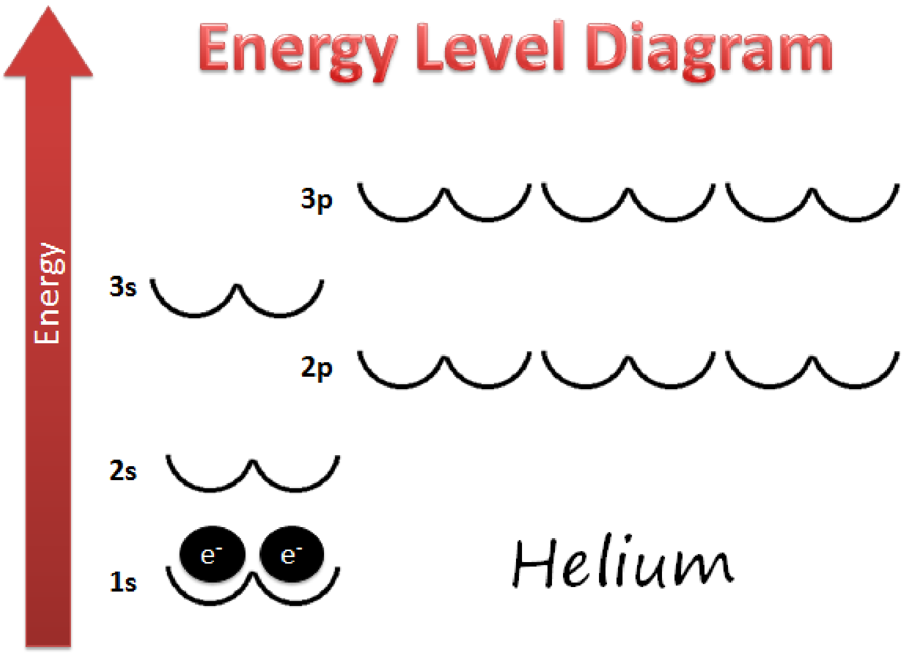
Now instead of memorizing the order of these orbital names (that would be a little tricky), let me show you a map you can use to follow through on your practice questions at the end of this chapter.



By following the arrows in the diagram, you can easily write out the electron configurations for every possible element you will study!

There is another way to visualize these orbitals which may help you "see" how these energy levels work. This tool is known as the **energy level diagram.**

Let's use the three former examples to visualize how energy level diagram can help you fill up each orbital. Helium is first in line:



The half-circles found within each level represent the maximum number of “seats” that could possibly be filled with electrons. In the case of helium, only two electrons are present so they fill up the first energy level (1s).

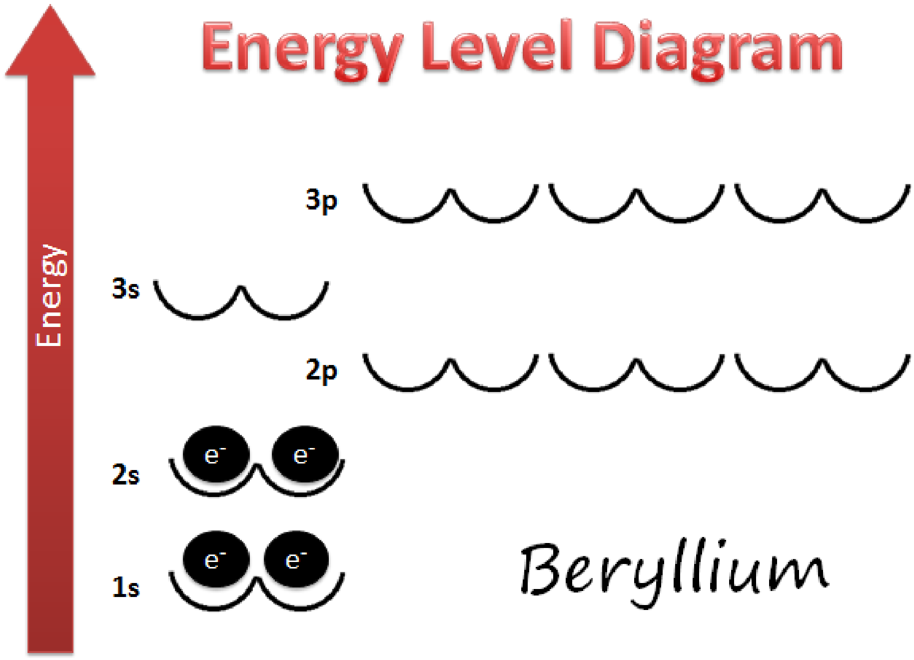
Why don’t these electrons fill up the 3p energy level on top of the diagram?

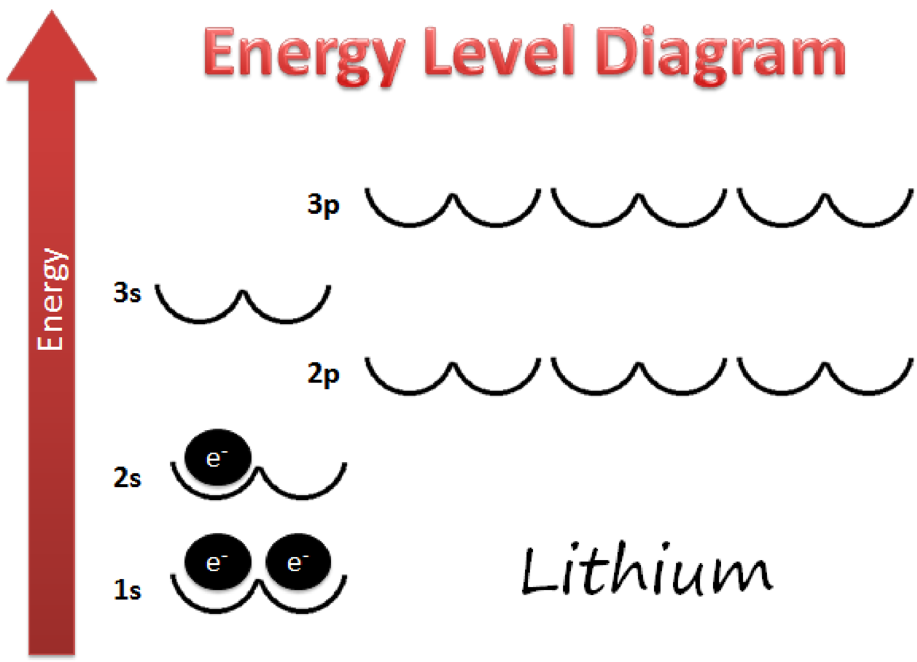
Remember, the fundamental rule from last week:

Everything in the universe tends to move towards the lowest energy possible.

Since the 1s orbital is closest to the nucleus, it has the lowest possible energy out of all the orbitals! The 1s orbital has to be filled up first, followed by the 2s, then the 2p and so on.

Let's look at beryllium and lithium's energy level diagrams:





*You might be thinking why this energy level diagram stops at the 3p orbital. Trust me, it only stops here because we'd quickly run out of room if I put all of the orbital's within this diagram!

If you ever get a little confused as to how or why the energy levels are so important, try to remember this little trick.

Go into the kitchen and fill up a tall drinking glass about half-full with ice cubes. Now I have to ask you, which ice cubes have the lowest amount of energy?

The ice cubes on the bottom of the glass have the lowest energy because they are closest to the Earth!

Therefore, the ice cubes on top have the most energy. But what about the accessibility of the ice cubes? Which ice cubes are easier to reach, the ones on top or those on bottom?

The ice cubes on the top are more accessible!

This is very similar to what goes on within an atom. These electrons are more accessible to other atoms and are known as **valence electrons.** The outer orbitals with their valence electrons and possible "empty seats" are very important in the bonding of atoms. But we'll look at that concept in later chapters.

You may not realize it yet, but you have nearly all of the tools you need to understand the one, the only...

Periodic Table

Don't worry. You don't have to start memorizing all of the information within this amazing resource. Once you understand how the table is put together and why it works in this manner, identifying the individual elements will be a breeze! You begin your journey next week. See you soon!

Electron Configurations Practice

In the space below, write the electron configurations of the following elements:

1) sodium
____

2) magnesium
____

3) iron
____

4) potassium
____

5) selenium
____

Determine what elements are denoted by the following electron configurations:

6) $1s^2\ 2s^2 2p^6 3s^2 3p^4$ ______________________

7) $1s^2\ 2s^2 2p^6 3s^2 3p^1$ ______________________

8) $1s^2\ 2s^2 2p^6$ ______________________

9) $1s^2\ 2s^2 2p^6 3s^2 3p^5$ ______________________

Explain what is wrong with the following electron configurations:

10) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^6$

11) $1s^2 2s^2 2p^6 3s^3 3d^5$

12) $1s^2 2s^2 2p^6 3s^1 3p^1 3s$

13) $1s^2 2p^6 2s^2 3s^2 3p^6$

Use the following clues to identify the element. Show any figuring in the space below:

14) This element has a 3p sublevel that contains 3 electrons.

15) This element has a 4s sublevel with 2 electrons for its outermost electrons.

16) This element has 1 electron in its 3d sublevel.

17) This element has 5 electrons in its 5p sublevel

18) This element has a completely filled 3p sublevel for its outermost electrons.

19) This element has 2 electrons in its 6p sublevel.

Write out the electron configurations for the atoms which contain the following numbers of electrons:

20) 10 ____________________________________

21) 24 ____________________________________

22) 13 ____________________________________

23) 3 ____________________________________

24) 5 ____________________________________

Chapter 9

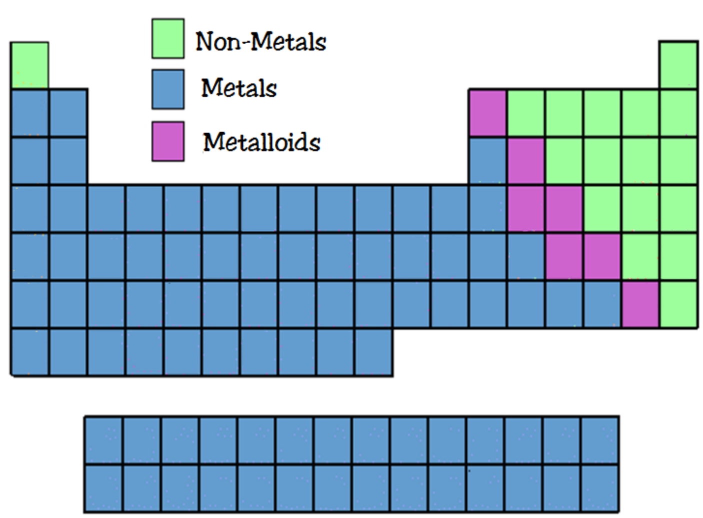
This chapter may seem to be a little different than the others you have studied so far. In this chapter you will explore several of the broad characteristics that identify elements within the periodic table. I wouldn't be surprised at all if you are constantly flipping back to this chapter as a reference for your future studies. **That is good!** You may also want to flip forward too as you will find an actual periodic table of the elements at the end of this chapter.

As I said last week, there's no need to start memorizing the information within this resource if you understand **HOW** it is put together. And that is what we are going to do right now!

Let's start off by identifying the three main categories of elements within the periodic table:

Non-Metals, Metals, and Metalloids

Non-Metals



Generally speaking, most non-metals are brittle, poor conductors of heat and electricity, have a very low density, and have lower melting and boiling points. In fact, most non-metals are <u>gases</u> at room temperature!

It is very difficult to describe groups of elements with the words "all" or "every" because there are always exceptions to these rules!

*Yes, I know the irony of that last statement, but it is true. There are ALWAYS exceptions!

Metals

Metals make up the majority of all the elements as you can see from the picture. Most metals (not "all" or "every") are hard, shiny, malleable (able to be pounded into sheets), ductile (can be drawn into wires), good conductors of heat and electricity, and have high densities, melting points, and boiling points.

Metalloids

Metalloids are stuck between the metals and non-metals and share the properties of both these types of elements. Most have high melting points, boiling points, and densities like metals. They also tend to be good conductors of heat and electricity like metals. However, most metalloids are brittle like non-metals. Again, there are always exceptions to these characteristics.

How to "look" at the periodic table

At first glance, all of the information you see on a typical periodic table can be intimidating. But if you know HOW it is put together, I know you will not be afraid of it at all!

By far, the easiest pattern to observe within the periodic table is the sequential numbering of the elements within each box (i.e. 1, 2, 3, 4…) Check this out on the periodic table located at the end of this chapter. These are the atomic numbers of the elements. You learned in Chapter 5 that the atomic number of an element is the number of protons found within its nucleus.

In addition to being placed in sequential order of their number of protons, the elements are also arranged in horizontal rows called **periods** and vertical columns called **families** or **groups**.

Elements within each period share few (if any) properties with each other with one small exception:

Elements in the same period have valence electrons in the same energy levels (orbitals) as each other.

This may sound important, but it really doesn't affect their properties too much. However, valence electrons are a VERY important factor within the elements of each family.

Families of elements within each column tend to have very similar physical and chemical properties. What causes this to happen?

Elements within each family have similar electron configurations!

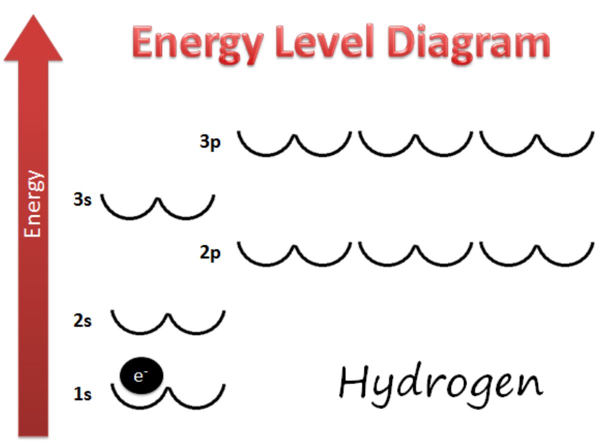
Let's look at a few examples of electron configurations in the first family to see what I'm talking about here:

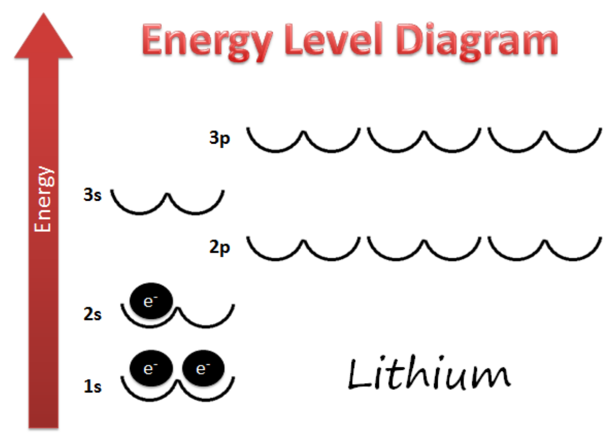
Element	Electron configurations
Hydrogen	$1s^1$
Lithium	$1s^2 2s^1$
Sodium	$1s^2 2s^2 2p^6 3s^1$

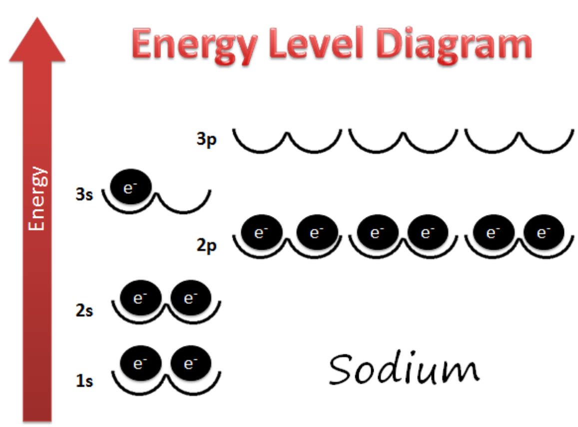
Do you see a pattern with these elements? If you said that they all end with one electron in their outermost s orbital you are absolutely correct!

Elements within each family all share the same number of valence electrons!

In case you need a more visual way of seeing how this works, here are the energy level diagrams for each of these elements:







All of these elements have one valence electron, as do the remaining elements within the first family of the periodic table. In all, there are eight families that we will be focusing upon throughout most of the book. These elements make up what are called the **main block elements**.

Family Name	Column number
Alkali Metals	1
Alkaline Earth Metals	2
Boron Family	13
Carbon Family	14
Nitrogen Family	15
Oxygen Family	16
Halogens	17
Noble Gases	18

Alkali metals

The alkali metals are not the type of elements you want to mess around with in their purest forms. Here are some of their properties:

- Very reactive with other elements (as we will see in future chapters.)
- Flammable in air and water
- Low melting and boiling points
- Soft
- Low density
- 1 valence electron

Alkaline earth metals

The properties of these guys are similar to the alkali metals, but they're not as extreme! For example, the alkaline earth metals are reactive with other elements – but not as reactive as the alkali metals. Their remaining properties follow in the same manner:

- Alkaline earth metals react in air and water (but not as violently as the alkali metals.)
- Alkaline earth metals have low melting and boiling points (but they are higher than the alkali metals.)
- Alkaline earth metals are soft (but harder than – you guessed it, the alkali metals.)
- And alkaline earth metals have low densities (but higher than those of the alkali metals.)
- Alkaline earth metals all have 2 valence electrons

Boron family

- Most members of this family are metals with the exception of boron itself, which is a metalloid
- Elements in the boron family have 3 valence electrons

Carbon family

- Members of the carbon family are made up of metals, metalloids, and non-metals (of which carbon is a non-metal)
- Elements in the carbon family have 4 valence electrons

Nitrogen family

- Members of the nitrogen family are made up of metals, metalloids, and non-metals (nitrogen itself is a non-metal)
- Elements in the nitrogen family have 5 valence electrons

Oxygen family

- Members of the oxygen family are made up of metals, metalloids, and non-metals (of which non-metal is a non-metal)
- This family is quite reactive with other elements (their reactivity is comparable to the alkaline earth metals)
- Elements in the oxygen family have 6 valence electrons

Halogens

- All halogens are non-metals
- These elements are very reactive (much like the alkali metals)
- Halogens have 7 valence electrons

Noble gases

- Noble gases are the least reactive elements on the periodic table
- Noble gases have 8 valence electrons

What about the elements between these eight families?

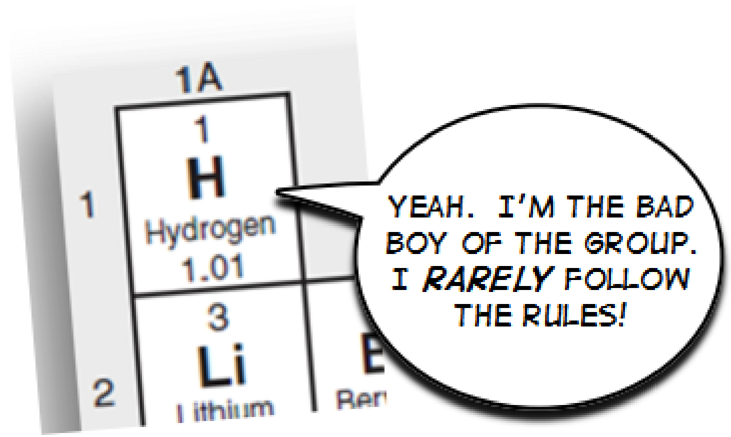
The elements between the eight families we have just looked at are called the transition elements. The elements within this group share many characteristics with each other; however, there are so many different rules that could be listed about each of them, it will be best to describe them individually in future chapters.

Despite their varying properties, most transition elements tend to be unreactive, hard, and maintain relatively high melting and boiling points.

There is one LARGE exception to the rule...

Hydrogen

Hydrogen is the first element on the periodic table. Although it is located in the first column, it is not an alkali metal! Hydrogen has properties unlike any other element on the periodic table. Unlike any other element, hydrogen can bond with most elements in the periodic table. It also plays a key role within all acids and bases as you will learn. I guarantee that you will become very well acquainted with hydrogen in the near future.



As you have discovered, families, periods, and atomic numbers are three patterns that exist within the periodic table. Next week, we are going to look at a few more trends that exist within this amazing resource. See you then!

Periodic Table of the Elements

Key

11 — Atomic number
Na — Element symbol
Sodium — Element name
22.99 — Average atomic mass

	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1	1 **H** Hydrogen 1.01																	2 **He** Helium 4.00
2	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
3	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
4	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
5	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
6	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)
7	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)									

Lanthanides	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
Actinides	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)

Periodic table practice

For this activity, you will need some markers or crayons!

1) Define these terms, and explain what elements within both of the following categories have in common with each other:

   - family/group:

   - period:

2) Identify where each of the following sections of the periodic table are and identify the main properties of the elements in these areas:

   - alkali metals:

   - alkaline earth metals:

   - transition metals:

   - lanthanides:

   - actinides:

   - halogens:

   - noble gases:

3) Color each of the groups in problem 3 on the blank periodic table.

Homemade Periodic Table

Identify twenty or so items within your home. As scientists, it's your job to group these items into families and periods, based on their observable properties. Your periodic table must fit within the grid given below!

1) What was the main property you used to classify the elements into groups? Explain why you chose this property and not another.

2) What did the elements in each period have in common with one another? How did his property change as you moved down each period of your periodic table? Explain how you decided which element should go at the top of each column and which should go at the bottom.

3) How does this exercise give you any insight as to why it may have been difficult to invent the first periodic table? Explain.

4) Your periodic table contains five blank boxes. Based on the properties and characteristics you used when arranging your periodic table, what objects do you believe would fit well in those blank spaces? Explain your answer.

Chapter 10

The vertical families, horizontal periods, and numerical ordering by atomic numbers all identify patterns found within the periodic table. But these are not the only patterns that can be observed! With a little imagination and a few well-known facts, you will be able to see three more trends in its structure. These trends are:

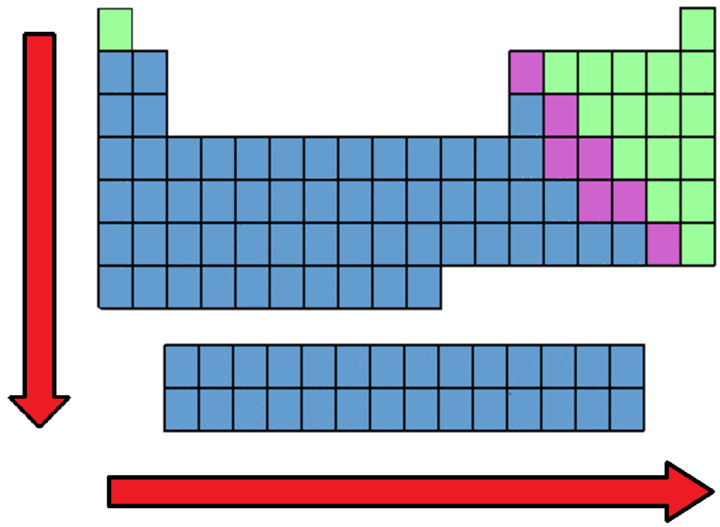
Atomic radius
Ionic radius
and Electronegativity

The radius of an object (most likely a spherical object) is the distance from its centermost point to the outside wall of its structure. Imagine a line being drawn from the core of an apple to its skin.

Although we have never actually seen an atom, it is highly unlikely that its shape is that of a perfect sphere – especially with electrons buzzing around its nucleus! Nevertheless, the **atomic radius** is a measurement of the relative length of an atom from its nucleus to its outermost orbital.

Atomic radius trends

From left to right, the atomic radius decreases; and, from up to down it increases.



This may need a little explanation, so let's look at both of these trends in a little more detail.

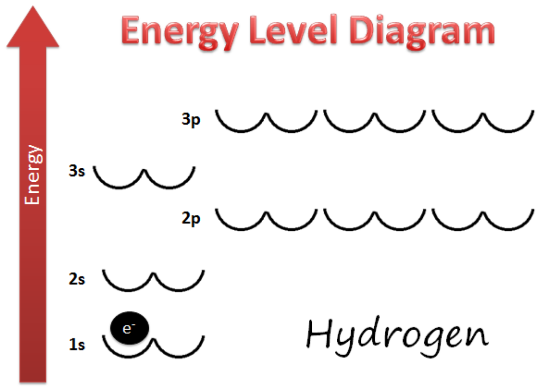
"From left to right, the atomic radius decreases."

This may seem a little peculiar. If the number of protons in each element is increasing from left to right, you may think that its size would increase as well. But this is not the case.

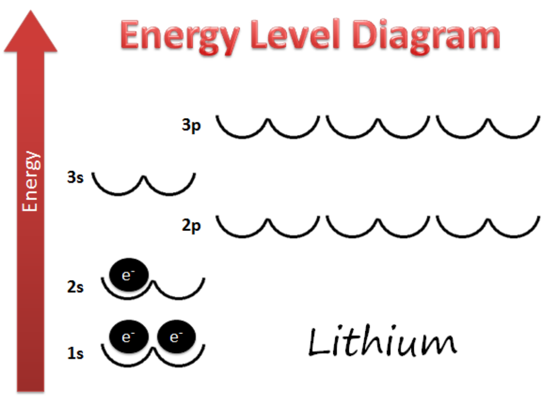
As you increase the number of protons in each element, their nuclei becomes more positively charged. This increased charge adds to the **electron affinity** of each element. Basically, the increased number of positively charged particles pulls the electrons a little closer to the nucleus, causing the overall size of the atom (atomic radius) to decrease a little bit.

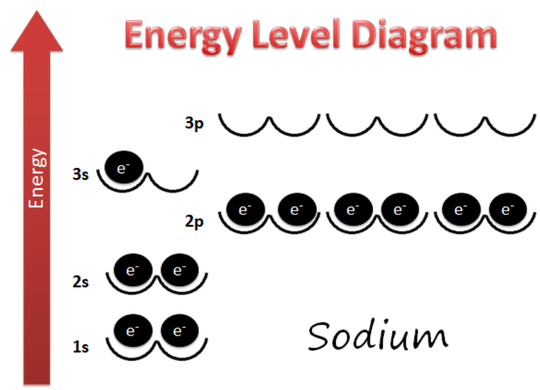
"From up to down, the atomic radius increases."

If you recall our study of the energy level diagrams from last week's reading, the elements within each family contain the same number of electrons in their outermost orbital. However, as you move from element to element down one period, the number of energy levels increases!



As you move down through the alkali metals, hydrogen's outermost electrons are in the 1s orbital, lithium's are found within the 2s orbital, and sodium's are within the 3s orbital. This pattern suggests that as the number of orbitals of an element increases, so does its atomic radius.





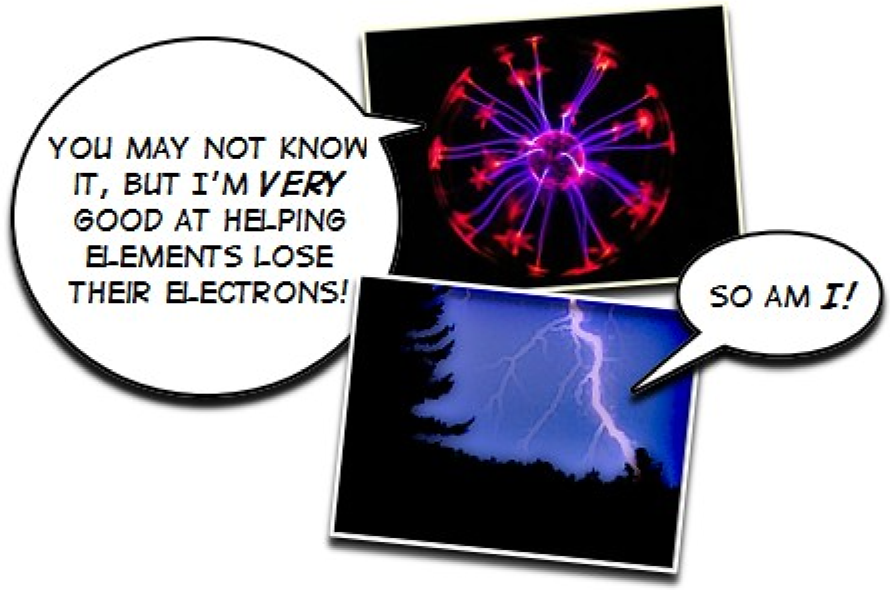
Ionic radius trends

Back in Chapter 5 you learned that it was safe to assume that an atom contains an equal number of protons and electrons. This is true with atoms that have a neutral charge; however, atoms do not always remain neutral. Sometimes they gain or lose one or more electrons in order to satisfy their #1 rule:

Everything in the universe tends to move towards the lowest energy possible.

Sound familiar? I thought so.

When an atom gains or loses one or more electrons it is called an **ion**. When an atom loses an electron, it removes some of its negative charge from the atom and leaves a more positive charge behind. This positively charged ion is known as a **cation**.



When the opposite happens, and an electron gains one or more electrons, it becomes more negatively charged and is then called an **anion**.

The relationship between cations and anions are very important when we start binding elements together.

Okay... back to the trends within the periodic table!

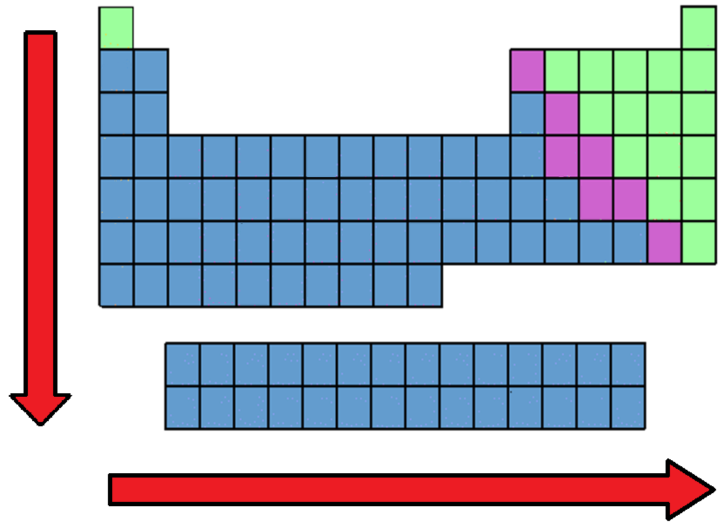
Cations have a smaller ionic radius than the atoms in which they were created because their fewer number of electrons takes up less space.

Anions, on the other hand, have gained extra electrons which causes their ionic radius to increase.

Electronegativity trends

Since we are starting to look at ions and how they can bind together, it is important to see another trend within the periodic table called **electronegativity**.

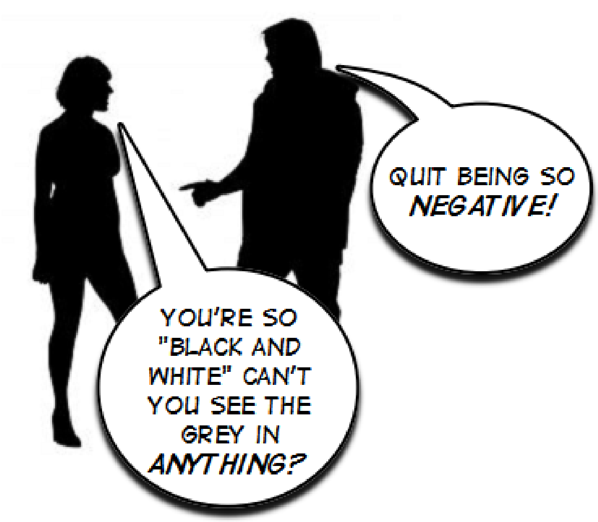
Electronegativity is a measurement of how much an atom can pull electrons away from other atoms. The higher the difference in electronegativity between two atoms, the more likely that electrons will be pulled away from each other.



If you look at the periodic table, the following trend can be seen:

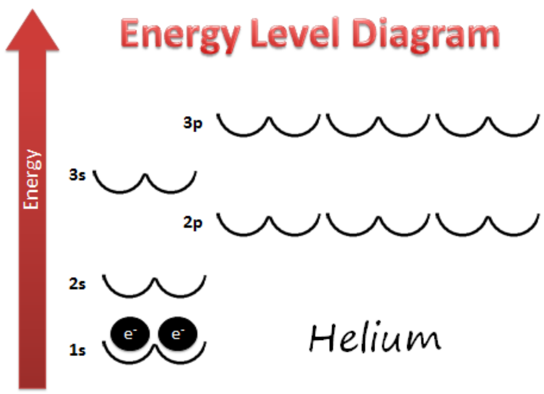
From left to right, the electronegativity increases; and, from top to bottom it decreases.

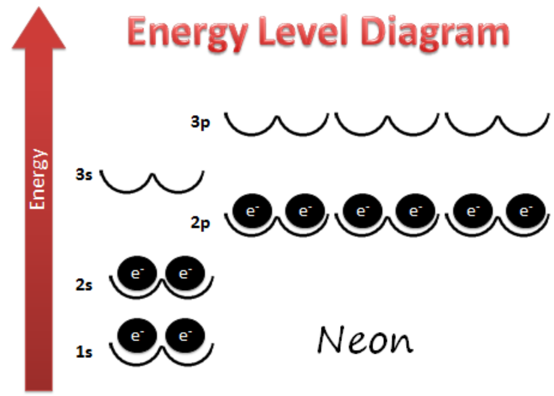
The only exceptions to the rule in this trend can be found within the noble gases – they have no electronegativity at all. Strange? Not at all! You see, there's another very important rule for you to remember:

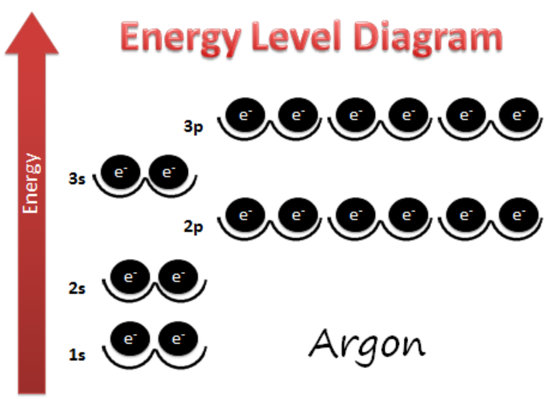


All elements on the periodic table want to be like the noble gases.

What's so special about the noble gases? Why are they so cool? Well, in order to answer that question you need to look at the energy level diagrams for a couple of these guys.







Helium, neon, and argon are the first three noble gases on the periodic table. Take a look at their energy level diagrams. What patterns do you see?

That's right! All of their outermost energy levels are completely filled up with electrons! There are no empty seats in any orbital.

Without going to into too much detail, whenever an orbital is filled with electrons, the atom is considered to be very stable and in the lowest possible energy situation it can reach.

This is why every element wants to be like a noble gas!

Scientists have created a rule for this phenomenon called the **octet rule**. The "oct-" prefix refers to the number eight which just happens to be how may valence electrons are in every noble gas.

(With the exception of helium which only has two valence electrons – I told you there are always exceptions to the rule!)

If you are counting the valence electrons within our examples above, you may be having trouble counting their number of valence electrons in neon and argon. No worries! Simply remember that we count the number of electrons within the outermost s and p orbitals to get the amount of valence electrons!

Neon has two electrons in its 2s orbital and 6 electrons in its 2p orbital. This adds up to eight electrons!

Argon has two electrons in its 3s orbital and 6 electrons in its 3p orbital. This adds up to eight electrons as well!

Now that you understand what electronegativity is, let's look back at its trends:

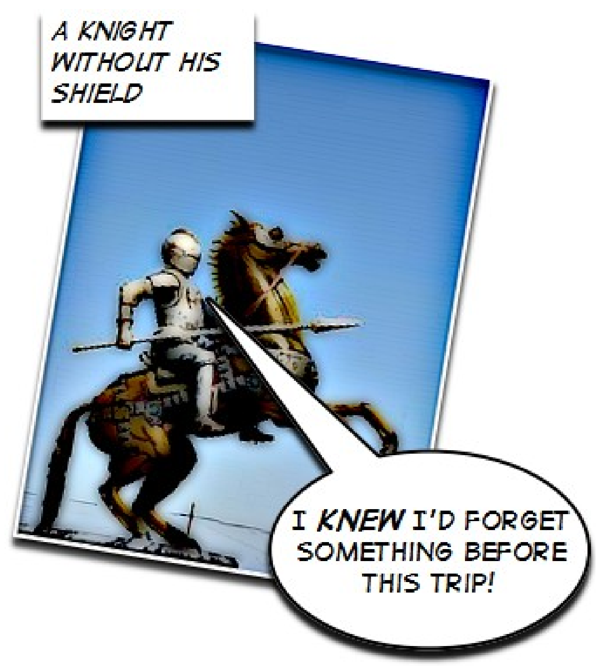
"From left to right, electronegativity increases."

This should make sense to you now that you understand why elements want to be like noble gases. Elements on the right side of the periodic table want to gain electrons in order to be like the noble gases; meanwhile, elements on the left side are trying to give away their electrons to be like noble gases.

Its much easier for the alkai metals, alkaline earth metals, and the boron family elements to give away their electrons to have a filled outer orbital. The nitrogen and oxygen families along with the halogens are attempting to take as many electrons as they can to fill up their outermost orbital shell and be like the noble gases!

"From top to bottom, electronegativity decreases."

This trend can be explained due to what is known as the **shielding effect**. Valence electrons are found in the outermost areas of the atom. Their negative charge is attracted to the strong positive charge within its nucleus. However, the farther down the periodic table you go, more electron orbitals are added as well. The distance between the valence electrons and the nucleus grows farther. As this happens, the increased numbers of negatively charged electrons that have filled up the inner orbitals of the atom begin to repel the valence electrons.



In essence, the inner electrons act as a shield between the nucleus and the valence electrons. This shielding effect lowers the electronegativity of the atom by reducing its ability to pull electrons away from other atoms.

All of these trends from the past two chapters are vital to your understanding of how atoms bond together. In the next unit, you will be looking at how ions play a role in the complicated world of atomic bonding!

Periodic trends practice

Electronegativity chart of the Elements

1 **H** 2.1																	2 **He**
3 **Li** 1.0	4 **Be** 1.5											5 **B** 2.0	6 **C** 2.5	7 **N** 3.0	8 **O** 3.5	9 **F** 4.0	10 **Ne**
11 **Na** 0.9	12 **Mg** 1.2											13 **Al** 1.5	14 **Si** 1.8	15 **P** 2.1	16 **S** 2.5	17 **Cl** 3.0	18 **Ar**
19 **K** 0.8	20 **Ca** 1.0	21 **Sc** 1.3	22 **Ti** 1.5	23 **V** 1.6	24 **Cr** 1.6	25 **Mn** 1.5	26 **Fe** 1.8	27 **Co** 1.9	28 **Ni** 1.9	29 **Cu** 1.9	30 **Zn** 1.6	31 **Ga** 1.6	32 **Ge** 1.8	33 **As** 2.0	34 **Se** 2.4	35 **Br** 2.8	36 **Kr** 3.0
37 **Rb** 0.8	38 **Sr** 1.0	39 **Y** 1.2	40 **Zr** 1.4	41 **Nb** 1.6	42 **Mo** 1.8	43 **Tc** 1.9	44 **Ru** 2.2	45 **Rh** 2.2	46 **Pd** 2.2	47 **Ag** 1.9	48 **Cd** 1.7	49 **In** 1.7	50 **Sn** 1.8	51 **Sb** 1.9	52 **Te** 2.1	53 **I** 2.5	54 **Xe** 2.6
55 **Cs** 0.7	56 **Ba** 0.9	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.9	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn** 2.4
87 **Fr** 0.7	88 **Ra** 0.9	89 **Ac** 1.1	104 **Rf**	105 **Ha**	106 **Sg**	107 **Ns**	108 **Hs**	109 **Mt**	110 **Uun**	111 **Uuu**	112 **Uub**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

1) In each of the following pairs, circle the species with the <u>larger</u> atomic radius:

a) Mg or Ba

b) S or S^{-2}

c) Cu^{+2} or Cu

d) He or H^{-}

e) Na or Cl

2) Circle the best choice in each list:

a) largest radius: S^{-2} or Cl

b) highest electronegativity: As, Sn, S

c) smallest radius: Na, Li, Be

3) Rank the following elements by increasing atomic radius:
carbon, aluminum, oxygen, potassium

4) Rank the following elements by increasing electronegativity:
sulfur, oxygen, neon, aluminum

5) Why do elements in the same family generally have similar properties?

6) Within a period, does the size of the atomic radius increase or decrease with increasing atomic number?

7) Within a family, does the size of the atomic radius increase or decrease with increasing atomic number?

8) When metallic atoms lose electrons do they form smaller or larger ions?

9) When nonmetallic atoms gain electrons do they form smaller or larger ions?

10) From each of the following pairs, circle the atom with the largest radius:
   a) Na or Li
   b) Br or I
   c) Cs or Ba
   d) Ne or Ar

11) List the three lightest members of the noble gases.

12) Within a group, what happens to the atomic radius as you go down the column? Why does that happen?

13) Within a period, what happens to the atomic radius as the atomic number increases? Why does this occur?

14) How are the shielding effect and the size of the atomic radius related?

15) How are neutral atoms converted into cations?

16) How are neutral atoms converted into anions?

17) When an atom becomes an anion, what happens to its radius?

18) When an atom becomes a cation, what happens to its radius?

Chapter II

If you recall from last week, an ion is the name of an atom that has gained or lost one or more electrons. Cations are positively charged ions which have lost one or more electrons and anions are negatively charged ions that have gained one or more electrons.

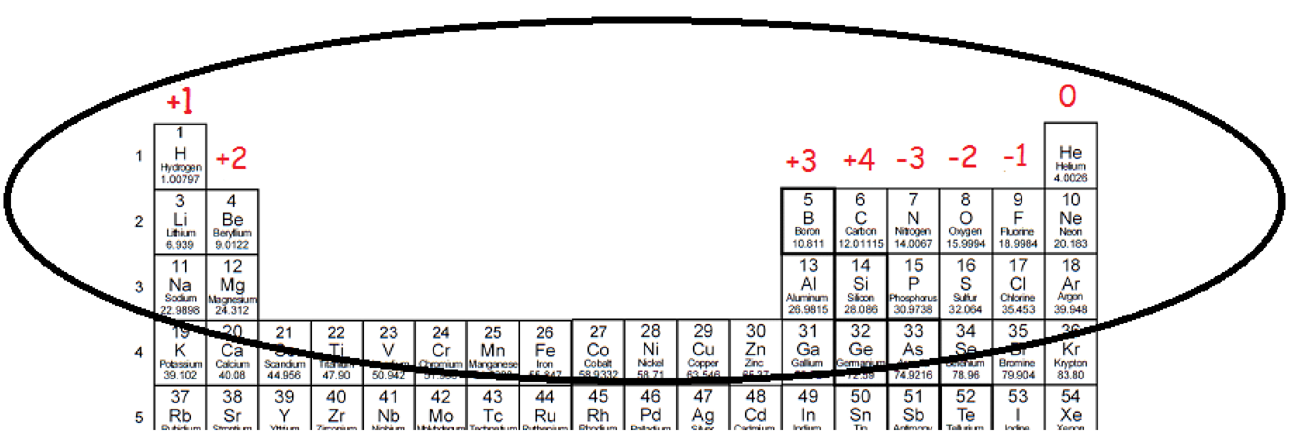
This week, we are going to use this information to learn how cations and anions bond together. It shouldn't be much of a surprise when you learn that the name of this type of bond is...

The Ionic Bond

I know the name isn't very original, but it gets the job done – right?

Before we begin our study of ionic bonds, we need to look at the periodic table to find yet another trend that exists within its families. All eight of the families we have explored have the potential to create ions. And with a few exceptions, we can calculate how many electrons each family's element will give away or take from other elements.

Within the first 18 elements of the periodic table, the trend for how many electrons are given away or taken can be found within this diagram.



Lithium (atomic number 3) is within the first family and can give away one electron. As it gives away one negatively-charged particle, it becomes an ion with a +1 charge. The proper way to write out the lithium ion is:

$$Li^{+1}$$

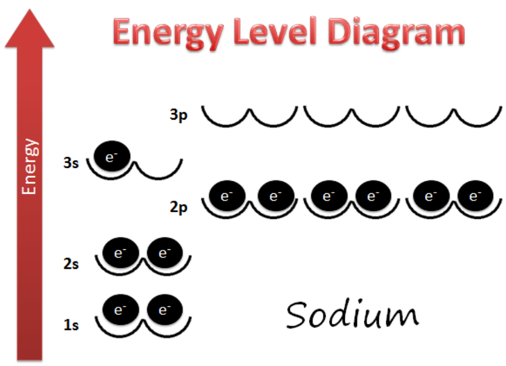
Oxygen (atomic number 8) will receive two electrons. As it receives two negatively-charged particles, it becomes an ion with a -2 charge. You would write its ion like this:

$$O^{-2}$$

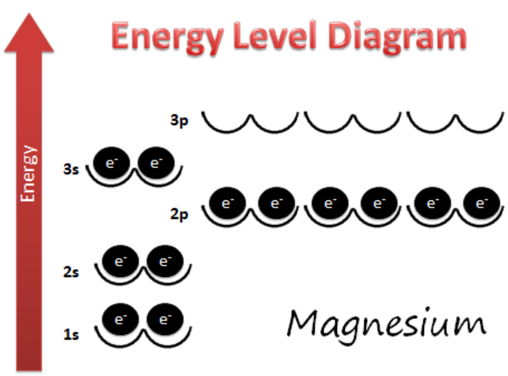
I believe it is best you understand WHY some elements prefer to "pitch" their electrons while others tend to "catch" them. To be honest, you should already know why these elements pitch and catch their electrons. Do you remember this little phrase…

All elements on the periodic table want to be like the noble gases.

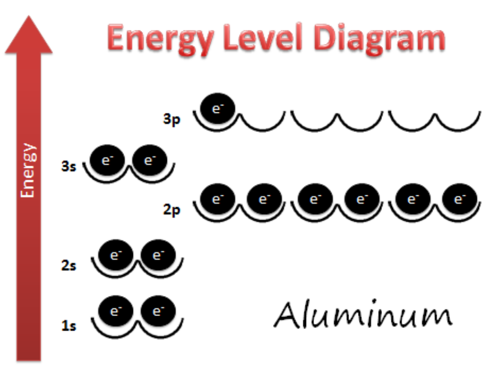
With the exception of helium, all noble gases have filled their outer orbitals with eight valence electrons (remember the octet rule!) Therefore, in order to achieve the lowest energy state possible, all of the elements want to have their outermost orbital filled up! Let's look at the energy level diagrams of a few elements within the third period in the periodic table. I believe this will help you identify our pitchers and catchers.



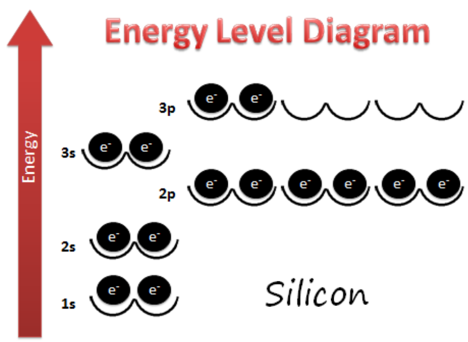
The sodium ion (Na^{+1}) has one empty seat in its 3s orbital. It could gain one electron and fill up its 3s orbital; or, it could lose one electron and allow its 2p orbital to be the outermost orbital. Since a filled 2p orbital is closer to the nucleus than a filled 3s orbital, sodium will "pitch" its electron to another atom.



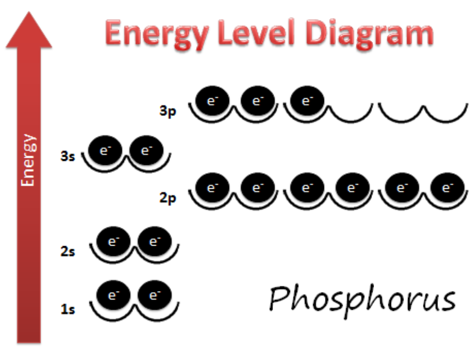
Magnesium has two valence electrons and needs six more (in the 3p orbital) to fill up its eight possible seats. It would be much easier to give away both of its electrons in the 3s orbital, thus creating a magnesium ion (Mg^{+2}).



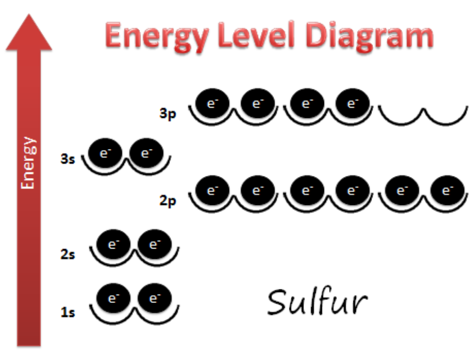
Aluminum has three valence electrons (remember that the valence electrons is the SUM of the outermost electrons in the s and p orbitals.) Much like magnesium and sodium, aluminum will pitch three of its electrons to create an aluminum ion Al^{+3} because pitching three electrons is easier than catching five to fill up its s and p orbitals.



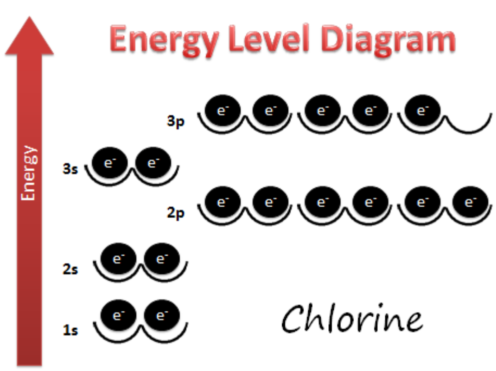
Silicon is in the carbon family and it tends to be caught in the middle of being a pitcher and catcher. It could either pitch OR catch four electrons to find its lowest energy state. This means its ion charge could be either Si^{+4} or Si^{-4} depending on what it is binding to.



Phosphorus is a nonmetal and, as you can tell from the diagram, it would require less energy to acquire three electrons than to pitch five. Therefore, phosphorus will gain three electrons to create the phosphorus ion P^{-3} and fill up its 3p orbital.



Sulfur is in the same position as phosphorus. The only difference is that it only needs two electrons to fill all eight of its valence seats. Sulphurs ion is S^{-2}.



Chlorine is a halogen, and like all halogens it only needs one electron to fill its outer orbital. This makes chlorine very reactive because it only needs one electron to reach its lowest energy state. With one more electron, all of its eight valence seats will be filled up; therefore, chlorine's ion is Cl^{-1}.

Naturally, argon (our noble gas in this period) has not been listed in these previous two pages. All of its s and p orbitals are filled up, making it the most stable element in this period. One general rule of thumb to remember within these ion charges is this:

Metals are pitchers and become cations (+)

Non-metals are catchers and become anions (-)

With a little imagination, I think you can figure out how cations and anions bind together to form molecules. Sometimes one of our pitchers gets close to one of the catchers and they exchange one or more electrons. When this happens, both of the atoms become ions and have opposite charges. Since opposite charges naturally attract each other, the two ions bind together.

This is how ionic bonds are formed!

In order for the ions to be bound together, they must have equal and opposite charges. For example, let's look at sodium and chlorine. When these two atoms get too close to each other, chlorine pulls one of sodium's electrons away forming a sodium ion (Na^{+1}) and a chlorine ion (Cl^{-1}).

Since these two ions have equal charges, they bind together to form NaCl.

$$Na^{+1} + Cl^{-1} = NaCl$$

(NaCl is the chemical formula of table salt.)

You need to notice two things here: First, once the two ions bind together, you do not need to write out the ion charges anymore; second, the cation is always written first within the chemical formula.

Don't forget! Ionic bonds are made between metals and non-metals AND their ionic charges must be equal! Let's look at an unequal example to give you some practice:

Combining magnesium with chlorine will not be as easy as it is with sodium? Why? Because their charges are not equal! Therefore, you have to bind two chlorines with one magnesium ion to balance the charges on both sides.

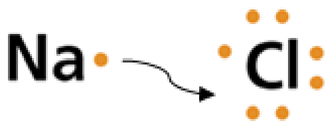
$$Mg^{+2} + Cl^{-1} + Cl^{-1} = MgCl_2$$

Like I've said before, sometimes you need to SEE how something works with pictures in order to help understand a certain concept. Therefore, let me show you how ionic bonding occurs using another little tool called the **electron dot structure**.

The electron dot structures for each of the elements within the period we just studied would be:

Notice that the number of dots that surround each element is identical to the number of valence electrons they contain.

Let's see what happens when you take the electron dot structure for sodium and pair it with chlorine:



The first thing that happens is chlorine takes one electron away from sodium.

Ions are formed from the transfer of electrons. This sets up equal and opposite charges. In addition, chlorine now has a full octet of electrons – it now has the same electron configuration as a noble gas!

The opposite charges on both ions cause them to naturally attract to each other, forming the ionic bond.

One last thing before we take a break for the week! The products of an ionic bond are always known as **salts**. Table salt (NaCl) is a very common and well-known salt. $MgCl_2$ may not be known to a lot of people by its chemical name, but it is commonly used as a road de-icer in wintery weather. That's right! $MgCl_2$ is a very good type of rock salt.

You are not going to practice drawing too many electron dot structures this week. Typically, we do not draw these types of structures for ionic compounds; however, what you will learn about next week will give you all the practice time you'll need!

Good job! Ionic bonds are something you will encounter throughout the rest of the book. In order to feel comfortable with this type of bond, it's important that you get a lot of practice. Therefore, let's try out some of the practice problems for the week.

Ionic bonding practice

How many valence electrons do each of the following elements have?

1) carbon ____________________
2) selenium ____________________
3) xenon ____________________
4) potassium ____________________

Which of the following ions are likely to be formed? Circle one:

5) N^{+5}	yes	no	8) Al^{+2}	yes	no
6) He^{+}	yes	no	9) P^{-3}	yes	no
7) F^{-1}	yes	no	10) Mg^{+2}	yes	no

Write the formula which result from combining the following ions:

11) K^{+1} and Br^{-} ____________________

12) Li^{+1} and O^{-2} ____________________

13) Ca^{+2} and Cl^{-1} ____________________

14) Al^{+3} and S^{-2} ____________________

15) Fe^{+2} and S^{-2} ____________________

16) Na^{+1} and NO_3^{-1} ____________________

17) Na^{+1} and SO_4^{-2} ____________________

18) NH_4^{+1} and CO_3^{-2} ____________________

19) Fe^{+3} and CrO_4^{-2} ____________________

20) Al^{+3} and PO_4^{-3} ____________________

Determine the formulas and charges of the cations and anions in the following ionic compounds.

		Cation	Anion
21)	CsCl	________	________
22)	CaF_2	________	________
23)	NaOH	________	________
24)	$Ca(OH)_2$	________	________
25)	KNO_3	________	________
26)	$Fe(NO_3)_2$	________	________
27)	Fe_2O_3	________	________
28)	$Al_2(SO_4)_3$	________	________
29)	$CaCO_3$	________	________
30)	FeO	________	________

31) For each of the following pairs of elements, use the periodic table to decide the charge on both the cation and anion and determine the formula of the compound(s) formed in each case. When writing the formulas put the cation first.

Elements			Compound formula	
Mg	Br			
K	S			
Cl	Al			
S	Cu^{+1}	Cu^{+2}		
F	Zn^{+2}			
O	Co^{+2}	Co^{+3}		
aluminum	oxygen			
calcium	iodine			

32) Explain why oxygen is a fairly reactive element while neon is not.

33) Explain why beryllium loses electrons when forming ionic bonds, while sulfur gains electrons.

34) Explain why fluorine and chlorine have similar reactive qualities (the word "valence" should be somewhere in your answer!)

Chapter 12

Last week, you learned how metals and non-metals get together to form ionic bonds. But how do metals bind to each other? And what about nonmetals - do they bind to each other? Oh yes they do! These bonds are the topics of study this week and therefore I give you...

Covalent bonds and Metallic bonds

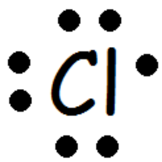
Let's look at **covalent bonds** first. What can you predict is going on within a "co-valent" bond? The prefix "co-" tends to mean some type of sharing is going on and the remaining "-valent" sounds a lot like the root word "valence." So if you are guessing that valence electrons will be shared within a covalent bond, you are absolutely correct!

Covalent bonds exist between two nonmetals. For example, hydrogen and chlorine are covalently bonded (yes... hydrogen is a non-metal) as are two chlorine atoms as well.

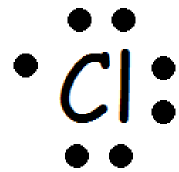
In covalent bonding, atoms still want to achieve the lowest energy level possible (which means filling up all eight of their valence seats just like a noble gas.) But rather than losing or gaining electrons, atoms now **SHARE** an electron pair.

Let's make this easier for you with some electron dot structures...

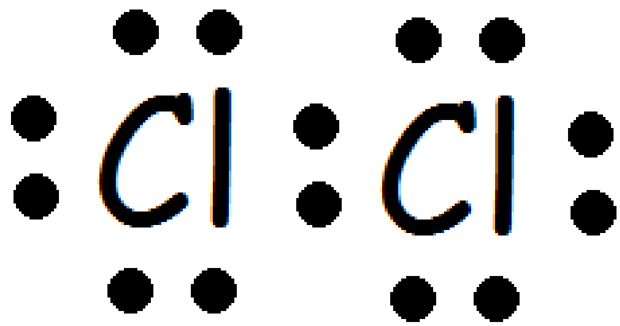
The electron dot structure of chlorine is below:



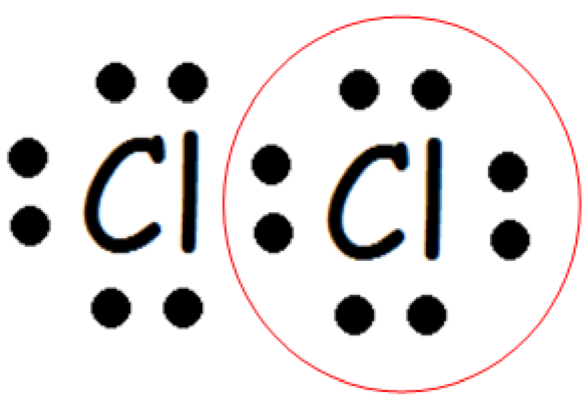
Each of these chlorine atoms wants to gain one electron to fill all of its eight seats.



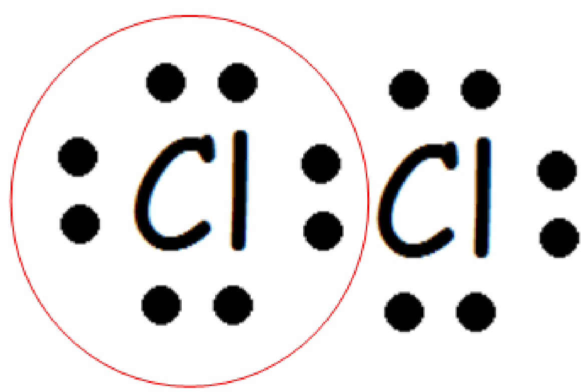
Therefore, if these two atoms get close together, they will each share one of their electrons an bind together by sharing an electron pair. Here's what it looks like:



You can see that both chlorine atoms fill up their eight valence seats thus fulfilling the octet rule:



Eight valence electrons



Eight valence electrons

The octets are achieved by each atom sharing the electron pair in the middle.

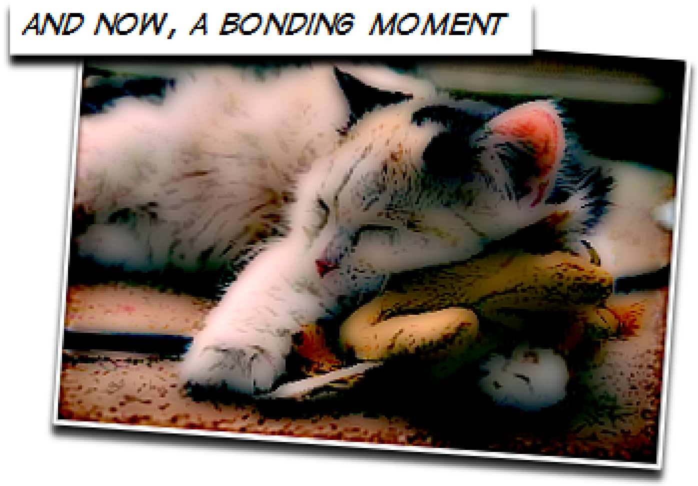
This shared pair is known as a **bonding pair** and since it is only made up of two electrons, it is also known as a **single bond**.

This single bond can also be abbreviated with a dash:

Chlorine forms a covalent bond with itself to form the molecule:

$$Cl-Cl$$

$$Cl_2$$

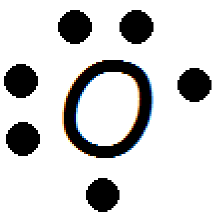


The bonding of chlorine to itself is an excellent example of how two atoms bind covalently within a single bond. However, it is possible for atoms to share more than one pairs of electrons. Double and triple bonds can be created as a precise number of bonds MUST be created as can be seen in the following table:

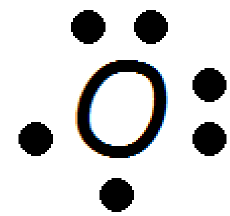
Family Name	Column number	Total number of bonds that must be formed
Alkali Metals	1	1
Alkaline Earth Metals	2	2
Boron Family	13	3
Carbon Family	14	4
Nitrogen Family	15	3
Oxygen Family	16	2
Halogens	17	1
Noble Gases	18	0

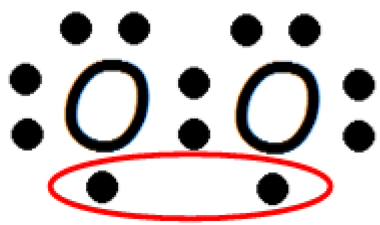
* Please note that a quadruple bond cannot exist within the carbon family. Each carbon atom can form four single bonds, two singles and a double, a triple and a single, or two double bonds.

Let's look at an example of a double bond with the formation of oxygen gas.

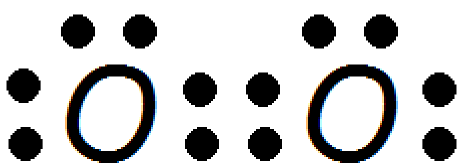


When two oxygen atoms get near to each other a single bond may first appear; however, neither of these atoms will have completed the octet rule as both have unshared electrons.





These two electrons need to be shared, therefore...



...a double bond is formed!

For convenience, the double bond may be shown as two dashes.

Always check the chart on the previous page for the total number of bonds that each element MUST create! This will help you identify where double and triple bonds will need to be located within your electron dot structures.

**The last two pages of this week's reading will help you draw these electron dot diagrams! Be certain to check it out!

Now it's time to look at the bonds that hold metals together!

Metallic bonds are much different than ionic or covalent. First of all, metals are pitchers – they give away their electrons. So how to positively charged particles bind together? This takes a little imagination...

Imagine an "island" of metal cations with their nuclei all stacked up like a collection of M&M candies. Now imagine this "island" being surrounded by a free-flowing , constantly moving "sea" of electrons. This "**sea of electrons**" is exactly what scientists call the force that holds the island of metal cations together. Much like the M&M's sealed in a baggie that can move around, the constant movement of the surrounding sea of electrons allows the cations to move around as well within the metal.

Before we move on, let's take a minute to review some of the concepts you have learned so far. This will help you to see some of the different properties of matter with ionic, covalent, and metallic bonding.

- A noble gas configuration is very stable and, therefore, is in the lowest energy state possible. Since noble gases have eight valence electrons, (nearly) all atoms in a compound need eight valence electrons to be equally stable. This is the reasoning behind the Octet Rule!
- All elements want to be like the nearest noble gas because everything in the universe prefers to be in a lowest possible energy state.
- In order to be like the nearest noble gas, atoms bond with each other to form chemical compounds.

Ionic bonds	Covalent bonds	Metallic bonds
Electrons are exchanged	Electrons are shared	Electrons move freely around nuclei
Ions are formed which "stick together" magnetically to form molecule	No ions are formed	Positively charged island of metal cations
Formed between metals and nonmetals	Formed between two nonmetals	Formed between two metals

The properties of matter that are made of ionic, covalent, and metallic bonds are all unique due to how they are held together. On the following pages, you will find a description of these common properties.

Ionic bond properties

Forms crystals (ordered arrangements of ions)

Cations and anions stack neatly together to minimize the distance between each other.

Conduct electricity when dissolved or melted

As solids, salts do not have the ability to move charged particles around which is needed to conduct electricity. However, in liquid form (called an **electrolyte**) the positive and negatively charged ions allow for the movement of electrons.

High melting/boiling points

With electrons being held together by strong nuclear forces between atoms, it takes a lot of heat to physically separate these atoms together to change their state.

Hard and brittle

The bonds that hold ionic compounds together are very strong which makes the substance hard; however, if enough energy is applied, many of the bonds will be shattered over large areas causing the substance to be brittle.

Rarely burn

Ionic compounds rarely involve the elements hydrogen and carbon which are needed for burning to take place.

Covalent bond properties

Low melting and boiling points

Unlike the magnetically bound ionic compounds, covalent compounds have very weak forces (called **Van der Waals** forces) holding them together. It does not take much energy to cause these substances to melt/boil.

Typically soft and squishy

Unlike ionic compounds, which are locked together like building blocks, covalent compounds are more like rubber balls thrown together in a box. The atoms can easy be moved around and molded into different arrangements, unlike the rigid crystal structure of ionic compounds.

Usually do not dissolve in water as well as ionic compounds

Water does a good job at surrounding ionic compounds and pulling them apart. This does not happen as easily with covalent compounds. In essence, water cannot hold onto covalent compounds very well so they do not dissolve as well.

Don't conduct electricity

Without the presence of moving electrons or charged particles, electricity cannot be conducted through covalent compounds in any state of matter.

Sometimes burn

Only covalent compounds that contain the non-metallic hydrogen and carbon atoms will burn.

Metallic bond properties

High melting and boiling points

To cause metals to melt, enough energy has to be added to the metal to overpower the sea of electrons which holds the island of cations together and allow them to move.

Malleable and ductile

Much like moving parts of an island around, it is possible to move the cations around in various different shapes. Whichever shape you create, the sea of electrons will continue to surround the nuclei and hold them in place.

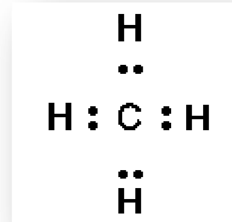
Good conductors of heat and electricity

In order to have electricity, you need moving electrons and there is no other model than through metallic bonds that allows for the movement of electrons.

Shiny

Light tends to bounce off of the sea of electrons quite nicely thereby making metals shiny.

How to draw a electron dot diagram for CH_4



(and everything else for that matter!)

#1 - Count the total number of valence electrons in the molecule

C: 4 electrons x 1 atom = 4 valence electrons
H: 1 electron x 4 atoms = 4 valence electrons
Total = 8 valence electrons

#2 - Find the number of "octet" electrons for the molecule.

C: 8 octet electrons x 1 atom = 8 octet electrons
H: 2 octet electrons x 4 atoms = 8 octet electrons
Total = 16 octet electrons

Rules: Hydrogen always has 2 octet electrons
Beryllium always has 4 octet electrons
Boron wants 6 for neutral molecules (8 if it's an anion)

#3 Subtract the number of valence electrons from the number of octet electrons to find the number of <u>bonding electrons</u>.

16 – 8 = 8 bonding electrons

#4 Divide the number of bonding electrons by 2 to find the number of bonds in the molecule.

8 / 2 = 4 bonds in CH_4

Why 2? Because there are two electrons in every covalent bond!

#5 Draw an arrangement of atoms that has the number of bonds you found in step 4. Of course, there are more rules…

Hydrogen and Halogens = bond <u>once</u>!
Oxygen's family and Beryllium bond <u>twice</u> in neutral molecules
Nitrogen's family and Boron bond <u>3 times</u> in neutral molecules
Carbon's family = bonds <u>4 times</u>!

More hints…

If you bond everything together and you have bonds left over, look for double/triple bonds.

The atom nearest to the left side of the periodic table is probably in the middle of the molecule.

Electron dot practice problems

Draw the electron dot structures for the following compounds:

CCl_4	HCl
C_2H_6	H_2O
BH_3	NH_3

HF	C_2H_2
PCL_3	N_2H_2
PBr_3	CO_2
N_2	CCl_2F_2

A few more practice problems…

Describe whether the following compounds are likely to be ionic or not ionic based on the properties given. Explain your reasoning.

1) Compound 1 has a melting point of 545 degrees Celsius and dissolves well in water.

2) Compound 2 is a brittle material that is used to melt road ice during storms.

3) Compound 3 melts at 85 degrees Celsius and catches fire when heated to 570 degrees Celsius.

4) Circle the substances that are covalent in the list below:

CH_4 $LiCl$ C_2H_4 $CoCl_2$ CO_2 KBr

5) Explain why ionic compounds have such high melting and boiling points when compared with covalent compounds.

6) **Why do ionic compounds conduct electricity when dissolved in water, but not when the solid state?**

7) **You have been given a substance in an unmarked beaker in the chemistry laboratory and asked to determine if it is an ionic compound or not. You tap the crystals gently and they shatter but still retain their sharp edges. You heat the substance gently at first and then harder and after two to three minutes of heating it does not melt. It dissolves in water and the water solution conducts electricity. Is the substance an ionic or a covalent compound? Explain**

8) Which of the following elements/compound(s) would have the highest melting points? Why did you choose these compounds? Defend your answer below:

Ca Br_2 PCl_3 CO_2 SiC   NaF SiH_4 CO H_2S HF

The following questions will require you to use your knowledge of electronegativity to explain the bonding properties you have learned so far. You may need the following chart to help you out:

Electronegativity chart of the Elements

1 **H** 2.1																	2 **He**
3 **Li** 1.0	4 **Be** 1.5											5 **B** 2.0	6 **C** 2.5	7 **N** 3.0	8 **O** 3.5	9 **F** 4.0	10 **Ne**
11 **Na** 0.9	12 **Mg** 1.2											13 **Al** 1.5	14 **Si** 1.8	15 **P** 2.1	16 **S** 2.5	17 **Cl** 3.0	18 **Ar**
19 **K** 0.8	20 **Ca** 1.0	21 **Sc** 1.3	22 **Ti** 1.5	23 **V** 1.6	24 **Cr** 1.6	25 **Mn** 1.5	26 **Fe** 1.8	27 **Co** 1.9	28 **Ni** 1.9	29 **Cu** 1.9	30 **Zn** 1.6	31 **Ga** 1.6	32 **Ge** 1.8	33 **As** 2.0	34 **Se** 2.4	35 **Br** 2.8	36 **Kr** 3.0
37 **Rb** 0.8	38 **Sr** 1.0	39 **Y** 1.2	40 **Zr** 1.4	41 **Nb** 1.6	42 **Mo** 1.8	43 **Tc** 1.9	44 **Ru** 2.2	45 **Rh** 2.2	46 **Pd** 2.2	47 **Ag** 1.9	48 **Cd** 1.7	49 **In** 1.7	50 **Sn** 1.8	51 **Sb** 1.9	52 **Te** 2.1	53 **I** 2.5	54 **Xe** 2.6
55 **Cs** 0.7	56 **Ba** 0.9	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.9	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn** 2.4
87 **Fr** 0.7	88 **Ra** 0.9	89 **Ac** 1.1	104 **Rf**	105 **Ha**	106 **Sg**	107 **Ns**	108 **Hs**	109 **Mt**	110 **Uun**	111 **Uuu**	112 **Uub**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

9) Which particles may be gained, lost, or shared by an atom when it forms a chemical bond?

a) protons
b) electrons
c) neutrons
d) nucleons

10) Which symbol represents a particle that has the same total number of electrons as S^{-2}?

a) O^{-2}
b) Si
c) Se^{-2}
d) Ar

11) Which of these elements has an atom with the most stable outer electron configuration?

a) Ne
b) Cl
c) Ca
d) Na

12) Based on electronegativity values, which type of elements tends to have the greatest attraction for electrons in a bond?

a) metals
b) metalloids
c) nonmetals
d) noble gases

13) Which element has atoms with the greatest attraction for electrons in a chemical bond?

a) beryllium
b) fluorine
c) lithium
d) oxygen

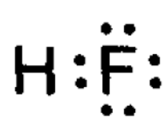
14) Which of the following elements has the greatest ability to attract electrons?

a) Li
b) Be
c) Na
d) Mg

15) Which type of bonding involves positive ions immersed in a sea of mobile electrons?

a) ionic
b) covalent
c) metallic

16) Given the electron dot diagram:



The electrons in the bond between hydrogen and fluorine are more strongly attracted to the atom of

a) hydrogen, which has the higher electronegativity
b) fluorine, which has the higher electronegativity
c) hydrogen, which has the lower electronegativity
d) fluorine, which has the lower electronegativity

17) An element with an electronegativity of 0.9 bonds with an element with an electronegativity of 3.1. Which phrase best describes the bond between these elements?

a) mostly ionic in character and formed between two nonmetals
b) mostly ionic in character and formed between a metal and a nonmetal
c) mostly covalent in character and formed between two nonmetal
d) mostly covalent in character and formed between a metal and a nonmetal

18) Electronegativity is a measure of an atom's ability to

a) attract the electrons in the bond between the atom and another atom
b) repel the electrons in the bond between the atom and another atom
c) attract the protons of another atom
d) repel the protons of another atom

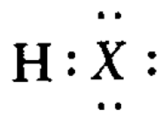
19) Which type of bonding is usually exhibited when the electronegativity difference between two atoms is 1.1?

a) ionic
b) covalent
c) metallic
d) network

20) Two atoms with an electronegativity difference of 0.4 form a bond that is

a) ionic, because electrons are shared
b) ionic, because electrons are transferred
c) covalent, because electrons are shared
d) covalent, because electrons are transferred

21) Given the electron dot formula:



Which atom represented as *X* would have the *least* attraction for the electrons that form the bond?

a) F
b) Cl
c) I
d) Br

22) Which substance contains bonds that involved the transfer of electrons from one atom to another?

a) CO_2
b) NH_3
c) KBr
d) Cl_2

23) Which type of bond is formed when electrons are transferred from one atom to another?

a) covalent
b) ionic
c) hydrogen
d) metallic

24) In which compound do atoms form bonds by sharing electrons?

a) H_2O
b) Na_2O
c) CaO
d) MgO

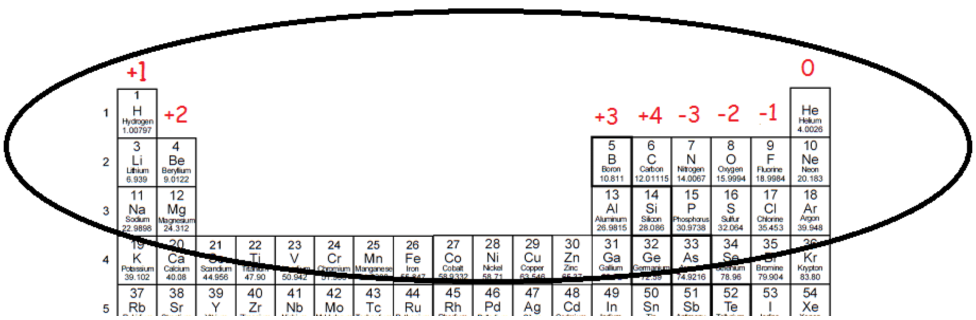
25) Which factor distinguishes a metallic bond from an ionic bond or a covalent bond?

a) the mobility of electrons
b) the mobility of protons
c) the equal sharing of electrons
d) the unequal sharing of electrons

Chapter 13

In the past two chapters, you learned how to bind elements together in different ways to obtain the molecules that exist within the natural world.

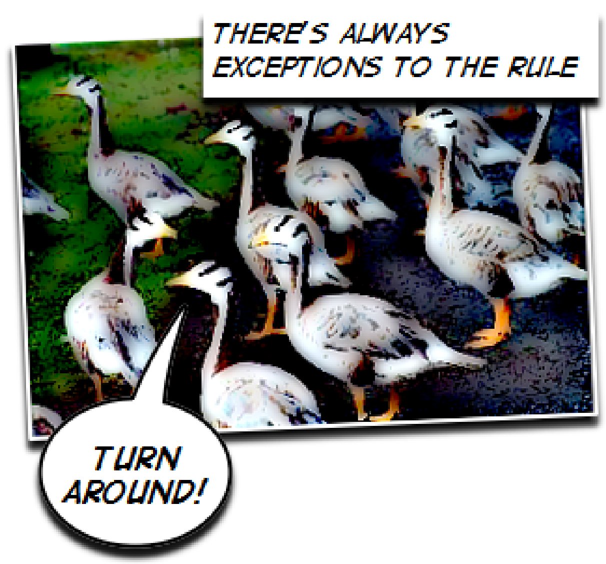
Now it is time for you to learn the language of how scientists name these molecules! The rules are not complicated, but there are a few of them to study. First of all, let's look back at a table we studied in Chapter 11:



This diagram will be very helpful for us this week as you start to learn how to name ionic compounds.

And as expected, there are ALWAYS exceptions to the rule. For example, two elements within the carbon family (tin and lead) do not follow the +4 charge of carbon and silicon. They have a unique set of charges that we are going to learn about right now along with that huge number of elements trapped between our eight families:

The Transition Metals



So far, we have been working on creating ions from within the eight families of the periodic table (Okay – only seven of them are really making ions! But who's counting, right?)

The transition metals have so many exceptions that it is difficult to memorize all of their unique characteristics. This is why scientists rely upon charts and other reference materials to help them out. In fact, some of these elements (like tin and lead) have the ability to pitch more than one amount of electrons at a time! Let's look at a list of these elements:

List of common elements with more than one possible ion charge

Cr^{+2}	chromium (II)
Cr^{+3}	chromium (III)
Co^{+2}	cobalt (II)
Co^{+3}	cobalt (III)
Cu^{+1}	copper (I)
Cu^{+2}	copper (II)
Fe^{+2}	iron (II)
Fe^{+3}	iron (III)
Pb^{+2}	lead (II)
Pb^{+4}	lead (IV)
Mn^{+2}	manganese (II)
Mn^{+3}	manganese (III)
Mn^{+4}	manganese (IV)
Hg_2^{+2}	mercury (I)
Hg^{+2}	mercury (II)
Sn^{+2}	tin (II)
Sn^{+4}	tin (IV)

Look at the first two examples within this chart: chromium (II) and chromium (III)

The element chromium has the ability to pitch either 2 or 3 electrons per atom in order to achieve the lowest energy level possible.

This is advantageous for elements as it allows for increased chances for bonding in the natural world. Unfortunately, it does make things a little more complicated for those of us needing to organize these elements.

But there is no reason to worry!

You will be a master at naming these compounds by the end of this chapter!

Okay – one more chart for you to look at before we begin naming these guys. Until now, all of our ions contain only one atom. These are known as **monatomic ions**. There are times, however, when molecules themselves can gain or lose some of their electrons and turn into ions as well.

We call these molecules **polyatomic ions**. Check out the list below to see some of the most common polyatomic ions in the natural world:

Common Polyatomic Ions

Cation	
ammonium	NH_4^{+1}

Anions	
acetate	$C_2H_3O_2^{-1}$
bromate	BrO_3^{-1}
chlorate	ClO_3^{-1}
chlorite	ClO_2^{-1}
cyanide	CN^{-1}
hydrogen carbonate	HCO_3^{-1}
hydroxide	OH^{-1}
hypochlorite	ClO^{-1}
iodate	IO_3^{-1}
nitrate	NO_3^{-1}
nitrite	NO_2^{-1}
permanganate	MnO_4^{-1}
perchlorate	ClO_4^{-1}
thiocyanate	SCN^{-1}
carbonate	CO_3^{-2}
chromate	CrO_4^{-2}
dichromate	$Cr_2O_7^{-2}$
oxalate	$C_2O_4^{-2}$
peroxide	O_2^{-2}
sulfate	SO_4^{-2}
sulfite	SO_3^{-2}
phosphate	PO_4^{-3}
phosphite	PO_3^{-3}
arsenate	AsO_4^{-3}

Now let's get down to business! It's time to start naming ionic compounds.

All ionic compounds have a two-word name with the cation written in front of the anion. Sodium chloride (table salt) is an excellent example. The first word "sodium" is the cation. The cation is always written the same as you find it on the periodic table (unless the cation is NH_4^{+1} aka- ammonium).

The second word in an ionic compound is changed a bit. For example, you are not able to find the word "chloride" on the periodic table – only the element "chlorine."

When you are writing out ionic compounds and there is only one atom as the anion, you remove the last few letters of the element's name and replace it with "-ide."

If the element has a polyatomic anion, you simply add the name of the anion after the cation and you are done!

Here are some examples for you:

Magnesium sulfate

$MgSO_4$

Potassium iodide

KI

Calcium bromide

$CaBr_2$

Naturally, you are going to want to check your work to be certain you placed the correct name with the compound. To do this, check for multiple ion charges for the cation. One rule of thumb for this is:

All transition metals (except for silver, gold, and zinc) have more than one possible ion charge.

With the exception of lead and tin, all cations from the alkali metal, alkaline earth metal, boron, and carbon families have set ionic charges.

If your cation does have more than one possible charge, you must indicate which ion is present within the chemical name (not the formula) by placing a Roman numeral between the cation and anion. For example...

iron (III) chloride	$FeCl_3$
copper (I) chloride	CuCl
tin (IV) fluoride	SnF_4
lead (II) chloride	$PbCl_2$
iron (III) sulfide	Fe_2S_3

In the above examples, iron (III) chloride's cation is Fe^{+3} and its anion is Cl^{-1}. **The Roman numeral tells you the charge of the cation!**

Writing out the chemical formulas for ionic compounds that contain polyatomic ions may need a little explanation here. When you write out the polyatomic ion and you are going to need more than one of them to balance the charge between the cation and anion, you have to add parentheses to indicate the number ions present. For example...

calcium carbonate $CaCO_3$

*Only one polyatomic anion is needed here so there is no need for parentheses! However, the next two examples will need them...

calcium bicarbonate $Ca(HCO_3)_2$

calcium phosphate $Ca_3(PO_4)_2$

That's all for this week! Good job! Now it's time to practice what you learned with a few dozen problems. Don't worry. I'm certain you will do just fine!

Polyatomic ionic formulas practice

Write out the chemical formula for each of the pairs within this table.

IONS	nitrate	sulfate	carbonate	phosphate	hydroxide	chromate
sodium	$NaNO_3$	Na_2SO_4				
silver						
ammonium						
mercury(I)						
zinc						
calcium						
magnesium						
copper(I)						
lead(II)						
aluminum						
manganese(III)						
cobalt(III)						
copper (II)						
iron (III)						
lead (IV)						
potassium						
barium						

Chapter 14

If you thought naming ionic compounds was easy, wait until you see how we name covalent compounds! It is so simple!

First of all, remember that covalent compounds occur between two non-metals that share their electrons. Unlike ionic compounds, atoms that are bonded covalently are all monatomic in nature. And since no ions are being shared, you have to check the total number of bonds each atom must have in order to place the correct number of atoms together.

Now let's start naming these compounds!

Much like ionic compounds, all covalent compounds contain two words. The first word is the same as the name of the first element in the formula. The second is the same as the second element in the formula AND its last few letters are replaced with the suffix "-ide."

The major difference between the naming of these two compounds can be found in the identification of multiple atoms within the formula. Prefixes are used in front of either or both of the words which identify how many of each atom is present within the compound.

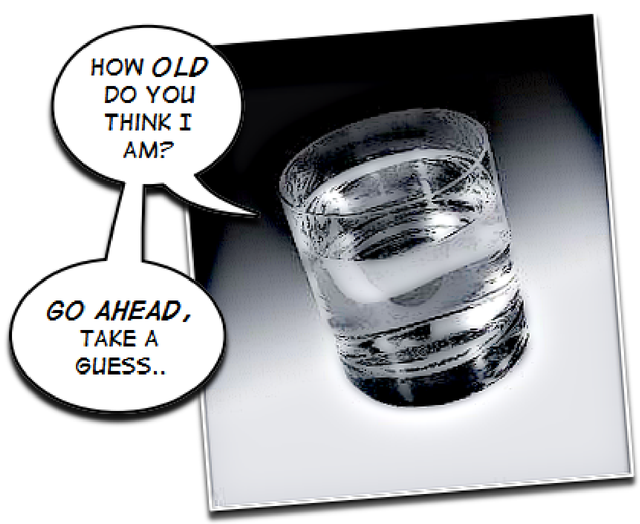
Number of specific atoms within the compound	Prefix
1	mono- (only used for oxygen)
2	di-
3	tri-
4	quad-
5	penta-
6	hex-
7	hept-
8	oct-
9	ennea-
10	dec-

For example, here are some common covalently bonded chemical formula and names:

Carbon monoxide - CO
Dinitrogen trisulfide - N_2S_3
Phosphorus tribromide - PBr_3

Naturally, there are always some exceptions to the rule. Most of these exceptions have existed for so long, that when scientists started using this method of naming the covalent compounds, they didn't change the names of some of the oldest compounds such as:

Water - H_2O
Ammonia - NH_3
Methane - CH_4

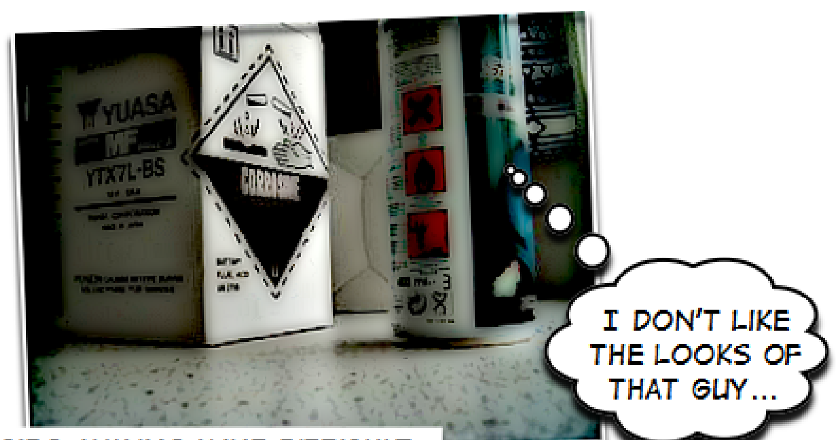


One more thing to look at – if a molecule only contains one type of element its name is the same as the element. For example…

F_2 - Fluorine
P_4 - Phosphorus

Even though we have not yet talked about the role of acids within chemistry, it is important that you learn how to name these important compounds. If you are really curious about how these guys work, feel free to jump ahead a few chapters and learn about these amazing molecules. For now, let's see how to properly name them.

Acids are compounds that begin with an H. This is rather important for the formation of the acid and it will be explained in future chapters. Even though all acids begin with an H, they also are categorized by whether or not they are made up of the element oxygen.

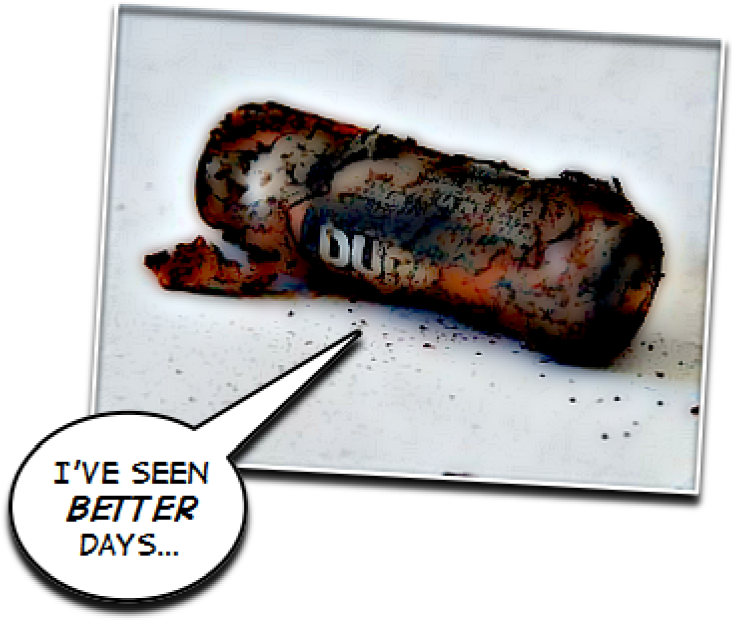


ACIDS ALWAYS HAVE DIFFICULT TIMES MAKING NEW FRIENDS

For example, acids without the presence of oxygen, the name is written as: Hydro(anion)ic acid

Hydrosulfuric acid - H_2S
Hydrobromic acid - HBr
Hydrocyanic acid - HCN

However, if the acids do contain oxygen, the method of naming them changes. If the anion ends in "-ate", the suffix that is added to the chemical name is "-ic".



Nitric acid HNO_3
Sulfuric acid H_2SO_4
Phosphoric acid - H_3PO_4

If the anion ends in "-ite", the suffix is "-ous".

Nitrous acid HNO_2
Sulfurous acid H_2SO_3

All in all, there are three different kinds of formulas that scientists use to describe all of the chemical compounds in the natural world...

Molecular formulas
Empirical formulas
and
Structural formulas

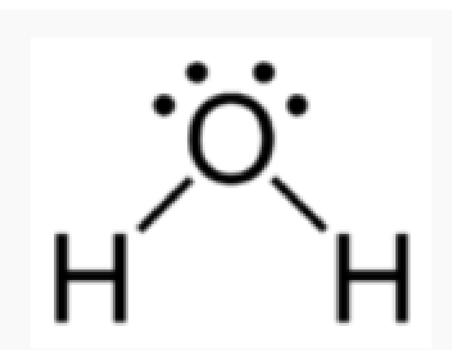
Molecular formulas tell you how many of each type of atom are present within the compound. For example, hydrogen peroxide is written as H_2O_2. This compound contains 2 atoms of hydrogen and 2 atoms of oxygen.

Empirical formulas only tell you the ratios of elements to each other. For example, the example of hydrogen peroxide (H_2O_2) above would be written as: HO

Although we will not be using empirical formulas in our studies, they are commonly used by chemists who specialize in counting the relative amounts of each element in a chemical compound.

Structural formulas are the most specific methods of identifying compounds as they tell you how many of each type of atom are present AND where they are located as well.

In essence, structural formulas are pictures of the elements much like those found within the electron dot structures. The structural formula for water would be:



This type of structural formula (known as a **Lewis structure**) is very similar to the electron dot structure we learned about previously. The only difference that can be found is that the valence electrons in lone pairs are represented as dots, but they also contain lines to represent shared pairs in a chemical bond (single, double, triple, etc.)

Naming compounds practice

Give the name of the following ionic compounds:

1) Na_2CO_3 ____________________
2) $NaOH$ ____________________
3) $MgBr_2$ ____________________
4) KCl ____________________
5) $FeCl_2$ ____________________
6) $FeCl_3$ ____________________
7) $Zn(OH)_2$ ____________________
8) Be_2SO_4 ____________________
9) CrF_2 ____________________
10) Al_2S_3 ____________________
11) PbO ____________________
12) Li_3PO_4 ____________________
13) TiI_4 ____________________
14) Co_3N_2 ____________________
15) Mg_3P_2 ____________________
16) $Ga(NO_2)_3$ ____________________
17) Ag_2SO_3 ____________________
18) NH_4OH ____________________
19) $Al(CN)_3$ ____________________

Write the formulas for the following covalent compounds:

20) antimony tribromide ____________________

21) hexaboron silicide ____________________

22) chlorine dioxide ____________________

23) hydrogen iodide ____________________

24) iodine pentafluoride ____________________

25) dinitrogen trioxide ____________________

26) phosphorus triiodide ____________________

Write the names for the following covalent compounds:

27) P_4S_5 ____________________

28) O_2 ____________________

29) SeF_6 ____________________

30) Si_2Br_6 ____________________

31) SCl_4 ____________________

32) CH_4 ____________________

33) B_2Si ____________________

34) NF_3 ____________________

Chapter 15

Welcome back! It's time we started to look at how we can use all of this information you have been studying. In order to do that, you need to learn how chemists measure the chemicals in a lab.

Until now, you should already know that the mass of an object is measured in grams. But if we are to be as precise as possible, we need to have a way to measure each individual atom within a gram of matter. How is this possible with the unbelievably tiny size of the atoms? It's easy! All you need is a simple conversion factor known as...

The Mole

When scientists say they have a mole of something, they are saying they have 6.02×10^{23} particles of some kind of matter. By particles, I am referring to atoms, molecules, or ions.

1 Mole = 6.02×10^{23} particles (atoms, molecules, or ions)

By the way, another name to use for "the mole" is **Avogadro's number**. You may see both of these terms throughout this chapter!

I hope this isn't confusing because using a conversion factor like the mole is no different than saying we have a dozen eggs, or a gross of bottle rockets, or a trio of singers. The words "dozen," and "gross," and "trio" all are conversion factors for the numbers 12, 144, and 3 respectively.

Remember, we could say that we have a dozen eggs, or a dozen shovels, or a dozen candy bars, etc. It would always mean that we have 12 items. The use of a "mole" is no different whatsoever:

It's just that the number represented by a mole is much larger!

Let's take a look at a few examples:

1 mole C = 6.02×10^{23} C atoms
1 mole H_2O = 6.02×10^{23} H_2O molecules
1 mole NaCl = 6.02×10^{23} Na^+ ions and 6.02×10^{23} Cl^- ions

As you can tell from these examples, a mole of one element equals the same number as a mole of molecules!

Since we have a new conversion factor to play with, let's plug it in to a few dimensional analysis problems.

If I asked you to solve the following problem:

What is the number of atoms in 0.500 moles of Aluminum?

Hopefully, you would do something like this:

$$\frac{0.500 \text{ mole Al}}{} \Bigg| \frac{6.02 \times 10^{23} \text{ Al atoms}}{1 \text{ mole Al}}$$

Remember! The purpose of dimensional analysis is to remove our unwanted units. Since we have "mole Al" in the numerator and the denominator, we can remove them from the equation like this:

$$\frac{0.500\ \cancel{\text{mole Al}}}{} \left| \frac{6.02 \times 10^{23}\ \text{Al atoms}}{1\ \cancel{\text{mole Al}}} \right.$$

Now we simply multiply the numerators together and divide it from the denominator to get our answer:

$$\frac{3.01 \times 10^{23}\ \text{Al atoms}}{1} = \mathbf{3.01 \times 10^{23}\ Al\ atoms}$$

How about one more example?

What is the number of moles of sulfur if a sample of sulfur contains 4.50×10^{24} sulfur atoms?

$$\frac{1.8 \times 10^{24}\ \cancel{\text{S atoms}}}{} \left| \frac{1\ \text{mole S}}{6.02 \times 10^{23}\ \cancel{\text{S atoms}}} \right.$$

$$= \mathbf{3.0\ moles\ S\ atoms}$$

Counting the number of atoms in a sample is pretty easy. In fact, we can also calculate the amount of atoms of a sample if we are only given the mass of the sample!

To do this, we need the periodic table to help us out. If you remember, the mass number of an element is the mass of the protons **AND** neutrons within an element's nucleus which is measured in amu's.

This is very important to remember because it is no coincidence that...

... the mass number of an element is the total number of grams per mole of that element!

For example, the mass number of carbon is 12 amu. This means that if you had 12 grams of pure carbon in your hand, it would contain one mole (6.02×10^{23}) of carbon atoms. And since the mass number of gold is 197 amu, if you were lucky enough to have 197 grams of gold in your hand, you would have 6.02×10^{23} atoms of gold.

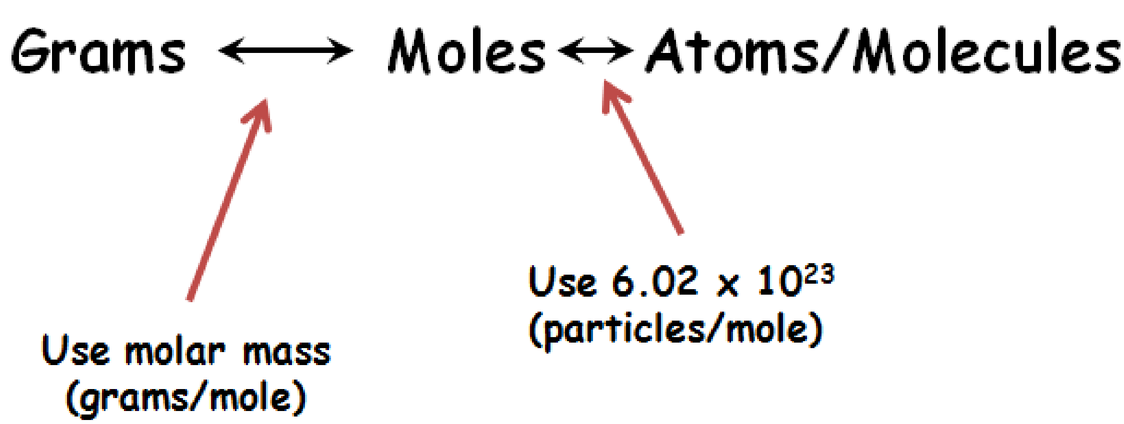
Scientists call this conversion factor the **molar mass** of a particle and it applies to compounds as well. The molar mass of a compound is equal to the sum of its element's atomic masses. For example, let's look at calcium chloride: $\mathbf{CaCl_2}$

(1 mole Ca x 40.1 g/mole) + (2 moles Cl x 35.5 g/mole) = 111.1 g/mole $CaCl_2$

And here's another example: $\mathbf{N_2O_4}$

(2 moles N x 14.0 g/mole) + (4 moles O x 16.0 g/mole) = 74.0 g/mole N_2O_4

With your newfound knowledge of molar mass, it is now possible to convert between grams to particles (atoms/molecules). Here's a little graphic that may help you out:



This graphic states that if you need to convert between grams and moles, you will need to use the molar mass of your sample. And, if you are converting between moles and particles you will need to use Avogadro's number.

Let's practice this with a sample problem:

How many grams of Aluminum are in 3.00 moles of Aluminum?

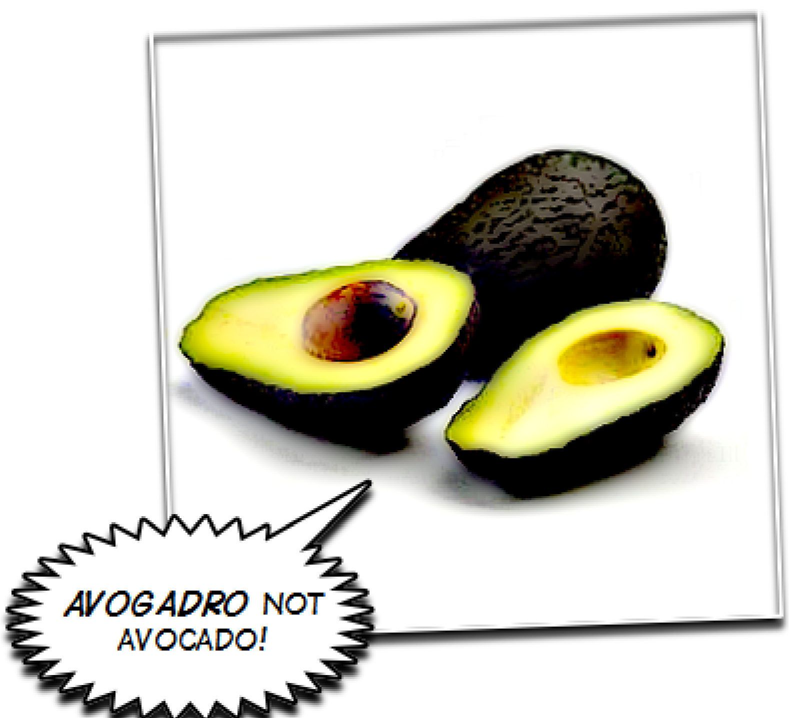
After looking at the periodic table, we know the molar mass of Al to be 27.0 grams/mole. So we now set up our dimensional analysis:

$$\frac{3.00\ \cancel{\text{moles Al}}}{} \left| \frac{27.0\ \text{g Al}}{1\ \cancel{\text{mole Al}}} \right. = \mathbf{81.0\ grams\ Al}$$

Let's take this sample problem a little further:

How many atoms of Aluminum are in 17 grams of Aluminum?

This problem is asking us to convert grams into atoms. After looking at our diagram on the previous page, we are going to need to use both the molar mass conversion factor AND Avogadro's number in this problem. So let's set it up!



17 grams Al	1 mole Al	6.02×10^{23} atoms
	27.0 grams Al	1 mole Al

In this conversion, we can cross off two sets of units: grams AND moles!

17 ~~grams Al~~	1 ~~mole Al~~	6.02×10^{23} atoms	= **3.8×10^{23} atoms Al**
	27.0 ~~grams Al~~	1 ~~mole Al~~	

How about one more problem?

How many grams are there in 7.4×10^{23} molecules of $AgNO_3$?

The molar mass of $AgNO_3$ is 169.8 grams/mole which can be found with the following equation:

(1 mole Ag x 107.8 g/mole) + (1 mole N x 14.0 g/mole) + (3 moles O x 16.0 g/mole)

= **169.8 g/mole $AgNO_3$**

And now we set up the dimensional analysis:

7.4×10^{23} ~~molecules $AgNO_3$~~	1 ~~mole $AgNO_3$~~	169.8 grams $AgNO_3$
	6.02×10^{23} ~~molecules~~	1 ~~mole $AgNO_3$~~

= **2.1×10^2 grams $AgNO_3$**

That wasn't so hard, was it? The only way you will become good at all these conversions is by practicing them a few times. Good luck with your practice problems this week and be certain to look back within the chapter to help you out!

Molar conversions practice

1) Define “mole”.

2) How are the terms “molar mass” and “atomic mass” different from one another?

3) How many moles are present in 27 grams of $Cu(OH)_2$?

4) How many moles are present in 2.45×10^{24} molecules of CH_4?

5) How many grams are there in 3.4×10^{24} molecules of NH_3?

6) How much does 2.4 moles of $Ca(NO_3)_2$ weigh?

7) What is the molar mass of MgO?

8) How many molecules are there in 41 grams of FeF_3?

9) How many molecules are there in 225 grams of Na_2SO_4?

10) How many grams are there in 2.3×10^{23} atoms of silver?

11) How many grams are there in 7.4×10^{23} molecules of $AgNO_3$?

12) How many grams are there in 3.2×10^{23} molecules of H_2SO_4?

13) How many molecules are there in 107 grams of $Cu(NO_3)_2$?

14) How many grams are there in 9.4×10^{6} molecules of H_2?

15) How many molecules are there in 751 grams of $CoCl_2$?

16) How many molecules are there in 2.3 grams of NH_4SO_2?

17) How many grams are there in 3.3×10^{5} molecules of N_2I_6?

18) How many molecules are there in 5 grams of CCl_4?

19) How many grams are there in 1.98×10^{21} molecules of BCl_3?

20) How many grams are there in 4.5×10^{22} molecules of $Ba(NO_2)_2$?

21) How many molecules are there in 19.14 grams of LiCl?

22) How many grams do 1.2×10^{21} molecules of UF_6 weigh?

Chapter 16

Imagine pulling a handful of change out of your pocket and placing the following coins on the table:

1 Quarter
7 Dimes
1 Nickel

I would guess that you could easily calculate what percentage of each coin you have as it pertains to their total worth? Right?

All of this change adds up to $1 which is the equivalent of 100 pennies. Since you already know the value of each of your coins in pennies, you should be able to easily tell me that 25% of your change is found in quarters, 70% is in dimes, and 5% is found in the nickel.

That should have been pretty easy for you! But what would you do if I asked you to calculate the percent of each element within the compound sodium sulfate - **Na_2SO_4**

Although it may not appear to be as easy as it sounds, I promise you will be a master of this procedure by the time you finish your reading this week!

Sometimes a chemist needs to know how much of a particular element is present within compound. You might be thinking this would be rather easy...

...and it is!

But it is not as easy as simply counting the number of elements within a compound and identifying what percentage exists! Remember – much like every coin has a different value, each element is made up of different numbers of protons and neutrons which alters its mass. This means that every element takes up a different amount of space within a compound. So how can we easily determine the size of each element and then calculate its percentage within a compound?

You use the molar mass!

That's right! Everything you used last week is going to be used again in order to determine what chemists call the...

Percent Composition

Even though the number of new concepts is pretty small this week, it is vital that you understand HOW and WHY you go through certain steps to calculate the percent composition of a compound. So let's get started!

The steps required to calculate the percent composition are pretty easy to follow:

1. Calculate the molar mass of the compound.
2. Multiply the atomic mass of each element within the compound by the number of atoms of that element that is present in the compound.
3. Divide each of these masses (Step 2) by the molar mass (Step 1) and multiply by 100%.

Now that you know the steps, let's practice on the same compound you heard about on the first page – sodium sulfate, $\mathbf{Na_2SO_4}$

Step 1: Calculate the molar mass of the compound.

The molar mass of sodium sulfate is:

(2 moles Na × 23.0 grams/mole)

(1 mole S × 32.0 grams/mole)

\+ (4 moles O × 16.0 grams/mole) = 142g/mol

142 grams Na_2SO_4/mole

Step 2: Determine the atomic mass of each element within the compound

Out of each mole of the Na_2SO_4 compound:

2 × 23.0 grams = **46.0 grams are sodium**

1 × 32.0 grams = **32.0 grams are sulfur**

4 × 16.0 grams = **64.0 grams are oxygen**

Step 3: Divide each quantity from Step 2 by the molar mass of sodium sulfate (142g)

46.0 grams Na/142 grams Na_2SO_4 x 100 = **32.4% sodium**

32.0 grams S/142 grams Na_2SO_4 x 100 = **22.5% sulfur**

64.0 grams O/142 grams Na_2SO_4 x 100 = **45.1% oxygen**

There is a final step I haven't mentioned yet, but I'm assuming you already know this one...

You need to check your work!

If you add up the percent compositions of all three elements within sodium sulfate they should add up to 100%

32.4% sodium
22.5% sulfur
+ 45.1% oxygen

100%

How about one more brief example?

What is the percent of carbon in $C_5H_8NO_4$ (MSG - monosodium glutamate), a compound used to flavor foods and tenderize meats?

Step 1: The molar mass of MSG is 146.0 grams/mole.

Step 2: The mass of 5 carbon atoms within MSG is 60.0 grams/mole.

Step 3: The percent composition of carbon within MSG is calculated by:

60.0 grams carbon/146.0 grams MSG x 100 = **41.1% carbon**

If you wanted to check your work, you would have to calculate the percent composition of all the elements within MSG; however, by looking at the compound you can see that roughly half is made up of carbon and the other have would be nitrogen or oxygen. Since hydrogen takes up so little space, you could assume that carbon would make up roughly half (41.1%) of the total mass of MSG.

Understanding how many individual atoms you are working with is vital to our studies in the next chapter - balancing chemical equations. I'm certain after a few practice problems you will be more than ready to jump into this new concept! See you next week!

Percent composition practice

Find the percent compositions of all of the elements in the following compounds:

1) $CuBr_2$

2) NaOH

3) $(NH_4)_2S$

4) N_2S_2

5) What is the percent of carbon in $C_5H_8NO_4$ (MSG monosodium glutamate), a compound used to flavor foods and tenderize meats?

6) Calculate the percent composition for each element present in sodium sulfate, Na_2SO_4

7) What is the percent composition of all elements in water?

8) Calculate the percent composition of potassium chromate, K_2CrO_4.

9) Calculate the percent composition of propane, C_3H_8

Up to know, you've been practicing finding the percent composition of molecular compounds. Now let's work backwards to find the chemical formulas from their percent composition. Here is a "roadmap" to help you through the next few problems:

A. **Assume that you have 100g of the unknown compound.**
   For example, if you assume that you have 100g of a compound composed of 60.3% magnesium and 39.7% oxygen, you know that you have 60.3g of magnesium and 39.7g of oxygen.
B. **Convert the assumed masses from Step 1 into moles by using their molar masses.**
C. **Identify the lowest mole calculation you found in Step 2. Divide all of the moles you calculated in Step 2 by this lowest mole calculation.**
   This answer gives you the mole ratios of the elements of the compound.
D. **If any of your mole ratios aren't whole numbers, multiply all numbers by the smallest possible factor that produces whole number mole ratios for all the elements.**
   For example, if there is 1 nitrogen atom for every 0.5 oxygen atom in a compound, the chemical formula cannot be N1O0.5. However, if you multiply all of the mole ratios you found in Step 3 by two, you would receive all mole ratios in whole numbers. Here's how it works - if you multiplied the mole ratios within our example (1:0.5) by 2 you would receive a 2:1 ratio of nitrogen to oxygen. This would make the chemical formula for this scenario N_2O.

Don't worry too much if your calculations are a little off when doing these problems. Remember, there is always a little error in all our measurements! If your mole ratios in Step 3 turn out to be 2.3, you can pretty much guess that the answer should be 2.

E. **Write the chemical formula by attaching these whole-number mole ratios as subscripts to the chemical symbol of each element.**

10) What is the empirical formula of a substance that is 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen, by mass?

11) Which chemical formula best fits the following percent composition?

50.00% carbon, 5.59% oxygen, 44.41% hydrogen

a) $C_5H_6O_4$
b) C_3HO_2
c) $C_3H_4O_2$
d) $C_4H_3O_2$
e) $C_3H_5O_2$

12) Calculate the empirical formula of a compound with a percent composition of 40.0% sulfur and 60.0% oxygen.

Chapter 17

I know you understand how to identify elements within a compound, but you may need a little review on what all those numbers attached to the chemical symbols mean. For instance, what does the "2" in H_2O really mean, anyway?

If you are going to master the concept of chemical equations, let's first review an important device we have been using for several weeks:

Subscripts

As you may know, **subscripts** are the numbers written after individual atoms that identify how many atoms of that element exist within the molecule. For example, within the compound glucose ($C_6H_{12}O_6$) the following number of atoms exists:

There are 6 atoms of carbon
There are 12 atoms of hydrogen
There are 6 atoms of oxygen

If there is not a subscript listed, it is understood that there is only one atom of the individual element. For example, within sodium chloride (NaCl)...

There is one atom of sodium
and
There is one atom of chlorine

There are times you will see a compound that contains polyatomic anions within parenthesis. For example:

$$Pb(NO_3)_2$$

The coefficient "2" after the parenthesis indicates there are two sets of anions within the parenthesis. The polyatomic anions stick together! Another way to view this would be:

$$Pb(NO_3)_2 = Pb(NO_3)(NO_3)$$

So, if you were to count the atoms you would have the following:

There is 1 atom of Lead
There are 6 atoms of Oxygen
and
There are 2 atoms of Nitrogen

Now let's explore a more realistic view of the chemical world. It is highly unlikely that you would find a single molecule like sodium chloride (table salt) hanging out all by itself. It is highly probable that where you find one compound, many more will also exist! Therefore, we need a way to calculate multiple amounts of individual compounds which exists in the form of...

Coefficients

A **coefficient** is a number placed in front of a compound to identify how many compounds (as a whole) exist. Let's look at two molecules of sulfuric acid:

$$2H_2SO_4$$

The number "2" in front of the H_2SO_4 is a **coefficient** which tells you there are 2 molecules of sulfuric acid. Another way to view this would be:

$$2H_2SO_4 = H_2SO_4 + H_2SO_4$$

By counting all of the atoms that exist in these molecules, you would have the following amounts:

There are 4 atoms of Hydrogen
There are 2 atoms of Sulfur
and
There are 8 atoms of Oxygen

Why are we spending so much time simply counting the numbers of atoms within compounds? Because that is precisely what you must do when you attempt to...

Balance chemical equations!

A chemical equation involves the combination of various molecules (called **reactants**) which produce different molecules (called **products**.) This doesn't mean that the new products are made up of different types of atoms!

The mass of all the reactants (the substances going into a reaction) must equal the mass of the products (the substances produced by the reaction).

This is what it means to BALANCE a chemical equation!

An example of a balanced chemical equation would be:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Reactants **Products**

A balanced chemical equation will follow the **Law of Conservation of Matter** which states that matter cannot be created or destroyed, only rearranged into different positions.

Did you notice how subscripts and coefficients were used in this example?

One molecule of methane (CH_4) reacts with 2 molecules of oxygen gas (O_2) to produce one molecule of carbon dioxide (CO_2) and 2 molecules of water (H_2O).

The presence of subscripts may exist to identify the proper number of atoms needed for the existence of a particular molecule; however, it may become necessary for us to add the coefficients within the equation to make it balanced.

The general rule for the use of subscripts and coefficients is this:

If you change a subscript, you change the molecule. If you change the coefficient, you change the number of that molecule.

Therefore, when balancing a chemical equation, you can only change the coefficient of the equation!

There are many different methods you can use to add the proper amounts of coefficients within an equation. One method (called **MINOH**) identifies the order in which you add coefficients to the type of particles within the equation:

Add coefficients first	M - Metals	Balance metals such as Fe or Na first.
Second	I - Ions	Looks for polyatomic ions (such as PO_4^{-3} or SO_4^{-2}) that cross from reactant to product unchanged. Balance polyatomic ions as a group within its parentheses.
Third	N - Non-metals	Chlorine (Cl) or Sulfur (S) are common non-metals to look for!
Fourth	O - Oxygen	Remember, oxygen by itself is O_2
Fifth (last)	H - Hydrogen	Remember, hydrogen by itself is H_2

If this method seems too technical, you can also try the **grid method**. This method sets up a table for each chemical equation to count the number of atoms within the reactants and the products. Take the following equation:

$$\_\_H_2 + \_\_O_2 \longrightarrow \_\_H_2O$$

Element	Number of atoms in reactants	Number of atoms in products
H	2	2
O	2	1

To balance the number of atoms on both sides of the equation you simply add coefficients into the equation until the number of atoms on both sides are equal.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Element	Number of atoms in reactants	Number of atoms in products
H	~~2~~ 4	~~2~~ 4
O	2	~~1~~ 2

Sometimes, the grid method will require you to simply keep adding coefficients until you finally reach a balanced equation. This may take several steps but be patient! It will work! Here's another example that may require several steps using the grid method:

$$\_\_C_3H_8 + \_\_O_2 \longrightarrow \_\_CO_2 + \_\_H_2O$$

Element	Number of atoms in reactants	Number of atoms in products
C	3	1
O	2	3
H	8	2

Can you work it out? I'm certain you can! Here's the completed equation and the table that shows the number of times the number of atoms were altered through the addition of multiple coefficients:

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2\ 4H_2O$$

Element	Number of atoms in reactants	Number of atoms in products
C	3	~~1~~ 3
O	~~2~~ 10	~~3~~ ~~7~~ 10
H	8	~~2~~ 8

Balancing chemical equations may take a few steps to complete, but they are very important in the field of chemistry! Remember - it is very rare that any individual atom will not have an opportunity to react with some other atom in the natural world. This means that chemical reactions are taking place all day long throughout the universe!

Luckily, most of these reactions can be classified into a few groups which are what you will be studying next week. Hang in there and I'll see you next week!

Balancing chemical equations practice

Balance the equations below:

1) _____ N_2 + _____ H_2 → _____ NH_3
2) _____ $KClO_3$ → _____ KCl + _____ O_2
3) _____ $NaCl$ + _____ F_2 → _____ NaF + _____ Cl_2
4) _____ H_2 + _____ O_2 → _____ H_2O
5) _____ $Pb(OH)_2$ + _____ HCl → _____ H_2O + _____ $PbCl_2$
6) _____ $AlBr_3$ + _____ K_2SO_4 → _____ KBr + _____ $Al_2(SO_4)_3$
7) _____ CH_4 + _____ O_2 → _____ CO_2 + _____ H_2O
8) _____ C_3H_8 + _____ O_2 → _____ CO_2 + _____ H_2O
9) _____ C_8H_{18} + _____ O_2 → _____ CO_2 + _____ H_2O
10) _____ $FeCl_3$ + _____ $NaOH$ → _____ $Fe(OH)_3$ + _____$NaCl$
11) _____ P + _____O_2 → _____P_2O_5
12) _____ Na + _____ H_2O → _____ $NaOH$ + _____H_2
13) _____ Ag_2O → _____ Ag + _____O_2
14) _____ $S8$ + _____O_2 → _____ SO_3
15) _____ CO_2 + _____ H_2O → _____ $C_6H_{12}O_6$ + _____O_2
16) _____ K + _____ $MgBr$ → _____ KBr + _____ Mg
17) _____ HCl + _____ $CaCO_3$ → _____ $CaCl_2$ + _____H_2O + _____ CO_2
18) _____ HNO_3 + _____ $NaHCO_3$ → _____ $NaNO_3$ + _____ H_2O + _____ CO_2
19) _____ H_2O + _____ O_2 → _____ H_2O_2

20) _____ NaBr + _____ CaF_2 → _____ NaF + _____ $CaBr_2$

21) _____ H_2SO_4 + _____ $NaNO_2$ → _____ HNO_2 + _____ Na_2SO_4

22) _____ C_6H_6 + _____ O_2 → _____ H_2O + _____ CO_2

23) _____ NaI + _____ $Pb(SO_4)_2$ → _____ PbI_4 + _____ Na_2SO_4

24) _____ NH_3 + _____ O_2 _____ NO + _____ H_2O

25) _____ $Fe(OH)_3$ → _____ Fe_2O_3 + _____ H_2O

Chapter 18

Mr.Q's Mother's Most Amazing Soup

1 pound ground chuck/sirloin (cooked)
3 cans Campbell's Minestrone soup
1 can Rotel chili peppers
1 can Ranch Style beans
1 soup can of water (if needed)

Add all contents to a pot and let simmer for 30-40 minutes. Fast, easy, delicious!

Can you pick out the reactants and products from this recipe?

Every time you cook a meal you are preparing a chemical equation in which the ingredients (reactants) are "reacting" with each other to produce an entirely different meal (product). In this case, the various cans of beans, minestrone, and peppers (reactants) are "reacting" together to form probably the easiest and most delicious soup in the world. This soup would be the product of the reaction.

I used the words "reacting" and "reaction" in that last paragraph for a reason. This week we will be studying the ever-popular...

Chemical reaction!

Chemical reactions are processes in which the atoms within reactants are rearranged to form new compounds (products).

Since we don't have the ability to see individual atoms moving around, chemists have created a list of "things" to look for which may tell you that a chemical reaction has taken place:

Temperature changes	If the reaction gives off energy it is known as **exothermic**; however, if it absorbs energy it is known as **endothermic**
Color change	Think of the color changes when an object rusts, rots, or burns (this is not the same as when you add a packet of powdered fruit punch to water!)
A solid forms	When you mix two liquids together and a solid forms, the solid is known as a **precipitate**
Formation of a gas	The presence of bubbles is a good indication that a chemical change has taken place!

Much like recipes, it is very important to know the quantities of the reactants we will need in order to figure out how much product we are going to make! This is why you learned how to balance chemical equations last week.

Out of the seemingly infinite ways that atoms can be rearranged within an equation to form new products, there are six basic types of chemical reactions you need to know:

Synthesis	Double Displacement
Decomposition	Acid-Base
Single Displacement	Combustion

Synthesis

Synthesis reactions occur when two or more reactants combine to form a more complicated molecule. Here's is a general equation to show you what it looks like:

$$A + X \longrightarrow AX$$

Example:

$$Pb + O_2 \longrightarrow PbO_2$$

Decomposition

Decomposition reactions are the opposite of synthesis reactions. In this type, a larger molecule breaks down to form two or more simpler products:

$$AX \longrightarrow A + X$$

Example:

$$2H_2O \longrightarrow 2H_2 + O_2$$

Single Displacement

A single element switches places with another element in a chemical compound.

$$A + BX \longrightarrow AX + B$$

Example:

$$Ni + 2AgNO_3 \longrightarrow 2Ag + Ni(NO_3)_2$$

Nickel replaces the metallic ion Ag^+. The silver stays by itself and the nickel becomes the nickel(II) ion.

Double Displacement

The cations of two compounds exchange places with each other.

$$AX + BY \longrightarrow AY + BX$$

Example:

$$2NaOH + CuSO_4 \longrightarrow Na_2SO_4 + Cu(OH)_2$$

The Na^+ and Cu^{2+} switch places. Na^+ combines with SO_4^{2-} to form Na_2SO_4.
Cu^{2+} combines with OH^- to form $Cu(OH)_2$

Acid-Base

Acid-base reactions are similar to double displacement reactions which take place when an acid and base react with each other; however, unlike double displacement reactions, water is one of the products of this reaction:

$$\underset{\text{Acid}}{HA} + \underset{\text{Base}}{BOH} \longrightarrow H_2O + BA$$

An example would be the reaction of hydrobromic acid (HBr) with sodium hydroxide (NaOH):

$$HBr + NaOH \longrightarrow NaBr + H_2O$$

The H^+ ion in the acid reacts with the OH^- ion in the base, causing the formation of an ionic salt (NaBr) and water.

Combustion

When a compound containing carbon and hydrogen combines with oxygen gas, the products formed are heat, water, and carbon dioxide.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

I know this may seem like a lot of information to absorb all at once. However, I think I have something that will help you out. When you are asked to identify which type of reaction is taking place in a chemical equation, go through the following steps in order:

1) Does the chemical equation contain oxygen gas, carbon dioxide, and water? If it does, it's a **combustion reaction**.
2) Do simple molecules combine to form a more complex molecule? If they do, it's a **synthesis reaction**.
3) Does a complicated molecule break apart to form two or more simpler ones? If it does, it's a **decomposition reaction**.
4) Are there any chemicals anywhere in the equation that consist of only one element (Fe, Na, H2, etc.)? If so, it's a **single displacement reaction**.
5) Is water formed during this reaction? If not, it's a **double displacement reaction**.
6) If water is formed, it's an **acid-base reaction**.

Try these steps out on a few practice problems and I'm certain you will be able to identify what type of reaction is taking place in every equation!

Chemical reactions practice

Balance the following reactions and indicate which of the six types of chemical reaction are being represented:

1) _____ NaBr + _____ $Ca(OH)_2$ → ___ $CaBr_2$ + _____ NaOH

Type of reaction: ___________________________________

2) _____ NH_3+ _____ H_2SO_4 → _____ $(NH_4)_2SO_4$

Type of reaction: ___________________________________

3) _____ C_5H_9O + _____ O_2 → _____ CO_2 + _____ H_2O

Type of reaction: ___________________________________

4) _____ Pb + _____ H_3PO_4 → _____ H_2 + _____ $Pb_3(PO_4)_2$

Type of reaction: ___________________________________

5) _____ Li_3N + _____ NH_4NO_3 → _____ $LiNO_3$ + ___ $(NH_4)_3N$

Type of reaction: ___________________________________

6) _____ HBr + _____ $Al(OH)_3$ → _____ H_2O + _____ $AlBr_3$

Type of reaction: ___________________________________

7) What's the main difference between a double displacement reaction and an acid-base reaction?

Here's a few more for you to work on:

8) $Na_3PO_4 + 3\ KOH \rightarrow 3\ NaOH + K_3PO_4$

Type of reaction: ______________________________

9) $MgCl_2 + Li_2CO_3 \rightarrow MgCO_3 + 2\ LiCl$

Type of reaction: ______________________________

10) $C_6H_{12} + 9\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$

Type of reaction: ______________________________

11) $Pb + FeSO_4 \rightarrow PbSO_4 + Fe$

Type of reaction: ______________________________

12) $CaCO_3 \rightarrow CaO + CO_2$

Type of reaction: ______________________________

13) $P_4 + 3\ O_2 \rightarrow 2\ P_2O_3$

Type of reaction: ______________________________

14) $2\ RbNO_3 + BeF_2 \rightarrow Be(NO_3)_2 + 2\ RbF$

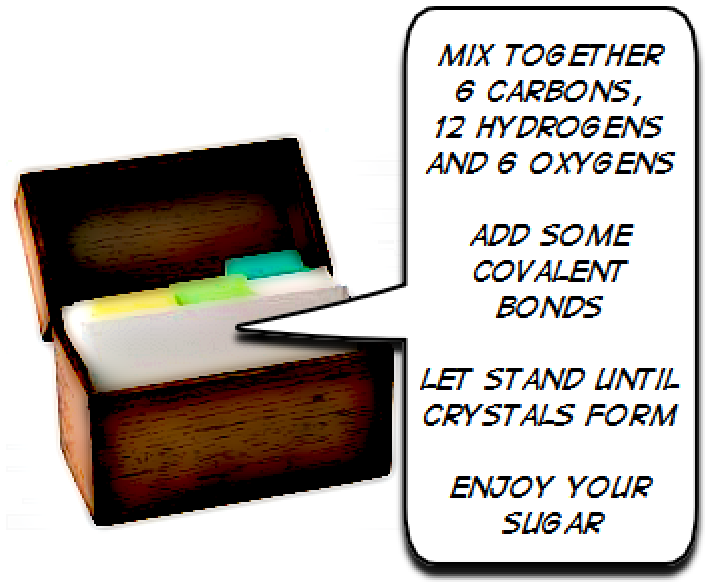
Type of reaction: ______________________________

15) $2\ AgNO_3 + Cu \rightarrow Cu(NO_3)_2 + 2\ Ag$

Type of reaction: ______________________________

Chapter 19

In the previous chapters, you have looked at how to calculate the molar mass of a sample so as to determine its concentration of atoms/molecules. You have also learned how to place elements and molecules together to create a balanced equation. And, you have learned that these balanced equations create specific products through chemical reactions.



Let's put these three concepts into a more practical concept by exploring how they are similar to a recipe:

The ingredients of a recipe are the same as the molar masses of each compound within a reaction. When the correct amounts of each ingredient are written out in the form of a recipe, the recipe itself would be equal to a balanced equation. This recipe tells you how much of each ingredient you need and how it is to be balanced together in order to create a particular product.

Naturally, when you place these separate ingredients together and heat them up, an amazing transformation of flavors takes place. These are the chemical reactions which give you a distinct (and quite tasty) product!

These three concepts may tell you how to prepare the meal, but there are still some preparations we need to do!

First of all, you wouldn't go to the store and buy 25 pounds of meat and 20 cans of beans if you wanted to follow this recipe:

1 pound ground chuck/sirloin (cooked)
3 cans Campbell's Minestrone soup
1 can Rotel chili peppers
1 can Ranch Style beans
1 soup can of water (if needed)

Add all contents to a pot and let simmer for 30-40 minutes. Fast, easy, delicious

Buying this many ingredients would be wasteful! There has to be a better way!

It is now time to learn a new "tool" in our studies of molar mass and balanced chemical equations. The process of calculating the amount of reactants we need for a certain amount of product is something that chemists call...

Stoichiometry

("stoy-kee-ah-meh-tree")

Don't let the fancy world scare you! The methods we will be using today are no different than what you have already learned in previous chapters. So relax... this is going to be fun!

If you look at the recipe on the previous page, you can see that in order to make our soup, you are going to need the following canned goods:

3 cans Campbell's Minestrone soup **1 can Ranch Style beans**
1 can Rotel chili peppers **1 soup can of water (if needed)**

After looking at this equation (recipe) we can see how many cans would be needed in order to make one batch of soup. If we needed to feed twice as many people we could easily double the recipe, right?

Well, the same is true for chemical reactions. Take the following as an example:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

If we want to make 2 moles of water, we will need 2 moles of hydrogen gas (H_2) and 1 mole of oxygen gas (O_2).

And if we need to double this recipe, we would need 4 moles of H_2 and 2 moles of O_2

In chemistry, the use of stoichiometry is important because (like groceries) the cost of "ingredients" is not always very cheap. Pure elements and compounds can be very expensive and hard to store for long periods of time if too much is purchased at once. Therefore, we use stoichiometry to measure precisely how much reactant we need to make a precise amount of product. Got it?

Now let's look at a more realistic example:

If you start with 45 grams of ethylene (C_2H_4), how many grams of carbon dioxide (CO_2) will be produced in the following reaction?

$$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$$

To solve this problem, we are going to use the following diagram to help guide us along. You may recall a similar diagram back in Chapter 15 that helped us calculate between grams and atoms/molecules. Here's the diagram:

Grams (known) ↔ **Moles** (known) ↔ **Moles** (unknown) ↔ **Grams** (unknown)

Use molar mass of known (grams/mole)

Use the Mole ratio

Use molar mass of unknown (grams/mole)

The **mole ratio** is a conversion factor that exists between the moles of the product and reactants from the chemical equation. Here's how you would set up this problem:

45 grams C_2H_4	1 mole C_2H_4	2 moles CO_2	44 grams CO_2
	28.1 grams C_2H_4	1 mole C_2H_4	1 mole CO_2
Grams (known)	Molar mass (known)	**Mole Ratio**	Molar mass (unknown)

You might be asking yourself, "How do you calculate the mole ratio?" Well, to get this conversion factor you simply need to look at the balanced equation:

$$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$$

*The **mole ratio** can be identified as the coefficients for the known and unknown compounds! In this example, there is one mole of C_2H_4 for every two moles of CO_2.*

To finish this problem, we simply have to cancel out all of the units and do our math:

$45\ \cancel{\text{grams } C_2H_4}$	$1\ \cancel{\text{mole } C_2H_4}$	$2\ \cancel{\text{moles } CO_2}$	$44 \text{ grams } CO_2$
	$28.1\ \cancel{\text{grams } C_2H_4}$	$1\ \cancel{\text{mole } C_2H_4}$	$1\ \cancel{\text{mole } CO_2}$

= **141 grams CO_2**

Easy, right? Even though the number of potential problems which exist for this method is massive, don't panic when you encounter one that seems a little different. For example, take a look at the following problem:

How many grams of carbon dioxide will be made when 100 grams of CH_4 burn in oxygen?

$$\mathbf{CH_4 + 2O_2 \longrightarrow CO_2 + H_2O}$$

The reference to the burning of methane (CH_4) in this problem should not confuse you. Simply set up the same dimensional analysis chart as you always have using a known mass of 100 grams CH_4 and solve for the unknown mass of carbon dioxide (CO_2).

By the way, the answer to this question is 275 grams. Is that what you got?

(If you want to see how to solve this problem, take a peek in the next chapter!)

Stoichiometry practice

1) Lithium hydroxide reacts with hydrobromic acid to produce lithium bromide and water. If you start with 6.2 grams of lithium hydroxide, how many grams of lithium bromide will be produced?

2) Ethylene (C_2H_4) reacts with oxygen gas to produce carbon dioxide and water. If you start with 17.6 grams of ethylene, how many grams of carbon dioxide will be produced?

3) Given the equation: $2HCl + Na_2SO_4 \rightarrow 2NaCl + H_2SO_4$
If you start with 62 grams of hydrochloric acid, how many grams of sulfuric acid will be produced?

4) Given the following equation: $LiOH + KCl \rightarrow LiCl + KOH$
If you start with 1.9 grams of potassium chloride, how many grams of potassium hydroxide will be produced?

5) Given the following equation: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
How much C_3H_8 would you need to produce 53.3 grams of carbon dioxide from this reaction?

6) $Be + 2\ HCl \rightarrow BeCl_2 + H_2$
How much hydrochloric acid would you need to produce 12.5 grams of beryllium chloride from this reaction?

7) Given the reaction: $2NaCl + CaO \rightarrow CaCl_2 + Na_2O$
How much sodium oxide would you produce if you used 58.3 grams of sodium chloride in this reaction?

8) $FeBr_2 + 2\ KCl \rightarrow FeCl_2 + 2\ KBr$
How much potassium chloride would you need to produce 125 grams of iron (II) chloride from this reaction?

Chapter 20

Last week, you learned how stoichiometry can be used to determine the amount of a product you will create within a balanced equation. This process can easily be reversed too! If you are given the amount of product within a reaction, you can determine the amounts of reactant needed for its creation.

Unfortunately, there is one small problem with stoichiometry...

It rarely ever is 100% accurate!

We are human and tend to make mistakes. Lots of them! In fact, it is pretty safe to say that every single measurement that is ever made by humans contains some amount of error.

Experimental error is to be expected in EVERY chemical reaction. Think about how this relates to cooking. A recipe for 3 dozen cookies sometimes only makes 33 cookies. Why?

Perhaps the size of the cookies was not the same.
Perhaps three cookies were burnt and had to be thrown away.
Perhaps someone ate some of the dough while you weren't looking.
(This happens all the time in my home!)

In any case, your theoretical yield of 36 cookies was not reached because of some kind of error. Errors are a part of everyday life. Because of this fact, chemists have created a simple procedure to measure how much error exists within a chemical reaction. This measurement is known as:

The Percent Yield

The procedure for calculating the percent yield of a chemical reaction is simple:

$$\text{Percent yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Let's look back at the last example you worked on during the last chapter to see how the percent yield works. The question asked:

How many grams of carbon dioxide will be made when 100 grams of CH_4 burn in oxygen?

$$CH_4 + 2O_2 \longrightarrow CO_2 + H_2O$$

Setting up the dimensional analysis should have looked something like this:

$$\frac{100\ \cancel{\text{grams } CH_4}}{} \left| \frac{1\ \cancel{\text{mole } CH_4}}{16\ \cancel{\text{grams } CH_4}} \right| \frac{1\ \cancel{\text{mole } CO_2}}{1\ \cancel{\text{mole } CH_4}} \left| \frac{44 \text{ grams } CO_2}{1\ \cancel{\text{mole } CO_2}} \right.$$

= 275 grams CO_2

However, what if after running this chemical reaction you DON'T end up with 275 grams of CO_2? What if you only created 225 grams of CO_2?

In order to measure the percent yield of your chemical reaction, you set up the following equation:

$$\frac{225 \text{ grams } CO_2 \text{ (actual yield)}}{275 \text{ grams } CO_2 \text{ (theoretical yield)}} \times 100\% = \mathbf{81.8\%}$$

Naturally, you want the percent yield to be as high as possible since that means your experiment was completed very accurately!

Now let's assume you want to use your cookie recipe mentioned earlier to make 9 dozen cookies. You gather up all of your ingredients to start cooking but you realize that you do not have enough eggs. For every dozen cookies you make, you need one egg – and you only have 4 eggs left. Therefore, you can only make a total of 4 dozen cookies.

Look back at your S'more lab last week. What if you only had three marshmallows but had several chocolate bars and a few boxes of graham crakers? Despite having a huge surplus of chocolate and graham crackers, you could only make 3 s'mores because of the smaller amount of marshmallows!

This situation happens all the time in the chemistry lab. The amount of product you can create is limited by the amounts of reactants you have to begin with. Within any chemical reaction, the reactant that runs out first is known as the **limiting reagent or limiting reactant.** In the above examples, your eggs and marshmallows would be the limiting reactant.

In order to solve problems which involve limiting reactant problems, follow the following steps:

1. Calculate the amount of product you can make from your known masses of each reactant.
2. Identify the smallest amount of product created from each of the dimensional analysis problems in Step 1. The reactant that created this smaller amount of product is known as the limiting reactant.
3. To measure how much excess reactant (the non-limiting reactant), use the following equation:

$$\text{Original amount of nonlimiting reactant} - \text{Original amount of nonlimiting reactant}\left(\frac{\text{Amount of product made}}{\text{Amount of product from the nonlimiting reactant}}\right)$$

How about an example? Let's say we are going to react hydrochloric acid with calcium to form hydrogen gas and calcium chloride. If we start with 100 grams of each reactant, how much hydrogen will be formed and how much of the excess reagent will be left over?

$$2HCl + Ca \rightarrow H_2 + CaCl_2$$

hydrochloric acid | calcium | hydrogen gas | calcium chloride

Step 1: Calculate the amount of product you can make from your known masses of each reactant.

$$\frac{100\ \cancel{\text{grams HCl}}}{} \left| \frac{1\ \cancel{\text{mole HCl}}}{36.5\ \cancel{\text{grams HCl}}} \right| \frac{1\ \cancel{\text{mole } H_2}}{2\ \cancel{\text{moles HCl}}} \left| \frac{2.0\ \text{grams } H_2}{1\ \cancel{\text{mole } H_2}} \right.$$

= 2.74 grams H_2

$$\frac{100 \cancel{\text{grams Ca}}}{} \left| \frac{1 \cancel{\text{mole Ca}}}{40.1 \cancel{\text{grams Ca}}} \right| \frac{1 \cancel{\text{mole } H_2}}{1 \cancel{\text{mole Ca}}} \left| \frac{2.0 \text{ grams } H_2}{1 \cancel{\text{mole } H_2}} \right.$$

= 5.0 grams H_2

Step 2: Identify the smallest amount of product created from each of the dimensional analysis problems in Step 1. This is your limiting reactant.

Since HCl produces the smallest amount of product, HCl is the limiting reactant.

Step 3: Measure how much excess reactant (the non-limiting reactant) will remain.

$$100\text{g Ca} - 100\text{g Ca}\left(\frac{2.74\text{g } H_2}{5.0\text{g } H_2}\right) = 45.2 \text{ grams of Ca in excess}$$

How about you work on another problem all by yourself?

How many grams of water can be formed if 10 grams of hydrogen react with 32 grams of oxygen? How much excess reactant will be left over?

$$2H_2 + O_2 \longrightarrow 2H_2O$$

If you said that 36 grams of water would be created and 6 grams of H_2 would be left over, you are doing just fine!

Spend some time this week working through the practice problems. You will begin to feel very confident in calculating these problems after doing a few on your own!

Limiting reactants and percent yield practice

1) Balance the following equation:

$$\_\_\_\_ \ C_6H_{10} + \_\_\_\_ \ O_2 \rightarrow \_\_\_\_ \ CO_2 + \_\_\_\_ \ H_2O$$

Using this equation, answer the following questions:

2) If I do this reaction with 4.6 grams of C_6H_{10} and 67 grams of oxygen, how many grams of carbon dioxide will be formed?

3) What is the limiting reagent for problem 2?

4) How much of the excess reagent is left over after the reaction from problem 2 is finished?

5) If 11.25 grams of carbon dioxide are actually formed from the reaction in problem 2, what is the percent yield of this reaction?

All of the questions below involve the following reaction: When copper (II) chloride reacts with sodium nitrate, copper (II) nitrate and sodium chloride are formed.

6) Write the balanced equation for the reaction given above:

7) If 15 grams of copper (II) chloride react with 20 grams of sodium nitrate, how much sodium chloride can be formed?

8) What is the limiting reagent for the reaction in #7?

9) How much of the nonlimiting reagent is left over in this reaction?

10) If 11.3 grams of sodium chloride are formed in the reaction described in problem #2, what is the percent yield of this reaction?

Chapter 21

Way back in Chapter 6 you learned about the different ways we classify the different forms of mixtures that exist in the natural world. If you recall, they are known as:

Homogeneous and Heterogenous

("hah-mah-jen-us") ("het-tur-oh-geen-us")

Heterogeneous mixtures are made up of two or more different kinds of molecules that are not bound together; and, due to their differences in size and shape, are easy to separate.

Homogenous mixtures (also known as solutions) are made up of two or more different kinds of atoms that are not bound together as well; however, these types of mixtures are not easy to separate.

In this chapter, you are going to look closely into the WHY things dissolve, WHO is involved, and HOW to change the speed in which something dissolves. Let's get to work!

The chemistry of solutions will be very important to your understanding of the natural world! This is because most of the reactions that take place around us every day have something to do with a solution containing water.

All solutions are made up of two different parts:

Solvent

This is what does all of the dissolving in the solution.

Solute

This is what gets dissolved in the solution.

Typically, when you mix two things together and let this mixture sit for a few hours and it doesn't separate into two different layers or it doesn't remain cloudy in appearance, you have a solution. Examples of solutions in the kitchen would be vinegar, food coloring, and tap water (there are chemicals dissolved in it too!)

But solutions do not have to be liquids! They can be solids and gases too!

The air we breathe is a homogenous mixture of many different gas particles; and, the stainless steel silverware in your kitchen is a mixture of the elements iron, carbon, chromium, and nickel!

Before we get too deep into the calculations of solutions, let's look a little at the topic of dissolving. The ability of a solute to be dissolved by a solvent is a chemical property known as **solubility**. We'll take a closer look at solubility in a little bit. For now, here's one little question to ponder:

What causes a solute to be dissolved by a solvent?

First of all, you learned back in Chapter 8 that...

Everything in the universe tends to move towards the lowest energy possible.

Basically, solvents dissolve solutes when it is more stable to form a solution than it would be if they were separated. Another way to say this is:

Dissolving a solute into a solvent will occur ONLY if it creates the lowest possible energy situation.

Okay. You know WHY solutes are dissolved by solvents; however, what you probably don't know is HOW to determine when a solute will be dissolved. One easy way to determine the solubility of a solute is to look at our old friend...

Electronegativity

We learned in Chapter 10 that electronegativity can help determine if electrons will be shared or stolen during chemical bonding.

Electronegativity can also be used to determine if a molecule is considered **polar** or **nonpolar**. To be considered **polar**, one of the atoms within the molecule must "pull" on its electrons a little stronger than the other atom(s) that are bonded together.

Since these electrons are found closer to one atom than another, a buildup of negative charge can be found around this atom.

The absence of electrons around the remaining atoms within the molecule creates a slightly positive charge as well.

Water is by far the most widely known polar molecule in chemistry!

The oxygen atom tends to pull on its electrons more than its two bonded hydrogen atoms. This makes the oxygen atom a little more negatively charged while the hydrogen atoms remain a little more positive.

Here's how you can determine if a molecule is polar or not:

Subtract the electronegativities of both atoms. If the difference between the two measurements is between 0.3 and 1.7, the compound is polar. If it is at or below 0.3, it is nonpolar. If it is greater than 1.7, the compound is bound by an ionic bond and ionic compounds are ALWAYS nonpolar.

Electronegativity chart of the Elements

1 **H** 2.1																	2 **He**
3 **Li** 1.0	4 **Be** 1.5											5 **B** 2.0	6 **C** 2.5	7 **N** 3.0	8 **O** 3.5	9 **F** 4.0	10 **Ne**
11 **Na** 0.9	12 **Mg** 1.2											13 **Al** 1.5	14 **Si** 1.8	15 **P** 2.1	16 **S** 2.5	17 **Cl** 3.0	18 **Ar**
19 **K** 0.8	20 **Ca** 1.0	21 **Sc** 1.3	22 **Ti** 1.5	23 **V** 1.6	24 **Cr** 1.6	25 **Mn** 1.5	26 **Fe** 1.8	27 **Co** 1.9	28 **Ni** 1.9	29 **Cu** 1.9	30 **Zn** 1.6	31 **Ga** 1.6	32 **Ge** 1.8	33 **As** 2.0	34 **Se** 2.4	35 **Br** 2.8	36 **Kr** 3.0
37 **Rb** 0.8	38 **Sr** 1.0	39 **Y** 1.2	40 **Zr** 1.4	41 **Nb** 1.6	42 **Mo** 1.8	43 **Tc** 1.9	44 **Ru** 2.2	45 **Rh** 2.2	46 **Pd** 2.2	47 **Ag** 1.9	48 **Cd** 1.7	49 **In** 1.7	50 **Sn** 1.8	51 **Sb** 1.9	52 **Te** 2.1	53 **I** 2.5	54 **Xe** 2.6
55 **Cs** 0.7	56 **Ba** 0.9	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 1.7	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.2	78 **Pt** 2.2	79 **Au** 2.4	80 **Hg** 1.9	81 **Tl** 1.8	82 **Pb** 1.9	83 **Bi** 1.9	84 **Po** 2.0	85 **At** 2.2	86 **Rn** 2.4
87 **Fr** 0.7	88 **Ra** 0.9	89 **Ac** 1.1	104 **Rf**	105 **Ha**	106 **Sg**	107 **Ns**	108 **Hs**	109 **Mt**	110 **Uun**	111 **Uuu**	112 **Uub**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

So what does electronegativity have to do with the concept of solubility?

That is a great question! Here's how we can use this information:

Polar solutes dissolve in polar solvents:

Partial charges on the solvent molecules will pull on the partial charges of the solute molecules.

Polar solutes don't dissolve in nonpolar solvents:

The solute molecules would rather stick to each other (because of the opposite charges) than the solvent molecules (which don't have them).

Nonpolar solutes don't dissolve in polar solvents:

The polar solvent molecules would rather stick to each other than to the nonpolar solutes.

Nonpolar solutes dissolve in nonpolar solvents:

Even though nonpolar solvent molecules don't pull on nonpolar solutes, there's also not much holding the solute together.

The easiest way to remember all of this is to memorize one little phrase:

Like Dissolves Like

As you can see, polar solvents dissolve polar solutes and nonpolar solvents dissolve nonpolar solutes!

Now that you understand WHY things dissolve and WHO is involved in this process it's time to learn HOW to change the rate of solubility. You have three ways to increase the speed in which a solute is dissolved:

#1 Grind the solute particles and make them smaller

Because solutes only dissolve at the point where their molecules hit the surface of the liquid, smaller particles will dissolve much faster.

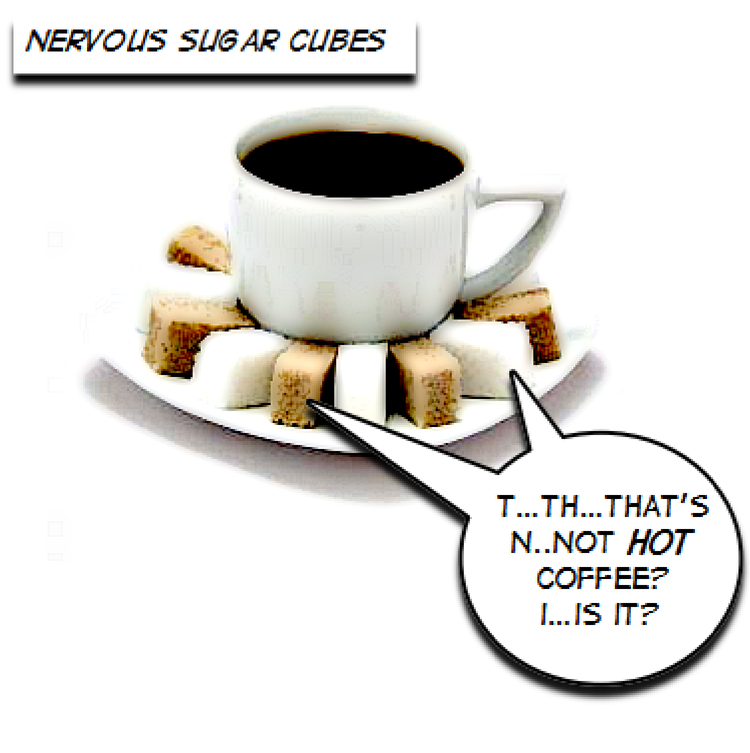
#2 Stir it

When you dissolve a solute the area immediately around the undissolved solute is in contact with very concentrated solution. By stirring it, you put the solute in contact with unconcentrated solution which makes it dissolve more quickly.

#3 Increase the temperature

Solids are usually more soluble in hot solvents than in cold solvents because the solvent molecules have more energy to pull them apart.

Gases are usually less soluble in hot solvents than in cold solvents because when energy is added to the gas molecules dissolved in a liquid, they are more likely to escape into the air than to remain in the liquid.



Now that you have a good idea about what solutions are and how they are formed, it's time to start learning how to use this information! In the next chapter, you will learn how to create solutions using some very simple calculations. See you then!

Solubility practice

Which of the following compounds is more polar (based upon their electronegativities). Defend your answer.

1) Water or hydrogen sulfide

2) Nitrogen triiodide or carbon disulfide

Identify if the following compounds are polar or nonpolar:

3) PH_3 ____________________
4) H_2Se ____________________
5) NF_3 ____________________
6) SBr_2 ____________________
7) NaCl ____________________
8) TeCl ____________________
9) SO_2 ____________________
10) NH_3 ____________________
11) BrF_5 ____________________
12) SeF_4 ____________________
13) ClF_3 ____________________

Using your knowledge of the "like dissolves like" principle of solubility, determine which compound will dissolve best in each of the following liquids.

14) Liquid: methanol (CH_3OH)
    Compounds: carbon diselenide and oxygen dichloride

15) Liquid: carbon disulfide
    Compounds: methanol (CH_3OH) and lithium chloride

16) Liquid: water
    Componds: sulfur dibromide and ammonia (NH_3)

17) Liquid: isopropanol (C_3H_8O)
    Compounds: water and methane

Chapter 22

In the last chapter you learned about the basics of forming a solution by using the following equation:

Solute + Solvent = Solution

The <u>amount</u> of solute we place into a solution is known as its **concentration**. Scientists and cooks have used the concentration of solutions to create everything from complex mixtures used in industry to something as simple as a lollipop.

There are two different ways in which we can measure the concentration of a solution. A **qualitative** method would be to simply keep adding an unmeasured amount of solute into the solvent until no further solute can be dissolved. There are three different qualitative methods we need to briefly explore:

1) *A solution is said to be* ***unsaturated*** *as it continues to dissolve more solute. Dissolving a pinch of sugar into a gallon of water would create an unsaturated solution.*
2) *Once the solvent can no longer dissolve any more solute, it is said to be* ***saturated****. This would be similar to adding an additional 5 pounds of sugar to our gallon of water!*

3) *It is possible to add more solute to a solution than should exist. This happens through the heating of the solution during the dissolving process.*
   As the molecules of solution speed up when heat is added they spread out, allowing for more solute to "slip into the gaps." When cooled, this mixture contains more solute that is usually possible and creates a solution known as ***supersaturated****. This happens very easily with sugar and water; however, it doesn't happen too often in the natural world.*

Unsaturated, saturated, and supersaturated solutions are qualitative methods of creating mixtures. These are typically used during cooking and when precise measurements are not needed to create a desired solution.

However, chemists typically want a much more reliable method of creating a particular solution. And by "reliable" I mean they want a procedure which will produce the same amount and type of product every time!

Think of it this way – you wouldn't read the directions for creating a pitcher of fruit punch and find the following instructions:

...add one packet of powdered mix to 3-5 cups of water.

No way! You want something much more reliable that that! Therefore, we need a **quantitative** method of creating solutions.

In order to measure the concentration of a solution, chemists have created a measuring scale which is very easy to calculate called **molarity**. The molarity of a solution is a numerical way to telling us the concentration of a solution, which is exactly how much solute is dissolved by a solvent.

Here's how we calculate molarity:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

This seems pretty easy, right? It should! You already know how to calculate the number of moles of solute by using the molar mass of a compound.

(Remember - the molar mass of a compound is equal to the sum of its element's atomic masses!)

Let's look at a sample problem to help you out:

What's the molarity of a solution that contains 0.5 moles of NaCl dissolved to make 1.5 L of solution?

$$\text{Molarity (M)} = \frac{\text{0.5 moles of NaCl}}{\text{1.5L solution}}$$

= 0.33 M NaCl solution

This is by far the easiest type of problem you could ever receive when calculating the molarity of a solution. Unfortunately, this type of calculation rarely ever exists in real life. Think about it – how realistic is it to look at the side of a container of salt and read the following:

"Contains 3.011×10^{23} molecules (0.5 moles) of NaCl"

If this is a regular occurrence for you, I would love to know where you do your grocery shopping!

I am willing to guess that 99.9999% of the time you would see that any measurement of salt (NaCl) would be measured in grams (or ounces.) Therefore, let's look at a molarity problem that is a little more realistic:

If I have 30 g LiOH dissolved in 300 mL of solution, what's the molarity?

The first thing to do is convert 30 grams of LiOH to moles of LiOH...

30 grams LiOH	1 mole LiOH	= 1.25 moles LiOH
	24 grams LiOH	

...then convert 300mL to liters with the following conversion:

300 mL solution	1 L solution	= 0.3 L solution
	1000 mL solution	

Now that you have both the moles of solute AND the liters of solution for this problem, you can calculate the molarity of the solution:

$$\text{Molarity (M)} = \frac{\text{1.25 moles of LiOH}}{\text{0.3L solution}}$$

$$= \text{4.2 M LiOH solution}$$

That wasn't so hard, now was it? Now you know how to calculate the amount of grams of solute needed to make a specific concentration of solution! With this information, you need to know how to actually create a solution on your own. Here's what you need to do:

Step 1: Using the equation M = mole/L, determine how much solute will be required. Place the solute into the container that will hold the solution.

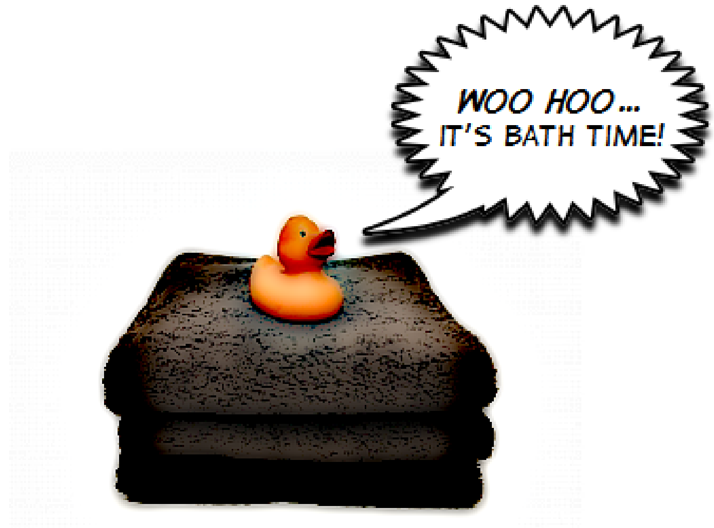
Step 2: Add solvent to the container until the solution is at the desired volume.

Step 3: Stir until the solute is completely dissolved.

Step 4: Add more solvent until the solution is at the desired volume.

You might be wondering why you can't just measure out the correct amount of solute, <u>then</u> the final amount of solvent, and mix them together to get your desired solution? I'll answer this question with a brief story...

What would happen if you filled up your bathtub to the very top and then stepped inside? It would overflow, right? If you measured out the correct amount of solvent and then added your solute to the fluid, you would create more solution than you needed!



There's another "little" problem that takes place when you dissolve a solute into a solvent:

The volume of the solution will decrease when the solute dissolves!

At first thought, you might be thinking I've completely lost my mind. But this is a true statement! Here's why:
As the solute dissolves, it "fills in the gaps" between the solvent molecules. This will lower the volume of solution that exists in your container. The problem is – you can never really know how much it is going to drop! Therefore, you have to follow the previous steps to be as precise as possible when creating solutions!

One more thing this week…

It's handy to take a very concentrated solution of a compound and dilute it so it's less concentrated. Why? Well, for one thing, it's much cheaper to do that than to purchase exactly what you need. To calculate these types of dilutions, use this equation:

$$M_1V_1 = M_2V_2$$

This is the **KNOWN molarity** and **volume** within the word problem.

This is the **FINAL molarity** and **volume** of the solution.

Let's look at a couple sample problems:

If I dilute 5mL of 0.15M NaCl solution to a final volume of 2mL, what is the final concentration of the NaCl solution?

$$(0.15M\ NaCl)(5mL\ NaCl) = (M)(25mL)$$

= 0.03 M NaCl solution

You can redefine this process with the following statement:

You diluted a (more concentrated) 0.15M solution of NaCL to a 0.03M solution of NaCl (less concentrated.)

Here's another sample problem that will take a little more effort:

If I add 25mL of water to 125mL of a 0.15M NaOH solution, what will the molarity of the diluted solution be?

$$(0.15\ M)(125\ mL) = M(150\ mL)$$

$$= 0.125\ M$$

Remember! This is the **FINAL volume** of the solution. You started with 125mL and ADDED another 25mL for a total volume of 150mL.

You diluted a (more concentrated) 0.15M solution of NaOH to a 0.125M solution of NaOH (less concentrated.)

A good way to check your answer when doing dilutions is by looking at the values of M_1 and M_2. The amounts for these variables should always decrease from M_1 to M_2 since you are diluting (watering down) the original concentration.

You are doing a wonderful job! Now you have the tools to create the correct solutions (literally) for any chemistry problem you encounter! It's time you practiced these new skills with a few problems on your own. Good luck!

Molarity and dilutions practice

Calculate the molarities of the following solutions:

1) 5.4 moles of sodium chloride in 0.45 liters of solution.

2) 4.3 moles of calcium carbonate in 1.22 liters of solution.

3) 0.15 moles of sodium sulfate in 12 mL of solution.

4) 1.05 moles of lithium fluoride in 65 mL of solution.

5) 0.18 moles of magnesium acetate in 5 liters of solution.

6) 12 grams of calcium nitrite in 240 mL of solution.

7) 101 grams of sodium hydroxide in 2.2 liters of solution.

8) .98 grams of hydrochloric acid in 25 mL of solution.

9) 27grams of ammonia in 0.75 L of solution.

Explain how you would make the following solutions.

10) 2 L of 3 M HCl

11) 1.5 L of 6 M NaOH

12) 0.75 L of 0.87 M Na_2SO_4

13) 45 mL of 0.45 M sodium carbonate

14) 250 mL of 0.12 M lithium nitrite

15) 56 mL of 1.9 M iron (II) phosphate

16) 6.7 L of 6.8 M ammonium nitrate

17) 4.5 mL of 0.01 M magnesium sulfate

Chapter 23

You've spent a lot of time learning about concentrations, solutions, and how to dilute solutions to a desired molarity. With all of this information swimming around in your head, it shouldn't be too surprising if I told you that the properties of a solution change when you alter its concentration.

That should make sense, right? If you were to add 1 gallon of water to your fruit punch mix instead of the recommended 2 cups, its properties would be considerably different (and much less tasty as well.)

These changes in properties will be the topic of discussion for the next couple of chapters. Chemists have named any property of a solution that changes with its concentration:

Colligative properties

Before we dig into the specifics of **colligative properties**, I think a little discussion is needed. First of all, if you have ever watched a weather report you probably have heard the meteorologists talk about high and low air pressures. Air pressure is the force exerted by the molecules of air (mostly nitrogen and oxygen gases) in any given area on the planet. Therefore, when you hear about a "high pressure system" moving into your area by your favorite TV weather person, they are saying that a larger than normal collection of gas molecules are moving in your direction.

There's no need to worry! These increased molecules will not cause any serious damage whatsoever – although they may bring with them an unfavorable storm to your area!

You can feel this change in air pressure every time you fly in an airplane. Your body is "trained" to press outward with an equal amount of pressure that it normally feels from the air pressure where you live. It is a fact that the amount of air molecules decreases the higher you travel in the atmosphere. Therefore, your ears will pop within an airplane as the higher pressure exerted by your body is greater than the air pressure pushing in on your eardrums.

What does all of this have to do with colligative properties?

Since the air is all around us, it is reasonable to assume that all of these air molecules are slamming into every object on the planet. Right? This would include the surface of a pot of water on a stove!

The air pressure on the surface of a pot of water pushes the water into the container and prevents it from vaporizing (turning into water vapor) throughout the atmosphere. However, if you turn on your stove, the water molecules will begin to absorb a lot of heat and move around a lot faster.

The slamming of molecules against each other causes an increase of pressure within the water. This increased motion causes the pressure within water (known as **vapor pressure**) to increase as well. When the vapor pressure of the water is greater than the air pressure forcing it into the container, the water molecules break away from each other and change into a gas.

But what if there were not as many water molecules on the surface of the water to vaporize?

This brings us to the first colligative property we will be studying:

Boiling point elevation

The **boiling point elevation** states that as you increase the concentration of a solution (by adding solute into a solvent), the boiling point of the solution increases as well. Why?

The greater amount of solute molecules (specifically on the surface of the solution) reduces the amount of solvent molecules that can vaporize out of the solution!

Since the solute molecules (which are typically solids) will not absorb heat energy and vaporize as easily as the solvent molecules, the vapor pressure of the solution will be lower. Since a solution's vapor pressure is lower than the air pressure, the solution needs more heat energy to raise its vapor pressure and begin to boil!

Chemists can calculate the boiling point of solutions of various concentrations using the following equation:

$$T_{b\,(solution)} - T_{b\,(pure\ solvent)} = K_b m$$

T_b = boiling point
(The $T_{b(pure\ solvent)}$ for water is 100°C)

K_b = 0.52°C/m for water
(All solvents have a different K_b value; however, we will be using only water as a solvent throughout this book. And trust me – the method of calculating this constant value is well beyond the scope of this book!)

m = the effective molality of the solute
The **effective molality (m) of the solute is much like molarity but with one small difference – molality is a measurement of the number of moles of solute per kilogram of solvent. It's equation looks like this:

$$\text{molality (m)} = \frac{\text{moles of solute}}{\text{kilogram (kg) of solvent}}$$

Since we are looking at how the amount of solute molecules affects the boiling point of a solution, we need to understand how many particles will be floating around inside our solutions, right?

This is where we use the molality of our solute to determine the boiling point of various solutions. To use the boiling point elevation equation properly, simply follow these two rules:

If the solute is a covalent compound, use its regular molality as the effective molality.

For example, if you have a 0.5 m solution of $C_6H_{12}O_6$, its effective molality is 0.5 m.

The reason this number does not change is because covalent compounds do not break apart within a solution – they only separate from other covalent compounds! This is completely different from ionic compounds that easily separate into ions within solution. Therefore…

If the solute is an ionic compound, multiply its regular molality by the number of ions within the compound.

For example, if you have a 0.8 m solution of $MgCl_2$, the effective molality would be 2.4 m (0.8 x 3 ions).

Okay. Let's practice with a couple problems to get you more comfortable:

What is the boiling point of a 1.25 m solution of $C_6H_{12}O_6$?

$$T_{b\,(solution)} - T_{b\,(pure\ solvent)} = K_b m$$

$$T_{b\,(solution)} - 100 = 0.52\ (1.25)$$

$$T_{b\,(solution)} - 100 = 0.65$$

$$T_{b\,(solution)} \cancel{- 100 + 100} = 0.65 + 100$$

$$T_{b\,(solution)} = 100.65\ °C$$

What is the boiling point of a 1.0 m solution of $Ca(NO_3)_2$?

$$T_{b\,(solution)} - T_{b\,(pure\ solvent)} = K_b m$$

$$T_{b\,(solution)} - 100 = 0.52\ (3.0)$$

Since this is an ionic compound, we multiply the 1.0 m regular molality by three (3) which is the number of ions in $Ca(NO_3)_2$

$$T_{b\,(solution)} - 100 = 1.56$$

$$T_{b\,(solution)} \cancel{- 100 + 100} = 1.56 + 100$$

$$T_{b\,(solution)} = 101.56\ °C$$

Remember! The addition of solute to solvent causes the boiling point of the solution ($T_{b(solution)}$) to be **greater** than the boiling point of the solvent, which for water is 100°C. This is how the colligative property of boiling point elevation works!

Boiling point and molality calculation practice

1) For a 1.87 m solution of sucrose ($C_{12}H_{22}O_{11}$), what is the boiling point?

2) What is the normal boiling point of a 0.65 m solution of KI?

Now that you're warmed up to calculate the boiling point elevations of various compounds, let's see how well you do if the molality is not provided for you. In the following problems, be certain to calculate the molality of your solution first before determining the boiling points. For all these problems, you will need to know that 1 mL H_2O = 1 gram H_2O

Do not forget the molality equation as well:

$$\text{molality (m)} = \frac{\text{moles of solute}}{\text{kilogram (kg) of solvent}}$$

3) 40g of sodium chloride is added to 1,500g of water. What is the molality of the resulting solution?

4) How many grams of magnesium iodide must you add to 2540g of water to make a solution with a molality of 0.42?

5) How many grams of sodium chloride must you add to 850g of water to make a solution with a molality of 2.35?

6) What is the boiling point of a solution containing 46.3 g $MgSO_4$ and 299 g H_2O? (Hint: First convert grams of solute to moles; then, convert grams of solvent to kg)

7) Which solution will have a higher boiling point: A solution containing 105 grams of sucrose ($C_{12}H_{22}O_{11}$) in 500 grams of water or a solution containing 58 grams of sodium chloride in 500 grams of water?

Chapter 24

The colligative property of boiling point elevation is only one of four properties you will master in these two chapters. Since the boiling point of a solution is a property that changes with its concentration, it should make sense that another phase change will contain a colligative property as well! Therefore, this week we will be looking at the second colligative property known as:

Freezing point depression

Freezing point depression states that solutions freeze at lower temperatures than pure solvents because the extra solute molecules "get in the way" of the solvent molecules. These extra solute particles keep the liquid from freezing at the normal freezing point of solvents.

The equation we use to determine the freezing point of solutions is (almost) identical to what you learned last week:

$$T_{f\,(\text{pure solvent})} - T_{f\,(\text{solution})} = K_f m$$

T_f = freezing point

(Take note that the T_f for the pure solvent and for the solution have been reversed from the boiling point elevation equation. The $T_{f(\text{pure solvent})}$ for water is 0°C)

K_f = 1.86°C/m for water

(Much like I said last week, all solvents have different K_f values; however, we will only be using water as our solvent.)

m = the effective molality of the solute

You can follow the same rules for calculating the effective molality of the solute from last week:

- **For covalent compounds** - use its regular molality as the effective molality.
- **For ionic compounds** - multiply its regular molality by the number of ions within the compound

Since you have a good idea how this equation works from your practice during the last chapter, let's run through a couple quick sample problems to get you comfortable. <u>Don't forget about the couple changes in this new equation!</u>

What is the freezing point of a 1.25 m solution of $C_6H_{12}O_6$?

$$T_{f\ (\text{pure solvent})} - T_{f\ (\text{solution})} = K_f m$$
$$0 - T_{f\ (\text{solution})} = 1.86(1.25)$$
$$0 - T_{f\ (\text{solution})} = 2.3$$
$$-T_{f\ (\text{solution})} = 2.3$$
$$T_{f\ (\text{solution})} = -2.3\ ^\circ C$$

What is the freezing point of a 1.0 m solution of $Ca(NO_3)_2$?

$$T_{f\ (pure\ solvent)} - T_{f\ (solution)} = K_f m$$
$$0 - T_{f\ (solution)} = 1.86(3.0)$$
$$0 - T_{f\ (solution)} = 5.6$$
$$-T_{f\ (solution)} = 5.6$$
$$T_{f\ (solution)} = -5.6°C$$

The third colligative property we will be looking at is known as:

Osmotic pressure increase

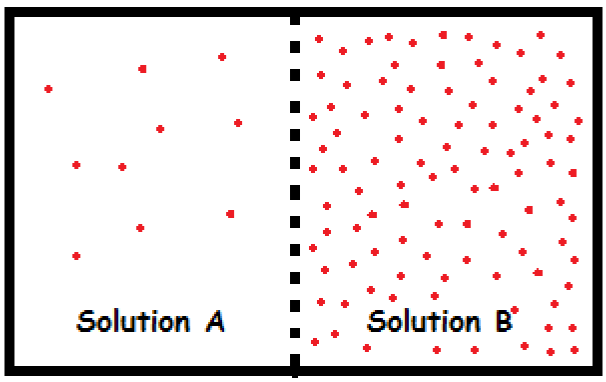
Osmotic pressure is the pressure which needs to be applied to a solution to keep water from moving across a semipermeable membrane (osmosis).

A **semipermeable membrane** is any "barrier" that will allow solvents (like water) to move through them; however, this "barrier" will not allow solutes to pass through. The best example of a semipermeable membrane would be the lining of your intestines or a plant cell wall.

The easiest way to understand **osmosis** is that water flows from the solution with the lower solute concentration into the solution with higher solute concentration. This flow will continue until the concentrations on both sides are equal.

Here's an example of how osmosis works:

Imagine having two containers of equal volume separated by a semipermeable membrane. Solution A has only 10 molecules of NaCl while Solution B has 100 molecules of NaCl. Which direction will the solvent (water) flow?



If you said the water would flow from Solution A into Solution B you would be correct! Solution A contains the lower solute concentration and therefore would allow water to pass through the membrane!

Think of it this way – within Solution B, some of the pores within the membrane get plugged up with NaCL molecules; however, this does not happen as easily in Solution A. Therefore, water is more likely to flow from Solution A into Solution B.

So... an increase in concentration (like you find in Solution B) will increase the osmotic pressure of the solution by **keeping** water from moving across a semipermeable membrane!

Almost done! Only one more colligative property to go...

Increased conductivity of electricity

Solutions which contain dissolved ionic compounds are able to conduct electricity due to their unbalanced numbers of electrons. The common name for these conductive compounds is **electrolytes**.

Generally speaking, the more concentrated a solution of electrolytes, the better it is at conducting electricity.

The extra amount of ions within the solution increases the ability to move electrons and charges around the solution, thus making it more conductive!

That wraps up our study of solutions for now! Don't forget the four colligative properties of solutions you have learned about...

Boiling Point Elevation
Freezing Point Depression
Osmotic Pressure Increase
and
Increased Conductivity of Electricity

...are all easily changed by increasing or decreasing their concentration of solutes. In the next unit, we will be moving into a new phase of matter that surrounds us all day long:

Gases!

More colligative properties practice

1) If I add 22 grams of sodium chloride to 500 grams of water, what will be the melting and boiling points of the resulting solution?

2) 22 grams of salt (NaCl) is added to 170 mL of water. What are the new freezing and boiling points?

3) What is the change in freezing point of a solution containing 132 g $C_{12}H_{22}O_{11}$ and 250 g of H_2O?

4) If I add 45 grams of magnesium chloride to 500 grams of water, what will the melting and boiling points be of the resulting solution?

5) Each kilogram of seawater contains roughly 35g of dissolved salts. Assuming that all of these salts are calcium chloride, what is the freezing point of seawater?

Chapter 25

It's time we started looking at one phase of matter which has only been briefly discussed until now:

Gases

Way back in Chapter 6, we explored a concept that is going to be very important for the next couple of weeks:

Pressure

Pressure is the force of gas molecules as they hit the sides of the container in which they are stored. Air pressure is measured in units called atmospheres (atm); and, 1 atm is considered to be the "normal" atmospheric pressure.

Another unit to measure pressure is the kilopascal (kPa). The conversion factor for kPa to atom is: 101.325 kPa = 1 atm

In addition to pressure, you also studied another concept as it relates to the various phases of matter: **Temperature**

Temperature is a measurement of the amount of energy found within molecules. You may need to know that "room temperature" is considered in the scientific world to be 298°K or 25°C

To convert degrees Celsius (°C) to Kelvin (°K), simply add 273 to your Celsius measurement.

Since we are on the topic of temperature and gases, it is very important that you follow this simple rule:

All temperatures within calculations involving gas must be in Kelvin units!

Why do we need to do this? We'll get to that in a minute. For now, let's look at a couple more concepts which are going to be important this week:

Volume

Volume is the amount of space in which a fluid is enclosed. And yes, not only can liquids be considered a "fluid" - so can gases! The unit we use in calculating volume is the liter (L).

The only other term that you will probably see from time-to-time when doing gas calculations is the term **STP**. STP stands for "standard temperature and pressure" which is considered to be 273°K (0°C) and 1.00 atm.

Okay! Now that you have all the background information you need, let's start working through some gas calculations.

First up, we have:

Boyle's Law $P_1V_1 = P_2V_2$

Mathematically, Boyle's law states that as a gas goes through some kind of change, the product of its initial pressure and volume (P_1V_1) will equal the product of its final pressure and volume (P_2V_2).

So what this means is...

Increasing the pressure on a gas makes its volume decrease; and, decreasing the pressure will increase its volume.

Imagine squeezing a balloon until it pops. As you squeeze the balloon, you are decreasing the volume. If the volume is decreasing, its pressure is increasing. Once the balloon can no longer take all those air molecules slamming into its walls... POP!

Let's see how this works with a sample problem:

If you have 50L of gas at a pressure of 2 atm and you double its pressure, what will the new volume of the gas be?

$$P_1V_1 = P_2V_2$$

$$(2 \text{ atm}) \times (50\text{L}) = (4 \text{ atm}) \times (V_2)$$

$$100 \text{ atm} \cdot \text{L} = (4 \text{ atm}) \times (V_2)$$

$$\frac{100 \cancel{\text{atm}} \cdot \text{L}}{4 \cancel{\text{atm}}} = \mathbf{25\ L}$$

Now that we know the relationship between pressure and volume, let's look at a different relationship within a new gas law:

Charles's Law $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Charles's law states that as you increase the temperature of a gas, its volume increases too!

This should make sense to you. Think about it... temperature is the measurement of energy found within molecules. When you add heat to molecules they absorb this energy and start moving around a lot more, right? This movement is the reason for the melting of solids and the evaporation of liquids! The gas molecules simply speed up and continue to move farther away from each other.

Just remember the rule that was mentioned earlier:

All temperatures within calculations involving gas must be in Kelvin units!

The reason for using Kelvin in gas law calculations is simple to understand. The Kelvin measurements scale contains the most extreme measurement known as absolute zero. This scientifically (and mathematically) determined "official zero" helps with our calculations a great deal. For example, if you double the Kelvin temperature of a gas, its volume doubles as well. This doesn't happen when you use the Celsius scale (it doesn't have an "absolute zero" in its range of temperatures.)

With this knowledge, calculating a practice problem should be easy:

If you heat a 2.5 L balloon from a temperature of 25°C to 50°C, what will the new volume of the balloon be?

$$\frac{2.5\ L}{25°C} = \frac{???}{50°C}$$

Convert to Kelvin: $25°C + 273 = 298°K$
$50°C + 273 = 323°K$

$$(2.5\ L) \times (323°K) = (298°K) \times (V_2)$$

$$807.5\ L \cdot °K = (298°K) \times (V_2)$$

$$\frac{807.5\ L \cdot \cancel{°K}}{298\ \cancel{°K}} = \mathbf{2.7\ L}$$

You first looked at the relationship between pressure and volume with Boyle's law. Now you understand the relationship between volume and temperature with Charles' law. There's only one more way to pair up volume, temperature, and pressure. Therefore, now it is time to look at the relationship between pressure and temperature with...

Gay-Lussac's Law $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

The Gay-Lussac law states that as you increase the temperature of a gas, its pressure goes up too!

This shouldn't surprise you! You wouldn't think about tossing a can of soda into a fire, would you? I hope not! As the heat of the can would increase, so would its gas pressure until the can explodes.

Practice problems for this law are very similar to those found with Charles' law. For example:

If you have a can of soda at a pressure of 20 atm at room temperature and put it into the fireplace which has a temperature of 1200°C, what will the pressure inside the can just before it explodes?

$$\frac{20 \text{ atm}}{25°C} = \frac{???}{1200°C}$$

Convert to Kelvin: $25°C + 273 = 298°K$
$1200°C + 273 = 1473°K$

$$(20 \text{ atm}) \times (1473°K) = (298°K) \times (P_2)$$

$$29460 \text{ atm} \cdot °K = (298°K) \times (P_2)$$

$$\frac{29460 \text{ atm} \cdot \cancel{°K}}{298\ \cancel{°K}} = \mathbf{98.9 \text{ atm}}$$

It's time you spent a little time practicing the relationships between volume, temperature, and pressure within the gas laws. Next week, you'll learn how these three simple equations can be used in a much more practical way.

See you then!

Gas law practice

1) If I have 5.6 liters of gas in a piston at a pressure of 1.5 atm and compress the gas until its volume is 2.3 L, what will the new pressure inside the piston be?

2) I have added 18.2 L of air to a balloon at sea level (1.0 atm). If I take the balloon with me to Denver, where the air pressure is 0.85 atm, what will the new volume of the balloon be?

3) I've got a car with an internal volume of 12,000 L. If I drive my car into the river and it implodes, what will be the volume of the gas when the pressure goes from 1.0 atm to 2.4 atm?

4) If I have 29.3 liters of helium in a balloon at 25°C and increase the temperature of the balloon to 55°C, what will the new volume of the balloon be?

5) Calcium carbonate decomposes at 1200°C to form carbon dioxide and calcium oxide. If 82 liters of carbon dioxide are collected at 1200°C, what will the volume of this gas be after it cools to 25°C?

6) I have 130 liters of gas in a piston at a temperature of 250°C. If I cool the gas until the volume decreases to 49.5 liters, what will temperature of the gas be?

7) A commercial airliner has an internal pressure of 1.00 atm and temperature of 25°C at takeoff. If the temperature of the airliner drops to 10°C during the flight, what is the new cabin pressure?

8) If divers rise too quickly from a deep dive, they get a condition called "the bends" which is caused by the expansion of very small nitrogen bubbles in the blood due to decreased pressure. If the initial volume of the bubbles in a diver's blood is 18.2 mL and the initial pressure is 12.75 atm, what is the volume of the bubbles when the diver has surfaced to 1.00 atm pressure?

Chapter 26

Last week you learned how the concepts of volume, temperature, and pressure are interconnected. Boyles', Charles's, and Guy-Lussac's Laws all describe how these three important concepts affect the laws of gases throughout the natural world.

But what if I told you there was another law that combined all of these three earlier laws together. Would that make life a little easier for you? I bet it would!

Now don't get angry because you spent all of last week calculating problems that can be simplified this week. You needed to learn how volume, temperature, and pressure affect gases within those laws. Now it's time to start looking more practically at how these interconnected concepts within the world. To do this, you need to learn about an entirely new law:

The Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The combined gas law is an amazing tool to measure changes in pressure, volume, or temperature of a gas. In fact, this equation can be used even if you do not have all three of these variables mentioned inside your problem! Simply remove the P's, V's, or T's from the equation and keep on working!

Let's put this law into action:

If I have 30mL of a gas at a pressure of 3.0 atm and a temperature of 400°K, what will the pressure become if I raise the temperature to 450°K and decrease the volume to 15mL?

$$\frac{(3.0\text{ atm})(30\text{mL})}{400°\text{K}} = \frac{(P_2)(15\text{mL})}{450°\text{K}}$$

$$.225 = \frac{(P_2)\ (15\text{mL})}{450°\text{K}}$$

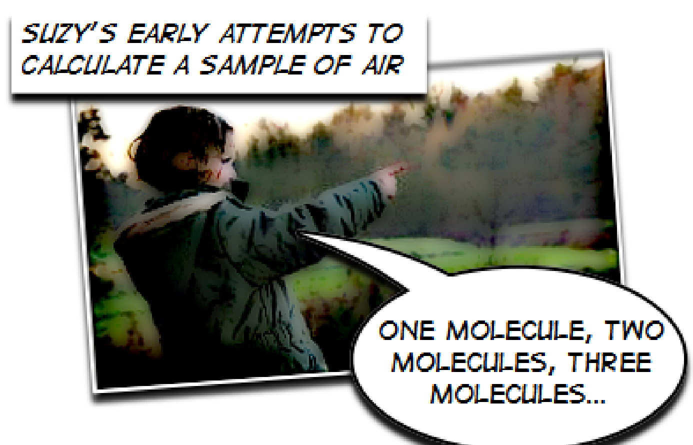
$$101.25 = (P_2)\ (15\text{mL})$$

$$6.75\text{atm} = P_2$$

Even though the Combined Gas Law seems to be very helpful, there is another factor we have to consider when looking at the volume, temperature, and pressure of a gas…

The number of gas particles

Since the increase of gas particles within a sealed object can affect its pressure, the number of gas particles must also be considered when calculating these equations. Think of it this way, when you run over a tire with a nail, the number of gas particles inside the tire begin to lower because they are escaping into the atmosphere. At the same time, your tire is getting flat (losing volume) because of the lowering of air pressure.



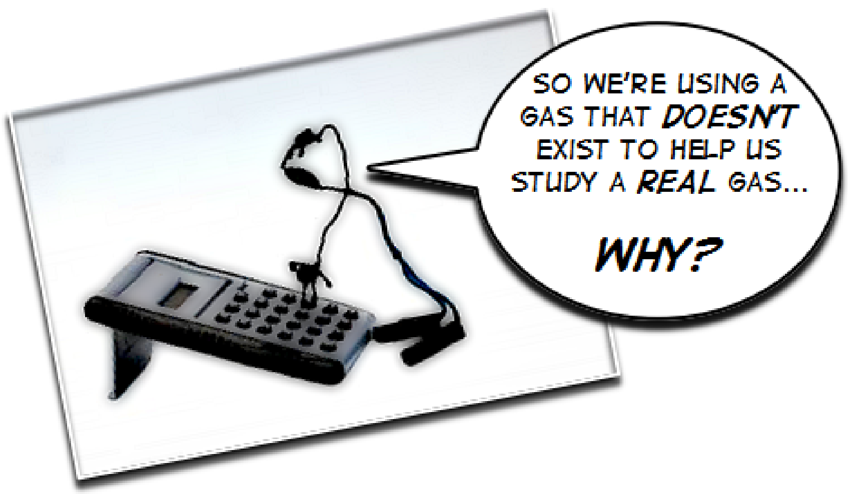
Besides, if you are going to be calculating how gases truly behave, the number of these particles will have to be measured in some way, right?

In order to measure these gases, scientists have created the idea of "ideal gases" which can be used in their calculations. Even though these hypothetical gases do not really exist, they are a useful tool to the chemist who is studying the all measurements within chemical reactions involving gases. This ideal gas, when kept at a constant temperature, obeys all of the gas laws we have learned about so far. Its importance to the chemist can easily be found within a single, very important law known as..

The Ideal Gas Law PV=nRT

Let's break down the Ideal Gas Law equation:

P = pressure (atm)
V = volume (L)
n = number of moles
T = temperature (°K)
R = ideal gas constant
(0.08206 L · atm/mole · °K) or
(8.314 L · kPa/mole · °K)



On many occasions, a standard temperature and pressure (STP) is expected within your calculations. Unless the problem says otherwise, you must assume a gas at STP has a pressure of 1 atm and a temperature of 273°K.

Now let's try an example:

If you have ever read the instructions on a container of any aerosol can, you will notice a huge warning not to place the can into an open flame. This is because the increased heat causes the atoms in the can to speed up, subsequentially increasing its pressure to the point of explosion. If you have a 2.0 liter canister which holds 4 moles of gas, and the temperature of the fire you place the can is 1400°C, what is the pressure inside the canister?

PV=nRT

P = ?
V = 2.0 L
n = 4 moles
R = 0.08206 L · atm/mole · °K
T = 1400°C = 1673°K
(Remember to convert all temperatures to Kelvin! So... T = 1673°K)

$$PV=nRT$$

$$(P)(2.0L) = 4\text{ moles}(0.08206\text{ L} \cdot \text{atm/mole} \cdot \text{K})(1673°\text{K})$$

$$(P)(2.0L) = 549.1\text{ L} \cdot \text{atm}$$

$$P = 274.6\text{atm}$$

Remember!

There are many more interactions between gas particles that cause them to move and bounce around each other. That is why we are looking at equations that describe an "ideal gas." In perfect situations, these gas laws would apply beautifully; however, as with all measurements done by humans, there is always some amount of error involved! They may not be perfect, but the gas laws you have learned about so far are pretty accurate for our purposes.

I need to share with you one more concept which pertains to "ideal situations." Unless you are purchasing a tank of pure gas (like hydrogen gas or oxygen gas), most of the gases we work with are a mixture of several compounds. Since each of these compounds has different kinds and amounts of atoms, each has a different amount of mass.

Each of these different masses has the ability to cause different amounts of pressure as they slam into the sides of a container, right? Therefore, we need a method of measuring the pressure of individual components within a mixture of gases. This is known as the **partial pressure** of a gaseous mixture. And so, I give you...

Dalton's Law of Partial Pressures

Dalton's Law of Partial Pressures simply states that the total pressure of a mixture of gases equals the sum of its individual partial pressures. To calculate Dalton's Law, all you need are the partial pressures of each gas within a mixture:

$$P_{total} = P_1 + P_2 + P_3 + \ldots\ldots$$

Yes! It is that easy! Let's give this a try:

A container is filled with a mixture of four gases. The partial pressures of these gases are as follows: Carbon dioxide = 0.45 atm, Oxygen gas = 0.50 atm, Methane = 0.25 atm, and Nitrogen gas = 0.60 atm. What is the total pressure inside the container?

$$P_{total} = 0.45 \text{ atm} + 0.50 \text{ atm} + 0.25 \text{ atm} + 0.60 \text{ atm} = \mathbf{1.8\ atm}$$

Now that you have a good understanding of how to calculate the relationships of gases to volume, temperature, and pressure – it's time to apply a little stoichiometry to your studies.

Since chemical reactions involve not just solids, but liquids and gases as well, you are going to need to learn how to use stoichiometry to figure out how much products will be produced from a certain volume of reactants.

There is only one simple conversion factor you need to know in order to complete this process! If you mastered your stoichiometry problems a couple of chapters ago this will be a breeze!

First of all, liquids and gases tend to be measured in liters (L) - not grams (g). **And the only conversion factor you need to learn is when you are converting between liters and moles at STP is...**

1 mole = 22.4 L

That's it! No kidding! With this simple conversion factor, all you really need to do differently in your conversion factors is to replace the molar mass components (grams/mole) with a molar volume (22.4 L/mole). <u>Just remember, this only works at STP!</u>

Let me show you how it works:

How many liters of water can be made from 30 liters of oxygen at STP from the following reaction?

$$2H_2 + O_2 \longrightarrow 2H_2O$$

$$\begin{array}{c|c|c|c} 30 \text{ L } O_2 & 1 \text{ mole } O_2 & 2 \text{ moles } H_2O & 22.4 \text{ L } H_2O \\ \hline & 22.4 \text{ L } O_2 & 1 \text{ mole } O_2 & 1 \text{ mole } H_2O \end{array}$$

Since this reaction is taking place at STP, you replace the molar masses with this new conversion factor.

Solving this problem is just like every other stoichiometry problem:

$$\begin{array}{c|c|c|c} 30 \cancel{\text{ L } O_2} & 1 \cancel{\text{ mole } O_2} & 2 \cancel{\text{ moles } H_2O} & 22.4 \text{ L } H_2O \\ \hline & 22.4 \cancel{\text{ L } O_2} & 1 \cancel{\text{ mole } O_2} & 1 \cancel{\text{ mole } H_2O} \end{array} = \mathbf{60 \text{ L } H_2O}$$

But what if either the reagents or the products are NOT measured in liters? How do we solve this problem?

Simple… We use the Ideal Gas Law PV=nRT

Let's try this method with a practice problem:

Using the same chemical formula for the production of water in the previous example, how many liters of water can be made from 200 grams of hydrogen at STP?

$$\frac{200\text{grams } H_2}{} \left| \frac{1 \text{ mole } H_2}{2.0 \text{ grams } H_2} \right| \frac{2 \text{ moles } H_2O}{2 \text{ moles } H_2} = \mathbf{100 \text{ moles } H_2O}$$

Now convert moles of water to liters of water:

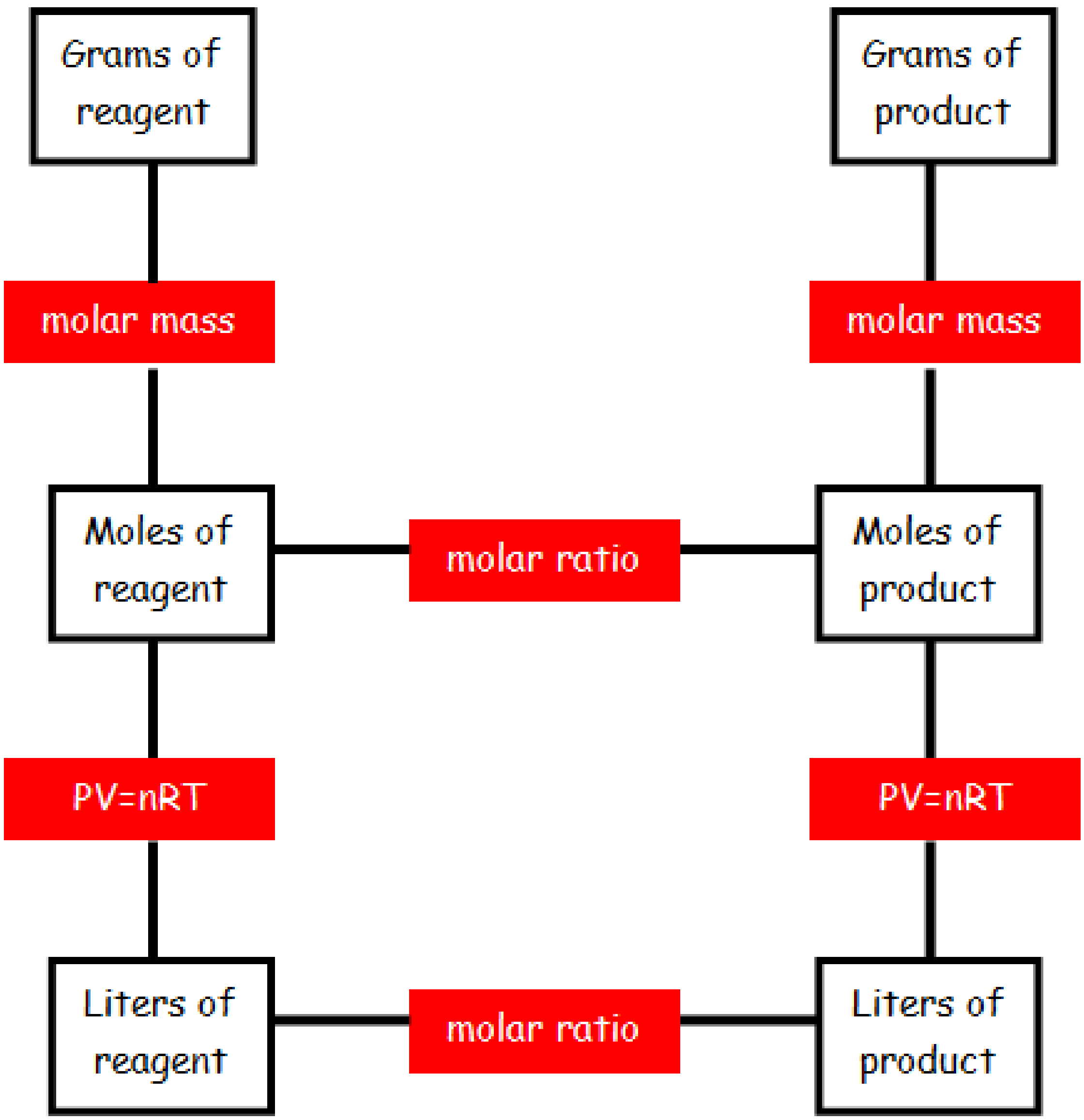
$$PV=nRT$$

$$(1\text{atm})(V)=(100\text{moles})(0.08206)(273°K)$$

$$\mathbf{V= 2{,}440.2\ L}$$

With a little practice, I am certain you will become very comfortable with these problems! Keep working and you'll do fine!

The "road map" below will help you determine which path to take when working through the word problems this week:



Combined gas law practice

1) A child has a toy balloon with a volume of 1.80 liters. The temperature of the balloon when it was filled was 20°C and the pressure was 1.00 atm. If the child were to let go of the balloon and it rose 3 kilometers into the sky where the pressure is 0.667 atm and the temperature is -10°C, what would the new volume of the balloon be?

Ideal gas law practice

2) How many moles of gas does it take to occupy 120 liters at a pressure of 3.2 atmospheres and a temperature of 340 K?

3) If I have a 62.5 liter container that holds 45 moles of gas at a temperature of 200°C, what is the pressure inside the container?

4) It is not safe to put aerosol canisters in a campfire, because the pressure inside the canisters gets very high and they can explode. If I have a 1.89 liter canister that holds 2 moles of gas, and the campfire temperature is 1400°C, what is the pressure inside the canister?

5) How many moles of gas are in a 25 liter scuba canister if the temperature of the canister is 300 K and the pressure is 200 atmospheres?

6) I have a balloon that can hold 100 liters of air. If I blow up this balloon with 3 moles of oxygen gas at a pressure of 1 atmosphere, what is the temperature of the balloon?

Dalton's Law of Partial Pressures Worksheet

7) If I place 2 moles of N_2 and 9 moles of O_2 in a 35 L container at a temperature of 25°C, what will the pressure of the resulting mixture of gases be?

8) Two balloons are connected with a sealable valve. The first balloon has a volume of 5 liters and contains nitrogen gas at a pressure of 0.75 atm. The second balloon has a volume of 8 L and contains oxygen gas at a pressure of 1.25 atm. When the valve between the balloons is opened and the gases are free to mix, what will the pressure be in the resulting mixture?

Gas stoichiometry practice

For this problem, assume that the reaction is being performed at a pressure of 1.0 atm and a temperature of 298 K.

9) Calcium carbonate decomposes at high temperatures to form carbon dioxide and calcium oxide:

$$CaCO_3 \rightarrow CO_2 + CaO$$

How many grams of calcium carbonate will I need to form 4.15 liters of carbon dioxide?

Chapter 27

Okay. By now you have a very good understanding of what atoms can do, how they move around, and how they bond with each other to create so many cool things! But all of this movement needs some kind of energy, right? We can't simply get all of this motion without some kind of energy. Therefore, it is time we start looking at the study of energy which chemists like to call...

Thermodynamics

Thermodynamics is the study of energy; and, for our purposes, we will be studying how energy is stored within molecules, how it can be transferred from one substance to another, and how this process is able to produce heat.

Since we are talking about heat, let's clear up one thing...

Heat and Temperature are not the same things!

Heat is the movement of energy from one thing to another through the motion of molecules.

Temperature is the measure of energy found within an object after heat has been transferred.

It is very common for people to say, "Turn up the temperature! It's cold in here." Unfortunately, this is entirely wrong. What you really need when your house starts to get cold is more heat – a movement of energy! This is a common mistake people make but don't be too hard on them. Not all of them have been lucky enough to study thermodynamics, right?

Since we are on the topic, let's get into the units for calculating energy. The main unit of energy is known as the **calorie (cal).** This unit is that amount of energy required to heat 1 gram of water by 1°C.

This "calorie" is not to be confused with the calories in our food. Yes, the calories within our food are a measure of the amount of energy it contains; however, food is actually measured in units of 1000 calories (a kilocalorie), which is commonly known as the Calorie (Cal).

*** You noticed the difference in the capitalization between the calorie and the Calorie, right? Keep an eye out for that difference within the labels of your preprocessed foods!

Even though the calorie (and the Calorie) is typically used in our everyday life, chemists tend to use the metric system when measuring energy. The metric unit of energy is known as the **joule (J).** The accepted conversion rate between joules and calories is this:

1 calorie = 4.184 J

Typically you will find yourself measuring large amounts of energy; therefore, be on the lookout for measurements in kilojoules (kJ) which is equal to 1000 joules of energy.

Let's get back to our study of energy transfer…

Everything in the universe is ruled by the transfer of energy from one source to another. And, luckily for us, this energy transfer takes place through chemical reactions!

Every single chemical reaction requires some form of energy change. The reactants of a reaction must rearrange themselves to form new products. Naturally, this "rearrangement" of atoms requires some form of energy. And, the formation of new bonds within the products of a reaction releases some of this energy.

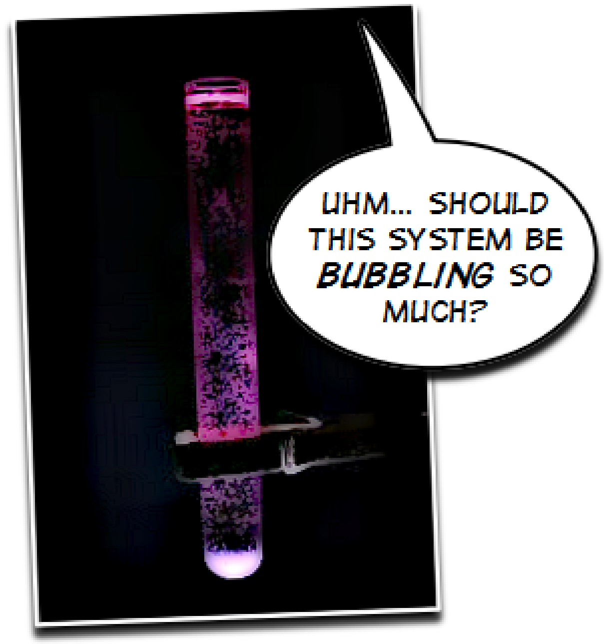
However, the amount of energy required by the reactants is not always the same that is released from the products. Take a look at the following charts:

Exothermic Reaction	Reactants	→	Products
	Bonds broken		Bonds formed
	Energy absorbed	<	**Energy released**

When the amount of energy absorbed by the reactants is less than (<) the energy released by the products, the reaction is said to be **exothermic**. These types of chemical reactions release energy in the form of heat or light into its surrounding area.

Endothermic Reaction	Reactants	→	Products
	Bonds broken		Bonds formed
	Energy absorbed	>	**Energy released**

Chemical reactions that tend to absorb more energy than the energy released by its products are known as **endothermic reactions.**



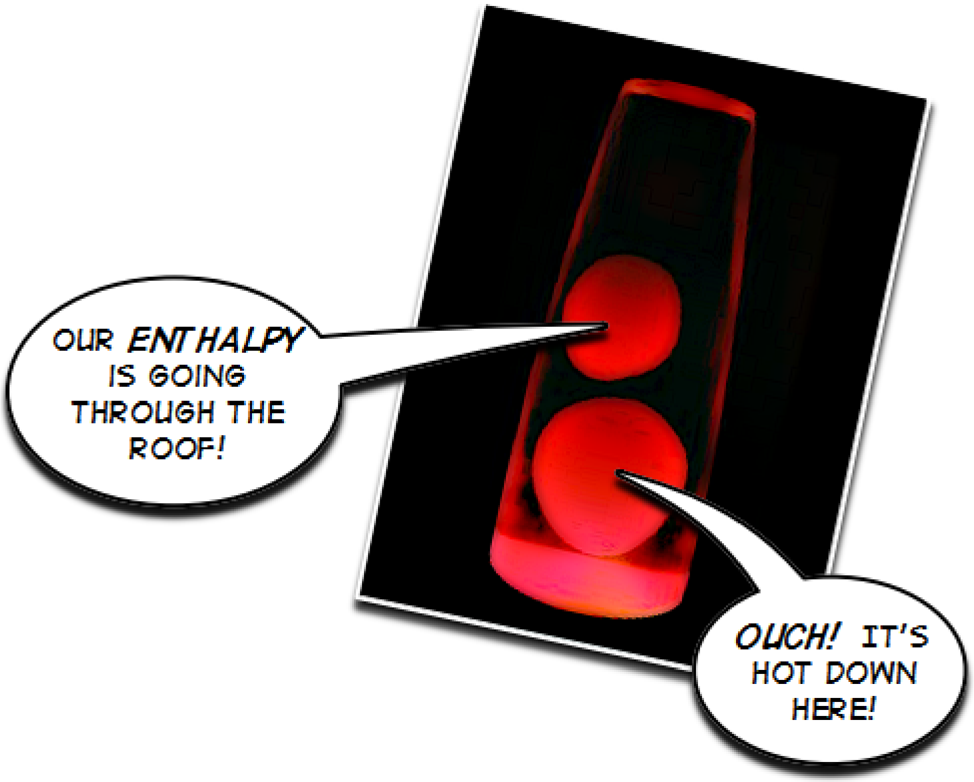
When they are talking about energy transfer, chemists like to call chemical reactions a **system**. The environment that surrounds these systems is known as the **surroundings**. We'll be using these terms throughout the next few chapters.

Throughout this chapter and the next we will be looking at two laws which apply to the study of energy and energy transfer. These laws are known as the **1st and 2nd Laws of Thermodynamics**. There is a lot of information to absorb in these two chapters and I'm certain you are going to do just fine. Let's get started!

1st Law of Thermodynamics

This law is also known as **The Law of Conservation of Energy**. It states that energy is never lost or gained, it only changes form. During an exothermic reaction, excess energy from the system is being released into the surroundings. Even though the total energy of the system decreases, the energy of the surroundings increases. During an endothermic reaction, the opposite takes place: the total energy of the system increases as it absorbs energy from its surroundings.

The total amount of heat a molecule contains is known as **enthalpy (ΔH)**. Unfortunately, it is impossible for chemists to be able to calculate how much enthalpy a molecule actually has! The transfer of heat from one system to another depends upon a huge amount of variables – too many to count!



Therefore, chemists study how much of a system's enthalpy <u>changes</u> when heat is removed or added to it.

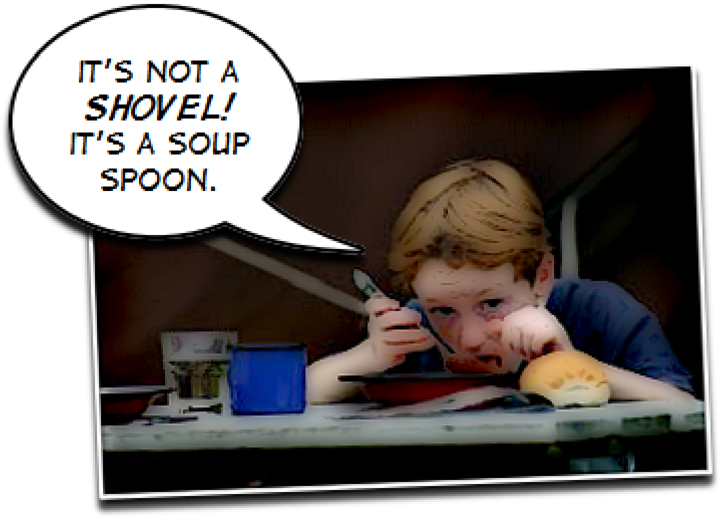
By far, the easiest way to change the amount of enthalpy in a system is by heating or cooling the system itself.

But how can you calculate how much energy is needed to heat or cool something to a certain temperature? That is a good question!

If you want to heat up a bowl of soup on the stove, you probably didn't pour out tiny amounts of soup into several tiny pots and heat them all up at the same time, right? (I hope not... that would seem a little strange.) Even though this procedure would allow you to heat all of the soup much faster, it would result in a massive clean-up operation of several tiny pots. The size of the pot will affect the rate in which your soup will heat up. A larger bowl will require a longer time in order to heat up. But the price of this "convenience" is that it takes longer for the stove to heat your soup!

What does all of this have to do with measuring the change in enthalpy?

Well, in order to measure the amount of energy a system has you need to understand a few things about the system... just like when you try to heat up your soup.



First, you need to know the size of the container you are going to use in order to heat the soup. This mass (m) is measured in grams.

Second, you probably should know the temperature of your final product. This requires you to know the change in temperature of the system from start to finish (ΔT).

And finally, you have to know what exactly how much heat can be absorbed by your item (C_p).

If you put all of these symbols together, you can calculate the change in enthalpy while heating a pure substance:

$$\Delta H = mC_p\Delta T$$

ΔH = the change in enthalpy ($+\Delta H$ is for endothermic and $-\Delta H$ is for exothermic)
m = the mass of the object that is being heated in grams
C_p = the **specific heat (heat capacity)** needed to heat the object by 1°C; every object you heat or cool will have its own unique C_p value
ΔT = the change in temperature in °C

Okay! Let's try out this new equation:

How much energy is needed to raise the temperature of 50g of methanol (C_p = 2.51 J/g°C) from 20°C to 70°C?

$$\Delta H = mC_p\Delta T$$
$$\Delta H = (50 \text{ grams})(2.51 \text{ J/g°C})(50°\text{C})$$
$$\Delta H = 6275 \text{ J}$$

Let's do one more:

If 245J is required to change the temperature of 14.4g of chromium by 38°C, what is the specific heat capacity of chromium?

$$\Delta H = mC_p\Delta T$$

To solve for C_p we rearrange the equation like this:

$$C_p = \frac{\Delta H}{m\Delta T}$$

$$C_p = \frac{245 \text{ J}}{(14.4\text{g})(38°\text{C})}$$

$$C_p = 0.448 \text{ J/g°C}$$

Remember! We are only calculating the change in enthalpy of pure substances. The specific heat of mixtures would be far too difficult to describe in this book!

Heating or cooling a pure substance to change its enthalpy is easy! But what about the transfer of energy during a chemical reaction? In order to calculate this, we are going to need a couple of "tools" to help us out:

Heat of reaction (ΔH_{rxn})

Measures the change in enthalpy during a chemical reaction

Heat of formation (ΔH^0_f)

This is the change in enthalpy when a compound is formed from its elements

You can calculate the **heat of reaction** by using the **heat of formation** of the products **and** reactants. Here's how:

$$\Delta H_{rxn} = \Sigma\Delta H^0_f \text{ (products)} - \Sigma\Delta H^0_f \text{ (reactants)}$$

Before we start using this equation, let me explain a couple of things...

First, the Σ symbol within this equation means "the sum of." So, what this equation is really telling us is that the heat of reaction is equal to the difference between all the heats of formation of the products and all of the heats of formation of the reactants.

Next, it is important for you to know that <u>all substances have their own unique heat of formation.</u> The only exception is with elements in their standard state (like O_2, N_2, and Cl_2). Elements in their standard state have a heat of formation equal to zero (0).

Perhaps a sample problem will make all of this a little clear:

Calculate ΔH_{rxn} of the following reaction using its heats of formation:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_20$$

In order to solve this problem, you need the heats of formation for each molecule:

C_3H_8 = -103.8 kJ/mole
O_2 = 0 kJ/mole (remember O_2 is in its standard state so the value is zero)
CO_2 = -393.5 kJ/mole
H_20 = -242.0 kJ/mole

Now all you need to do is plug these heats of formation into our equation:

$$\Delta H_{rxn} = \Sigma\Delta H^0_f \text{ (products)} - \Sigma\Delta H^0_f \text{ (reactants)}$$

$$\Sigma\Delta H^0_f \text{ (products)} = [(3 \times -393.5 \text{ kJ/mole}) + (4 \times -242.0 \text{ kJ/mole})]$$

$$= 212.5 \text{ kJ/mole}$$

$$\Sigma\Delta H^0_f \text{ (reactants)} = [(1 \times -103.8 \text{ kJ/mole}) + (5 \times 0 \text{ kJ/mole})]$$

$$= -103.8 \text{ kJ/mole}$$

$$\Delta H_{rxn} = 212.5 \text{ kJ/mole} - (-103.8 \text{ kJ/mole})$$

$$\Delta H_{rxn} = 316.3 \text{ kJ/mole}$$

That wasn't so hard, was it? You've done an excellent job! Now spend a little time practicing your newfound skills before you tackle this week's lab activity!

Enthalpy practice

1) Calculate the amount of heat needed to increase the temperature of 127g of water from 20°C to 46°C.

2) Calculate the specific heat capacity of copper given that 294.75 J of energy raises the temperature of 15g of copper from 25°C to 60°C.

3) 419 J of energy is required to raise the temperature of aluminum from 15°C to 35°C. Calculate the mass of aluminum.
(C_P of aluminum = 0.90 J/g · C)

4) The initial temperature of 71g of ethanol was 22°C. What will be the final temperature of the ethanol if 3240 J was needed to raise the temperature of the ethanol? C_P of ethanol is 2.44 J/g · C

5) Paraffin wax is sometimes incorporated within sheetrock to act as an insulator. During the day, the wax absorbs heat and melts. During the cool nights, the wax releases heat and solidifies. At sunrise, a small hunk of solid paraffin within a wall has a temperature of 288K. The rising sun warms the wax, which has a melting temperature of 354K. If the hunk of wax has 0.257g mass and a specific heat capacity of 1.80J/g ·K, how much heat must the wax absorb to bring it to its melting point?

6) At some point, all laboratory chemists learn the same hard lesson: Hot glass looks just like cold glass. Heath discovered this when he picked up a hot beaker someone left on his bench. At the moment Heath grasped the 413K glass beaker, 567J of heat flowed out of the beaker and into his hand. The glass of the beaker has a heat capacity of 0.57J/g ·K If the beaker was 400K the instant Heath dropped it, then what is the mass of all the pieces of broken beaker now littering the lab floor?

7) The temperature of a piece of Metal X with a mass of 95.4g increases from 25.0°C to 48.0°C as the metal absorbs 959 J of heat. What is the specific heat of Metal X?

8) When 435 J of heat is added to 6.2 g of olive oil at 21°C, the temperature increases to 85°C. What is the specific heat of the olive oil?

Chapter 28

2^{nd} Law of Thermodynamics

The second law of thermodynamics states that energy, if concentrated in an area, will spontaneously spread out unless a force keeps it from doing so. A classic example of the 2^{nd} Law in action would be heat energy always moving from a warmer area to a colder area.

I'm certain you have known about this since you were a child. If you touch a hot object, its heat transfers into you and causes a burn. This applies to all situations involving heat transfer. For example, the next time your mom says, "Close that door! You're letting in the cold," you will know this to be impossible. It is the heat within your house that would move towards the cold air outside.

(I recommend waiting until she is in a good mood before you inform her of her mistake!)

Chemists can measure the 2^{nd} Law using a "tool" that is known as:

Entropy

Entropy (ΔS) is a measure of the amount of energy that spreads throughout the surroundings of a system. Much like enthalpy, there are far too many variables to determine a system's actual entropy; however, you can easily measure the change in entropy of a system. The units for calculating entropy are joules per temperature (°K) which looks like J/K. You will be needing this in a little bit...

You may have noticed that the word "spontaneously" was underlined when we were discussing the 2nd Law of Thermodynamics. Chemists refer to **spontaneous processes** as any naturally occurring actions that exist without the addition of any extra energy. For example, an odor moving through a room through diffusion, ice melting in lukewarm water, iron rusting, and objects moving downhill are all spontaneous processes.

As you can guess, not all processes within the world occur naturally. These **non-spontaneous processes** require some form of effort to make them occur. The heating of the metal filaments within the light bulb and the production of chemical energy from photosynthesis are examples of non-spontaneous processes because energy is required for these to occur.

The cool thing about these two different types of processes is this:

Spontaneous processes cause non-spontaneous processes to take place!

Imagine a water wheel attached to a generator on the banks of a river. The water moving downstream is spontaneous while the electrical energy generated by the moving water wheel would be non-spontaneous.

This is where we bring back an old familiar phrase...

Everything in the universe tends to move towards the lowest energy possible.

Since spontaneous processes occur naturally, they tend to require less energy than non-spontaneous processes. Therefore, nature prefers spontaneous process over non-spontaneous processes.

Now let's use this knowledge to study how spontaneity affects chemical reactions!

There are two reasons that chemical reactions will occur spontaneously:

The products of spontaneous chemical reactions have less energy than the reactants.

(Remember – all physical systems want to be in the lowest energy state possible!)

The energy within the products of a spontaneous chemical reaction is spread out more.

(Remember the 2nd Law – energy will spontaneously move from high to low concentrations causing an increase in entropy. This means that exothermic reactions are favored over endothermic reactions.)

You may be asking yourself why exothermic reactions are favored in the natural world?

Since nature favors chemical reactions that result in the lowest energy state possible, systems which undergo exothermic reactions release energy into their surroundings. This reduces the total energy of the system while following both the 1st and 2nd Laws of Thermodynamics.

All of this leads into another question:

How can you predict if a process is going to be spontaneous?

Prediction #1:
The 2nd Law of Thermodynamics states that energy will flow spontaneously throughout a system unless a force acts to stop it. Since entropy is used to measure how much a system's energy will spread throughout its surroundings, entropy will always increase when a spontaneous process happens. In short, when you calculate the amount of entropy within a system and ΔS ends up being a positive number, the process is spontaneous.

Prediction #2:
Phase changes that result in energy being released results in a positive ΔS. This means during melting or boiling, ΔS will be positive (since extra energy is being released into the surroundings); and, during freezing or condensing, ΔS will be negative (since energy is being absorbed primarily into the system.)

Prediction #3:
Dissolving solids and liquids results in a positive ΔS. In order for solids and/or liquids to be dissolved, they have to absorb energy in order to break the bonds that hold them together. This increase of energy allows these particles to move freely through the system while spreading their energy around as well.

Prediction #4:
Whenever a gas is dissolved into a liquid, like when carbon dioxide is dissolved into soft drinks, a lot of the energy that once existed within the gas particles is now being used to keep the liquid molecules together. Therefore, energy is not spreading out and ΔS will be negative.

Prediction #5:
During a chemical reaction, if a gas is created from non gases, ΔS is positive. More energy is being spread throughout the surroundings along with the energized gas particles!

The opposite is true as well – if non gases (like liquids or solids) are created from a reaction of gases during a chemical reaction, ΔS will be negative. This is because more energy is needed to hold the new products together.

Prediction #6:
If the amount of moles of gas during a chemical reaction is decreased - much like in the following reaction $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, ΔS will be negative.

The (g) after each gas means that the H_2 and O_2 are in a gaseous state while the (l) identifies H_2O as being in a liquid state.

The opposite is true for this prediction as well - If the number of moles of gas produced is increased, ΔS will be positive.

A brief review so far…

- Nature favors exothermic reactions because they result in the lowest energy state possible.
- Nature also favors an increase in entropy as this follows the 2nd Law of Thermodynamics.
- Enthalpy is how much energy is contained within a system. If the amount of enthalpy is negative ($-\Delta H$), the system is releasing energy; if the enthalpy is positive ($+\Delta H$), the system is absorbing energy.
- Entropy is the amount of energy that spreads out from a system. If the amount of entropy is negative ($-\Delta S$), the system is not releasing energy into the surroundings; if the entropy is positive ($+\Delta S$), the system is releasing energy into its surroundings.

Even though we may know how much energy is contained within a system (enthalpy ΔH) and the amount of energy that spreads out from a system (entropy ΔS) there still is one little problem: <u>You can't determine if a process is going to be spontaneous if you only know the enthalpy or entropy of a system.</u>

You must use BOTH of these measurements to determine if a process is going to be spontaneous!

Entropy and enthalpy <u>have</u> to work together and the simplest way to connect these two measurements within a system is to calculate the **free energy (ΔG)** that occurs during a reaction. Free energy is the amount of energy that is available by a system to do some form of work. It is calculated using the following equation:

$$\Delta G = \Delta H - T\Delta S$$

Let's break down this equation:

$$\Delta G = \Delta H - T\Delta S$$

The total amount of energy available within a system (ΔG) is equal to the total amount of energy available (ΔH) minus the amount of energy released to the surroundings (ΔS) which is based upon its temperature (T).

Remember! The purpose of calculating ΔG is to determine if the reaction is spontaneous or not. Therefore, if...

$-\Delta G$ = spontaneous

$+\Delta G$ = not spontaneous

Here's a sample problem for you:

Is the following reaction spontaneous at 450°K? (ΔH = -91.8kJ and ΔS = -197 J/K)

$$N_2 + 3H_2 \rightarrow 2NH_3$$

As we work through this problem, we are going to need to be certain that all of our units are the same!

$$\Delta H = -91.8kJ = -91,800\ J$$

$$T\Delta S = (450°K)(-197\ J/K) = -88,650\ J$$

Now that we have our variables converted to their proper units, all you need to do is plug them into your equation like this:

$$\Delta G = \Delta H - T\Delta S$$

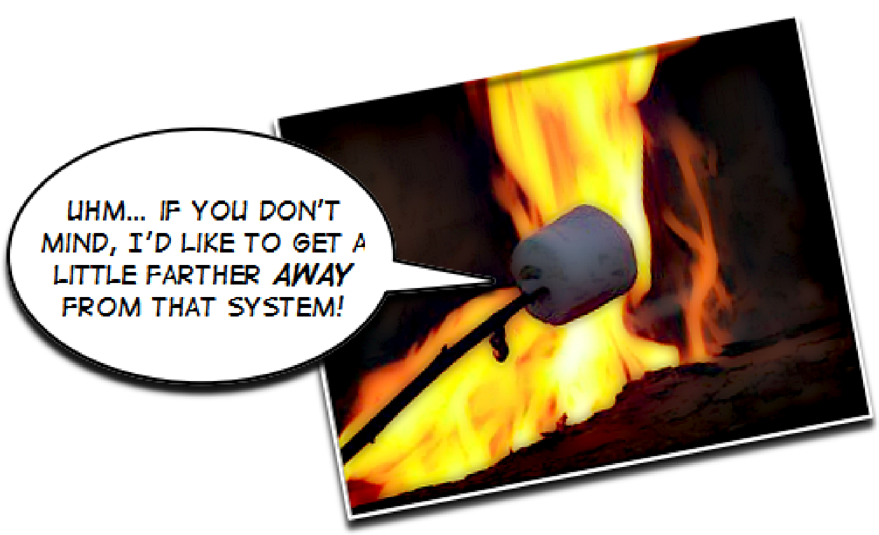
$$\Delta G = -91,800\ J - (-88,650\ J) = \mathbf{-3,150\ J}$$

Since ΔG is a negative amount, the reaction is spontaneous at 450°K

It is possible for a system to be releasing energy ($-\Delta H$) but not spreading out its energy within the system ($-\Delta S$) <u>because the energy that is being released is being absorbed by its surroundings!</u>

This follows the 1st Law of Thermodynamics -

Energy cannot be created or destroyed, but it can be transferred from one place to another!



The following table may help you out when practicing your free energy problems:

	$+\Delta H$	$-\Delta H$
$+\Delta S$	$-\Delta G$; reaction is spontaneous at high temperatures	$-\Delta G$; reaction is always spontaneous
$-\Delta S$	$+\Delta G$; reaction is never spontaneous	$-\Delta G$; reaction is spontaneous at low temperatures

Entropy and free energy practice

1) Predict which of the following reactions has a **negative** entropy change:
   a) $2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$
   b) $Ba^{+2}(l) + SO_4^{-2}(l) \rightarrow BaSO_4(s)$
   c) $2H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g)$

2) Predict which of the following two reactions has a **positive** entropy change:
   a) $2\ N_2(g) + O_2(g) \rightarrow 2\ N_2O(g)$
   b) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   c) $Zn(s) + 2\ HCl(l) \rightarrow ZnCl_2(l) + H_2(g)$

3) Consider the reaction below:

$$CO_2\ (g) + H_2O\ (g) \rightarrow HCOOH\ (l) \qquad \Delta H = -150\ kJ$$

Which of the four conditions below best represents the reaction being spontaneous.
   a) The reaction would be spontaneous at all temperatures.
   b) The reaction would be spontaneous at lower temperatures but not at higher temperatures.
   c) The reaction would be spontaneous at higher temperatures but not at lower temperatures.
   d) The reaction would not be spontaneous at any temperature.

For each of the following reactions, indicate whether you would expect the entropy of the system to increase or decrease, and explain why.

4) $CH_3OH(l) \rightarrow CH_3OH(g)$

5) $N_2O_4(g) \rightarrow 2NO_2(g)$

6) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2\ (g)$

7) $2NH_3(g) + H_2SO_4(l) \rightarrow (NH_4)_2SO_4(l)$

8) How does entropy increase or decrease with increase in temperature?

9) Metallic iron is produced by a high temperature process through a series of reactions, one of which is shown below. Find the various parameters requested using the data which can be found below.

$$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$$

What is the standard entropy change (ΔS) for the reaction?

$$\Delta S = (\text{sum of } \Delta S_{\text{products}}) - (\text{sum of } \Delta S_{\text{reactants}})$$

Use the following table:

Species	**S (J/K·mol)**
$CO_2(g)$	213.7
$CO(g)$	197.5
$H_2O(g)$	188.7
$Fe_3O_4(s)$	145.3
$H_2O(l)$	69.94
$Fe_2O_3(s)$	87.40
$CH_4O(l)$	127
$C_2H_6O(l)$	161
C_8H_{18} (l)	329.3
C (diamond)	2.439

Chapter 29

You spent the last couple of chapters learning if a chemical reaction would occur spontaneously in nature. With this knowledge you can now learn how fast these reactions will take place AND how we can speed them up! This branch of chemistry is known as the study of...

Kinetics

Kinetics is the study of the rates of chemical reactions and explains how we can speed them up. The rate of a reaction is known as the speed in which the reactants are converted into products.

In order to calculate the rate of a chemical reaction you need to identify the amount of reactant that is used within a reaction and the amount of time it took for this to happen. Dividing these two pieces of data will give you the reaction rate. Here's what it looks like:

$$\text{Rate of reaction} = \frac{\text{amount of reactant used}}{\text{time for this reactant to be used}}$$

Typically, we measure the amount of reactant used in molarity (M) per second (M/s). This calculation seems harmless, right? Well it is! Let's try an example:

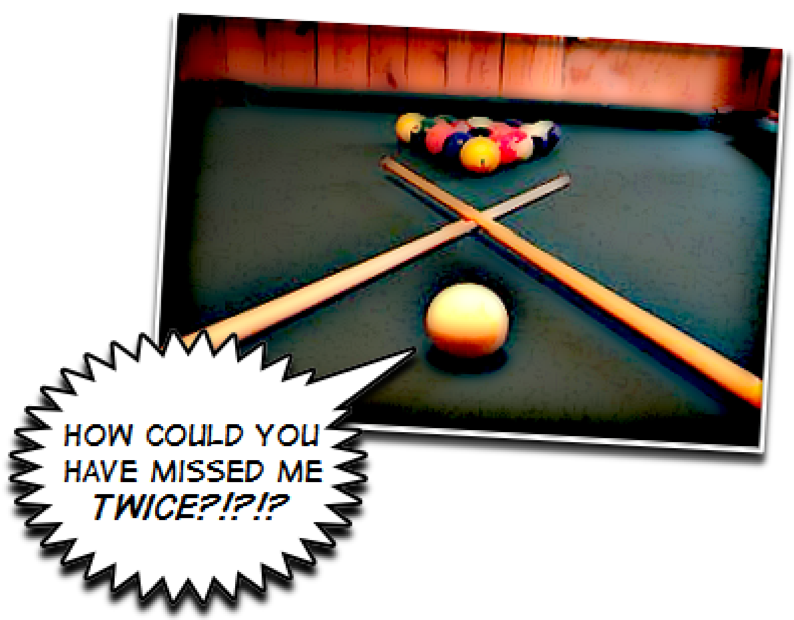
When starting a reaction, the concentration of one of the reagents was 1.80 M. After two minutes, the concentration of this reagent was 0.65 M. Using this data, determine the average rate of this reaction in M/sec.

Rate of reaction = $\dfrac{1.15 \text{ M}}{120 \text{ seconds}}$ ← The amount used is 1.80M - 0.65M

= 0.0096 M/s

So what is causing this reaction rate anyway?

The rate of a chemical reaction is based upon a well-tested idea by chemists called **Collision Theory**. This theory states that the speed in which a chemical reaction takes place is based solely on how hard the reactants slam into each other and the directions they are traveling when this happens.

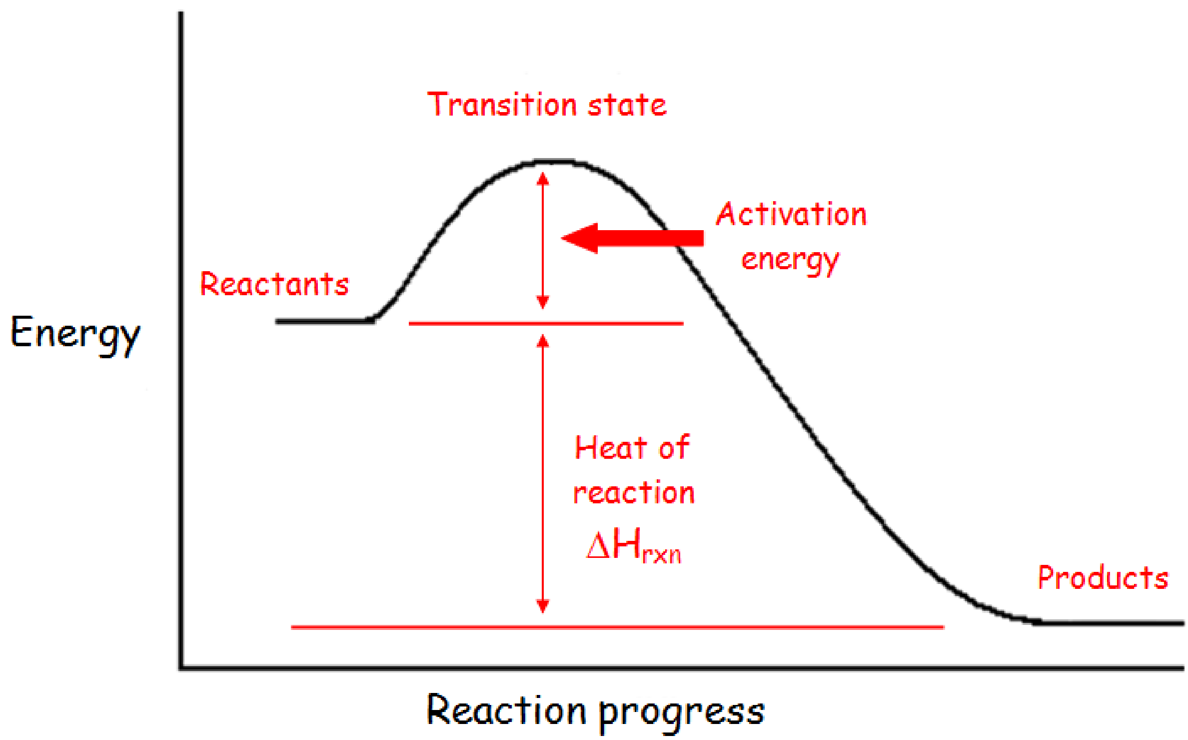


This makes sense, doesn't it? Reactants have to bump into each other in order for their energies to be transferred to break the bonds that hold them. And if they don't bump into each other in the proper way, the desired products will not be formed.

It's almost like bowling. If you hurl your bowling ball down the lane as fast as you can, its energy will likely create an "explosion" of pins! BUT… if you don't hit the center pin at the correct angle, it will be very difficult to knock over all of them at once.

Of course, you don't want the "pins" in your chemical reaction to go flying all over the place! You want them to stick together in new ways to create new products!

A more "professional" way to think about collision theory is by using something called an energy diagram:



This is an energy diagram for an exothermic reaction. If the reactants start slamming into each other in the right directions and with the right amount of energy (known as **activation energy** as seen in the graph), the reaction will begin to form a new product.

The point in which the new product is formed is known as the **transition state**. Once the transition state has been reached, new products will continue to form. As you learned in the last chapter, the **heat of reaction** ΔH_{rxn} is the change in enthalpy between the reactants and the products. You can see that the total energy of the products is much lower than that of the reactants within this graph.

So how do you speed up these chemical reactions?

You learned about three ways to speed up a chemical reaction back in Chapter 21. Now it's time to look at these in a little more depth. Let's look at these original three and add a couple more to the list:

Increase the temperature
This should be easy to understand. The more energy a fast moving particle contains, the harder it will slam into other particles. Heating the particles will give them all the energy they need!

Increase the surface area of the reactants
Most reactants are very large groups of molecules bound together. Think about a single grain of salt. There are close to 6.0×10^{17} molecules of NaCl in a single grain of salt! That's 600,000,000,000,000,000 molecules of NaCl! If you break apart these massive groups of molecules (by grinding them into a fine powder), you will increase the amount of reactants that could possibly react with each other.

Increase the concentration of the reactants

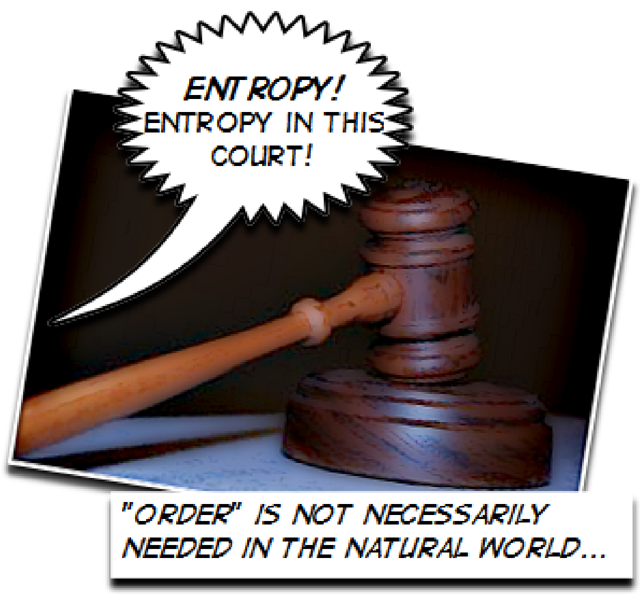
This is an easy one too! If you increase the number of reactants, it is likely they will slam into each other more frequently thus speeding up the reaction!

Use unstable reactants

Unstable reactants such as halogens and alkali metals tend to move their electrons around much easier than other elements which make them much more reactive. The more likely an element is to react with other elements, the faster a reaction may take place.

Use a catalyst

A catalyst is anything that can be used to create the products of a reaction without being broken apart in the process. This helps to speed up a reaction as the amount of catalyst will never be reduced.



In many reactions, the rate of reaction changes as the reaction progresses. At first, the rate of reaction is pretty fast; however, after a short burst of rapidly creating a product, the rate of reaction decreases to zero (at which point the reaction is complete). In order to know what your reactants are doing within a reaction, chemists have created **reaction rate laws** which help to determine how the rate of reaction varies as the reaction progresses.

The **general rate law** for a chemical reaction that follows a simple formula like A + B → C looks like this:

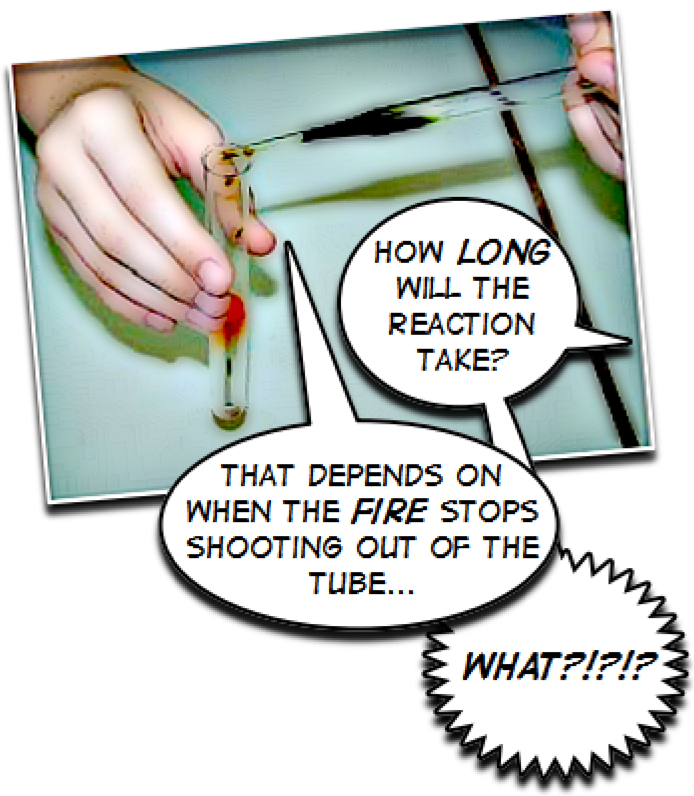
$$\text{Rate} = k[A]^x[B]^y$$

[A] and [B] = the molarities (M) of the compounds A and B

k = the **rate constant** (this number is unique for each reaction and is based upon the temperature of the reaction)

x and y = the **reaction orders**; these exponents define how the rate is affected by the reactant concentration (you'll see how this works very quickly!)

Setting up a general rate law for a reaction requires that you actually run the reaction and calculate its speed!



Let's see how you can set up one of these equations using a set of data on three experiments ran with the following reaction:

$$O_2 + 2NO \rightarrow 2\,NO_2$$

Experiment	Initial concentration of O_2 (M)	Initial concentration of NO (M)	Reaction rate (M/sec)
1	0.10 M	0.10 M	2.00×10^{-3}
2	0.20 M	0.10 M	4.00×10^{-3}
3	0.10 M	0.20 M	4.00×10^{-3}

Step 1: Write the general rate law:

$$\text{Rate} = k[O_2]^x[NO]^y$$

Step 2: Calculate the x and y exponents:

You have to experimentally determine the reaction rate in order to determine the rate law! That is why we have to look at the above data to determine our rate law.

Below is a chart to help you calculate the x and y exponents:

Order of Reactant	Change in rate of reaction when concentration <u>is doubled</u>
First order	Rate doubles
Second order	Rate quadruples
Zero order	Rate remains unchanged

The order of NO in this reaction:

Look at the sample data in the above example. The reaction rate stays the same in experiments 2 and 3 when the concentration of NO doubles from 0.10 M to 0.20 M. Therefore, NO is considered to be zero order in this reaction.

Take special note of the concentration of O_2 in experiments 2 and 3. The concentration of O_2 remains constant while NO doubles. This is what you want to look for during your practice problems this week!

The order of O_2 in this reaction:

The reaction rate doubles in experiments 1 and 2 when the concentration of O_2 doubles from 0.10 M to 0.20 M. Therefore, O_2 is considered to be first order in this reaction.

*If the reaction rate quadruples when the concentration of the reactant is doubled (and the remaining reactants remain constant), the value for the exponent would be considered to be second order. This does not happen within our set of data (but it could in future problems!)

The rate law for this equation would be:

$$\text{Rate} = k[O_2]^1[NO]^0$$

If you were to describe the reaction order of this process you would say:

The reaction is first order in O_2, and zero order in NO; therefore, by adding up both of the exponents, <u>the reaction is first order overall</u>.

So why do we need to know all of this?

Well, chemists tend to be asked a lot of technical questions like:

"How long will it take for 12 moles per liter of reactant 'A' to be used up?"
and
"What is the concentration of "A" after 2 minutes of the reaction?"

In order to answer these types of questions, all we need to do is integrate the amount of time it takes for these reactions to occur (rate constant; k) into the equation and we can solve these types of problems! Cool, huh?

Here's an example:

The reaction A + B → C is first order in A and second order in B. Given that the rate constant of this reaction is 2.48×10^{-4} $/M^2 \cdot sec$, what is the rate of reaction when the concentration of A is 0.04M and the concentration of B is 0.07M?

** I hope you will forgive me for not explaining why the units for the rate constant seem a little odd. To be honest, the rate constants for zero-, first-, second-, and third-order reactions all have different units! The reason for this would require way too much explanation for now!

<u>Step 1: Write the general rate law:</u>

$$Rate = k[A]^x[B]^y$$

<u>Step 2: Calculate the x and y exponents:</u>

This information is given to you within the problem. [A] is first order and [B] is second order; therefore, the general rate law will look something like this:

$$Rate = k[A]^1[B]^2$$

<u>Step 3: Enter the values and calculate the reaction rate:</u>

$$Rate = (2.48 \times 10^{-4} \ /M^2 \cdot sec)\ (0.04M)^1(0.07M)^2$$

$$= 1.2 \times 10^{-6} \ M/sec$$

Remember! This reaction rate is only happening at one particular time within the reaction because...

The concentration of the reactants change over time throughout the reaction!

Now that you are a master at calculating how fast a reaction will take place, let's see how you do when we turn your world upside down. Stay tuned!

Kinetics practice problems

Answer questions 1-8 using the following information:

$$2ICl + H_2 \rightarrow I_2 + 2HCl$$

Experiment	[ICl]	[H_2]	Initial Rate
1	0.10 mol/L	0.01 mol/L	0.002 mol/L · s
2	0.20 mol/L	0.01 mol/L	0.004 mol/L · s
3	0.10 mol/L	0.04 mol/L	0.008 mol/L · s

1) Examine the [ICl] concentration between experiment 1 and 2. What happens?

____

2) What happens to the [H_2] between experiment 1 and 2?

____

3) What happens to the initial rate between experiment 1 and 2?

____

4) What happens to the [H_2] between experiments 1 and 3?

____

5) What happens to the [ICl] between experiments 1 and 3?

____

6) What happens to the initial rate between experiments 1 and 3?

____

7) Calculate the rate law of this reaction:

____

8) Determine the rate law constant (k) for this reaction:

____

9) What is the overall reaction order for the following rate law:

$$Rate = k[A]^1[B]^1$$

a) 1 b) 2 c) 3 d) 4

10) Write the rate constant from the information below:

$$A + B \rightarrow Products$$

Experiment	Initial [A]	Initial [B]	Initial Rate
1	0.01 mol/L	0.03 mol/L	2.4×10^{-4} mol/L · s
2	0.03 mol/L	0.03 mol/L	7.2×10^{-4} mol/L · s
3	0.01 mol/L	0.06 mol/L	2.4×10^{-4} mol/L · s

11) NO reacts with chlorine in a gas phase reaction to form nitrosyl chloride, NOCl. From the following experimental data, determine the form of the equation that describes the relationship of reaction rate to initial concentrations of reactants.

$$2NO + Cl_2 \rightarrow 2NOCl$$

Experiment	**Initial [NO]**	**Initial [Cl_2]**	**Rate of reaction**
1	0.50 M	0.35 M	1.14 M/hr
2	1.00 M	1.00 M	9.12 M/hr
3	1.00 M	0.35 M	4.56 M/hr

a) Rate = $k[NO]$
b) Rate = $k[NO][Cl_2]$
c) Rate = $k[NO]^2$
d) Rate = $k[NO]^2[Cl_2]$
e) Rate = $k[NO]^2[Cl_2]^2$

Chapter 30

You have done an excellent job at figuring out the different kinds of reactions, how many different atoms and molecules are involved in each one, whether or not they will occur spontaneously or would require energy, and how fast they can take place. But have you ever considered the possibility that...

...chemical reactions can take place in the reverse direction?

That's right! Some chemical reactions can take place in both the forward and reverse directions. For example, the decomposition reaction of ammonium chloride (NH_4Cl) breaks down into ammonia (NH_3) and hydrogen chloride gas (HCl):

$$NH_4Cl \rightarrow NH_3 + HCl$$

However, when this system is cooled, ammonium chloride is formed:

$$NH_3 + HCl \rightarrow NH_4Cl$$

How does this work? At first, lots of NH_3 and HCl are being created in this forward direction. But this doesn't last for long! As more NH_4Cl is breaking down, the reverse reaction begins to speed up. Why? Because the reactants for the reverse reaction are increasing!

Eventually, the system will come to a point where the forward and reverse reactions will take place at the same rate. When this happens, the reactions are said to be in **equilibrium**.

As you know, there are always a number of rules about chemical reactions. Equilibrium reactions are no exception! Here they are:

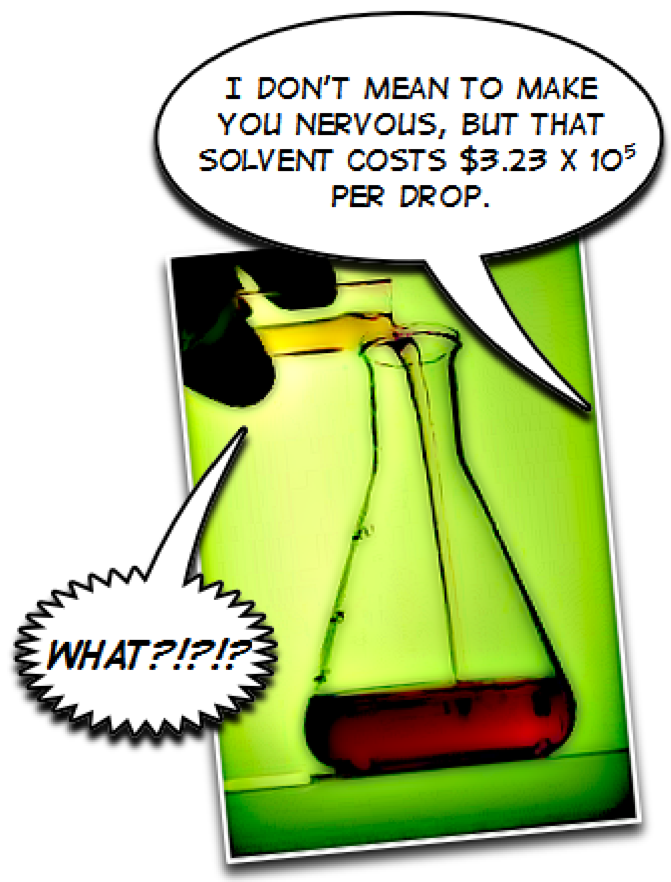
<u>Equilibrium can only happen with reversible reactions.</u>
This is a no-brainer. How could you be in equilibrium if you are not experiencing both forward AND reverse reactions?

<u>During equilibrium, products and reactants are still being formed.</u>
The reactions do not stop when they reach equilibrium. They are still reacting together to form new product; however, both reactions are occurring at the same <u>rate</u>.

<u>The concentrations of the reactants and products remain the same.</u>
This means for every product being made, the same amount of reactant(s) is being made as well!

So how can we use this "equilibrium" to help us out?

That is the #1 question for most chemists! Since the reagents used to create products can be a little expensive at times, chemists do not want to waste any of their money. Therefore, in order to get as much product as possible out of a reaction, chemists use a little thing called...



Le Châtlier's Principle

(It's pronounced "leh sha-tel-lyey" if you are interested!)

This principle states that whenever you mess with a reaction in equilibrium, the equilibrium will change to accommodate whatever you did. There are four scenarios that can be followed with this principle:

Scenario #1

If you add extra reactants to a reaction in equilibrium, the reaction will begin to produce more products.

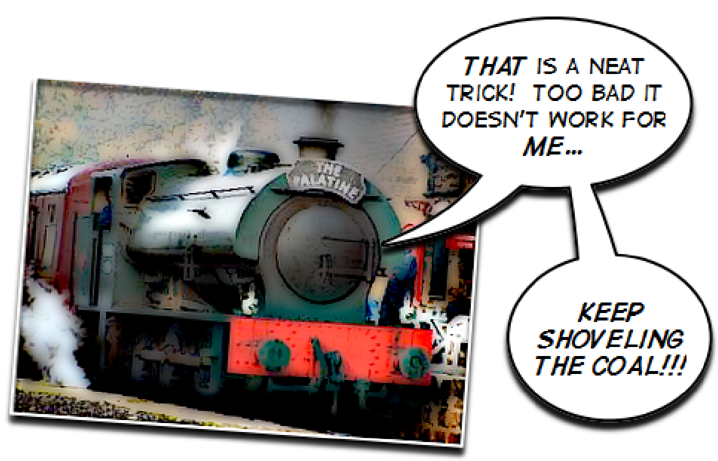
This should make sense to you. Think about the following scenario and imagine if a chemist wants to create as much product (compounds C and D) as possible.

If the following reaction is in equilibrium...

$$A + B \leftrightarrows C + D$$

...and the compound A is pretty cheap but compound B is very expensive, a chemist wouldn't want to waste too much of compound B!

Therefore, according to Le Châtlier's principle, if you start adding a large amount of compound A to the reaction, it will begin to form much more products (compounds C and D). There's no need of worrying about running out of compound B since compounds C and D will go through another reaction to form additional compound B! Cool, huh?



By using Le Châtlier's principle AND monitoring the reaction very carefully, the chemist can reduce the amount of compound B yet still obtain the needed amount of compounds C and D. This would be economically good for the chemist!

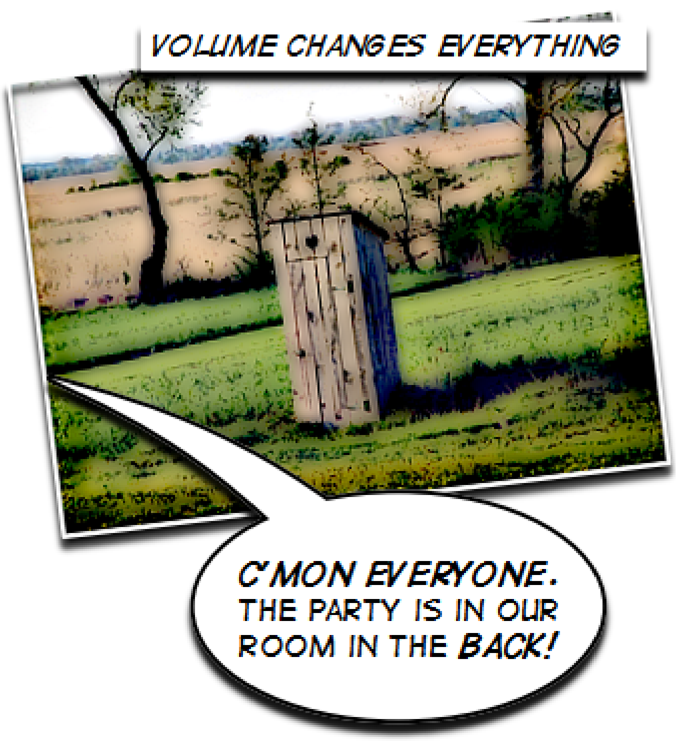
Scenario #2

If you remove the products of a chemical reaction, you will quickly run out of reactants!

At times a chemists wishes to use up the reactants within a chemical reaction. To do this, she needs to keep the reaction moving in one direction continuously. A reverse reaction is not what is wanted! Therefore, if the products of a reaction are continually removed, there will be no way for a reverse reaction to take place.

Scenario #3

If you change the volume of the container in which the reaction is taking place, the amount of products can be increased or decreased.



By decreasing the volume of a container (especially for gases), you are encouraging more collisions of molecules with each other as the pressure increases. Imagine having all of your friends over for a party and then moving everyone into a closet. The number of times you are going to bump into each other will increase greatly, won't it?

Each time a collision occurs, a new product can be formed. And, since you have already learned that...

Everything in the universe tends to move towards the lowest energy possible.

...the lowest amount of energy possible within these reactions would be products that contain the fewest amount of compounds.

Let's imagine the following reaction was taking place:

$$3A \leftrightarrows 2B + C$$

If you were to decrease the volume of the container, you would create more compound A since only one compound is being formed (which results in fewer collisions and a lower pressure.)

If, however, you increase the volume of the container, the reaction will tend to create more compounds B and C as the reaction attempts to reestablish equilibrium. Another way to look at this is, the reaction will create more products (B and C) to fill up a larger volume; but, it will also create fewer products (A) if the volume of its container is reduced.

Scenario #4

If you change the temperature of an exothermic or endothermic reaction, you can control which products will likely form.

Since endothermic reactions absorb energy, its equation would look something like this:

$$\text{energy} + A + B \leftrightarrows C + D$$

If you were to add energy to an endothermic reaction your result would be the same as we learned in Scenario #1. Just think about "energy" as another component of the reaction. If you add more of this component to the reaction, it will tend to create more of the products C and D.

On the other hand, exothermic reactions release energy, so its equation will look a little different:

$$A + B \leftrightarrows C + D + \text{energy}$$

Adding energy to an exothermic reaction would drive the reaction to create more compounds A and B!

All of these scenarios provide the chemist with a way of controlling which products they want from a reversible chemical reaction. As you can see, Le Châtlier's Principle has many practical applications for chemists!

Le Châtlier's Principle practice

For the following reactions at equilibrium:

$$H_{2(g)} + I_{2(g)} \leftrightarrows 2\ HI_{(g)}$$

1) Predict the shift in equilibrium when more HI(g) is added to the system.

2) How will the concentration of I_2 change?

For the reaction below, predict the direction the equilibrium will shift given the following changes. Temperature and volume are held constant.

$$2\ NO_{2(g)} + 7\ H_{2(g)} \leftrightarrows 2\ NH_{3(g)} + 4\ H_2O_{(g)}$$

3) Addition of ammonia (NH_3)

4) Removal of nitrogen dioxide (NO_2)

5) Removal of water vapour

6) Addition of hydrogen

Write reversible reactions for each of the following situations (be sure to balance your equations):

7) Hydrogen iodide gas (HI) decomposes into its elements.

8) Hydrogen and nitrogen gases combine to form ammonia gas, NH_3

If the system represented by the following equation is found to be at equilibrium at a specific temperature, describe why the following statements are true or false:

$$H_2O_{(g)} + CO_{(g)} \leftrightarrows H_{2(g)} + CO_{2(g)}$$

9) All species must be present in the same concentration.

10) The rate of the forward reaction equals the rate of the reverse reaction.

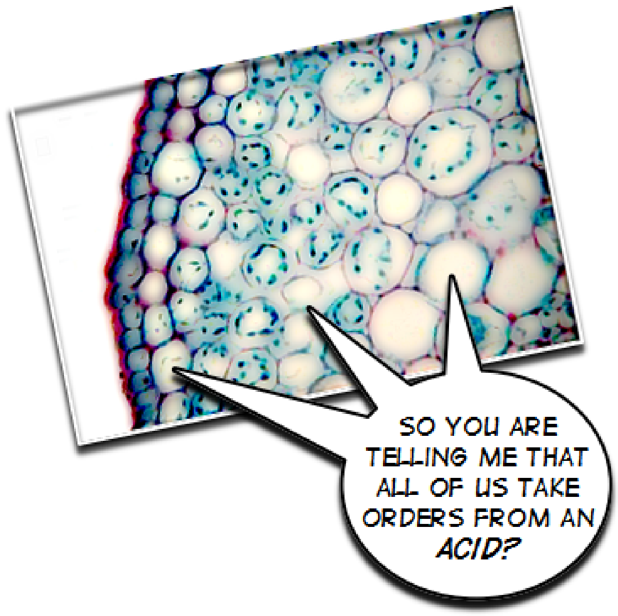
11) We can measure continual changes in the reactant concentrations.

Chapter 31

Acids always get a bad reputation. Typically, acids are seen as dangerous fluids that cause huge amounts of damage to every single thing it touches.

However, in reality, you have probably consumed a good deal of acids already today! The sour taste of citrus fruits and of vinegar are due to the presence of weak acids. In fact, what do you think the "A" in DNA stands for anyway? Deoxyribonucleic–**acid!**

But not all acids are so kind to our bodies. It is true that strong acids can cause some serious damage. Their counterparts – bases, can cause just as much damage as well! In the next couple of chapters, we are going to take a closer look into acids and bases. Let's get started...



Solutions that are known to be "**acidic**" contain a high concentration of acidic molecules while "**basic**" solutions contain a high concentration of basic molecules.

But what makes a molecule acidic or basic?

In order to answer this question, we have to look at two different chemists who discovered how these molecules work. The first chemist we will be looking at is **Svante Arrhenius**. Through careful experimentation, Arrhenius discovered that acids are compounds that give off H^+ ions when dissolved in water and bases are those which give off hydroxide ions OH^- when dissolved in water. This was a very helpful discovery at the time! After all, you can't tell the difference between an acid and a base by simply looking at them!

Because of his discovery, Arrhenius acids could be more easily identified as they nearly always begin with the letter "H" in their chemical formulas. For example:

Hydrochloric Acid HCl
Nitric Acid HNO_3
Nitrous Acid HNO_2
Sulfuric Acid H_2SO_4

The hydrogen within these chemicals is the source for the H^+ ions that are dissolved into the water solution.

As you may expect, Arrhenius bases tend to have "OH" within their chemical formulas which are the source of the OH^- ions:

Calcium Hydroxide $Ca(OH)_2$
Magnesium Hydroxide $Mg(OH)_2$
Potassium Hydroxide KOH
Sodium Hydroxide $NaOH$

Arrhenius' definitions of acids and bases have been widely accepted for many years – and they continue to be useful today! However, chemists knew there had to be more to learn about the nature of acids and bases after Arrhenius' discovery. Why?

Because many bases do not give off OH^- ions when dissolved in water!

This fact drove many chemists to find a more complete definition of acids and bases. A pair of researchers were the first to create another definition of acids and bases to take care of this OH^- problem – Johannes Brønsted and Thomas Lowry. The **Brønsted-Lowry method** of defining acids and bases adds to the Arrhenius method by stating:

Acids are compounds that give H^+ ions to other compounds; while bases are the compounds that accept these H^+ ions.

The best way to identify Brønsted-Lowry acids and bases is to watch the movement of H^+ ions within the reaction. For example, take a close look at the following reaction:

$$Na_2CO_3 + 2H_2O \rightarrow H_2CO_3 + 2NaOH$$

Two hydrogens from the water in this reaction switches places with the sodium of the Na_2CO_3 to form one mole of carbonic acid (H_2CO_3) and two moles of sodium hydroxide (NaOH).

According to the Brønsted-Lowry definition, the water would be considered the acid in this reaction because it donated its H^+ while Na_2CO_3 would be the base since it receives the H^+.

But there is more to the Brønsted-Lowry definition - the products of an acid-base reaction are known as **conjugate acids and bases**. The conjugate acid would be H_2CO_3 as it is formed when the base in the above reaction (Na_2CO_3) accepts the H^+ ion. NaOH would be the conjugate base as it is formed when the acid in the above reaction (H_2O) loses a H^+.

But what is the importance of knowing the conjugate acids and bases?

The answer to this question has to do with the strength of the acids and bases. An additional discovery within the Brønsted-Lowry definition of acids and bases says:

Weak acids always form strong conjugate bases and vice versa.

and

Strong acids always form weak conjugate bases and vice versa.

Arrhenius and Brønsted-Lowry did an excellent job at identifying the difference between acids and bases; however, chemists needed a method of measuring these substances. Therefore, two different scales were created which calculated the concentrations of H^+ or OH^- ions in solution. These scales are known as the **pH/pOH scales**.

Both scales range from 0 to 14. All measurements closer to 0 signify a solution with a higher acidity (pH) and those closer to 14 are more basic (pOH). In the center of these scales, a measurement of 7 identifies a neutral solution.

In order to calculate the pH of a solution, chemists use the following formula:

$$pH = -\log[H^+]$$

*The "log" function can be easily found within any scientific calculator. I would be happy to explain this function to you; however, I'm afraid its description may put you to sleep!

The H^+ that is located within the brackets $[H^+]$ identify the concentration of the hydrogen ions by their molarity (moles per liter).

pOH is measured with a nearly identical formula:

$$pOH = -\log[OH^-]$$

A solution cannot be both acidic and basic. This makes our calculations a little easier when we are trying to figure out the pH and pOH levels of a solution. Since both pH and pOH scales have a range of 0-14, once you know either the pH or pOH of the solution, you can easily calculate the other! Below is the equation that can be used to calculate either pH or pOH:

$$pH + pOH = 14$$

Without getting into a lot of mathematical discussion, you can use a "log" to figure out the pH and pOH of a solution very easily with the following equations:

$$[H^+] = 10^{-pH}$$

$$[OH^-] = 10^{-pOH}$$

The exponents of these equations tell you the pH and pOH of the solution.

Let's put these equations to work for us!

Calculate the pH and pOH of a solution with an $[H^+]$ of 1×10^{-8}. Is the solution acidic or basic?

Answer: You have been given the H^+ concentration, so first solve for the pH by plugging $[H^+]$ into the formula for pH:

$$pH = -\log[H^+]$$
$$pH = -\log[1 \times 10^{-8}] = 8$$
$$pH = 8$$

Now that you know the pH of this solution is 8, now plug it into the following equation:

$$pH + pOH = 14$$
$$8 + pOH = 14$$
$$pOH = 6$$

Since the acidity/basicity of this solution is pH=8 and pOH=6, <u>the solution is a base</u>.

A pH of 8 indicates that this solution is only slightly basic. I hope you noticed that the exponent of the H^+ concentration is equal to the pH. This is true whenever the coefficient of the H^+ concentration is 1.

But what if the coefficient is NOT 1?

This isn't going to be a problem at all! Let's take a look at another sample problem:

Calculate the pH and pOH of a solution with an [OH^-] of 2.3×10^{-11}. Is the solution acidic or basic?

The first thing we do to determine the pH and pOH of this solution is to pull out our calculator – we're going to need it! Now, since we know the [OH^-] of this solution is 2.3×10^{-11} we can calculate the pOH as follows:

$$pOH = -\log[OH^-]$$
$$pOH = -\log[2.3 \times 10^{-11}] = 10.6$$
$$pOH = 10.6$$

Knowing the pOH helps us to easily determine the pH of the solution with the following equation:

$$pH + pOH = 14$$
$$pH + 10.6 = 14$$
$$pH = 3.4$$

A solution with a pH of 3.4 is very acidic! And what if you are given the pH or pOH of a solution and are asked for the concentration?

These types of problems are just as easy as calculating the pH and pOH. All you need to do is look for the inverse log button on your calculator. It should look like this: 10^x

What is the [H^+] of a solution that has a pOH 0f 10.21?

Since the problem is asking us to find $[H^+]$ and we are given pOH, the first thing you need to do is find out the pH of this solution:

$$pH - pOH = 14$$
$$pH - 10.21 = 14$$
$$pH = 3.79$$

**You would skip this step if the problem asked you to find $[H^+]$ and you were given the pH value!

Now, let's convert our equation for pH and solve for $[H^+]$

$$pH = -\log[H^+] \text{ is converted into...}$$
$$[H^+] = 10^{-pH}$$
$$[H^+] = 10^{-3.79}$$
$$[H^+] = 1.62 \times 10^{-4}$$

Now that you know what acids and bases are, it's time you started looking at how these guys mix together in solutions. In the next chapter, you will be exploring how acids and bases two work together in real life! Stay tuned!

pH and pOH practice

1) **Identify the acid and base in the following reaction and label their conjugates.**

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

2) **Consider the following reaction. Label the acid, base, conjugate acid, and conjugate base, and comment on their strengths. How can water act as an acid in one reaction and a base in another?**

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

Use the Arrhenius definition to identify the acid or base in each reaction.

3) $NaOH + H_2O \rightarrow Na^+ + OH^- + H_2O$

4) $HF + H_2O \rightarrow H^+ + F^- + H_2O$

Determine the pH given the following values:

5) $[H^+] = 1 \times 10^{-13}$ ______________________________

6) $[H^+] = 1.58 \times 10^{-9}$ ______________________________

7) $[OH^-] = 2 \times 10^{-7}$ ______________________________

8) $[OH^-] = 1 \times 10^{-7}$ ______________________________

Determine whether the following are acidic, basic, or neutral:

9) $[OH^-] = 2.5 \times 10^{-7}$ ______________________________

10) $[OH^-] = 3.1 \times 10^{-13}$ ______________________________

11) $[H^+] = 4.21 \times 10^{-5}$ ______________________________

12) $[H^+] = 8.9 \times 10^{-10}$ ______________________________

Determine the $[H^+]$ from the following pH or pOH values:

13) pH = 3.3 ______________________________

14) pOH = 1.26 ______________________________

Chapter 32

Okay. You may be a master at calculating the pH of a solution but there still remains one little problem to solve...

How can you tell the concentrations of acids or bases of unknown strengths?

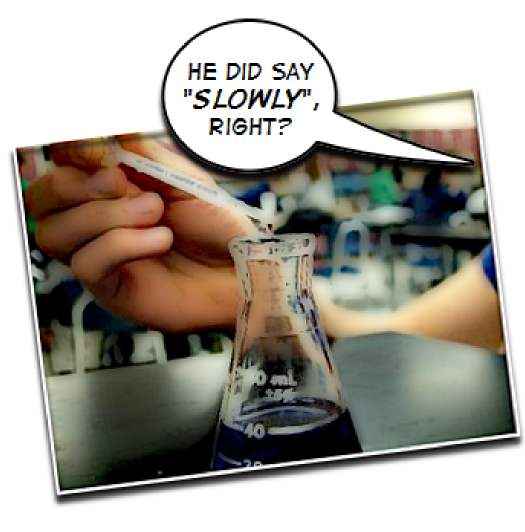
Think about it... 18M sulfuric acid (really nasty stuff) doesn't look any different than common vinegar! They are both clear and colorless liquids. And, if you had a container of 18M sulfuric acid in your kitchen, I certainly hope you wouldn't attempt to smell (bad idea) or taste (REALLY bad idea) the fluid.

So how do chemists determine the concentration of H^+ or OH^- ions in a solution?

Titrations

A **titration** is really a fancy word for a method used to neutralize an acid or a base (bring to pH of 7). Once neutralized, the fluid is generally safe to handle or dispose of.

For example, if you had a basic solution and wanted to neutralize it to pH7 you would slowly add a measured amount of acid until the solution reached pH7. Once pH7 was achieved you would simply measure out the volume of acid you added and this would tell you how much base was originally within the solution.

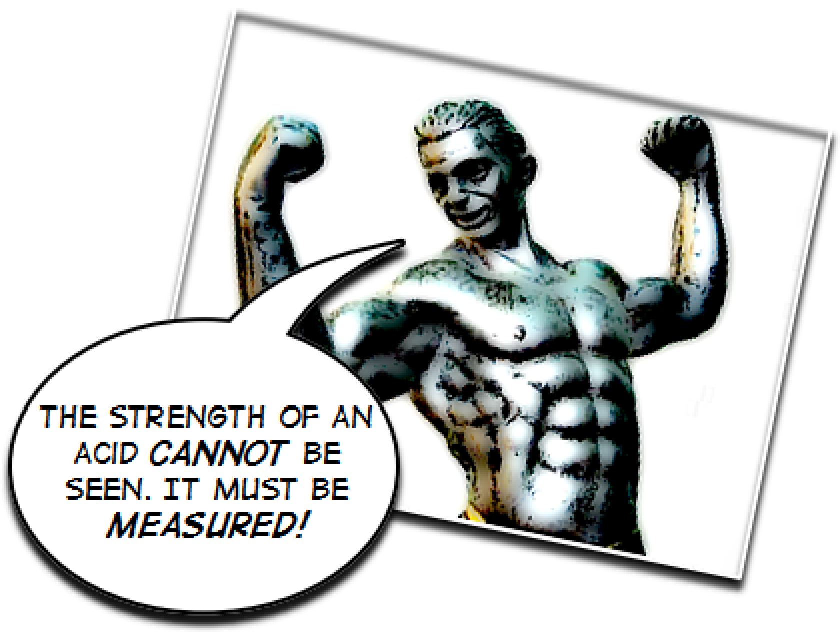


Why are these amounts the same?

In order to answer this question, you will need to follow three steps:

1) Calculate the molarity of the acid or base that was added for the neutralization to occur.
2) Convert the number of moles of the neutralizing substance to the number of moles of the neutralized substance.
3) Calculate the molarity of the acid or base that was neutralized.

Look back on our study of molarity in Chapter 22. You should remember that molarity (M) is a measurement of the amount of moles of solute per liter of a solution. This tells us the "strength" or concentration of the acid or base solution.



In the first and third steps, we are asked to calculate the molarity of the acidic and basic component of our solution. Therefore we use the molarity equation:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Here's a typical problem you will encounter while calculating the neutralization of an acid or base:

30mL of a 0.1M NaOH solution was added to neutralize a 25 mL solution of HCl. What was the concentration of the HCl?

Before we get started, we are going to need to write out a balanced equation for this reaction:

$$NaOH + HCl \rightarrow NaCl + H_2O$$

Step #1:

Calculate the molarity of the base that was added for the neutralization to occur.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

$$0.1M = \frac{\text{moles of NaOH}}{0.03L\ \text{NaOH}} = \mathbf{0.003\ moles\ NaOH}$$

Step #2:

Convert the number of moles of the neutralizing substance to the number of moles of the neutralized substance.

$$\frac{0.003\ \cancel{\text{moles NaOH}}}{} \left| \frac{1\ \text{mole HCl}}{\cancel{1\ \text{mole NaOH}}} \right. = \mathbf{0.003\ moles\ HCl}$$

Step #3:

Calculate the molarity of the acid or base that was neutralized.

$$M = \frac{0.003\ \text{moles HCl}}{25L\ \text{HCl}} = \mathbf{0.12M\ HCl}$$

Another way of completing these types of problems is to use the following simple equation:

$$M_A V_A = M_B V_B$$

$$(25mL)(?M)=(30mL)(0.1M)$$

$$= \mathbf{0.12M}$$

In the above example, a basic NaOH solution was used to neutralize an acidic HCl solution. However, you may be wondering one little thing...

How do you know when the pH level has reached neutrality? When do you know that the solution has a pH of 7?

This is a big problem! Remember... 18M HCl looks just like regular water! When you neutralize it to pH7, guess what it looks like...

Regular water!

Because of this, chemists have either created and/or discovered a variety of chemical **indicators** to help them identify the presence of acids or bases. Indicators are chemical compounds that change color in the presence of acids or bases.

Natural indicators like boiled rose petals and red cabbage are easily created in the kitchen. More recently, chemists have created their own types of pH indicators. These items include the high-end electronic pH indicators which very accurately measure the amount of H^+ in a solution to cheaper to easy-to-use items such as **Litmus paper** (which turns red in the presence of an acid and blue in a base.) Another manmade chemical, **phenolphthalein**, is a clear liquid while in the presence of an acid but turns pink when mixed with a base.

Since the concentration of H^+ within a solution can be increased or decreased rather easily, the use of indicators are very handy during titrations.

Furthermore, once added to a solution, phenolphthalein does not break down into simpler substances; therefore, it can change the color of the solution from pink to clear as many times as possible.

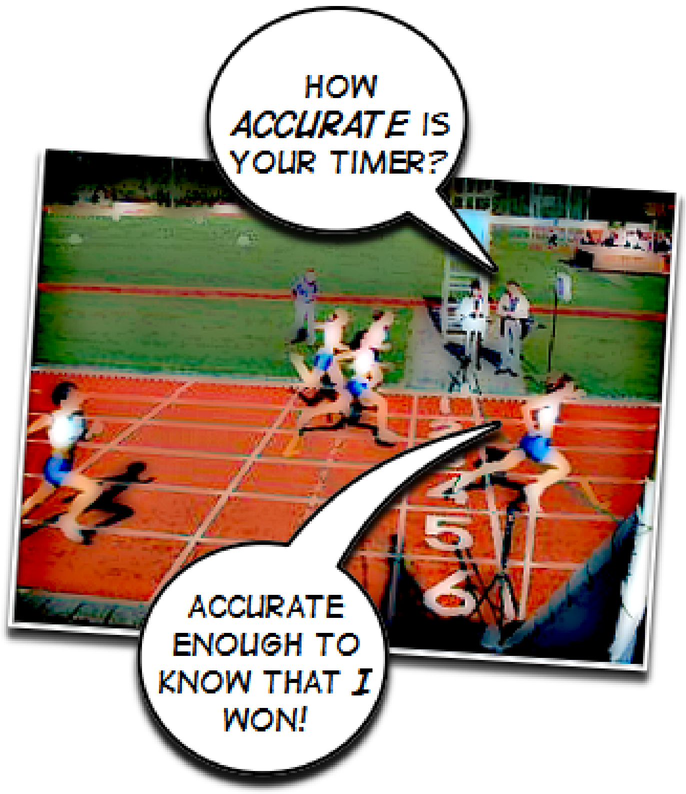
Let's say you have placed your indicator into a base of unknown concentration. You are carefully adding a measured amount of acid into the basic solution when suddenly the color of the base begins to change. What do you do?

First of all, the point in which the color change begins to appear is known as the **endpoint**. This is the point in which you should stop titrating!

You are probably thinking this is not the most accurate method of determining when a solution of unknown strength has reached a pH of exactly 7...

And you are right!

It is nearly impossible to reach the **equivalence point** (the point in which the volume of acid equals the volume of base) through titration when the endpoint is identified. Therefore, when chemists titrate they attempt to get as close to the equivalence point as possible and follow up with the most accurate pH meter they have in their lab.



If you haven't figured this out yet, science builds models and procedures that give us the most accurate and precise measurements about the natural world. Even though our results are not 100% perfect, we do seem to get pretty darn close most of the time!

Titrations practice

1) If it takes 26 mL of 0.1 M NaOH to neutralize 125 mL of an HCl solution, what is the concentration of the HCl?

2) If it takes 61 mL of 0.05 M HCl to neutralize 345 mL of NaOH solution, what is the concentration of the NaOH solution?

3) If it takes 62 mL of 0.5 M KOH solution to completely neutralize 130 mL of sulfuric acid solution (H_2SO_4), what is the concentration of the H_2SO_4 solution? (Hint: you will need to divide your answer by 2 since there are 2 hydrogen atoms to be given off with H_2SO_4)

4) Can I titrate a solution of unknown concentration with another solution of unknown concentration and still get a meaningful answer? Defend your answer:

5) How many mL of a 3M NaOH solution are required to completely neutralize 45.0 mL of 1.5M H_2SO_4?

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

(Hint: when using the $M_1V_1 = M_2V_2$ equation on these types of problems, you will need to multiply your M_1V_1 and M_2V_2 values by the number of moles of each compound. For example, in this problem you would multiply your M_1V_1 product by 1 for H_2SO_4 and you would multiply your M_2V_2 product by 2 for NaOH)

6) What is the molarity of a solution of $Ca(OH)_2$ if 15.0 mL of the solution is required to neutralize 37.2.0 mL of 2M HCl?

$$2HCL + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$$

7) What volume of 0.27M phosphoric acid is required to neutralize 42.0mL of 0.05M barium hydroxide?

$$2H_3PO_4 + 3Ba(OH)_2 \rightarrow Ba_3(PO_4)_2 + 6H_2O$$

8) Calculate the **mass** of aluminum hydroxide required to completely react with 13.5 mL of 0.45M HCl. (Hint: Use stoichiometry)

$$Al\ (OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$$

Glossary

1st Law of Thermodynamics	also known as "The Law of Conservation of Energy"; states that energy is never lost or gained, it only changes form
2nd Law of Thermodynamics	states that energy, if concentrated in an area, will spontaneously spread out unless a force keeps it from doing so
accuracy	how close a measured value is to the real value of the object
acid-base reactions	a chemical reaction similar to double displacement reactions; however, unlike double displacement reactions, water is one of the product
acidic solutions	solutions which contain a high concentration of acidic molecules
activation energy	energy required for a reaction to begin making a new product
alkali metals	elements within the first family of main block elements that are very reactive with other elements, flammable in air and water, have low melting and boiling points, are soft, have low densities, and contain 1 valence electron
alkaline earth metals	Second main block family of elements which react in air and water (but not as violently as the alkali metals), contain low melting and boiling points (but they are higher than the alkali metals), are soft (but harder than – you guessed it, the alkali metals), and have low densities (but higher than those of the alkali metals); alkaline earth metals all have 2 valence electrons
anion	a negatively charged ion
atomic mass	the average mass of an element and all of its isotopes

atomic mass unit (amu)	the mass of one proton (1.66 x 10^{-27} kg) which are used to determine the overall atomic mass of an element
atomic number	the total number of protons in the nucleus of an atom
atomic radius	a measurement of the relative length of an atom from its nucleus to its outermost orbital
atoms	tiny particles that make up the universe
base units	basic metric units such as seconds (time), meters (length), kilograms (mass/weight), liters (volume), Pascals (pressure), and Kelvin (temperature)
basic solutions	solutions which contain a high concentration of basic molecules
boiling	transition from a liquid phase to a gas phase
boiling point elevation	colligative property of a solution; as you increase the concentration of a solution (by adding solute into a solvent), the boiling point of the solution increases as well
bonding pair	a shared pair of electrons between covalently bonded atoms
boron family	third main block family of elements which contain mostly metals with the exception of boron itself, which is a metalloid; boron family elements all have 3 valence electrons
Boyle's law	gas law which states that as a gas goes through some kind of change, the product of its initial pressure and volume (P_1V_1) will equal the product of its final pressure and volume (P_2V_2)
Brønsted-Lowry method	method of identifying acids from bases in that acids are compounds that give H^+ ions while bases are the compounds that accept these H^+ ions

calorie	standard unit of energy
carbon family	fourth main block family of elements which are made up of metals, metalloids, and non-metals (of which carbon is a non-metal); elements in the carbon family have 4 valence electrons
cation	a positively charged ion
Charles's law	gas law which states as you increase the temperature of a gas, its volume increases too
chemical reactions	processes in which the atoms within reactants are rearranged to form new compounds (products)
coefficient	the number used within scientific notation between 1 and 9.9 to be multiplied by an exponent
coefficients	a number placed in front of a compound to identify how many compounds (as a whole) exist
colligative properties	properties of solutions that depend upon the amount of solute dissolved in a solution; boiling and melting points, osmotic pressure, and electrical conductivity are examples of colligative properties
collision theory	the speed in which a chemical reaction takes place is based solely on how hard the reactants slam into each other and the directions they are traveling
combined gas law	gas law which combines all of the relationships which exist between the Boyle, Charles, and Gay-Lussac gas laws
combustion reaction	a chemical reaction in which a compound containing carbon and hydrogen combines with oxygen gas; the products formed are heat, water, and carbon dioxide
compounds	pure substances that are made up of 2+ <u>different</u> atoms (or molecules) that are bonded together

concentration	the amount of solute we place into a solution
condensation	transition from a gas phase to a liquid phase
conjugate acids and bases	Additional discovery by Brønsted-Lowry; states that weak acids always form strong conjugate bases and vice versa; and, strong acids always form weak conjugate bases and vice versa
conversion factor	a shortcut that can be used to help with your conversion
covalent bond	a bond between two non-metals in which both atoms share one or more electrons
critical point	position on a phase change diagram where a substance can act like both a gas and a liquid
crystals	ordered arrangements of ions
Dalton's Law of Partial Pressures	the total pressure of a mixture of gases equals the sum of its individual partial pressures
decomposition reaction	the opposite of a synthesis reaction; a chemical reaction in which one large reactant is broken down into two or more smaller products
dimensional analysis	a method of converting one unit into another unit through the use of conversion factors
double bond	two pairs of covalently bonded electrons (4 electrons total)
double displacement reaction	a chemical reaction in which the cations of two compounds exchange places with each other
ductile	can be drawn into wires
effective molality	a measurement of the number of moles of solute per kilogram of solvent
electrolyte	a solution of dissolved ions

electromagnetic waves	energy which is released from an oscillating electron
electron affinity	the relative strength of an atom's charge
electron configuration	a method of describing electron orbitals
electron dot structure	graphical method of describing the number of valence electrons that surround an atom; contains up to eight dots surrounding the chemical symbol of an element
electronegativity	a measurement of how much an atom can pull electrons away from other atoms
elements	close to 120 unique atoms, each of which containing its own physical and chemical characteristics
empirical formulas	formulas which tell you the ratios of elements to each other within a compound
endothermic	chemical reaction which absorbs energy
endothermic reactions	chemical reactions that tend to absorb more energy than the energy released by its products
endpoint	the point in which the color change within a titration begins to appear
energy level diagram	visual tool used to identify how to write out the electron configurations of elements
energy levels (orbitals)	layers around a nucleus which contain specific numbers of electrons
enthalpy	the total amount of heat a molecule contains
entropy (ΔS)	a measure of the amount of energy that spreads throughout the surroundings of a system
equilibrium	time within a reversible reaction (system) where the forward and reverse reactions will take place at the same rate

equivalence point	the point in which the volume of acid equals the volume of base during a titration
exothermic	chemical reaction which gives off energy
exothermic reactions	a reaction in which the amount of energy absorbed by the reactants is less than the energy released by the products
exponent	a factor of ten; used within scientific notation to represent large or small numbers.
families (groups)	vertical rows within the periodic table
free energy (ΔG)	the amount of energy that is available by a system to do some form of work
freezing	transition from a liquid phase to a solid phase
freezing point depression	colligative property of a solution; solutions freeze at lower temperatures than pure solvents because the extra solute molecules "get in the way" of the solvent molecules
frequency	the rate of oscillations within electromagnetic waves
Gay-Lussac law	gas law which states that as you increase the temperature of a gas, its pressure goes up too
general rate law	used to help to determine how the rate of reaction varies as the reaction progresses; for a chemical reaction that follows a simple formula like $A + B \rightarrow C$ looks like this: $\text{Rate} = k[A]^x[B]^y$
grid method	method of adding coefficients to balance a chemical equation; uses a table to count the number of atoms within the reactants and the products in a reaction
halogens	seventh main block family of elements which are non-metals, are very reactive (much like the alkali metals), and each contain 7 valence electrons

heat	the movement of energy from one thing to another through the motion of molecules
heat of formation ($\Delta H^0{}_f$)	the change in enthalpy when a compound is formed from its elements
heat of reaction (ΔH_{rxn})	the change in enthalpy during a chemical reaction
heterogeneous mixtures	combination of two or more different kinds of atoms that are easy to separate (mixtures are not bound together)
homogenous mixtures (solutions)	combination of two or more different kinds of atoms that are not easy to separate (mixtures are not bound together)
ideal gas law	gas law which utilizes a fictional "gas" that obeys specific rules when kept at a constant volume, temperature, and/or pressure
indicators	chemical compounds that changes color in the presence of acids or bases
ion	an atom that has gained or lost one or more electrons
ionic bond	a bond between a metal cation and a non-metal anion in which one of the ions steals one or more electrons from the other, thus creating a magnetic bond between the two oppositely charged ions
isotope	an element which contains a different amount of neutrons despite having the same number of protons.
joule	metric unit of energy
Kelvin scale	temperature scale which sets its zero degree (0°) at the most extreme measurement known as absolute zero
kinetics	the study of the rates of chemical reactions and

	explains how we can speed them up
Law of Conservation of Matter	law of chemistry which states that matter cannot be created or destroyed, only rearranged into different positions
Le Châtlier's principle	principle which states that whatever changes are done to a reaction in equilibrium, the equilibrium will change to accommodate whatever was altered
Lewis structure	very similar to the electron dot structure; however, these drawings depict the valence electrons in lone pairs as dots and contain lines to represent shared pairs in a chemical bond (single, double, triple, etc.)
limiting reagent (reactant)	the reactant that runs out first within any chemical reaction
litmus paper	pH indicator which turns red in the presence of an acid and blue in a base
main block elements	eight families within the periodic table which contain similar chemical characteristics; alkali metals, alkaline earth metals, boron family, carbon family, nitrogen family, oxygen family, halogens, and noble gases
malleable	able to be pounded into sheets
mass number	the sum of the protons and neutrons within an atom
melting	transition from a solid phase to a liquid phase
metallic bond	the force of attraction between two or more metal nuclei which are surrounded by a sea of electrons
metalloids	elements found between the metals and non-metals on the periodic table which share the properties of both these types of elements
metals	the majority of elements which are typically hard, shiny, malleable, ductile, good conductors of heat and

	electricity, and have high densities, melting points, and boiling points
MINOH	method used to add coefficients to the type of particles within the equation for it to be balanced; follows a path by which metals are balanced first, followed by ions, non-metals, oxygen, and hydrogen
mixture	combination of 2+ atoms that are not bonded together
molar mass	sum of all atomic masses of elements within a compound
molarity	measuring scale used to determine the concentration of a solution
mole (Avogadro's number)	a conversion factor for anything in the universe which contains 6.02×10^{23} items
mole ratio	a conversion factor that exists between the moles of the product and reactants from the chemical equation; can be identified as the coefficients for the known and unknown compounds
molecular formulas	formulas which contain how many of each type of atom are present within a compound
molecule	combination of 2+ atoms that are bonded together
monatomic ions	ions which contain only one kind of atom
nitrogen family	fifth main block family of elements which are made up of metals, metalloids, and non-metals (nitrogen itself is a non-metal); elements in the nitrogen family have 5 valence electrons
noble gases	eighth main block family of elements which are the least reactive elements on the periodic table and each contain 8 valence electrons
non-metals	elements that are typically brittle, poor conductors of

	heat and electricity, have a very low density, have lower melting and boiling points, and are <u>gases</u> at room temperature
nonpolar molecule	phenomenon in which the electrons within a molecule are spread evenly throughout its structure; no single atom within the molecules "pulls" on its electrons with a greater force due to their similar electronegativity
non-spontaneous processes	any process that requires some form of effort to make it occur
nucleus	the center of an atom which contains its protons and neutrons
octet rule	the phenomenon by which all elements move towards the filling their eight possible positions within their outermost s and p orbitals with electrons
osmosis	the movement of a fluid across a semipermeable membrane
osmotic pressure	the pressure which needs to be applied to a solution to keep water from moving across a semipermeable membrane
oxygen family	sixth main block family of elements which are made up of metals, metalloids, and non-metals (of which non-metal is a non-metal); this family is quite reactive with other elements (their reactivity is comparable to the alkaline earth metals); elements in the oxygen family have 6 valence electrons
percent composition	the total amount (by mass) of each element within a compound
percent error	the accuracy of a measurement as determined by the difference between the accepted measurement and the experimental measurement divided by the

	accepted measurement and multiplied by 100.
percent error	procedure used to measure how much error exists within a chemical reaction
periods	horizontal rows within the periodic table
pH/pOH scales	two different scales (between the values 0-14) which were created to calculate the concentrations of H^+ or OH^- ions in solution
phase change diagram	a graph which depicts the phase change of a pure substance or mixture over time as either pressure or temperature is changed
phases	the three states of matter – solid, liquid, and gas
phenolphthalein	pH indicator which is a clear liquid while in the presence of an acid but turns pink when mixed with a base.
polar molecule	phenomenon in which one of the atoms within a molecule "pulls" on its electrons a little stronger than the other atom(s) that are bonded together
polyatomic ions	molecules which have gained or lost some of their electrons and have turned into ions
precipitate	solid product formed from the chemical reaction between two liquid reactants
precision	how close a series of measurements are to each other
pressure	the force of gas molecules hitting the sides of the container in which they are stored
products	compounds that are created from the reaction between reactants
pure substances	mixtures that cannot be physically separated very easily at all
quantum	branch of science in which scientists study the

mechanics	measurable (quantifiable) movements of atoms within orbitals
reactants	compounds which react together to form products
reaction orders	exponents (x and y) within a general rate law (Rate = $k[A]^x[B]^y$); used to define how the rate is affected by the reactant concentration
relative abundance	a measured percentage of each element that exists in nature was determined
salts	products of an ionic bond
saturated solution	a solution that is <u>not</u> able to dissolve more solute
scientific notion	method of writing large and small numbers in shorthand using a system of coefficients (1-9.9) multiplied by exponents by factors of 10. For example, 1224578 would be written as 1.224578×10^6
sea of electrons	a free-flowing collection of electrons surrounding a group of metal ions which act to force these positively charged ions together
semipermeable membrane	any "barrier" that will allow solvents (like water) to move through them
shielding effect	the phenomenon by which the inner electrons of an atom act as a shield between the nucleus and the valence electrons which lowers the electronegativity of the atom
significant figures (numbers)	any digits that give us some amount of precision about a set of data
single bond	a single pair of covalently bonded electrons (2 electrons total)
single displacement	a chemical reaction in which a single element switches places with another element in a chemical compound.

reaction	
solubility	the ability of a solute to be dissolved by a solvent
specific heat (heat capacity)	amount of energy needed to heat a substance by 1°C
spontaneous processes	any naturally occurring actions that exist without the addition of any extra energy
stoichiometry	the process of calculating the amount of reactants we need for a certain amount of product
STP	"standard temperature and pressure" which is considered to be 273°K (0°C) and 1.00 atm
subatomic particles	protons, neutrons, and electrons
subscripts	the numbers written after individual atoms that identify how many atoms of that element exist within the molecule
supersaturated solution	a solution that has dissolved more solute than it naturally should be able to contain
surroundings	the environment that surrounds systems (chemical reactions)
Svante Arrhenius	discovered that acids are compounds that give off H^+ ions when dissolved in water and bases are those which give off hydroxide ions OH^- when dissolved in water
synthesis reaction	chemical reaction in which two or more reactants combine to form a more complicated molecule
system	name given to chemical reactions and its energy transfer(s)
temperature	a measurement of the amount of energy found within molecules
temperature	the measure of energy found within an object after

	heat has been transferred
thermodynamics	the study of energy; more specifically, how energy is stored within molecules, how it can be transferred from one substance to another, and how this process is able to produce heat
titration	a method used to neutralize an acid or a base (bring to pH of 7)
transition state	the point in which a reaction begins to make a new product
triple bond	three pairs of covalently bonded electrons (6 electrons total)
triple point	single position on a phase change diagram where a substance can be either a solid, liquid, or gas
unsaturated solution	a solution that is able to dissolve more solute
valence electrons	electrons that make up the outermost s and p orbitals
Van der Waals	weak forces which hold covalent bonds together
vapor pressure	pressure which exists between water molecules in a liquid phase
volume	the amount of space in which a fluid is enclosed

Credits

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