Introductory Chemistry

Chapter 01: The Chemical World

Science

What is Science? **Science** is a body of knowledge concerning the laws of nature accumulated through observation and experimentation. Science has the following 5 characteristics:

1. It is **testable**; this means that an explanation for an observation can be proven or disproven by experimentation.
2. It is **reproducible**. Results obtained by one scientist or by one method of testing or analysis can be duplicated by other scientists or other methods of analysis.
3. It is **explanatory**. The observation, event or phenomenon can be described and a tentative explanation can be formulated and tested or verified through experimentation.
4. It is **predictive**. Knowledge obtained from past experiments can be used to forecast expected results based on similarities or patterns of behavior.
5. It is **tentative**. The explanation for the observation can be modified or discarded as more data are obtained.

**How Science Happens- The Scientific Method**

When scientists conduct investigations or research, they employ a generally accepted process known as the **scientific method**. This process is a logical series of steps that include the following:

1. **Observation**: The scientist describes the observation and attempts to explain it by developing
2. **hypotheses**. These are tentative or temporary explanations that must be tested by
3. **experimentation**. The results of the experiments either support or reject the hypothesis. The hypothesis can be modified and further experiments are conducted. After many experiments the investigator may be able to draw
   1. **conclusions** which summarize the data and their significance. The investigator must be able to communicate his results to other scientists. This is done through
   2. **submission** for publication in a scientific journal. Before publication the research is subjected to a process called
   3. **peer review** where the scientific merit of the investigation is judged by other scientists who are experts in the area being studied. Not all research submitted for publication is accepted. The reviewers may suggest further experiments before acceptance or may outright reject it.
   4. **publication.** If accepted the research is published and
   5. other scientists respond by either trying to disprove or validate the results of the investigation.
4. **Scientific theory**. When a large body of data has been accumulated through the work of numerous scientists and the explanations for the observation have become generally accepted and supported by the data, it becomes a **scientific theory**. You must differentiate a scientific theory from our everyday use of the word, which implies no proof or evidence is offered. The more appropriate term for such a case is hypothesis. What are some examples of a scientific theory? 1) The **atomic theory of matter**, which we will study in the next two chapters. 2) The **theory of evolution**, which explains how the diversity of life arose over a long period of time (geologic). Although a scientific theory is supported by considerable experimental evidence, it can be modified as more data are obtained.
5. **Scientific law**. You must also differentiate a scientific law from a scientific theory. A **scientific law** is a brief statement or mathematical equation summarizing a large body of data or observation or phenomenon. Some examples are the law of conservation of mass, which states that “matter is neither created nor destroyed in chemical reactions” and Boyle’s law, which states that “the volume of an ideal gas decreases as pressure is increased if the amount of gas and the temperature are held constant.” A scientific law describes an observation that is universally true under the specified conditions.

**Chemistry**

We will now turn our attention to **chemistry** a branch of science concerned with the study of matter and the changes it undergoes. **Matter** is anything that has **mass** and occupies space; it is the physical material of the universe. Everything in the universe is composed of matter. Furthermore, all matter is composed of atoms and molecules, which we will study in more detail as we progress through this course. Chemists try to understand how the universe and everything in it works on a large scale by studying how the atoms and molecules interact with each other on a submicroscopic scale.

**Succeeding in Chemistry**

**Be curious**! Chemistry is not obscure- **you do chemistry every day** and your very life depends on it every single day!

**Ask lots of questions**- especially why and how! In this class, use the discussion board on the Blackboard course site to ask questions about anything you might not understand from your reading. Most especially, ask the questions you might think would be “dumb”- these are usually the ones to which your classmates want answers also and they will appreciate your boldness in asking (as will your instructors!)

**Notice the details!** What do you see, hear, smell or feel as you are doing your labs? What are the units of the numbers you use to solve your homework problems? How good are the numbers?  
**Work at solving problems**- as with becoming a good athlete or a good musician, problem-solving takes a lot of practice! Don’t just read the example problems in the book- work along with the author in solving the problem. Be sure you ask someone if you don’t understand how the author gets from one step to the next.

**Commit to working on chemistry at least an hour or two every day**- or at least on the days that you eat! You will be much more successful if you take the material in small bites and take time to digest it, rather than rushing through several chapters because the assignment is due in a half hour!

**Workflow:** Working through the material in this course will definitely take up a chunk of your time. We suggest that you begin by reading through the chapters that are assigned for the week, taking your own notes in a notebook as you go. Highlighting points in your text works for some people, but for most, the act of writing out definitions and concepts is important in really understanding them and committing them to memory.   
  
One method of note-taking is called the Cornell method. In this method, you divide your notebook page by drawing a line down the page so that the left column takes up approximately one-third of the page and the right column takes up the remaining two thirds. Put a vocabulary word or concept in the left column and put the related definition or details on the right side of the page. This way as you review, you don’t have to read all the details every time and a glance at the left column will jog your memory as to what is important   
  
Once you have read over the chapters, view the powerpoint slides, and read over the lecture notes. Then begin the homework problems. Remember that the discussion board is available at any point along the way, to get clarification on specific points or help with problems. It pays to be specific about what you don’t understand in asking questions on the discussion board. If you can help answer another student’s questions, that is a good thing also. Explaining something to someone else is the very best way to learn it. Your instructor keeps a good eye on the discussion board to make sure no one is led astray!   
  
Finally, be aware that it is your responsibility to ask for help when you need it. If you need assistance with something beyond the scope of the discussion board, consult with your instructor by phone or by email. Don’t flounder along until the exam, expecting things to take care of themselves. They won’t!  
  
Please be aware that although some online courses have a reputation for being “easy,” this is not one of those courses. You will work hard in this course. However, we believe that most students have the ability to be successful in this course, if you follow our suggestions and study tips. You must take responsibility for your own success, but remember that you are not alone- we are here to help and guide you!

# Balancing a Chemical Reaction

## *Procedure*

 

1. Simple reactions can be balanced by inspection – using trial and error to find the combination of stoichiometric coefficients that will balance the reaction.

More complicated reactions may require a systematic approach to balancing as follows:  
**methane plus oxygen react to form carbon dioxide plus water**

1. Remember that you must **never change the formulas** of reactants or products by changing the subscripts, **only the coefficients may be changed.**
2. Write out the correct formula for each reactant and write the reactant(s) on the left side of the reaction arrow:



1. Write out the correct formula for each product and write the product(s) on the right side of the reaction arrow.



1. Record a list of all the elements present in the reactants below the left side of the equation. Record a similar list of all the elements present in the products below the right side of the equation.



C C

H H

O O

1. Inventory the number of atoms of each element on each side of the reaction **as it is written at the present time.**



C 1 atom C 1 atom

H 4 atoms H 2 atoms

O 2 atoms O 3 atoms

1. You can see that the numbers of H and O atoms are not balanced. Begin by balancing the atoms in the “most complicated” molecule in the reaction. Since the carbons are already balanced, let’s begin with the hydrogen atoms. We see that there are four hydrogens on the left side of the reaction and two hydrogens on the right side of the reaction. We can put a “2” as a coefficient in front of the water molecule to balance the hydrogens:



C 1 atom C 1 atom

H 4 atoms H 2 atoms

O 2 atoms O 3 atoms

1. Writing the “2” in front of the water molecule changes the inventory of atoms on the right side of the reaction, so we need to stop and take a new inventory: There is still one carbon atom, but now there are four hydrogens and four oxygens.



C 1 atom C 1 atom

H 4 atoms H X 4 atoms

O 2 atoms O X 4 atoms

1. When we look at the new inventory of atoms, we see that if the only thing left to balance is the oxygen. We can do that by placing a “2” in front of the oxygen molecule on the left side of the equation and then taking another new inventory of atoms.



C 1 C 1

H 4 H X 4

O X 4 O X 4

The reaction is now balanced.

**Procedure Summary for Balancing Reactions:**

1. Write the reaction placing formulas for reactants on the left and formulas for products on the right side of the reaction..
2. Inventory the atoms on the left and right side of the reaction.
3. Beginning with an element from the “most complicated” molecule, adjust one coefficient as needed to balance that element on both sides of the reaction, then stop and inventory the elements again.
4. If there are other elements in the”most complicated” molecule that remain unbalanced, balance those next, one at a time. Each time you make one change to the coefficients of the reaction, you should stop and make a new inventory of elements.
5. Balance common small molecules such as carbon dioxide, water, or diatomic elements last.

Note: In a reaction in which there are polyatomic ions that do not change from the left side of the reaction to the right side, you may balance the polyatomic ions as if they were monatomic ions or elements, rather than balancing the individual atoms. (Both ways will work- this way should be a little easier.)

Chapter 02: Measurements and Problem-Solving   
Part 1: Measurements

**Measurement of Matter: the Metric System**

All measurements consists of two parts: a **number** and **unit**. Keeping the number and the units together is very important. A number without an accompanying unit is meaningless. **Always** **(!!)** write out the units for the numbers you use in calculations and problem-solving. The cool thing about units is that you can do the same algebraic operations with units that you do with numbers and in this course, you will learn a method of problem-solving in which the units of the numbers will help lead you to the correct calculation.

**Scientific Notation**

Scientific notation is a way of expressing very large or very small numbers in terms of powers of ten, written as a number with a value between 1 and 10 (called the coefficient) which is multiplied by a power of 10 (called the exponent. ) For example, in the number 1.7 x 109, the exponent “9” tells us that the decimal point was moved nine places to the left from where it started in the original number. Therefore the actual number is 1,700,000,000, a very large number. Thus large numbers will have a positive exponent while very small numbers will have negative exponent. In the example 5.6 x 10-6, the exponent “–6” tells us that the decimal place was moved 6 places to the right from its place in the original number, so the regular number is 0.0000056. Practice converting numbers to scientific notation and vice versa. Study the examples in the chapter and in the powerpoint slides.

**Using Scientific Notation on Your calculator**.

Practice entering and doing calculations using scientific notation on your calculator. Most scientific calculators have a button named either “EE” or “exp” for entering exponents.   
  
Two common types of calculators used by students include the TI-83 or TI-84 series and the TI-30 series. Numbers in scientific notation are entered in a slightly different manner for these two types of calculator.

When using a TI-83 and trying to enter scientific notation, do not enter the "10" before the exponent or you will not get the correct answer. Often, when a student’s answer is off by a factor of ten, this is the detail that causes the answer to be off.  
  
Scientific Notation on a TI-83 or TI-84

To enter a number such as 1.7 x 10 -9 in scientific notation in the TI-83 calculator, enter “1.7”, then the yellow "2nd" button to shift, then the comma button (this gives the "EE" function because we previously hit "2nd") followed by the exponent. If the exponent is negative, push the (-) button (located in the bottom row of buttons next to the decimal point) before you enter the number for the exponent.  Do not enter the "10" before you enter the exponent: the “EE” function takes the place of “times ten to the.”  
  
To recap: To enter 1.7 x 10 -9 in a TI-83 calculator, enter 1.7, "2nd", "EE", (-)9 and then push the enter button and continue on with your calculation.

Scientific Notation on a TI-30XA

There are several different models of TI-30XA calculators that look a little different from each other.   
It appears that TI-30XA calculators all have an "EE" button, so you will not need to use the "2nd" function to use the "EE". Do not enter the "10" if you are using the "EE" button. Simply type in the coefficient, hit the EE button, then type the exponent. If it is a negative exponent, you should hit the change sign button (on the bottom row just to to left of the equals sign ) before you type in the number for the exponent.

To recap: To enter 1.7 x 10 -9 in a TI-30XA calculator, enter 1.7, "EE", (-)9 and then push the enter button and continue on with your calculation.  
  
Some calculators may use “EXP” instead of “EE”, either as a separate button or as a “2nd” function. If you are consistently off in your calculations when using scientific notation, ask for help! The math lab at any campus, the discussion board in Blackboard, or your instructor may be able to save you much frustration if you will ask for help!

**Exact Numbers vs Inexact Numbers**

All numbers are not created equal. It is important to differentiate between numbers that are exact and numbers that are inexact. An exact number is a number that is determined by counting or by definition. For example, the number of students in a particular class section can be determined exactly by counting them. We can determine the exact number of nuts or bolts in a bag, or the number of apples in a bowl by counting. By definition, there are exactly 12 eggs in a dozen eggs, exactly two socks in a pair, or exactly 144 pencils in a gross of pencils. All of these are examples of exact numbers.  
  
By contrast, an inexact number is a number that is measured, usually by using some sort of measuring tool such as a ruler, tape measure, scale, or measuring cup. Measured numbers are always inexact to some degree. The degree of uncertainty may depend on how good the measuring tool is, or how skilled is the person making the measurement. Examples of measured numbers would include the length of a piece ribbon, the weight of a chicken at the store, or the amount of sugar added to a recipe measured in cups or teaspoons. All would have some degree of uncertainty because a measurement is involved.

**Uncertainty in Measurements**

It is important to use the correct number of **significant figures** when reporting a measured number, just as it is important to report the units for the number. Significant figures apply only to measured numbers, and not to counted or defined numbers. To report the appropriate number of significant figures in a measured number, we report all the digits we know with certainty, plus one estimated digit.   
  
Sets of measurements can be characterized by the **precision** or the **accuracy** of the measurements. The precision of a set of measurements reveals how close the measurements in the set are to each other. The accuracy of a measurement involves how close that measurement is to a “known”, “true”, or “standard” value. Thus, precision and accuracy are quite different characteristics. It is possible for measurements to be very precise, but not accurate. They may be very precise and very accurate or they may be neither precise nor accurate.  
  
**Rules for Counting Significant Figures in a Number**

All nonzero digits are always significant.   
Zeros are always significant if they fall between two nonzero digits.

Zeros are never significant if they come before the first nonzero digit.

Zeros at the end of a number are only significant in a number with a decimal point.

**Rules for Rounding to the Correct Number of Significant Digits**Round a number by looking at the first digit after the last significant digit. If it is 0 to 4, round down to the next lower digit. If it is a 5 or greater, round up to the next larger digit. Thus, to round 755.4 to three significant digits, round down to 755. To round 755.5 to three significant digits, round up to 756.

**Rules for Significant Figures in a Calculation Involving Addition or Subtraction**Round the answer to the same number of **decimal places** as the number with the **fewest decimal places** used in the calculation.

**Rules for Significant Figures in a Calculation Involving Multiplication or Division**

Round the answer to the same number of **significant figures** as the number with the fewest **significant figures** used in the calculation.

**Rules for Significant Figures in a Calculation Involving both Addition/Subtraction and Multiplication/Division**

Do the operations in the order designated by the order of operations, rounding the digits at the intermediate steps.

**Units of Measurement**

In order for scientists to be able to compare their results a standard system of units was adopted internationally, known as the SI Units. In the SI system, the basic unit of length is the meter (abbreviated as “m”), the basic unit of mass is the kilogram (abbreviated as “kg”), the basic unit of time is the second (abbreviated as “s”), and the basic unit for temperature is the Kelvin (abbreviated as “K”). Although the SI system is based on the metric system, there are some differences. In the metric system, the gram (abbreviated as “g”) rather than the kilogram is the unit of mass and the liter (abbreviated as “L”) is the unit of volume. Note that there is no basic unit of volume in the SI system. You should memorize the basic units used in both the SI system and the metric system.   
  
**Length**

**Length** is a measure of a single linear dimension of an object, usually the longest dimension. The SI unit for length is the meter, which is slightly longer than the English unit “yard”. When measuring very long or very short lengths, it is common to use a prefix in front of the SI unit of meters for convenience.

**Mass vs Weight**

**Mass** is a measure of the amount of matter and it is the same anywhere, on earth or another planet. Mass should be differentiated from **weight**,which is a measure of the gravitational force exerted by the object. The weight of an object varies with location and depends on the force of gravity. For example the mass of astronaut Neil Armstrong was the same on the moon as it was on earth; however, his weight on the moon was only 1/6 that of his weight on earth because the force of gravity on the moon is 1/6 that of earth. In outer space where there is no gravitational force, an object is weightless but not massless. We use the term weight when we actually mean mass in the lab because most of the work we do involves comparing masses measured in the same location subject to the same force of gravity.   
  
**Prefixes in the Metric System**

In science, we may deal with very large or very small numbers. For example, the number of molecules in a drop of water is a very large number while the diameter of an atom is a very small number. In the metric system, **prefixes** are attached to the base units to increase or decrease the value of the base unit by factors of 10. You must memorize the following prefixes, their symbols and meaning and be able to convert from one unit to another: mega-, kilo-, centi-, milli-, and micro-. The symbols and meanings for these prefixes are in Table 2.2 in your textbook on page 24, as well as in the slides, and in a separate handout in the Chapter 2 folder under Course Documents. Keep them handy until you have them memorized!

**Volume**

As mentioned in the notes above, there is no basic unit for volume in the SI system as there is in the metric system. Rather, volume units are derived from the basic unit for length. Remember that units can be multiplied just like numbers can. To calculate the volume of a cube or a box, you multiply length times width times height carrying the units along in your calculation. Thus the volume of a box that is 1 meter tall by 2 meters wide x 8 meters long would be 16 cubic meters: 1 m x 2 m x 8 m = 16 m3.

**Common Units and Their Equivalents**

Often, it is necessary to convert measurements in one system to their equivalents in another system. To do this, it is necessary to have a conversion factor for doing the calculation.

Chapter 02: Measurements and Problem-Solving   
Part 2: Problem-Solving and Dimensional Analysis

**Units**

All measurements consists of two parts: a **number** and **unit**. Keeping the number and the units together is very important. A number without an accompanying unit is meaningless. **Always** **(!!)** write out the units for the numbers you use in calculations and problem-solving. The cool thing about units is that you can do the same algebraic operations with units that you do with numbers and in this course, you will learn a method of problem-solving in which the units of the numbers will help lead you to the correct calculation.

**Conversion Factors**

Often, it is necessary to convert measurements in one system to their equivalents in another system. To do this, it is necessary to have a conversion factor for doing the calculation.

**Using Conversion Factors to Convert Units**

A **conversion factor** gives the equivalent relationship between two units. For example, the equivalent relationship between a pound (lb) and grams (g) is defined as 1 lb = 453.6 g. This relationship can be written as a conversion factor in the form of a fraction in two ways:

**a)** **1 lb/453.6 g or b) 453.6 g/lb**

A general formula for converting a quantity from one unit to another is as follows: **multiply given quantity with conversion factor (written as a fraction):**

**Given quantity \* conversion factor = desired quantity**

To convert from one unit to another you choose the conversion factor that allows you to cancel the given unit and obtain the desired unit. Therefore, the desired unit must be in the numerator and the given unit in the denominator.

Example: Convert 500 g to lbs.

Step 1. Analyze: given unit is g; desired unit is lbs

Step 2. Choose conversion factor with lb in numerator and g in denominator [a) from above]

Step 3. Write given quantity with unit x conversion factor = desired unit

500 g \* (1lb/453.6 g) = 1.1 lb g cancels out

**Solution Maps**

Students often report that the thing they dread most is solving a word problem or a “story problem” and yet everybody solves routine “word problems” every day: these problems are presented in every day terms like figuring how much of something you need to buy to make a meal, how to find an alternate route to work to avoid construction, or how much gasoline the money in your wallet can purchase. Although students can often figure out answers to these problems without much difficulty, they freeze up when presented with a problem with “chemistry units” in it.

The most difficult aspect of solving such a problem is often just figuring out where to begin. It can be very helpful to map out a potential solution to a problem using an outline form that indicates what needs to be done in each step without worrying about the exact numeric values to be used in making the conversion. These solution maps give a visual guide to follow in actually working the problem. When doing a unit conversion problem, the solution map will focus on the units and will indicate what conversion factors are needed to be able to convert from one unit to the next.   
  
Making certain that you routinely follow a systematic approach to doing unit conversion and dimensional analysis will benefit you by giving you a place to start and it will simplify the process of getting to the final solution. One reason this textbook was selected for this course is that presents problem-solving in a very clear manner. Your instructors hope you will take advantage of this. The main difference between an expert problem-solver and a beginning student is that an expert follows a systematic process. Learn the **process** of solving problems rather than trying to memorize formulas and with practice, you will find word problems much less daunting.

**Systematic Approach to Solving Problems**

The process presented in the textbook and the powerpoint slides is similar to using a road map to plan a trip across the country.

1. *Begin by figuring out where you are now.* What do you already know? This includes the quantities you are given in the problem; be sure you include the units **always**!
2. *Where do you want to go?* What quantity is the problem asking you to find? What are its units? It is important to figure this out early so you can figure out the shortest way to get from where you already are to your final destination.
3. *What cities would you like to pass through on your way from here to there? Do you want the most direct route or the scenic route?* For solving unit conversion problems, this would be similar to identifying what equivalence statements might be useful. You will want to find equivalence statements that have units in common with the value you are given to start with and the number you want to end up with. Each of your equivalence statements can be converted to two possible conversion factors.
4. *Lay out the route.* Take the conversion factors and write out your solution map. In doing this, begin with the units of the value you are given and figure out which conversion factor can cancel the units of the number you started with. Remember that in order to cancel, the unit must appear in the numerator of one of the numbers and in the denominator of the other one.
5. *Figure out distances and times and finalize your plans.* Once you have all your conversion factors lined up, it is time to do the calculations. Multiply all the terms in the numerators together and then divide by each bottom term. (A common mistake students make is that they divide by the first term on the bottom and then multiply by the rest. That does not work!)
6. *Double-check your plans.* Make sure they take you where you want to go. Is your answer reasonable? Does it make sense? If you have just calculated that a person is 73 feet tall or that a factory makes only 4 tiny ball bearings per day, you have probably made a mistake somewhere along the way.

You must practice problem-solving to become good at it., just as you must practice a musical instrument or a sport if you want to excel. A little practice every day is better than one long practice a week. Ask any music teacher or coach!   
  
We recommend that you **work through the example problems** in the powerpoint slides and in the textbook. Don’t just read them! Get out a calculator, pencil, and paper and work the practice problems! Just reading a practice problem is a little like trying to learn to swim or figure skate by watching the Olympics on TV! You might learn a few things, but you probably won’t be ready to compete on your own. If the first time you try to work a problem on your own is on the exam, I guarantee you will not like the results!

**Density**

Why does a cork float while a penny sinks in water? This is because of the physical property termed density. Differences in density determine whether objects sink or float. We can measure the mass and volume of any object. However, just looking at those measurements separately does not give us an idea of how closely packed the particles are in the object. If we now compare the mass of the object to its volume, we obtain the relationship called density. We can therefore define **density** verbally as the mass of a substance per unit volume. Mathematically, we can write the formula or equation as:



To calculate the density of an object, you simply substitute the values (including units) for mass and volume and then divide. You must always give the units for your final answer. Without the units, the number is meaningless. The above equation can be rearranged so we can calculate an unknown quantity if any two are known. For example if we know the density and mass of an object we can calculate its volume by rearranging the equation to:

 or

calculate the mass of an object if we know the density and the volume:

*mass(m) = density (d) \* volume (v)*

Again, please work through the example problems presented in the powerpoint slides!

Chapter 03: Matter and Energy

**Matter**

As you have read previously, **matter** is anything that has **mass** and **occupies space**; it is the physical material of the universe. Everything in the universe is composed of matter. Furthermore, all matter is composed of atoms and molecules, which we will study in more detail as we progress through this course. Chemists try to understand how the universe and everything in it works on a large scale by studying how the atoms and molecules interact with each other on a submicroscopic scale.

**The Particles of Matter**

In chemistry, matter can exist as either one of two types of particles or building blocks. 1) An **atom** is the smallest particle of an element that retains the characteristics of the element, for example the hydrogen atom, symbol “H” or copper atoms, symbol “Cu.” 2) A **molecule** is made up of 2 or more atoms chemically combined. In a molecule, the atoms can be identical, or there can be two or more different kinds of atoms. For example the hydrogen molecule, symbol “H2” has two hydrogen atoms bonded together, but a water molecule is made up of two hydrogen atoms and one oxygen atom, and we write the formula for it as “H2O.”

**The States of Matter**

A sample of matter can be observed in any one of three different **physical forms or state**: solid, liquid or gas. A **solid** has a definite shape or volume because the particles making up the object are very close to each other and relatively strong forces of attraction between them exist. The particles can’t move very much except vibrate in place. A **liquid** has a definite volume but no definite shape because the forces of attraction between particles are weaker than they are in a solid. The particles in a liquid are a little farther apart thaan they are in a solid; this allows the particles to slide past each other. A liquid assumes the shape of the occupied part of its container. A **gas** has no definite volume or shape. The particles are so far apart from each other that there are few forces of attraction between them. Each particle in a gas moves freely and independently of each other. A gas will fill and assume the shape of its container.

**Classifying Matter**

Matter can be classified into **pure substances** or **mixtures** based on chemical composition.

**Types of Pure Substances**

A **pure substance** is matter made up of **one type of particle** with **definite composition and distinct characteristics.** A pure substance can be either an **element** or a **compound**. If you take several samples from a pure substance, they will have identical properties because all the particles have the same composition. For example, pure distilled water has only H2O molecules, no matter where the distilled water comes from.

An **element** is a substance that cannot be broken down into simpler substances by chemical reactions. It is the fundamental type of matter from which all other matter is composed. All elements known are shown in the Periodic Table in their symbolic notation. Examples are carbon, symbolized “C” and copper, symbolized “Cu.” **You must memorize the symbols for the elements in slide 19 of the powerpoint slide show for this chapter**. The study of chemistry is like learning a new language. The symbols of the elements make up the alphabet. If you don’t know the alphabet of a language, you cannot make words and sentences. If you don’t memorize the symbols, you will not be able to write chemical formulas (words) or chemical equations (sentences).   
  
 An **element** consists of **one type of atom** but it can exist as an atom or a molecule. Examples of elements that exist in nature as individual atoms include Na, C, Fe, Ca, etc. Some elements are more stable in nature as **diatomic molecules** (two atoms of the same type), such as O2, N2, Cl2, etc. In an element that exists naturally in the diatomic form, the two atoms are always identical.

A **compound** is made up of **two or more elements or two or more types of atoms**, chemically combined and therefore exists as molecules. Examples of compounds are water, H2O; sulfuric acid, H2SO4; carbon monoxide, CO.. Although there are two or more different types of atoms present, it is important to realize that a compound has a fixed composition. That means that in water, there are always two hydrogen atoms bonded to one oxygen atom, no matter where the water comes from. In glucose, there must be six carbon atoms, 12 hydrogen atoms, and six oxygen atoms combined in a specific way. If there are more carbon atoms, fewer hydrogen atoms, or if the atoms are arranged differently, the material is not glucose.

A **mixture**, on the other hand, is made up of **two or more types of particles**, which **retain their chemical identity** and **can be separated from one another by physical methods.** For example, you can compare two different brands of Italian salad dressing. Both will have oil and water as their main ingredients and both will have particles of herbs and spices floating in them but one would not expect the mixtures to be identical if they are made by different companies. You can separate the different components by physical means by allowing the mixture to separate into oil and vinegar layers and then filtering off the solids.

Mixtures can either be **homogeneous** or **heterogeneous.**

In a **homogeneous mixture**, samples appear to be uniform throughout. An example would be Kool-Aid dissolved in water. If it has been prepared properly, the last glass in the pitcher should look and taste exactly like the first glass poured from the pitcher. **Homogeneous mixtures** are also called **solutions**.

Chocolate chip cookie dough is an example of a **heterogeneous mixture**. It is easy to see that it is a **mixture** because the chocolate chips are definitely different from the surrounding dough. It is **heterogeneous** because it is difficult to guarantee that the last scoop you place on the cookie sheet will have exactly the same number of chocolate chips as the first scoop did.

**Chemical vs Physical Properties**

How do we differentiate a sample of matter from another? When we compare samples of matter, we look at the properties or characteristics that distinguish them from one another. The properties may be chemical properties or physical properties.

**Chemical properties** describe the **composition** (what is it made of) of a sample of matter and the way it may **change or react** with another material to form new substances (chemical reaction). Examples of chemical properties include elemental composition (for example, water is made of 11.2% hydrogen and 88.8% oxygen), flammability (whether something will burn), or inertness (lack of reactivity). **Chemical changes** are changes that result in formation of completely different substances. Examples of chemical changes include formation of rust or burning gasoline in a car engine. Once a chemical change occurs, it is usually not easy to return to the original material.

**Physical properties** describe characteristics that can be determined without changing the composition or chemical identity of the substance. Examples are boiling point, density, color, physical state, hardness or softness, etc. It is possible to melt ice to form water, boil the water to form steam, condense the steam back to liquid water, and freeze the water into ice. In this case the water never changes its composition, it just changes its physical form. Another example would be dissolving salt in water. The salt seems to “disappear” but if the water is allowed to evaporate, the salt can be recovered, virtually unchanged.

**Separation of Mixtures**

A mixture can be separated based on differences in the physical properties of the substances in the mixture. Any difference in physical properties can be the basis for a method of separation. For example, liquids that boil at different temperatures can be separated from each other by distillation. The liquid with the lower boiling point will evaporate more easily, resulting in separation. A method that we will study later in the course is paper chromatography. In paper chromatography, materials can be separated based on their affinity for being attracted to the cellulose molecules in paper compared to their solubility in a solvent.

**Law of Conservation of Mass**

Remember that a **scientific law** is a brief statement or mathematical equation summarizing a large body of data or observation or phenomenon. The Law of Conservation of Mass says that in a chemical reaction, matter is neither created nor destroyed. This law was first proposed by Antoine Lavoisier, who did many experiments in which he found that the masses of all the products in a chemical reaction equaled the total masses of all the reactants. This leads to the observation that matter is neither created nor destroyed in a chemical reaction.

**Energy**

Physical and chemical changes of matter are always accompanied by energy changes. How do we define energy? When I am very tired and asked to do something, I might say “no, I don’t have the energy to do that.” When you are walking, running, thinking, etc. you are using energy to do work which might involve changing the physical or chemical property of matter. We therefore define **energy** as the ability to do work or the ability to change matter, either physically or chemically. In earlier notes, we mentioned that chemistry is the study of the properties and interactions of matter. Energy is important to these interactions; the flow of energy determines when a reaction may or may not take place.

The **law of conservation of energy** says that energy is conserved, just as matter is. Energy can be changed from one form to another form, but it cannot be created or destroyed in a chemical reaction.

There are two main forms of energy, potential and kinetic energy. **Potential energy** is stored energy while **kinetic energy** is energy of motion. A boulder perched on the top of a hill has potential energy because of its position. If it is nudged, it rolls down the hill and potential energy is converted to kinetic energy. Water stored in a reservoir is potential energy that becomes kinetic energy as it goes over the dam. The food you eat has potential energy stored in its chemical bonds. It is converted to kinetic energy during biological activity such as muscle contraction. Energy can therefore be converted from one form to another. The ultimate source of energy on our planet is the sun. Specific forms of kinetic energy include electrical energy, heat or thermal energy, and light or radiant energy. Nuclear energy and chemical energy are specific examples of potential energy.

**Temperature**

When we want to know how hot or cold an object is, we measure its temperature. What makes an object hot or cold? Temperature is a measure of the kinetic energy of the atoms &/or molecules making up the object. The faster they are moving, the higher the kinetic energy and the higher the temperature is of the object. In science, we define absolute zero as the lowest possible temperature where there is no molecular or atomic motion whatsoever. There are three different temperature scales in use today; the Kelvin (K), Celsius (°C) and Fahrenheit (°F) scales. The Kelvin scale is an absolute scale because 0 K is set at absolute zero. The Celsius and Fahrenheit scales are based on the freezing and boiling points of water. In science, the Kelvin and Celsius scales are commonly used. 0 K = - 273 °C or 0 °C = 273 K, this is the conversion factor between these two scales. To convert Celsius to Kelvin: K = °C + 27

Chapter 04: Atoms and Elements

Atoms

In the previous chapter we learned that an **atom** is the smallest particle of an element that retains the characteristics of that element. Where did this term come from? Centuries ago Greek philosophers speculated about the nature of matter. The Greek philosopher Leucippus was the first to propose that substances could be subdivided into very small particles. His student, Democritus called these particles “*atomos*” which meant “indivisible” in Greek. The difference between the Greek philosophers and scientists who came later is that scientists used experiments to confirm and validate their hypotheses or to modify and refine them.

Is the atom really indivisible? In these chapters we will learn more about atoms, their structure and the development of the atomic theory of matter and we will learn more about whether atoms are “indivisible.”

Dalton’s Atomic Theory

Dalton proposed his atomic theory to explain the chemical laws of combination and stated four postulates as written in the slides.

**Postulate 1**, “all matter is composed of very small particles called atoms,” explains that atoms are the building blocks of matter.

**Postulate 2**, “all atoms of a given element are alike and they are different from atoms of another element,” explains that atoms of a given element are identical (no longer true) and are different from atoms of any other element.

**Postulate 3**, explains the **law of constant composition** – when **compounds** (pure substances made of two or more elements which have been chemically combined) form, the elements combine in fixed proportions. So when a compound is decomposed into its component elements, the elements will be found in the same proportions regardless of where the compound came from. Scientists observed that **each compound was always composed of the same elements in the same proportions, no matter where it came from**. The work of Joseph Proust provided convincing evidence for this observation. He demonstrated that when copper carbonate is decomposed into its component elements, it was always made of 51% copper, 39% oxygen and 10% carbon. From the same observations with many different compounds, he formulated the **law of constant composition**.

Each molecule of a compound will contain exactly the same types of atoms in the same numbers. This means that each compound can be represented by a chemical formula that describes the types and numbers of atoms in the compound. The law of constant composition is also sometimes called the **law of definite proportions**.

**Postulate 4,** explains the **law of conservation of mass**. Antoine Lavoisier, an 18th century French chemist who did many experiments studying chemical reactions made the following observation: **if a chemical reaction is carried out in a closed system, the total mass of the system remained constant**. Atoms are not created or destroyed in a **chemical reaction**; the atoms are simply rearranged forming new compounds. The total mass of products formed in a reaction equals the total of the starting materials (reactants) present before the reaction begins. For example: we write C + O2 → CO2 to mean that carbon reacts with oxygen to form carbon dioxide. No new atoms were created; no atoms were destroyed; the carbon and oxygen atoms were rearranged to form carbon dioxide. Scientists found the same result for many different reactions as long as the reaction was carried out in a closed system. (Remember that a **natural law** describes observations that hold true for many different systems!) Later, we will see that the law of conservation of mass is the basis for being able to balance chemical reactions.

This postulate also explains why it is not possible to turn lead into gold using a chemical reaction. To do so would require changing one type of atom into another type of atom. In an ordinary chemical reaction, this is not possible.

### Law of Multiple Proportions

One consequence of Dalton’s postulates is called the **Law of Multiple Proportions**. Understanding this law requires first that we remember that any compound made of two elements will have a constant ratio of element A to element B. For example, in **carbon monoxide** (a colorless, toxic gas which has the formula CO,) the mass of oxygen to carbon is **always** 1.33 grams of oxygen to each 1 gram of carbon. Carbon and oxygen can also form **carbon dioxide** (also a colorless gas with the chemical formula CO2 which is formed in the respiration process.) In **carbon dioxide**, the ratio of oxygen to carbon is always 2.67 grams of oxygen per 1 gram of carbon. If we compare the ratio for carbon dioxide to the ratio for carbon monoxide we get :   
2.67 : 1.33 which is a ratio of 2. Ultimately, this led to the understanding that atoms have to combine with other atoms in whole numbers. In other words, you can combine one oxygen atom with one carbon atom or two oxygen atoms to one carbon atom, but you can never make a compound that will have 1.6 oxygen atoms to one carbon atom.   
  
When you are comparing compounds to investigate the ratios of the elements, it is very important to remember that you must have the same two elements present. Thus, you can compare NO, NO2, and N2O4, but you can’t compare NO2 with NH3 and you can’t compare NO2 with HNO2.

Modern Evidence for the Atomic Theory

Early scientists based their beliefs in the existence of atoms on relatively crude experiments. Today modern instruments such as the scanning tunneling microscope provide images of atoms and molecules. Imagine how excited Dalton and other scientists of the nineteenth century would be to see this type of additional evidence supporting the work they did 150-200 years ago!

Mass of Atoms

Dalton performed many experiments in synthesizing and decomposing compounds to learn about the rules governing chemical composition. As a result, he developed a scale of the relative masses of different types of atoms. His scale was based on each hydrogen atom having a mass of 1 unit, which we call an **atomic mass unit** (**amu**). The modern scale for atomic masses is now based on carbon rather than on hydrogen. In the modern atomic mass scale, a single atom of carbon-12 has a mass of exactly 12 amu by definition.

Atomic Structure- the atom is actually divisible!

The development of the atomic theory of matter is a clear example of how scientists practice science by means of the scientific method. When John Dalton formulated his theory, he considered the atom to be indivisible, just as the Greek philosopher Democritus did. However, more experimental data began to accumulate indicating that the atom can be subdivided into smaller particles. Dalton’s theory has been modified to what is now the modern atomic theory.

Modification of Dalton’s atomic theory began in the late nineteenth century when work done by J.J. Thomson and others led to the discovery of tiny negatively-charged particles emitted from metal electrodes in a high voltage field and observed in a cathode ray tube (similar to old-fashioned curved TV or computer screens). These particles became known as **electrons**. Since atoms were known to be neutral, Thomson concluded that the charge of these negatively-charged particles must be balanced within the atom by positively-charged particles. Remember that the internal structure of an atom was not yet known, so Thomson proposed that atoms were spheres of positive charge dotted throughout with negatively charged electrons. This became known as the **plum-pudding model** of the atom; a more familiar modern image might be to imagine a muffin made of positive charges dotted throughout with negatively charged blueberries.

Rutherford and the Nuclear Atom

Rutherford’s famous experiment involving bombarding gold foil with alpha-particles (small positive particles) began as an experiment to confirm the plum-pudding model of the atom. The **hypothesis** was that if the plum-pudding model was correct, and charge and mass were evenly distributed throughout the atom, then the alpha-particles should pass right through the gold foil with little deflection. The results were completely astonishing! Some particles did pass through the gold foil with little deflection, some were deflected at larger angles, and some **bounced back** toward their source! To Rutherford, this was as believable as if “you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you.” The result is a major modification to the plum-pudding model and a revision to the prevailing atomic theory at that time. (This is a great example of the scientific method at work! Unexpected experimental results lead to modification of a hypothesis or theory.)

Rutherford needed to develop a new model to explain the results of his experiment. In particular, it seemed that the atom must be mostly empty space, since most of the alpha particles passed right through the foil without being deflected. However, the particles that were deflected must have met with some obstacles that kept them from passing through; that could be explained if the alpha particles encountered some kind of dense particles on their way through the foil. Furthermore, these dense particles might be positively charged, which would explain the large deflections of the positively charged alpha-particles, since the positive charges would repel each other.

As a result of Rutherford’s experiments, a new nuclear model for the structure of the atom was proposed. In this model, the atom contained a dense center called the nucleus. The nucleus contains most of the mass of the atom and it is positively charged. The negatively-electrons contribute little to the mass of the atom and they move around in the empty space surrounding the nucleus.

Subatomic Particles

Although the atom is the smallest particle of an element that retains its characteristics, the atom is actually made up of three types of smaller subatomic particles:

uncharged neutrons (n),

positively charged protons(p), each possessing a charge of +1 and

negatively charged electrons (e) each with a charge of –1.

The protons and neutrons are located in the **nucleus**, the central, very small and dense portion of the atom. The negatively charged electrons orbit around the nucleus. While these particles are very small, they do have masses. **The neutron and proton each has a mass of approximately 1 atomic mass unit (amu)**, while the mass of the electron is considered to be negligible (essentially zero) compared to the mass of the neutron and proton. Thus the total mass of an atom is mostly due to its protons and neutrons. On the other hand, the volume (space occupied) of an atom is mainly due to its electrons orbiting outside the very small nucleus. The atom is held together by the force of attraction between the positively charged protons in the nucleus and the negatively charged electrons orbiting around it.

Element Identity Based on Atomic Number

Each element has a characteristic number of protons, electrons and neutrons. The number of protons in an atom is called the **atomic number**, which identifies the element. All the elements known today are arranged on the **periodic table** according to **increasing atomic number** (refer to the periodic table in inside cover of your textbook). The number of electrons equals the number of protons in a **neutral atom**. The number of neutrons varies in the atoms of an element and from one element to another.

Periodic Table

As mentioned above, **elements are arranged in the periodic table according to their atomic number**. If you look at the periodic table inside the front cover of your book, the **atomic number** is the number printed right above the **element symbol**. The element symbol is a one- or two-character symbol, which we use as a shorthand way of indicating an element. For example, hydrogen, denoted by a capital “H” is the first element in the periodic table and has an atomic number of “1”. Helium is denoted by “He” and has an atomic number of “2”. Note that one-character element symbols are always capital letters and two-character symbols have one capital letter followed by a lower-case letter. (This is very important to avoid confusion in writing formulas!)

The periodic table is a very important tool for chemists, and provides much chemical information in a relatively compact space. For example, the location of an element in the periodic table indicates whether the element is a **metal** (generally located on the left side of the periodic table), **non-metal** (generally located on the right side of the periodic table) or a **metalloid** (occurring along a zig-zag line between the metals and the nonmetals.)

**Units of Measurement**

SI Base Units

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Physical Quantity Name of Unit SI Symbol**

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Length Meter m

Mass Kilogram kg

Time Second (s) Second (s)

Temperature Celsius (°C) Kelvin (K)

Amount of substance Mole mol

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# Metric Prefixes

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## Prefix Symbol Meaning Number Scientific Notation

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## Prefixes that increase the size of the unit

mega- M one million 1 000 000 106

kilo- k one thousand 1 000 103

## Prefixes that decrease the size of the unit

deci- d one tenth 0.1 10‑1

centi- c one hundredth 0.01 10-2

milli m one thousandth 0.001 10-3

micro- μ one millionth 0.000 001 10-6

nano- n one billionth 0.000 000 001 10-9

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## Converting Within the Metric System: Examples of Conversion Factors

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## Prefix Base Unit Symbol Unit Equivalence Conversion Factors

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Mega- meter Mm 1 Mm = 106 m  or 

Kilo- meter km 1 km = 103 m  or 

1 meter

Deci- meter dm 1 dm = 0.1 or 10-1 m  or 

Centi- meter cm 1cm = 0.01 or 10-2 m  or 

Milli- meter mm 1mm = 0.001 or 10-3 m  or 

Rules for Determining Significant Figures in Measurements

1. All nonzero numbers are significant.
2. Zeros may or may not be significant depending on their position in the number. Examples are shown on the table below.

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# Rule Measured # of Significant

# Number Figures

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. **A number is a *significant figure* if it is**
   1. A nonzero digit 6.5 g 2

132.34 m 5

* 1. A zero between non zero digits 305 m 3

2.056 kg 4

* 1. A zero at the end of a decimal number 50. L 2

28.0 cm 3

18.00 g 4

* 1. Any digit in a number written in 4.0 \* 105 m 2

scientific notation 6.70 \* 10-3 g 3

1. **A number is not *significant* if it is**
   1. A zero at the beginning of a decimal 0.0008 kg 1

number (between a decimal point and 0.0953 m 3

a nonzero digit)

A zero used as a placeholder in a large 750 000 km 2

number without a decimal point 1 430 000 mm 3

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

## Rules for Rounding Off

When doing calculations from measured numbers using a calculator, you get answers that give several digits. However, your answer cannot be more precise than your actual measurements. For example if you are calculating the area of a piece of cloth that measures 6.3 m by 3.4 m, the answer is 21.42 m2. Your measurements only have 2 significant figures; so all four digits cannot be significant. You must round off your final answer to two significant figures: 21 m2. There are two rules to remember when rounding off numbers:

1. If the first digit to be dropped is **4 or less**, it and the following digits are just dropped.

Ex. Rounding off 5.3132 to 3 significant figures = 5.31

1. If the first digit to be dropped is **5 or greater**, the last retained digit is increased by 1.

Ex. Rounding off 15.684 to 3 significant figures = 15.7

Metals

Metals are relatively easy to recognize when we encounter them. They are usually shiny solids, although you might be familiar with mercury, a metal that is liquid at room temperature. Metals conduct heat and electricity well (for example: aluminum cookware, copper wiring in a home.) In addition they can be hammered into shapes (jewelry, decorative ironwork, horseshoes), or drawn into thin wires. These are physical characteristics of metals that we can see easily.

On an atomic level, **metals are elements that easily lose electrons to form positively charged particles** we call cations (positive ions) in chemical reactions and in compounds.

Approximately ¾ of the elements known are metals.

Nonmetals

Although metals are usually solid, nonmetals may be solids, liquids, or gases. Unlike metals, they do not conduct electricity well. They have a very wide variety of physical characteristics such as color and form.

On an atomic level, nonmetals tend to gain electrons in chemical reactions. When a nonmetal gains one ore more electrons, it becomes a negatively—charged particle called an anion ( a negative ion.)

There are far fewer nonmetals in the periodic table than metals. They are generally located in the upper right corner of the periodic table.

Metalloids

Metalloids share some of the properties of both metals and nonmetals. They are a small group of elements that fall on either side of a zip-zag line toward the right side of the periodic table. Metalloids such as silicon conduct electricity well but do not conduct heat. They are also called semiconductors.

Periodic Table- Groups and Periods

In the periodic table, elements that are found in the same vertical column are called a group or family. Elements within a group or family have similar chemistry because they have similar patterns in their electron structure. For example, lithium, sodium, and potassium have similar chemistry because they all tend to want to lose one electron. Fluorine, chlorine, and bromine all want to gain one electron, so their chemistry is similar. Each group is designated by a number or number-letter combination.

Elements that are found in the same horizontal row or period of the periodic table will show a pattern of properties; for example, moving from left to right along a row, the atoms decrease in size. The pattern of those properties will be repeated in other horizontal rows (periods) of the periodic table.

Important Groupings of Elements

**Main group elements** are indicated in the periodic table in the front of the book by column designations containing an “A”. Main group elements are also called **representative elements.** They are located in the first two groups at the left of the periodic table and in the six groups at the right side of the periodic table.

**Transition metals** are designated by group numbers containing “B”. They are in the middle of the periodic table.

**Inner transition** or **rare earth elements** are found in the two separate rows at the bottom of the periodic table. The upper row are called **lanthanides** and the lower row are the **actinides**. These two rows really should be included in period 6 and period 7, respectively, but to do so would make the periodic table very long from left to right.

Although **hydrogen** appears above the first group at the left of the periodic table, it is quite unique in its properties. It really is a group unto itself. Unlike the other members of the group, hydrogen is a nonmetal rather than a metal. It is a colorless, **diatomic** gas, which means it occurs naturally in molecules consisting of two hydrogen atoms. It reacts with other nonmetals to form molecular compounds and reacts with metals to form hydrides. Ability to release hydrogen ions is an important characteristic of many acids, so we will study more about this element later in the course.

The group that is the furthest left in the periodic table (Group 1A) is called the **alkali metals.** These are very reactive, soft metals that are not found uncombined in nature. They react with water to form alkaline solutions (bases).

Group IIA, the **alkaline earth metals** are reactive, but less so than the alkali metals. They are harder, and denser than alkali metals.

Skipping to the right side of the periodic table, the elements of group VIIA are called **halogens**. All of the elements in this group are diatomic as elements- they exist in 2-atom molecules in their pure elemental forms. Although fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid. Like the alkali metals at the other side of the periodic table, the halogens are very reactive.  
  
The group furthest right in the periodic table, Group VIIIA is the **noble gas** group. The name “noble gas” comes from the fact that these elements are all very unreactive. Their atomic structures are full of electrons, so they have no need to give electrons up, accept electrons, or share electrons from other elements- in other words, no need to react with other elements. We will see that the other elements in the periodic table tend to react in ways that allow those other elements to achieve electron structures similar to those of the noble gases.

Ions

We learned above that an element can be identified by how many protons are in the nucleus of the atom. All atoms of a particular element will have the same number of protons. It is important to note that the number of protons in the nucleus does not change when an element is involved in a chemical reaction.

However, atoms (or molecules) can lose or gain electrons forming charged particles called **ions**. When an atom (or molecule) loses an electron, the atom becomes positively charged (+), and is called a **cation**. When it gains an electron(s), it becomes negatively charged and is called an **anion**. The atom F, becomes an anion by gaining 1 electron giving it an extra negative charge, forming the fluoride ion, F−. The atom Ca, becomes a cation by losing 2 electrons, giving it an excess of 2 + charges forming the calcium ion, Ca2+.

To determine the charge on an ion, subtract the number of electrons from the number of protons. Remember that in a neutral atom, the number of electrons is the same as the atomic number or number of protons in the nucleus.

If the **ion charge = # protons minus # electrons,** the charge on an ion will always be negative if the neutral atom or molecule has gained extra electrons. The charge will be positive, if the neutral atom loses electrons because then the number of electrons is smaller than the number of protons.   
  
**Metals form cations by losing electrons.** **To name a cation,** use the name of the metal from which it is formed followed by the word “ion”. A sodium atom can lose one one electron and become sodium ion, or a magnesium atom can lose two electrons and become a magnesium ion.

To **predict the charge on a cation among the representative elements only**, look at the group number for that element in the periodic table. Alkali metals will form ions with a 1+ charge. Alkaline earth metals will form ions with 2+ charges, and elements under aluminum in the group IIIA will form ions with 3+ charges.   
  
**Nonmetals form ions by gaining electrons**, so they form anions. **To name an anion**, change the ending of the element name to “-ide” followed by the suffix “ion”. Chlorine will become chlor**ide** ion, fluorine will become fluor**ide** ion, and oxygen will become ox**ide** ion.

To predict the charge on an anion, subtract 8 from the group number of the element. Halogens in group VIIA will therefore have a -1 charge as anions because 7-8= -1.

Isotopes

As we mentioned earlier, the discovery of subatomic particles required modification of Dalton’s atomic theory. Recall that his second postulate stated, “atoms of an element are alike but different from atoms of another element.” Years after discovery of subatomic particles it was found that atoms of a given element are **not all alike**. Atoms of a given element will have the same number of protons but may vary in the number of neutrons. Atoms with the same number of protons but different number of neutrons are called **isotopes**. Fro example, hydrogen has 3 known isotopes. They all have 1 proton and 1 electron. They differ in their number of neutrons as seen on the slide. Some isotopes of an element are unstable and are **radioactive** (spontaneously decay). The hydrogen isotope that has two neutrons is radioactive.

Isotopes are identified by their mass number, that is, the number of protons plus neutrons. Hydrogen itself has one proton and no neutrons so its mass number is 1. The isotope of hydrogen with one proton and one neutron has a mass number of 2, and the isotope of hydrogen with one proton and 2 neutrons has a mass number of 3. Those two isotopes of hydrogen are called deuterium and tritium, respectively.

Because reactivity in a chemical reaction does not depend on what is in the nucleus, all isotopes of a given element will act and react identically in a reaction.

Designating an isotope with Chemical symbols

Chemists use a system of symbols to represent an atom and its subatomic composition that identifies the specific atom. The previous notes indicate that the mass of an atom is essentially entirely due to its protons and neutrons. We define a quantity, called the **mass number**, which is the sum of the number of protons and neutrons in the nucleus of an atom. We use the mass number together with the atomic number and the element symbol to specify which isotope of an element is under consideration.

 

In this symbolic notation, “X” represents the symbol for the element, the atomic number (number of protons) represented as “Z” is written as a subscript on the front bottom of the symbol, and the mass number (no. of protons + no. of neutrons) represented as “A” is written as a superscript on the top front of the symbol. For example, the symbol on the right indicates that the element is neon, which has an atomic number of 10, and a mass number of 20, meaning there are 10 protons and 10 neutrons. Sometimes, the atomic number, which identifies the atom, is omitted because it is understood from the symbol for the element. When the atomic number is not given, it is easily found in the periodic table.

Determining Number of Subatomic Particles

The symbolic representation allows us to determine the number of subatomic particles in an atom when the symbol for the element and the mass number are given. In the first example:



we look in the periodic table and see that the atomic number of chromium, Cr, is 24 which means it has 24 protons. Since

Mass number = no. of protons + no. of neutrons  
 No. of neutrons = mass number – no. of protons = 52 – 24 = 28 neutrons

Therefore, 23Cr has 24 protons, 24 electrons and 28 neutrons.

Identifying an Element or Isotope from Subatomic Particles

Using the periodic table, an element or an isotope of an element can be identified from given subatomic particles.

Mass Number is not the same as Atomic Mass.

The mass number refers to the number of protons plus neutrons in one specific isotope of an element. The atomic mass refers to the weighted average of the masses of all naturally occurring isotopes of an element. Mass number will always be expressed as a whole number. Atomic mass will generally be expressed as a decimal number.

Chapter 5: Molecules and Compounds

In this chapter we will learn how compounds are formed, the types of chemical bonds in the compounds, how to write correct formulas and name the two types of compounds.

Molecules and Compounds

A molecule is a particle of matter in which there are two or more atoms combined together chemically.

In Chapter 3, we learned that pure substances could be classified either as elements or compounds. Elements may exist as individual atoms, or as molecules, depending on the element. Elements that exist as molecules have more than one atom of the same type chemically joined together.

A **compound** is made up of **two or more elements (two or more types of atoms)** which have been chemically combined and therefore exists as molecules. Examples of compounds are water, H2O; sulfuric acid, H2SO4; carbon monoxide, CO. **Compounds generally have completely different properties than the elements from which they are formed.**  For example, table salt, NaCl (sodium chloride), is commonly used to make our food taste better, but both sodium metal and chlorine gas can be quite harmful as individual elements.

Law of Constant Composition

In a compound, there are two or more different types of atoms present. However, it is important to realize that a compound has a **fixed composition,** whereas a mixture has variable composition. The Law of Constant Composition (sometimes called the Law of Definite Proportion) states that a compound will always be made up of the same elements combined in the same ratios by mass. For example, water will always have eight parts of oxygen for every part of hydrogen by mass.

Chemical Formulas

Chemical formulas provide us with a shorthand way of describing the makeup of a compound by listing the type of atoms present as well as the number of atoms in the smallest unit of the compound.

In a chemical formula, the atom types are represented by the element symbol from the periodic table. If more than one atom of a particular type is present, a subscript numeral to the right of the element symbol specifies how many are present. For example, the formula of the compound ammonia is NH3; one unit of ammonia contains one nitrogen atom and three hydrogen atoms. If there is only one atom of a particular type, the subscript “1” is not used. Use parentheses around a repeating group of atoms in a formula.

Elements can be atomic or molecular

As we stated previously, elements can exist as individual atoms (**atomic elements**) or as molecules (**molecular elements**.)

Molecular Elements: Rule of Sevens

There are seven common **molecular elements** that occur naturally as 2-atom molecules: H2, N2, O2, F2, Cl2, Br2, and I2. If you locate these elements on the periodic table, you will notice that beginning at nitrogen, six of these elements form a figure “7”. You must remember to add hydrogen into this group to account for all seven elements. This is called the rule of 7’s.

A two-atom molecule is called a **diatomic** molecule. These seven **molecular elements** are diatomic. You can also have **molecular compounds** that are diatomic .

Compounds can be molecular or ionic

Compounds can be classified as **molecular compounds** or as **ionic compounds**.

Ionic compounds are made up of **ions** (pronounced “eye′-ons”.) Ions are atoms or molecules that have an overall charge: positively charged ions are called **cations** (pronounced “cat′-eye-ons”) and negatively charged ions are called **anions** (pronounced “an′-eye-ons”). A molecule that carries an overall positive or negative charge is called a **polyatomic ion**. Ionic compounds can be recognized as compounds made up of a metal bonded to a nonmetal or as compounds containing polyatomic ions.

**Molecular compounds** are those containing only nonmetals bonded together and in which there are no polyatomic ions present.   
CaCl2 would be an ionic compound because it contains a metal bonded to a nonmetal.

NH4NO3 would be an ionic compound because it contains at least one polyatomic ion. (More about polyatomic ions later…)

NH3 and H2O would be molecular compounds because they contain only nonmetals bonded to each other and there are no polyatomic ions.

The smallest unit of a molecular compound is a molecule. In an ionic compound, there are no molecule units; rather, there is an array or latticework of repeating units of cations and anions, which we call a “**formula unit**.”

Writing Formulas of Binary Ionic Compounds

**Binary** **ionic compounds**, as the name implies, are composed of two elements. How do we write the correct formula for the ionic compounds formed between a metal and a non-metal? There are 3 steps to follow: you will remember these by practicing how to write formulas.

1. Write the symbol for the cation first, followed by the symbol for the anion. For the ionic compound formed between magnesium and chlorine, write Mg+2 Cl-1.
2. Cross over the charges so that the charge on the cation becomes the subscript for the anion and the charge for the anion becomes the subscript for the cation. Therefore in the magnesium chloride example, you would write Mg -1 Cl+2.Drop the positive and negative signs. For magnesium chloride, the formula would therefore be MgCl2. (If the subscript is 1, it is not written.)
3. Reduce and simplify so that subscripts are the lowest whole number ratios. For example, in the ionic compound formed between calcium and sulfur, Ca+2 and S-2 → Ca2S2 → CaS.
4. Although an ionic compound is composed of charged particles, they must combine in such a way that the compound has an overall charge of zero so that the compound is neutral.

Predicting Charges of Ions

The above instructions assume that you will know what the charge an ion will have. How will you know?

**If an atom loses one or more electrons,** there will be more protons than electrons in the atom; therefore, it will become positively charged, a cation. The positive (+) charge will be equal to the number of electrons lost.

**If an atom gains** **one or more electrons,** there will be more electrons than protons; the atom becomes negatively charged, an anion. The negative (–) charge will be equal to the number of electrons gained.

**Once more to help you remember, a atom loses electron(s) to form a cation; an atom gains electron(s) to form an anion.**

How do we know how many electrons an atom will lose or gain? For many elements, is relatively easy: look at its position in the periodic table. Main group elements (group 1A-7A) form only one charge, and the group number is the number of electrons in the outermost shell of the atom. The atoms will take the easiest route to having eight electrons in their outer shell; they want to lose or gain electrons to have the more stable electron configuration of a noble gas, an octet. An atom with less than four electrons in its outermost shell will lose electrons and an atom with more than four electrons in its outermost shell will gain electrons. Atoms with four outer electrons prefer to share rather than gaining or losing.

For main group (“A” group) elements, the group number is equal to the number of electrons in the outermost shell. Group 1A elements will lose their 1 valence electron and will always have +1 charge, group 2A will lose 2 electrons and form ions with +2 charge and group 3A will lose 3 electrons and form ions with +3 charge.

Group 4A elements do not have a strong desire to lose or gain electrons but prefer to share, so they don’t form stable monatomic ions.

Groups 5A-7A need to gain electrons to have the stable configuration of an octet. Group 5A needs 3 electrons, so will gain 3 electrons and form –3, Group 6A will form –2 and Group 7A will form –1 ions.

Will a group 1A element form +2 ion? No!! It will not be stable. Will a group 2 element form a +1 ion? No!! It will not form a stable ion.

Naming Ionic Compounds

Ionic compounds are composed of a cation and an anion. To name ionic compounds from their formulas, simply name the cation first, followed by the anion. Here are the rules for naming cations and anions.

If the **cation is a Type I or main group (“A” group) metal**, it will only have one possible charge, which is understood. There is no need to specify it in naming. So Mg2+  is simply magnesium ion.

If the **cation is a Type II metal**, (these are transition metals that can have more than one possible charge), the name of the cation is the name of the metal followed by its charge, expressed as a Roman numeral in parentheses. Thus in naming Fe3+ will be called iron(III) and Fe2+ will be called iron(II). You must determine the charge on the cation by looking at the anion it is paired with. In an ionic compound, the charges must add up to give zero so the compound will be neutral- it will have no overall charge.

If the **cation is a polyatomic ion**, you must know its name. The only polyatomic cation you need to be concerned about is the ammonium ion which is NH4+.

If the **anion is a nonmetal**, it is named by dropping the ending of the element name and adding “–ide” in its place. Thus, the anion formed by the element chlorine is called “chloride,” and the anion formed by oxygen is “oxide.”

If the **anion is a polyatomic ion**, you must know its name.

Examples of Formulas and Names of Binary Ionic compounds

The powerpoint slides include some slides to help you practice how to write the correct formulas and names for a few binary ionic compounds. Test yourself by choosing elements from groups 1 –3 A and form compounds with elements from groups 5-7. Practice writing their formulas and names.

Polyatomic Ions.

**Polyatomic ions** are ions made up of two or more elements. These ions are formed when **molecules** lose or gain electrons. Examples of polyatomic ions, their formulas, charges and electron-dot structures are illustrated in the slides. You should be able to recognize and name the following polyatomic ions: ammonium (the only polyatomic cation of importance in this course), acetate, carbonate, hydrogen carbonate (also known as bicarbonate), hydroxide, nitrate, nitrite, sulfate, and chlorate.

Writing formulas and Naming Ionic compounds with Polyatomic Ions

To write the formula of ionic compounds with polyatomic ions, follow the same procedure as previously described for monatomic ions. Write the cation and its charge first, then the anion and its charge then crossover the charges. If a subscript is written for the polyatomic ion, it must be taken as a unit and enclosed in parentheses, then the subscript is written. The subscript multiplies the subscript on each atom in the polyatomic ion. Examples are shown for compounds formed between Na and SO42– → Na2SO4, Fe3+ and NO3–→ Fe(NO3)3.

Naming Binary Molecular Compounds

Remember that molecular compounds are composed only of nonmetal elements; it is important to remember that molecular compounds will not contain polyatomic ions such as ammonium. If ammonium is part of the element name (or NH4+ is in the compound’s formula) it will be an ionic compound, even if there are no metals in the compound.

To name molecular compounds there are 3 simple rules to follow:

1. Name the first nonmetal in the formula, using the full name of the element.
2. Name the second nonmetallic element but change the ending of its name to –***ide***.
3. Use a prefix in front of each name to tell how many atoms of each element are in the compound as indicated by the subscripts in the formula. Use ***mono***- = 1; ***di***- = 2; ***tri***- = 3; ***tetra***- = 4, etc. (Exception: Never use “mono–“ in front of the first element name.) The prefixes for numbers one through 8 should be memorized.

Example: BF3 would be named boron trifluoride; P2O5 would be diphosphorus pentoxide.

Determining the Formula Mass

In the previous chapters, we designated the average mass of an atom in an element as the atomic mass in atomic mass units (amu). When we talk about compounds, we define the formula mass as being the sum of the atomic masses of each atom in a formula unit for an ionic compound or in the molecular formula for a molecular compound.

Chapter 6: Chemical Composition

Chemical Composition- Why is it important?

In the previous chapters we learned about chemical bonds, the two types of compounds (ionic and covalent or molecular) and how to write their formulas and name them. The chemical characteristics of matter vary according to what makes up the matter. Any business that produces any type of product realizes that it is important to know how to combine the raw materials efficiently to make the product with as little waste and as cost effectively as possible in order to make a profit. Chemists are very concerned with that as well; it is important to produce the product we are seeking as efficiently as possible with as little waste as possible. To do this, we must know the chemical composition of the starting materials and the desired products, and we investigate the best conditions for running a reaction.

We know that atoms are so tiny they can’t be seen or weighed individually in a most labs, much less on your pocket scale at home. In the next three chapters, we discover ways to discuss atoms and molecules and their reactions in a quantitative manner in spite of the limitations of common equipment. These three chapters are the real essence of what chemists do.

Study Tips

Although the three chapters are so closely related that we present them together, it is important in studying them that you take the material in small bites. We will cover the three chapters over the next two weeks. Pay attention to the syllabus in determining which sections we expect you to read. Read one or two sections and work through the examples and skillbuilder problems. Then try an end-of-chapter practice problem. Once you have that topic under control, move to the next. This way of digesting the material in small increments is much more effective than the all-night-Sunday-night method and is really important in learning the material presented here. Make good use of the discussion board, posting questions as they occur to you. Remember that your instructors want to help you understand; we remind you that the only “dumb question” is the one you don’t ask!

Cheaper by the Mole?

**Pair, dozen, gross, and ream** are all nouns that commonly represent a certain quantity of something. Socks, shoes, and mittens commonly are purchased in pairs; eggs and doughnuts are counted by the dozen. Schools buy pencils in boxes containing 144 pencils; this quantity is called a gross. A package of paper containing 500 sheets of paper is a ream. When the print shop at TCC orders paper for all the copiers and computer printers for all the campuses, it would make no sense to order a certain number of individual sheets. In fact, paper for TCC is ordered by cases, knowing that a case contains a certain number of reams, and each ream has 500 sheets.

Because atoms and molecules are so tiny, it makes sense to measure them collectively as well. A very small sample of an element or a compound contains a very large number of particles- either atoms or molecules. Chemists define a quantity called a “**mole**”; a mole of atoms contains 6.022 x 1023 atoms. This very large number is of a convenient size for measuring quantities of atoms and molecules, as we shall see.

The original definition of a **mole** is the **number of atoms in exactly 12 grams of carbon-12**, the most abundant naturally-occurring isotope of carbon. You might recall that by definition one atom of carbon-12 weighs exactly 12 atomic mass units (amu.) The definition of a mole is quite convenient then, since one atom of carbon-12 weighs 12 amu and one mole of carbon-12 atoms weighs exactly 12 g.

The number of atoms in exactly 12 grams of carbon- 12 is **6.022 x 1023 atoms**. We call this number **Avogadro’s number**.

Remember, however, that naturally occurring carbon contains a mixture of isotopes. The average mass of a carbon atom in a sample of naturally-occurring carbon is 12.01 amu. Therefore, a sample of naturally occurring carbon weighs 12.01 grams because it will contain a mixture of carbon isotopes.

Counting Atoms by the Mole is like Counting Nails by the Pound

The analogy of counting nails by the pound (or any other small piece of hardware for that matter) is a pretty good one. Would you like a job where all you did all day was to fill small plastic bags with precisely 8 screws, each ¾ long size #10? Probably not. Nor is it a job for which a manufacturer would want to pay you because it is one that can be done more cheaply and precisely by machine. By knowing the average mass of one such screw, a machine could be set to weigh out the desired number of screws and bag them.

In weighing out atoms or molecules, we can do a similar thing if we know the average mass of one atom or one molecule. As mentioned above, the average mass of one naturally occurring carbon atom is 12.01 amu and by definition, the average mass of one mole of carbon atoms (Avogadro’s number) is 12.01 grams. If we need 2 moles of carbon atoms, we can measure out 24.02 grams of carbon, or if we need ½ mole of carbon atoms, we can measure out just over 6 grams of carbon.

If a different element is needed, a similar operation will work, but it is necessary to know the average mass of one atom of that element. ( Similarly, if we want to switch the machine to weigh out 1 inch screws instead of ¾ inch screws, we need to know the average mass of the longer screws.) Luckily, in the world of elements, this information is readily available to us in the periodic table. For each element in the periodic table, the **atomic mass, the average mass of one atom of that element in amu**, is given and it is numerically equal to the average mass of one mole of atoms of that element in grams. We define the **molar mass as the mass of one mole of an element in grams**.

One mole of any substance will contain 6.022 x 1023 particles of that substance; the particles may be atoms, molecules, or any other type of particle, including #8 screws, basketballs, or people.

1 mole of any substance = 6.022 x 1023 particles of that substance

Each type of substance will have its own molar mass, which will depend on what the particles are.

For elements, the molar mass is numerically equal to the atomic mass given in the periodic table.

Use the **molar mass as the conversion factor to convert grams to moles or moles to grams**, as shown in the examples in the slide show. These are unit conversion or dimensional analysis problems. Learn to do them by letting the units guide you, rather than by memorizing whether you to need to multiply or divide to do a particular operation.

Molar Mass of Compounds

Just as the molar mass of an element is the mass of 6.022 x 1023 atoms of that element, the **molar mass** of a compound can be defined as 6.022 x 1023 **molecules** of that compound (if it is a molecular or covalent compound) or 6.022 x 1023 **formula units** of an ionic compound.

When determining the molar masses of compounds we use the formula weight for ionic compounds and molecular weight for covalent compounds. The **formula weight** of an ionic compound is the sum of the atomic weights of the elements present multiplied by the number of atoms of each element in the formula. The **molecular weight** of a covalent compound is equal to the sum of the atomic weights of the elements present multiplied by the number of atoms of each element in the molecular formula. **The mass of 1 mole of ionic compound is equal to its formula weight in grams and the mass of 1 mole of a covalent compound is equal to its molecular weight in grams.**

Thus, to find the molar mass of an ionic compound such as NaCl, you would look on the periodic table to find the atomic mass of sodium, Na, which is 22.99 and the atomic mass of Cl, which is 35.45. Add the two to get 22.99 + 35.45 = 58.44. The formula weight of NaCl would therefore be 58.44 grams/mole.

To find the molar mass of a covalent compound such as ammonia, NH3, you would find the masses for nitrogen, N, which is 14.01 plus 3 times the mass for hydrogen, H, which is 1.01 for a total of 17.04 g/moles. This is the molar mass or molecular weight of ammonia.

Just as with elements, the molar mass can be used to convert grams of a compound to moles of a compound or moles of a compound to grams of a compound. The importance of being able to do this will become more clear in the next chapter.

Percent Composition

The percent composition is the percentage of each element present in a compound by mass. To find the percent composition, you calculate how much of the molar mass is supplied by a particular element, divided by the molar mass, then multiplied by 100 to convert to a percentage.   
  
Here are two examples of this type of calculation using the examples above for calculating molar masses.   
  
The percent composition of NaCl would be as follows:



Notice that although the percentages should add up to 100%, these only add up to 99.99%, probably because of rounding.

The percent composition of NH3 would be as follows:



Chapter 7: Chemical Reactions

Evidence of Chemical Reactions

If you observe everything around you carefully, you should be able to find evidence that chemical reactions are occurring all around you and even in you constantly. In an earlier chapter, we described changes as being either physical changes or chemical changes. How can we tell the difference?

In a chemical reaction, atoms of the materials at the start of the reaction are rearranged to form new materials with different characteristics. No new matter is produced, but atoms are rearranged. By this definition, a change such as melting ice to form liquid water and then boiling the liquid water to form steam, would not be a chemical change, only a physical one. Why? Because it is easy enough to condense the steam back to liquid water and then cool it back to form ice. The water has changed its form, but not its chemical makeup. Whether it is present as ice, liquid water, or steam, all forms of water still have the chemical formula H2O.

On the other hand, if we burn wood in a fireplace, the main products are carbon dioxide, water, and some noncombustible wood ash. Once the reaction is complete, there is no way to easily bring the wood back. Its atoms have been rearranged to form new compounds.

Careful observation of changes in a process is necessary as we try to sort out evidence of a chemical reaction. Any of the five senses may provide clues to a chemical change: color change, formation of a solid precipitate, bubble formation, emission of light, cold or heat produced in a reaction, new odor produced, or sounds heard such as fizzing may all signify a chemical change. Simple observation is not enough, however. The only way to know for sure that a chemical change has occurred is by analyzing the starting materials and products from the reaction to see if the molecules themselves have been changed.

Chemical Equations

When chemical reactions occur, bonds are broken and new bonds are formed, resulting in rearrangement of atoms and production of new compounds. **Chemical equations** are symbolic descriptions of chemical reactions using the chemical formulas for the **reactants** (starting materials) and **products** of the reaction. From a chemical equation, we can determine the physical states (solid, liquid, gas, or aqueous solution) for the reactants and products. A balanced chemical equation can help us predict the relative numbers of molecules required for the reaction. The arrow in a chemical equation is somewhat similar to the equal sign in a math equation and may be read as “yields” or “react to form”.

For example let’s consider the combustion reaction between gasoline (octane, C8H18) and sufficient oxygen to produce carbon dioxide and water. The chemical equation is written as:

**2**C8H18(l) +  **25**O2(g)  → **16**CO2(g) + **18**H2O(l)

**Two possible ways to read the equation are to say:**2 molecules of liquid octane plus 25 molecules of oxygen gas will yield 16 molecules of carbon dioxide gas plus 18 molecules of liquid water.

or

2 moles of liquid octane plus 25 moles of oxygen gas will react to form 16 moles of carbon dioxide gas plus 18 moles of liquid water.

The starting materials are called **reactants** and written on the left and the and new compounds formed are called **products**, written on the right of the arrow pointing toward the right. Sometimes the physical states of the reactants and products are written using the first letter of the state: (l) for liquid; (s) for solid and (g) for gas. The abbreviation (aq) is used for solutions in which water is the solvent.

Balancing a Chemical Reaction

The **law of conservation of mass** was introduced in previous chapters. It says that matter cannot be created or destroyed in a chemical reaction. Atoms are simply rearranged. This means that the total mass of the reactants will equal the total mass of the products since all the atoms you start with will still be present when the reaction is complete. The law of conservation of mass is the basis for balancing a chemical reaction.

Once a chemical equation is balanced, the numbers of each type of atom will be the same on the left side and on the right side of the reaction. In the balanced equation given above for the combustion of octane, there are bold-faced numbers in front of each of the formulas of the reactants or products; those numbers are called **coefficients**. They represent the number of molecules or the number of moles of that particular substance, depending on how the equation is read. The **coefficients** are the only numbers used for balancing the chemical reaction; **never change the subscripts of any formula**. To do so would alter the identity of the compound.

Some reactions are simple enough to balance by inspection, but sometimes it is necessary to follow a procedure to make balancing easier. The steps are:

1. Write the unbalanced chemical equation making sure that the formulas are correct for each reactant and each product.
2. Inventory the atoms present in the starting materials and in the products. A simple way to do this is to make a list of the atoms under the reactants and make the same list under the products. To count the atoms present correctly, you must multiply the coefficient of the formula in which the atom appears by the subscript that tells how many of that atom is in that formula.
3. If the same polyatomic ion appears on both sides of the equation, you may count it as a unit, rather than counting the individual atoms in the ion.
4. Begin balancing by looking for an element that is present in only one formula. Change the coefficients of the compounds as needed to make that element balance. Stop and inventory the atoms on both sides again after each time you change the coefficients, since changing a coefficient may change the numbers for more than one element.
5. If there are elements present that are not combined in a compound, save balancing those for last.
6. Double check to be sure the numbers of atoms are the same on both sides of the equation, and then reduce coefficients to the lowest whole number ratios.

An example of this procedure is included as a separate handout in the Course Documents folder for these three chapters.

Chapter 8: Quantities in Chemical Reactions Stoichiometry

**Stoichiometry (stoy-key-om′-et-tree)**  is a big word that describes the relationships between quantities of reactants and products in a chemical reaction. It relies on our ability to apply the law of conservation of mass and procedures for balancing chemical equations. Stoichiometry allows us to calculate how much of each reactant to measure out at the beginning of a reaction and figure out how much product should be expected. Although it sounds formidable, stoichiometry is quite similar to using a recipe in the kitchen to make a favorite dish.

The lecture slides make a good analogy to making pancakes using 2 eggs, 1 cup of flour, and ½ teaspoon of baking powder to make five pancakes. If you are making pancakes for a crowd, you can figure out how much of each ingredient to use if you know how many pancakes you need to end up with and you don’t run out of anything. If you have a dozen eggs, you can make 30 pancakes, or if you only have half- a dozen eggs, you can make 15 pancakes.

A chemical equation can be read as if it was a recipe also. In chapter 7 above, there were two possible ways to read the chemical reaction. The formulas represent molecules of a substance. Because moles are a way of counting molecules, the formulas can also represent moles of a substance. Both ways are valid ways to read the chemical equation. However, when doing a reaction in the lab, we already learned that chemists use moles rather than molecules in order to have quantities that are measurable with common lab equipment.

Mole Ratios and Mole to Mole Conversions

The coefficients in the balanced equation give the mole ratios of the reactants and products. In the reaction: 2H2 + O2 → 2H2O, we can say that 2 moles of H2 is ≡ 1 mole O2 ≡ 2 mole H2O. The symbol “≡” means “stoichiometrically equivalent to”. The coefficients can be used as conversion factors for calculating the amounts of other two if amount of one is known.

**Example: 2H2 + O2 → 2H2O**

a) If you have 1 mole H2 how much oxygen can you react it with?

1 mol H2 x = 0.5 mole O2

b) how much water can be made?

1 mol H2 x= 1 mol H2O.

It is critically important that the ratios you use are always ratios of moles from the balanced chemical equation (never ratios of masses!)

Mass to Mass Conversions

Although we must use mole ratios for calculations in stoichiometry equations, we realize that there is no direct way to measure out a mole of substance. Like the analogy to nails at the hardware store, we usually rely on the measuring the mass of a substance to know how much of it we have. However, we cannot use the coefficients in the chemical equation to convert grams of one reactant to grams of another reactant or grams of a reactant to grams of product. **We must always(!!!) convert grams of a substance to moles before using any ratios from a balanced chemical equation.**

Example: 2H2 + O2 → 2H2O. Calculate how much oxygen you need to react with 32.3 grams of hydrogen to produce water.

1)The first step is converting the mass of hydrogen to moles of hydrogen.



2) The next step is calculating how many moles of oxygen are needed:



3) Moles of oxygen can then be converted back into grams of oxygen:



**Always(!!!) convert grams of a substance to moles before using any ratios from a balanced chemical equation. Use the mole ratios to calculate moles of the second substance, then convert back to grams of the second reactant.**

Limiting Reactants and Excess Reactants

At some time, you may have experienced this concept in your kitchen. You need to make a double recipe or triple recipe of some treat, but you are not sure you have the right amounts of ingredients. For example, in the pancake example, it takes 2 eggs to make 5 pancakes, so with 6 eggs you can make 15 pancakes. Making 15 pancakes also requires 3 cups of flour and 1.5 teaspoons of baking powder. It is important to be able to determine which ingredient will be the one that you run out of first. For example, if you have only 4 eggs, but you have 20 cups of flour and several tablespoons of baking powder, you will be limited by the number of eggs.

A similar situation may arise when figuring out ingredients for a reaction in the lab. If your reactants are not present in the exact stoichiometric ratios, the amount of product formed will be limited by the reagent that is completely used up.

For the reaction between Mg + S → MgS, the ratios say that 1 mole Mg reacts with 1 mole S producing 1 mole MgS.

If both reagents are present in the exact proportions as shown by the coefficient in the balanced equation, then you will get the amount of product predicted by the equation.

However, if they are not, as in the case when you have 1 mol Mg + 2 mol S, your product will be 1 mol MgS. Why? Because once all the Mg has been used up, the reaction stops. Even though you still have 1 mole of S, there is nothing to react with, so it is left over as S. In this case, Mg is the limiting reagent and S is the reagent in excess.

If you reversed the amounts, so that you react 2 mols Mg and 1 mol S, the product will still be 1 mol MgS, because this time you used up all the S and 1 mol excess Mg has nothing to react with.

Chapter 09: Electrons in Atoms and the Periodic Table

We learned in Chapter 4 that Rutherford’s gold foil experiment indicated that the atom had a small dense nucleus that was positively charged and that the electrons must therefore be in the empty space between the nuclei. As time went on, evidence mounted that the electrons were not randomly distributed in the empty space around the nuclei, but were arranged in a systematic fashion. In this chapter, we will explore models of the atom that explain how electrons are arranged in an atom and how those electron arrangements relate to how different elements react with each other.

Electromagnetic Radiation: Wave properties of light

**Electromagnetic radiation** is one form of energy with which you are familiar, although you may not know it by that name. **Visible light** is one form of electromagnetic radiation; others include **gamma rays, x-rays, ultraviolet radiation, infrared radiation, microwaves, and TV and radio waves.** Visible light is a very small portion of the whole range of electromagnetic radiation, which we call the electromagnetic spectrum.

Before the development of quantum mechanics, light was described in terms of its wave properties.

Characteristics of a Wave

Waves can be characterized by their **wavelength, amplitude, frequency and speed**.

**Wavelength** is the distance between successive peaks in a wave. For visible light, the color of the light depends on the wavelength. Violet blue light has a short wavelength, while red light has longer wavelengths.

**Frequency** indicates how many wave peaks pass by a given point in a given amount of time. The speed of the wave will be affected by the medium through which it is passing. In a vacuum such as space, light travels at a constant speed of 3.0 x 108 m/s.

**Wavelength and frequency are inversely related to each other.** The longer the wavelength of a particular wave, the smaller is its frequency and also, the lower is the energy of the wave. Shorter wavelengths and higher frequencies mean that a wave is of higher energy.

Photons: Particles of Light

In the early 20th century, the work of Albert Einstein and Max Planck led to an expansion of our understanding of the nature of light as having a particle nature. They described light as a stream of particles called **photons**. The **energy of a photon** was related to the wavelength of the light. The shorter the wavelength, the more energy a photon has. Thus a blue photon (shorter wavelength) has higher energy than a red photon (longer wavelength.)

Emission Spectra

Atoms can absorb energy in the form of heat, electricity, or light. When an atom gives off that energy again, it does so in the form of light. We see this when we look at a light bulb. In an incandescent light bulb, electricity makes the filament glow; the atoms that make up the filament are absorbing the electrical energy and emitting it again in the form of light. If the white light from a regular incandescent bulb is passed through a prism, it is separated into a rainbow of colors. This rainbow represents an **continuous spectrum.** As we saw above, the colors represents a different wavelength of light, but in a continuous spectrum, all the wavelengths are present.

Now consider a different kind of light source: an advertising light, commonly referred to as a “neon” light. This type of light doesn’t have a filament. Instead, a glass tube is filled with a gas such as neon, and a voltage is applied across the tube, causing the gas molecules to glow. However, in a real neon light, the light is not white, it looks red to the naked eye. (If the tube is filled with a gas other than neon, a different color is produced; for example, argon will produce a purple glow.) If you tried to separate the light from a neon light into a rainbow, you would find that there would not be a continuous spectrum, but instead there would be discrete lines of different colors at specific wavelengths with dark areas in between them. This is called an **emission spectrum** or a **line spectrum.** What would explain the difference in the types of spectra produced?

Electron Arrangements

The physicist, Niels Bohr, proposed the first model for how electrons are arranged around the nucleus of an atom. He proposed that electrons cannot just randomly move anywhere around the nucleus. Instead electrons occupy specified **energy levels** at certain distances or orbits (like the planets moving around the sun) from the nucleus. The lowest energy level is closest to the nucleus and the energy level increases as it moves farther from the nucleus. Electrons are limited to these energy levels and cannot exist in between specified energy levels. Think of these energy levels like a staircase where you can stand on the 1st, 2nd, 3rd, etc. steps but not on the spaces between them. This model became known as Bohr’s planetary model for electron arrangement. In addition, Bohr proposed that electrons can jump from one energy level to another by absorbing or giving off energy.

Ground and Excited States

Atoms normally have their electrons in the lowest energy levels possible. This lowest energy state is called the **ground state** and is the most stable form of the atom. When atoms absorb energy electrons jump to a higher energy level and are said to be in an **excited state**; they give off the same amount of energy when they return to the original energy level. Atoms prefer to be in the ground state. In atoms with many electrons, it becomes a problem because they can’t all go to the lowest energy levels.

Electron Energy Levels

Bohr presented his model for the hydrogen atom where its single electron occupies the first energy level. Further work by Bohr and other scientists showed that each energy level has a maximum number of electrons that it can hold. This number can be calculated from the formula 2*n*2, where *n* is the energy level 1, 2, 3, etc.

Building Atoms: Lowest Energy Levels Fill First

With the limitations imposed by the number of electrons that each energy level can hold, electrons follow an order of filling. The lowest energy levels are filled first. As an example, let us look at the distribution of electrons around the nucleus of the element carbon, C. The atomic number of C is 6, therefore it has 6 protons and 6 electrons. The first 2 electrons go to the first energy level. Since the 1st energy level can hold a maximum of 2 electrons, the next 4 go to the 2nd energy level. Similarly, in magnesium, Mg, with 12 electrons, 2 electrons go to the 1st energy level, 8 electrons can go in the 2nd energy level, which can hold a maximum of 8. The next 2 electrons then go to the 3rd energy level.

Orbitals

Bohr’s model worked well for explaining the spectrum of hydrogen. Unfortunately, it did not work well for predicting the spectra of elements having many electrons. Bohr’s model of electron arrangement was later refined as more data accumulated to show that there are sublevels or subshells called orbitals within the principal energy levels. The quantum mechanical model of the atom describes these orbitals.

These **orbitals** describe the 3-dimensional shape of the space around the nucleus where there is a high probability of finding electrons. There are 4 types of orbitals named ***s, p, d, f*** , each with a characteristic shape and energy associated with it.

Orbitals are not Orbits

It is very important to realize that the word “orbital” has to do with describing where you might find a particular electron within an atom. It does **not** specifically describe the movement of the atom in the way we think of planets having “orbits” around the sun. Orbitals are probability maps that describe where the electron might be found.

As an analogy, consider an elementary school at 10 am. Most of the students will be in their classroom working on their lessons. However, the 2nd and 3rd graders have morning recess at that time. If someone gave you a floor plan for the school and asked you to map out where the students are at 10 am, you could reasonably put 20 or 25 dots in all the classrooms, except the 2nd and 3rd grade classrooms. You would leave those empty because all those students are outside. To map those students, you would probably cluster most of the dots around the various pieces of playground equipment with a few dots here and there in other parts of the playground. You wouldn’t put any dots in the parking lot, because students are not allowed to play there. You can’t reasonably map exactly where each student is at any point in time…for one thing, they keep moving. They also do not move in specified pathways but move randomly around within certain constraints imposed by the teachers monitoring their activities. You know that the students who are in class have a high probability of being within the space described by the walls of their classrooms. There are several places you are likely to find students outside at recess but they shouldn’t be found in the parking lot.

Orbitals are similar to the map you would draw of the school. Orbitals have broader areas where electrons are most likely to be found. There are also places called nodes, where the electrons shouldn’t be at all.

Orbital Energies

Electrons are placed in orbitals according to the energy levels, sublevels, and orbitals from lowest energy to higher energy. The energy of the primary energy levels increases as you move out from the nucleus. The energy associated with the subshells and orbitals increases in order from ***s < p < d < f.*** As the energy level increases, the number of orbitals also increases.

Some orbitals have energy that overlap into the next higher energy level. For example, ***d*** orbitals in the 3rd energy level have higher energies than the ***s*** orbital of the 4th energy level. In each energy level, d-orbitals have higher energies than the s orbital of the next higher energy level.

Order of Filling

A chart is presented that will help you with the order of filling the orbitals. A similar chart is presented in the textbook at the top of page 279. Begin at the top of the chart, and follow the arrows down and to the left, looping back each time the s orbitals have been filled. Note that the chart leads you to automatically fill the 4 s orbitals after the 3p orbitals, and then to follow through with the higher energy 3 d orbitals.

Electron Configurations

The electron configuration of an element is the distribution of its electrons into the orbitals around the nucleus. How do we write electronic configurations? Distribute the electrons into the orbitals, specifying the principal energy levels in which the orbitals are found, the specific orbitals and the number of electrons in each orbital, keeping in mind the following rules: Each individual orbital can only have 2 electrons. An ***s*** orbital can only contain 2 electrons; a set of 3 ***p*** orbitals can contain a total of 6 electrons (2 per orbital x 3 ***p*** orbitals per sublevel) ; a set of ***d*** orbitals can have a total of 10 electrons (2 electrons per orbital x 5 orbitals per energy sublevel); a set of ***f*** orbitals can have a total of 14 electrons (2 electrons per orbital x 7 orbitals per energy sublevel).

When filling the orbitals within a particular subshell, keep in mind that each orbital should be occupied by at least one electron before pairing the electrons in any of the orbitals.

In reading or writing an electron configuration, large Arabic numerals represent the primary energy level, the small letters represent orbitals, and superscript numbers indicate how many electrons are in a particular sublevel.

**Shorthand methods** are used to write long electron configurations. As you move through the periodic table, each element has the same general electron configuration as the element before it, with the addition of one electron. In the shorthand method, the noble gases are recognized as having their energy levels completed filled, so these are used as landmarks. In the shorthand method, the core electron configuration of the noble gas in the previous period of the periodic table is indicated by using the element symbol of that noble gas in square brackets, followed by the additional electrons to complete the electron configuration.

For example, the electron configuration of phosphorus would be 1s2 2s2 2p6 3s2 3p3. Since the electron configuration of neon is 1s2 2s2 2p6, we can write the electron configuration of phosphorus as [Ne] 3s2 3p3.

Importance of Electron Configuration

Knowing the electron configuration of an element is very important. The physical and chemical (how they react) properties of an element are determined by its electronic configuration.

Valence Electrons

When elements react to form compounds, the electrons in the outermost **principal** energy level are the ones involved. These electrons are called **valence electrons**. The number of valence electrons of an element can be determined from its electronic configuration. For example. The electronic configuration of carbon, C is   
1***s*** 2 2***s*** 2 2***p*** 2 - 2 electrons in the 1st energy level and 4 electrons in the 2nd energy level, the outermost energy level. Therefore, the number of valence electrons of C is 4 or we say the valence of C is 4. In Chlorine, Cl the 3rd energy level is the outermost level with 7 valence electrons. The electrons in the inner energy levels are called **core electrons** and do not participate in chemical reactions.

The Periodic Table and Electronic Configuration

The importance of electronic configurations is illustrated by the fact that the arrangement of elements in the periodic table is in fact, based on electronic configuration. In the periodic table all known elements are arranged in increasing atomic number divided into horizontal rows and vertical columns. The horizontal rows are called **periods** and correspond to the energy levels. The vertical columns are called **groups,** which contain elements having similar properties and the same number of valence electrons. Referring to the Periodic Table in the inside cover of your book, the columns numbered 1A – 8A are called the main groups and the column number is in fact, the number of valence electrons of each element in the group. For example, all elements in group 1A have 1 valence electron, group 2A have 2, etc.

The Periodic Table, Orbital Filling and Valence Electrons

This slide shows the order of filling of orbitals with electrons as they are distributed around the nucleus of the atom of the element. We can see that s orbitals in the highest or outermost energy level are getting filled in groups 1A and 2A, p-orbitals in groups 3A-8A. The elements in the middle of the periodic table are called transition elements numbered Groups 1b –12B, whose higher energy d-orbitals are getting filled.

Determining Highest energy levels and Valence Electrons

Sodium, Na is in row or period 3; so highest energy level is 3. One electron is in s orbital, so valence electron is 1 (group 1a). In phosphorus, P, which still in row 3, so highest energy level is 3, but now we have the s orbitals completely filled (2 e’s) and 3 electrons in p orbitals, so valence electrons is 5 (group 5A). Do the same for Cl and Pb.

Chapter 10: Chemical Bonding Chemical Bonds

In the previous chapter, we learned to identify two main classes of compounds: ionic and molecular, and we learned to write formulas for them and name them. In this chapter, we will explore a model of chemical bonding.

**Chemical bonds** are the forces that hold atoms together in a molecule and ions in ionic compounds. The tremendous variety of substances that we have is due to the ability of atoms to bond or combine with other atoms. The type of bonds they form determines many properties of substances. For example, properties such as melting point and boiling point, solubility, hardness, shape of molecule, types of reactions etc. depend on the type of bonds present in the compounds.

Bonding Theories

Bonding theories are ways chemists look at the interaction between different atoms in a compound in order to explain the properties of the compound. By understanding how bonding occurs in compounds, chemists are in a better position to be able to design new compounds in which they can maximize desirable properties and minimize undesirable properties.

Lewis Electron Dot Structures

Since valence electrons, which are the outermost electrons, are the ones involved in chemical reactions, G. N. Lewis, an American chemist invented the electron-dot structure to help us visualize how they participate in chemical bonding. **Electron-dot structures** for an individual element represent valence electrons as dots around the symbol for the element. To write the electron-dot structures, distribute the valence electrons one at a time around the element symbol as though it is inside a square. Distribute the electrons, one at a time starting on one side of the square and moving around to each side of the square. Do not start pairing until each side has one electron. If an atom has more than 4 valence electrons, begin pairing the electrons after the first four are placed.

Examples of the electron-dot structures for the atoms in the second row of the periodic table are shown on slide three. Elements in the same group or column of the periodic table have the same number of valence electrons; their Lewis structure will look the same.

Why Do Chemical Bonds Form?

Elements combine to become more stable. Substances are more stable when they are at a lower energy state and are excited or less stable when at a higher energy state. Valence electrons are the ones involved in chemical reactions and therefore the ones involved in chemical bonding. Noble gases are very unreactive. What do all noble gases have in common? They all have 8 valence electrons or an “**octet**,” **except helium**, which has 2 because that is the maximum number of electrons that the first energy level can hold. This led scientists to conclude that the octet arrangement of valence electrons must confer great stability to the noble gases. The basis for the Lewis bonding theory is that elements other than the noble gases combine and form bonds to copy the electron structure of the noble gases. This is the “**octet rule**.”

Two Types of Chemical Bonds

There are two types of **intra**molecular bonds (intra = within, these are bonds **between atoms** in a molecule):

an **ionic bond** is formed by transfer of electrons, producing **ionic compounds**.

a **covalent bond** is formed by sharing electrons, producing **molecular compounds**.

Writing Lewis Structures of Ions

Remember that ions form as a result of elements gaining or losing electrons. When a metal loses its valence electrons to attain the electron configuration of the next lowest noble gas, its valence shell becomes empty. The Lewis structure for a cation will therefore show only the element symbol and the charge on the ion (written as a superscript); there will be no valence electrons shown.

Nonmetals prefer to gain electrons to attain the electron configuration of the next highest noble gas. Therefore the Lewis structure for an anion will show a complete octet. The convention is to enclose the element symbol of an anion in square brackets, showing the complete octet of electrons in the valence shell. The charge of the anion is indicated as a superscript to the right of the ion outside the square brackets.

Determining Whether a Bond Will Be Ionic or Covalent

Whether an ionic bond or a covalent bond will form between two elements depends on their **electronegativity**: their ability to attract electrons. A large difference in electronegativity between two elements will result in ionic bonds while, no difference or a small difference in electronegativity will result in covalent bonds as are found in molecular compounds.

What types of elements have low electronegativity and what types have high electronegativity? Metals have low electronegativities because they do not attract electrons; they prefer to give them up to achieve the electron configuration of the next lowest noble gas in the periodic table. Non-metals on the other hand, have high electronegativities because they prefer to gain electrons. Doing so, they get the “octet” configuration of the noble gas in the same row of the periodic table.

Is it possible to predict whether an element will have a high or low electronegativity? (Yes, the position of an element on the periodic table is a good predictor of electronegativity. **Electronegativity increases from left to right and from bottom to top of the periodic table**. From this trend can you predict which element is the most electronegative (exclude noble gases, why?)?

Ionic Bonds and Electron-Dot Structures

To form an ionic bond, typically a metal transfers electrons to a nonmetal. The metal becomes the positively-charged **cation** and the non-metal becomes the negatively-charged **anion**. Remember that ions are not subatomic particles: they are atoms and molecules that have gained or lost electrons.

Once electron transfer has taken place, the ionic bond results from the electrostatic attraction between the positively-charged cation and the negatively-charged anion. Since the ionic bond depends on attraction of opposite charges, stronger bonds will occur when ions have higher charges. Smaller ions that allow closer approach of ions will also result in stronger attraction between the ions.

Slides 10-11 show an example for writing electron-dot structures for the formation of an ionic bond by transfer of an electron from a Na atom to a Cl atom to form NaCl, sodium chloride, which is table salt. Don’t forget the brackets and charge for the ions.

Charges on Ions

Students often have a difficult time with the charges on ions. Remember that in a neutral atom, the number of protons equals the number of electrons. Losing one electron means that there will be one more proton than there are electrons, so the charge overall will be +1. Losing two electrons will produce a +2 charge.

Gaining one electron means the electrons will outnumber the protons by one, so the charge on the ion will be –1. Gaining two electrons means the charge will be –2.

Electron-Dot Structures of Molecular Compounds

Molecular compounds form **covalent bonds** when elements have similar electronegativity. Since neither element exerts a much stronger pull on the electrons than the other does, the elements share electrons rather than transferring electrons; this has the effect of allowing each element to attain an octet of electrons. This type of bond typically occurs between two nonmetals, and is quite strong. Formation of a single bond requires sharing of a pair of electrons between two atoms. Slide 18 shows two chlorine atoms, each with seven valence electrons. Each atom contributes its unpaired electron to form a shared pair with the other atom. Pairing the electrons improves stability.

Atoms may form more than one covalent bond with surrounding atoms as indicated in the Lewis structure of methane, CH4. The carbon atom shares a pair of electrons with each of four hydrogen atoms.

Double or triple bonds can also be formed by sharing two or three pairs of electrons. Single bonds are longer than double bonds, which are longer than triple bonds. By contrast, triple bonds are stronger than double bonds, which are stronger than single bonds.

Exceptions to the Octet Rule

There are numerous exceptions to the octet rule, but we will only consider a few types.

First and foremost, helium represents a noble gas that is satisfied with only two electrons in its valence shell. This is because the outermost shell of the atom can hold only two electrons; it can be said to follow a duet rule rather than an octet rule. Hydrogen which has one electron can share or gain one electron to attain the same electron configuration as helium. Lithium can lose one electron to attain the same electron configuration as helium. These two elements can also be expected to follow the duet rule.

Other exceptions to the octet rule occur with beryllium and boron. Beryllium is a group IIA metal and might be expected to lose two electrons, but it often shares two electrons to form two single bonds. Boron shares three electrons forming three single bonds. Because these elements are electron-deficient and do not conform to the octet rule, they form very reactive compounds.

Some elements are stabilized by sharing more than eight electrons. This can occur in elements in row 3 or lower in the periodic table. In this situation, empty “d” orbitals in an element can accommodate additional electrons to expand the bonding beyond the octet rule.

Finally, an important exception involves molecules that have an odd number of electrons. There is no way to write a Lewis structure for such a compound that will obey the octet rule for all the atoms. This is because the Lewis model is a fairly simple model of bonding; a more sophisticated model is required for some circumstances. Like beryllium and boron compounds, molecules containing unpaired electrons are very reactive. Presence of this type of compound in the human body or in the environment have severe consequences, which will be discussed later in the semester.

Molecular Geometry of Compounds

Although we are often limited to two-dimensional representations of molecules on paper when drawing Lewis structures, they are more properly thought of in three dimensions. (Think about it- if molecules were flat, wouldn’t we all look like cardboard cutouts?)

Shapes of molecules can be described in terms of geometric shapes or geometric solids. We imagine the atoms of a molecule to be located at the “corners” or vertices of a particular shape. The central atom may be located at the center of the body of the shape. These shapes have characteristic angles between the vertices which represent the bond angles of the compound.

Predicting Shapes of Molecules using VSEPR

Molecules have 3-D shapes. How can we predict what shape a molecule will assume in space? The **valence shell electron pair repulsion** **(VSPER)** theory is used to predict shapes of molecules by assuming that electron pairs repel each other and therefore will arrange themselves as far away from each other as possible.

You could approximate the shapes predicted by VSEPR by blowing up several round balloons and tying them closed. Tie the tails of two of the balloons together with a string and set them on a table. The two balloons will not try to lie side-by-side, but will point in opposite directions. In trying to be as far apart as possible, the two balloons will lie in a more or less straight line (180º angle) with the tails between them.

Add a third balloon to the pair and tie it in place with a string. Depending on the size of the balloons and how tightly you tie them, the three balloons will probably still lie fairly flat on the table, with the tails in the center. This shape is called a trigonal planar shape, and the balloons will be approximately 120º apart.

If you add a fourth balloon to the mix, the balloons will no longer lie flat on the table. Now, the most efficient shape allows three of the balloons to point downward, resting on the table, with the fourth pointing upward. This shape will resemble a tetrahedron, a four-sided figure with vertices at 109.5º apart.

Slide 28 in the powerpoint slides includes perspective drawings of these figures. These are the common molecular shapes when the central atom in a molecule has 2, 3, or 4 groups of electrons around it, respectively.

VSEPR is quite good at predicting the shape of simple molecules.

Consider the compound **methane,** **CH4** ,shown in slide 28, with its electron-dot structure. What shape will it assume so that the 4 electron pairs are as far apart from each other around the central C atom? By measuring bond angles it can be determined that the molecule has a **tetrahedral** shape. Imagine the C atom to be in the center of a tetrahedron (where the four tails of the balloons were tied together), the four electron pairs shared with H are represented by the balloons, and the hydrogen atoms would be at each vertex of the tetrahedral shape.

**NH3** **(ammonia)** has a **trigonal pyramidal** shape because the nitrogen atom has one unshared pair of electrons that strongly repels the other three. The unshared pair of electrons is at the apex and the other three pairs are at each corner or vertex of the base. Although the unshared pair is not “seen” if we build a model of ammonia, it still takes up a lot of space at one vertex of the tetrahedron. As a result, if we consider only the vertices with actual atoms present, the molecule looks a little like a three-legged stool.

**H2O(water) has a bent** shape because it has two unshared pairs of electrons and two shared pairs of electrons around the central oxygen atom. The two unshared pairs push the other two shared pairs. Again, if we consider only the atoms that are present, the shape of the molecule itself is bent because the unshared pairs of electrons are not considered to be part of the molecular shape.

Just as a straight line can be defined by two points, in a molecule where only **two atoms are bonded together**, the only possible shape is **linear.**

Molecules with **two groups of electrons around the central atom** would be also be **linear**, as we saw in the balloon model when we tied two balloons together. Carbon dioxide is an example of such a molecule. Although there are eight electrons around the central carbon, they are in two groups of four electrons each. There are no unshared pairs in the model of carbon dioxide (as there were in the model of water.) Thus, **carbon dioxide has a linear shape.**

Two Types of Covalent Bonds – Polar and Non-Polar

Two types of covalent bonds can form: polar bonds and non-polar bonds.

A bond is **non-polar** if the two atoms have identical or very similar **electronegativities** and therefore are sharing the electrons equally. (Think of it like a tug of war between two students of equal size and strength- both will be able to pull equally on a rope.)

**Polar bonds** are formed when one atom has a greater electronegativity (attraction for electrons) than the other atom so electrons are not shared equally. The more electronegative atom pulls the electrons more toward it therefore, they spend more time closer to the nucleus of the more electronegative atom than to the nucleus of the other atom. (This would resemble a tug-of-war between a fifth grader and a kindergartner.)

Polar bonds create partial positive and negative charges on the molecule. For example, in H2O, oxygen is much more electronegative than hydrogen. The oxygen pulls the electrons toward it creating a partial negative (-) charge on the O atom and partial positive (+) charge on the H atom. This makes the **H–O bond a polar bond.**

Electronegativity and Bond Polarity

We have mentioned the concept of electronegativity in discussing ionic bonds and also in terms of polar and nonpolar bonds. Electronegativity is the attraction that an atom has for shared electrons in a covalent bond. Electronegativity increases from left to right across the periodic table, and it decreases as you go down the periodic table. The greater the electronegativity differences between two atoms bonded together, the more polar the bond. Electronegativity values range over a scale from 0.8 to 4.0, with fluorine being the most electronegative element and cesium being the least electronegative element.

To determine how polar a bond is, find the electronegativities of the two elements from the periodic table in the book at the bottom of page 322. Subtract the smaller electronegativity from the larger electronegativity to find the difference. In the example we used for water, oxygen has an electronegativity of 3.5 and hydrogen has an electronegativity of 2.1. The difference between them is 1.4. According to table 10.2 on page 323 in the book, an electronegativity of 1.4 would fall in the polar covalent category.   
  
If the difference ranges from zero to approximately 0.4, the bond is “pure covalent” or “nonpolar.” If the difference falls in the range from 0.4 to 2.0, the bond is polar covalent. A large electronegativity difference means the bond is essentially ionic. These numbers are guidelines only will vary somewhat from one textbook to another.

Polarity of Molecules

The paragraphs above are concerned with the polarity of individual bonds. If a molecule has polar bonds, is it necessarily a polar molecule? The answer is no.

We mentioned that **water has polar O–H bonds,** and it is indeed a polar molecule. Not only are both of the O–H bonds polar, the pull of the electrons toward the oxygen as well as the unshared pairs of electrons on the oxygen make that end of the molecule considerably more negative than the hydrogen end of the molecule, making **H2O a polar molecule**.

On the other hand, **carbon dioxide also has polar bonds**. (Electronegativity for C=2.5 and for O=3.5.) Overall, a C=O bond would definitely be a polar covalent bond. However, as we said, carbon dioxide is a linear molecule. If both of the oxygen atoms exert identical pull on the electrons they are sharing with carbon, the pulls will cancel each other out. Carbon dioxide is symmetrical, and therefore **carbon dioxide is a nonpolar molecule**, even though it has polar bonds.

**To summarize, in order for a molecule to be polar, it must not only have polar bonds, but shape of the molecule must not have symmetry that would cause the pull on the electrons to be canceled (balanced) out.**

The final slide in the series shows another example. **Carbon tetrachloride, CCl4,**  is a tetrahedral molecule with four chlorine atoms arranged at the vertices of a tetrahedron with the carbon at the center. Carbon has an electronegativity of 2.5, and chlorine has an electronegativity of 3.0. The difference in electronegativity is 0.5, which means each bond is a polar covalent bond. However, in carbon tetrachloride, the chlorines pull equally on the electron density and the molecule is symmetrical in shape, so carbon tetrachloride is a nonpolar molecule.   
  
On the other hand **dichloromethane, CH2Cl2,** has two polar bonds between the carbon and the two chlorine atoms. It also has two nonpolar bonds between the carbon and the two hydrogen atoms. The molecule has some symmetry to it, but the two C–Cl bonds do not cancel each other out. Dichloromethane is a polar molecule.

**Guide to Molecular Weight and Stoichiometry**

**Molecular Weight:**

Molecular weight = sum of the atomic weights of all atoms in a molecule. Its units are g/mole.

To calculate molecular weight,

1) count how many of each type of atom there are in the molecule.

2) add up the atomic weights for all atoms (atomic weights are on the periodic table)

Example: N2O molecular weight = 2(14) + 16 = 44 g/mole

(The atomic mass of N is 14 g/mole, and the atomic mass of O is 16 g/mole.)

**Converting from grams to moles:** (of same compound)

1) You must first calculate the **molecular weight** (for molecular or covalent compound) or **formula weight** (for ionic) of the compound.

2) Set up a conversion:

number of grams you are converting from

given # grams 1 mole = # of moles

MW grams

the molecular weight of the compound

3) Notice: this means you will be dividing the grams given, by the molecular weight of the compound. The answer will be in moles, since grams cancel.

Example: If you have 100 g nitric acid (HNO3), how many moles is that?

MW= 1 + 14 + 3(16) = 63 g/mole

100 g HNO3 1 moleHNO3 = 1.59 molesHNO3

63 g HNO3

**Converting from moles to grams:** (of same compound)

1) You must first calculate the molecular weight of the compound.

the molecular weight of the compound

2) Set up a conversion:

number of moles you are converting from

given # moles MW grams = # of grams

1 mole

3) Notice: this means you will be multiplying the given number of moles by the molecular weight of the compound. The answer will be in grams, since moles cancel.

Example: If you have 2.00 moles H2SO4, how many grams H2SO4 is that?

MW = 2 + 32 +4(16) = 98 g/mole

2.00 moles 98 grams = 196 grams H2SO4

1 mole

**Using Chemical Equations: Stoichiometry**

A balanced chemical equation tells you the proportions by which the reactants combine. It also tells you the relationship between the quantity of the reactants and the amount of product.

3Fe + 2O2 → Fe3O4

The numbers in front of each compound/element are called “coefficients”. They tell the proportions by which compounds combine and are formed. If there is not a number shown (as for Fe3O4) then the coefficient= 1.

**Moles of one reactant/product to moles of another reactant/product:**

When given an amount for one reactant and asked how many moles of the other reactant will be needed.... If you are given an amount in grams, convert to moles (see reverse).

1) Set up a conversion, using the coefficients in the balanced chemical equation.

number of moles for the compound you know

the number in front of the compound your calculating for

given # moles known compound coefficient moles unknown compound = # of moles of

coefficient moles known compound “unknown” compound

the number in front of the known compound in the equation

2) So you will be multiplying the number of moles for one reactant times the coefficient for the other reactant and then dividing by the coefficient for the first reactant. Moles of “known” compound cancel, leaving you with moles of “unknown” compound.

Examples:

1) Use the chemical equation at the top to calculate:

If you have 4.75 moles Fe, how many moles of O2 will it take to react it?

4.75 moles Fe 2 moles O2 = 3.17 moles O2

3 moles Fe

2) Use the chemical equation: 2NH4Cl + Pb(NO3)2 → 2 NH4NO3 + PbCl2

to calculate how much lead nitrate [Pb(NO3)2 ] would be needed to react completely with 25 moles of ammonium chloride [NH4Cl]?

25 moles NH4Cl 1 mole Pb(NO3)2 = 12.5 moles Pb(NO3)2

2 moles NH4Cl

For a covalent compound, writing the formula from the name is super easy:

The name tells you how many of each atom.

# Prefix # nitrogen monoxide = NO

none 1 dihydrogen monosulfide = H2S

mono- 1

di- 2

tri- 3

tetra- 4

Naming an ionic compound from a formula is super-easy:

Just name each element, in order.

End the last one with –ide.

If you have a polyatomic ion, just use its name, and never change its ending to –ide.

NaCl = sodium chloride

KNO3 = potassium nitrate

Reduce if possible. The subscripts should give the proportions in the lowest numbers possible. Example:

Mg2O2 → MgO

Naming a covalent compound from its name is easy, too:

Name the elements in order; with the last one ending in “ide”

Add a prefix to tell how many atoms of each element.

Exception: drop any initial “mono-“

CBr4 = Carbon tetrabromide

H2O = dihydrogen monoxide

Writing formulas from a name of an ionic compound.

Decide what two ions are present.

If more than two elements are present, then look for polyatomic ions.

Cross-over charges to get subscripts (drop the sign, use the numbers)

Al+32 S-23 → Al2 S3

Write symbols for each ion, with the positive one 1st. Indicate the charge on each ion.

Al+3 S-2

Determine charge on each ion.

For elements, look at trend on periodic table. For polyatomic ions, look on list (lower right).

**Polyatomic Ions:**

**SO4-2 sulfate**

**NO3-1 nitrate**

**NH4+1 ammonium**

**PO4-3 phosphate**

**OH-1 hydroxide**

**CO3-2  carbonate**

Drawing an electron dot structure for a covalent compound:

Before you start drawing electrons, arrange atoms (element symbols). If there is just two atoms, this is not a big deal, but if there are more, follow this rule:

Put the atom that is “different” in the center.

Draw each atom with its valence electrons. (See periodic table for number of valence electrons.)

Place electrons one of each of four sides first, then pair up as necessary.

correct: C incorrect: C

Figure out which electrons can “pair-up” and be shared by both atoms. Arrange atoms and electrons accordingly.

Count to make sure all atoms have an octet (8 electrons). Remember, Hydrogen is happy with just 2. Be sure to count both shared electrons for both atoms sharing them.

**Drawing an electron dot formula for an**

**ionic compound:**

If you are give only a name, determine the

formula, using the flowchart on back.

If you have a formula, first write each atom

with its correct number of valence electrons. (Use the periodic table to determine this number.)

Decide which electrons will be transferred.

If an atom has 3 or less electrons, it will lose them all.

If an atom has 5 or more electrons, it will gain enough to have 8.

The goal is for each atom to show either no or 8 valence electrons.

Ca will lose 2 e’s

P will gain 3 e’s

Draw atoms with electrons transferred.

Bracket each atom, and write its charge.

If an atom gains electrons, it’s negative;

if it loses electrons, it’s positive.

Ca will lose 2 e’s → Ca+2

Br will gain 1 e → Br-1

So: [ Br ] -1 [ Ca ]+2 [ Br ]-1