WELCOME TO AP CHEMISTRY!

Even as we wrap up another successful school year, it is never too early to start thinking about the future. Next year, new and interesting challenges await you in Advanced Placement Chemistry. From its humble alchemical beginnings as a search to turn other metals into gold and to produce an elixir that grants endless life, chemistry has grown into a comprehensive study of the composition and behavior of the universe. In your first year chemistry course you were able to see how broad the spectrum of topics can be; and you learned not only the general chemical principles involved, but also how these topics fit together. It is the goal of the AP Chemistry course to delve deeper into each of those areas to discover new truths about the beauty and complexity of creation. Next year, you will gain a deeper understanding of the "what" and "how" that makes the world around you familiar and functional. Topics to be addressed include:

- Organic Nomenclature and Reactions
- Quantifying Chemical Reactions
- The Behavior of Gases
- Thermochemistry
- Submicroscopic Structure and Behavior
- Chemical Bonding and Periodicity
- Solids, Liquids, and Solutions
- Chemical and Nuclear Kinetics
- The Equilibrium Condition
- Acids & Bases
- Thermodynamics
- Electrochemistry

While it is certainly possible to appreciate chemistry from a purely qualitative perspective, true awe is derived (Ha! Get it?) from an understanding of mathematical relationships. You need to be very comfortable with algebraic manipulation and logarithmic functions. Concurrent enrollment in Pre-Calculus or higher is strongly recommended but not required.

Now to the point – there are not enough hours in the day and not enough days in the year for us to arrive at our destination if we delay our departure. You must hit the ground running next August, and that will require you to do some personal preparation beforehand. Attached you will find sample exercises to reacquaint yourself with fundamental concepts and skills from the first-year chemistry course. You are expected to make sure you are comfortable with these exercises before you attend the first class. Please email me for a key when you are ready to check your work, because ensuring mastery will solidify your Pre-AP Chemistry fundamental knowledge assessment we will have during the second week of school. Most - if not all - of this material is simple to do and to understand, but repetition is often the key to mastery. Skills addressed in this summer assignment include:

- Dimensional Analysis and Unit Conversion
- Naming and Writing Formulas for Compounds
- Writing and Balancing Equations
- The Mole and Stoichiometric Relationships

Please don't hesitate to send an email to me at djdotson@prosper-isd.net if you are having difficulty with this assignment, simply need a question or two answered, or would like to see the answer key. I will do my best to help you solidify this foundation on which we can build your understanding of chemistry next school year. You making the choice to do more in Chemistry makes you very important to me. Get comfortable e-mailing me when needed because I will be there for you! I'm excited about how much you will learn next year!

Regards,
Mrs. Deshaun Dotson

ATTACHMENTS:

Chapter One: Introduction to Chemistry	READ ME!
Chapter Two: Measurement	READ ME!
Scientific Notation & Significant Digits	**COMPLETE ME!**
Dimensional Analysis #1 and #2	**COMPLETE ME!**
More Dimensional Analysis Practice Problems	**COMPLETE ME!**
Naming and writing formulas for Ionic Compounds	READ ME!
Naming and writing formulas for Covalent Compounds	READ ME!
Naming and writing formulas for Acids	READ ME!
Common Polyatomic Ions ********	**MEMORIZE ME!**
Solubility Rules and Strong Acids and Bases********	**MEMORIZE ME!**
Periodic Table of the Elements	REFER TO ME!
Worksheet: Formulas and Nomenclature	**COMPLETE ME!**
Lots of Ionic Naming Practice Problems	**COMPLETE ME!**
Naming Covalent Compounds Worksheet	**COMPLETE ME!**
Naming Acids and Bases	**COMPLETE ME!**
Nomenclature – Practice Sheet	**COMPLETE ME!**
Balancing Equations Practice	**COMPLETE ME!**
Balancing Equations #1	**COMPLETE ME!**
Balancing Equations #2	**COMPLETE ME!**
Balancing Equations #3	**COMPLETE ME!**
Molar Mass Conversion Practice	**COMPLETE ME!**
Molarity	**COMPLETE ME!**
Stoichiometry Practice A	**COMPLETE ME!**
Stoichiometry Practice B	**COMPLETE ME!**
Let's Assign Oxidation Numbers	**COMPLETE ME!**

Online Resources:

www.khanacademy.org
A great website with video tutorials on specific problems. Click "Watch," then navigate to the Chemistry section.

http://www.sciencegeek.net/APchemistry/Powerpoints.shtml
Prepared Powerpoint® notes directly from our textbook, Chemistry, 5th edition, by Steven and Susan Zumdahl

CHAPTER ONE:
INTRODUCTION TO CHEMISTRY

1.1 THE CENTRAL SCIENCE

How often have you thought about chemistry today? The answer is likely "very little" or "not at all," but chemistry is all around you. From the processes that allow you to digest your breakfast to the many functions of the car you ride in – the engine, brakes, air conditioner, CD player – all of them can be explained with chemistry. Even the ability to read this page is dependent upon chemical signals and responses in your brain. Chemistry is often called "the central science" not because the course is between biology and physics, but because there is practically nothing that lies outside the realm of chemistry.

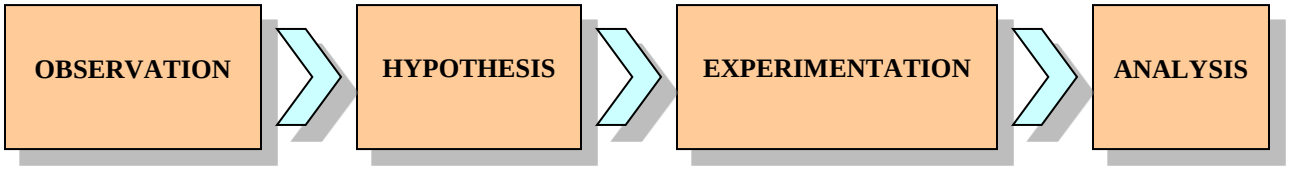
Chemistry is the study of matter and its interactions. Though there are many branches of chemistry, the differences lie only in the type of matter or the particular properties being studied. The Chemical Abstracts Service (CAS) Registry has catalogued the existence of over 27 million different substances, so it is impossible for one person to be an expert in every area of a field as broad as chemistry. Organic chemists, for example, only study substances that make up living matter. All living matter is made of compounds that contain the incredibly versatile element carbon in addition to other elements such as hydrogen, oxygen, nitrogen, and sulfur. Inorganic chemistry is the branch of science that studies all other substances. Although it may seem that inorganic chemistry is the larger field, approximately 85% of all compounds are classified as organic substances. The properties of carbon are such that it is capable of combining with other elements in nearly an infinite number of ways. Many of these compounds occur naturally, but a growing number are synthesized in a laboratory. Nearly 50,000 new substances are added to the CAS Registry every week.

Other branches of chemistry include biochemistry (the study of the chemical processes of living organisms), physical chemistry (the study of the properties of matter and the energy associated with it), and analytical chemistry, which is concerned with the identification and measurement of substances.

1.2 SCIENTIFIC METHOD

Table 1-1	Field of Study	Careers
ORGANIC CHEMISTRY	Compounds containing carbon	Pharmaceuticals, Plastics, Agro-chemicals
INORGANIC CHEMISTRY	Compounds without carbon	Environmentalist, Materials science, Metallurgy
ANALYTICAL CHEMISTRY	Identification and measurement of matter	Forensics, Food Quality, Manufacturing
PHYSICAL CHEMISTRY	Properties and Energy	Nanotechnology, Molecular modeling, Biosensors
BIOCHEMISTRY	Processes of living organisms	Medicine, Genetics, Pharmacy

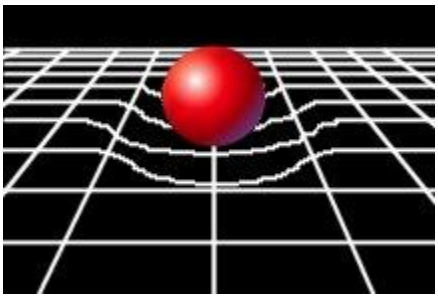
Each of these branches of chemistry focuses on matter, but what exactly is matter? Simply put, matter is anything that has mass and takes up space. Practically everything in the universe is matter, and to understand it better, chemists – and you, now that you are in this course – need a methodical and practical way to study it. Chemistry is an experimental science. Everything we know is a result of careful observation and a lot of testing. The framework for gaining knowledge by experiment is called the scientific method, and here we will consider the following basic model:



Observation is perhaps the most important tool we possess to help us understand the world around us. Observations can be either quantitative (measurements that involve a number and a unit) or qualitative. Sometimes an observation applies to many different systems and is formulated into a statement called a natural law. A common misconception is that a law is something that has been proven, but in fact a law is simply a summary of *what happens*. For example, when you drop a pencil, it falls to the floor. This is a qualitative observation. You might also say that a pencil dropped from a height of four feet takes exactly 0.5 seconds to reach the ground. This is a quantitative observation. In fact, you will find that a book, a nickel, a golf ball, or any object at all also takes 0.5 seconds to reach the floor from a height of four feet. This observation appears to apply to any system. The idea that all objects accelerate toward the earth is a natural law. There is nothing to prove, it is just an observation of a naturally occurring event.

Once an observation has been made, a hypothesis - a possible explanation for the observation - is formulated. Perhaps we may believe that objects fall to the ground because they are being pulled downward by tiny green men with ropes. It may sound absurd to you, but it is a hypothesis nonetheless. With our explanation in hand, it is time to perform an experiment in an attempt to verify or refute our hypothesis. Experimentation and the collecting of data form the basis for all scientific knowledge, and it is important for experimental data to be accurate and unbiased. So how will we test our hypothesis about little green men? Perhaps it will be as simple as looking for them: set up an array of high speed cameras, drop a pencil, and carefully examine every frame of recorded footage. But one trial is never enough to confirm or invalidate a hypothesis. The experiment must be repeated many times, sometimes by many different scientists, before the results can be claimed as truly valid. In our scenario, after recording the fall of a multitude of objects, there has been no evidence to support the existence of little green men with ropes. The gravitational attraction between objects must be caused by something else.

In 1920, a young patent clerk published an idea that appeared to experimentally account for the phenomenon of gravity, even the quite astounding observation that light can be “bent” by a strong gravitational attraction. This idea was firmly based upon years of observation, countless experiments, and a multitude of detailed calculations. The young man, Albert Einstein, suggested that space itself was distorted around very massive objects, and that what we see as “falling” is really just an object following the curves in space. This idea is called the General Theory of Relativity, and has been the accepted explanation for gravity for nearly a century. It is not a law, nor has it been “proven.” A theory is just an explanation of how or why something occurs and is backed up by experimentation. We assume this particular theory is true because there hasn’t been another explanation proposed that is supported by experimental evidence to the same degree. If observations are made which contradict the theory, General Relativity will have



to be modified, or even thrown out altogether. Science, along with the theories and models it produces, is constantly changing as new data is collected and analyzed.

1.3 EXPERIMENTATION

As you can see, experimentation is the focal point of science. It provides the evidence to support our conclusions about the both the visible and microscopic worlds. In an introductory chemistry class such as this one, experiments will be performed with one of three goals in mind:

- To test a hypothesis
- To confirm or demonstrate a natural law
- To gather and analyze data

It is important for you as a chemist to understand how an experiment is set up. As has been said before, experimental data must be accurate and unbiased. Any person who performs the experiment should be able to replicate the results of any other person. This can only be done when the focus of the experiment is explicitly defined – only one of the innumerable variables associated with the investigation can be changed by the researcher. This single variable is called the independent variable. To illustrate this concept, imagine that you are going to gather information about how plants grow when exposed to lights of different color. There can be only one independent variable – the color of the light – and everything else must remain the same for every trial. The type of plant used, the intensity of the light, the temperature, the amount of water, and the composition of the soil must be the same – constant - for every experiment. Although there can only be one independent variable, there can be numerous dependent variables which change as a result of the independent variable. For our experiment, several things might be affected by the color of light: the height of the plant, the number of leaves, the depth of the root system, etc. This type of experiment has one final component, the control. The control is a trial for which the independent variable is considered normal or unchanged. In this scenario, the control would be white light or sunlight. The control is used for comparison, serving as a baseline by which to evaluate the effectiveness of a change in the independent variable. In addition, multiple trials should be conducted for each change of the independent variable to prevent errors from skewing the results.

Collecting data from an experiment in an organized way allows a chemist to find information quickly and easily and often helps in evaluating its significance. There are many ways to organize gathered information, and the method used depends entirely on the type of experiment being performed. A data table is the most commonly used organizational tool, but every data table looks different. Two of the most common are shown below; the first for an experiment gathering qualitative data and the second for a quantitative investigation. Often a data table must be custom-made to fit the needs of a particular experiment, and when all else fails, a simple list can suffice.

1.4 ANALYSIS

Collecting data is only part of experimentation. Equally important is the analysis of the data that is collected. There are many ways to analyze data, but here we will look at three:

- compare and contrast
- patterns
- graphs

It is as important to look at what is similar as to examine the differences between things. For example, in Table 1-3 we see that each plant sprouted on the third day, was 0.5 cm tall, and had two leaves. This should lead us to the conclusion that light does not play an important role in the initial growth of a plant. We also see that there is always an even number of leaves. The differences are obvious; plants grew larger under white light than red light, and larger under red light than blue light. Therefore our conclusion might be that colored light hinders the growth of

Table 1-2

ACTION	OBSERVATION	INTERPRETATION
Put ice in water	Ice floats	Ice is less dense than water
Put ice in alcohol	Ice sinks	Ice is more dense than alcohol
Put alcohol in water	Liquids mix	Water and alcohol are miscible
Put oil in water	Liquids separate	Oil and water are immiscible

Table 1-3

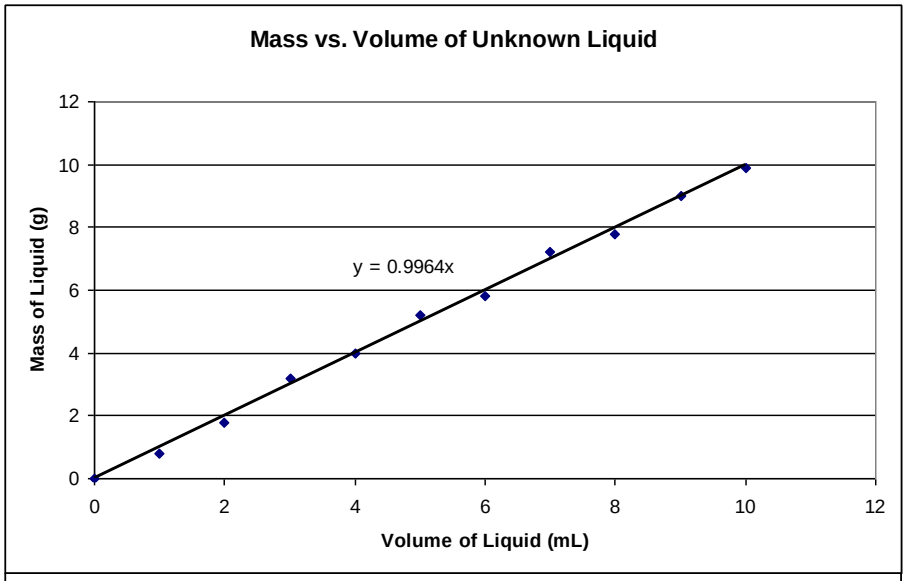
	White Light		Red Light		Blue Light	
Day	Height	Leaves	Height	Leaves	Height	Leaves
1	0 cm	0	0 cm	0	0	0
2	0 cm	0	0 cm	0	0	0
3	0.5 cm	2	0.5 cm	2	0.5 cm	2
4	1.0 cm	2	0.5 cm	2	0.5 cm	2
5	1.5 cm	2	0.5 cm	2	0.5 cm	2
6	2.0 cm	4	1.0 cm	2	0.5 cm	2
7	2.5 cm	4	1.0 cm	2	0.5 cm	2
8	3.0 cm	4	1.2 cm	2	0.5 cm	2
9	3.5 cm	4	1.2 cm	2	0.5 cm	2
10	4.0 cm	6	1.5 cm	2	0.5 cm	0

plants, but we might improve this experiment by testing more colors such as yellow, orange, or green.

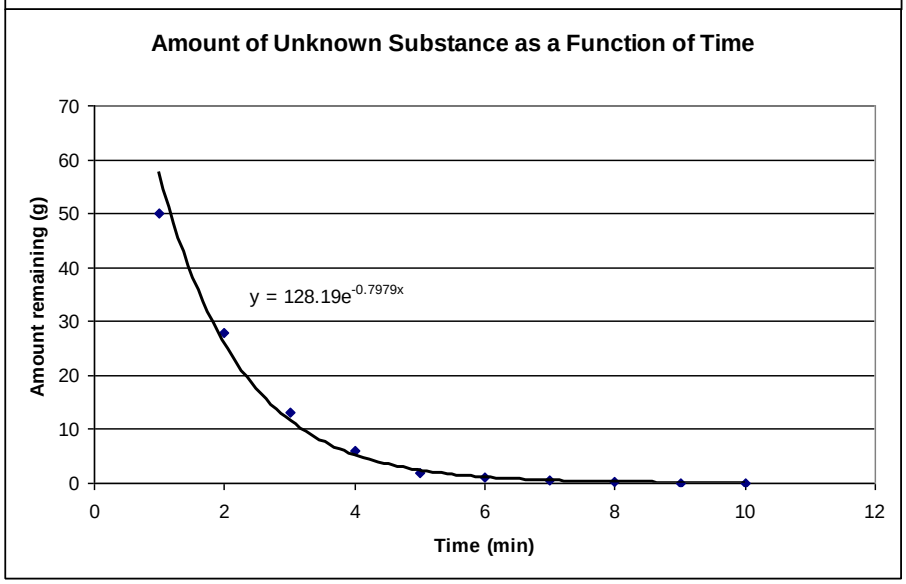
Patterns in data lead us to important conclusions. In Table 1-2, four substances are mixed together in various ways. We see that ice is less dense than water but more dense than alcohol. We can create a pattern out of this that lists the three substances in order of increasing density: alcohol, ice, water. The data for the fourth substance, oil, does not give us any indication as to its relative density. This experiment could be improved by examining the relationship of oil to ice and alcohol as well. In the experiment described in Table 1.3, the number of leaves is a pattern. There is always an even number of leaves, and each pair appears to develop after an additional 2.0 cm of growth. The last plant, however, lost leaves on the 10th day, leading us to conclude that the plant has died and its leaves have fallen off.

Perhaps the most useful way to analyze data is to create a graph – a model that shows the relationship between one variable and another in an experiment. For most graphs, the independent variable is plotted along the x-axis and the dependent variable along the y-axis. The data points are plotted and the mathematical relationship is determined to be linear, exponential, quadratic, etc. Since all experimentation involves a certain amount of error, it is important to create best-fit lines or best-fit curves. Examples are shown on the following pages. All graphs must be labeled appropriately – a title that describes the graph, labels on each axis that include the units of the measurements being utilized, and if necessary, the slope of the line or its equation.

Graph #1: A scatter plot with a best-fit line. In this experiment, the mass (dependent variable) of a liquid was measured at varying volumes (independent variable). The data points indicate a linear relationship, so a best-fit line was drawn that most closely approximates each point. Notice that it is not essential for every point to be on the line. The equation for the line is in slope-intercept form, y = mx + b, and the slope indicates the density of the liquid (d=m/V).

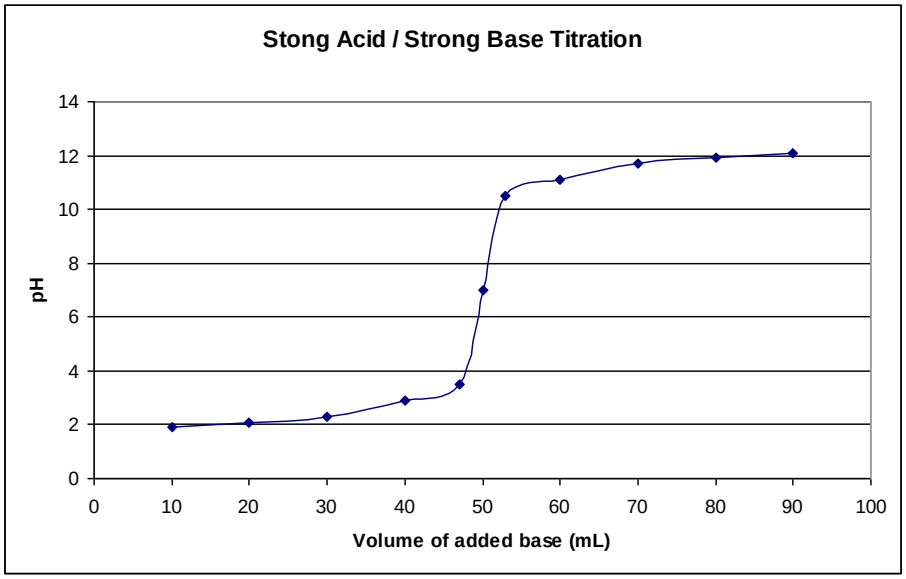


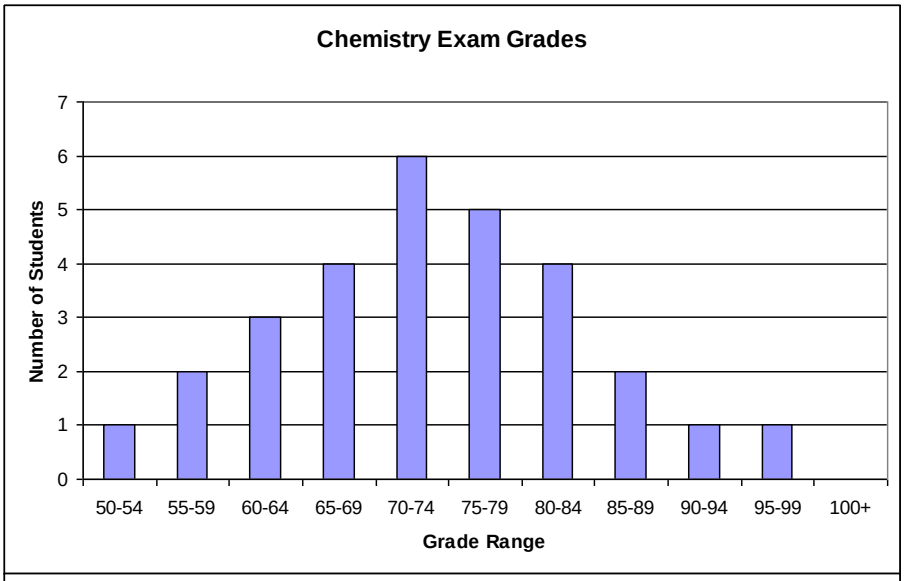
Graph #2: A scatter plot with a best-fit curve. In this experiment, the amount of substance remaining (dependent variable) was measured as time (independent variable) passes. The data indicates an exponential or logarithmic relationship, so a best-fit curve was drawn that most closely approximates each point. Notice that it is not essential for every point to be on the curve. The equation for the line is written in logarithmic form, $y = Ae^{-kx}$, where A is the original amount of sample and k is the rate of decay.



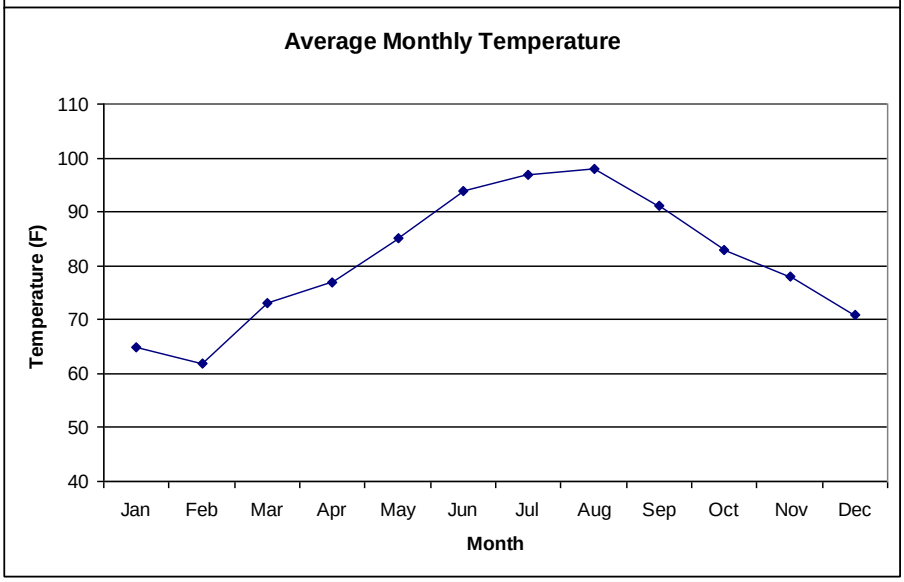
Graph #3: A scatter plot with a connecting smooth curve. In this experiment, the pH (dependent variable) of an acidic solution was measured as base is added (independent variable). The data indicates a complex polynomial relationship, so we connect all the data points with a smooth curve. No equation is given.

Scatter plots are used when we are certain that a mathematical relationship exists between the independent and dependent variables in an experiment.

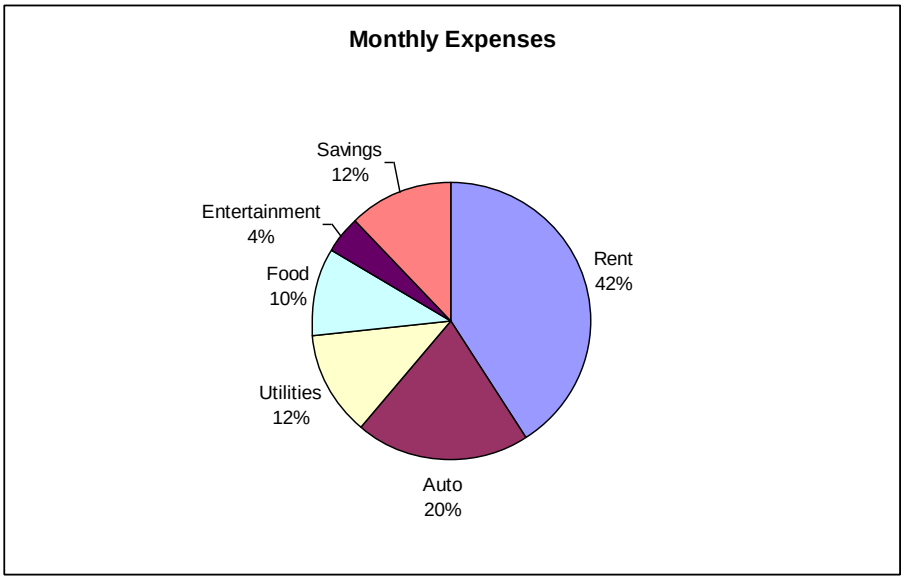




Graph #4: A Histogram. This graph indicates the grade distribution on a typical chemistry exam. *Histograms are useful for graphing the number of items in a category.*



Graph #5: A line graph. This graph plots the average temperature (dependent variable) for twelve months (independent variable). *Line graphs are almost exclusively used for graphing the change in a property over time.*



Graph #6: A pie graph. This graph shows the relative amounts of an average person's income that are spent in certain categories. *Pie graphs are almost exclusively used for showing percentages, though data in a histogram can also be shown in pie graphs.*

1.5 LABORATORY SAFETY

Since much of your work in this class will be done in the laboratory, it is important that you learn to follow certain safety guidelines to protect yourself and your classmates from serious injury.

General Safety Rules:

1. **Eye protection is required at all times in the laboratory.** Your eyes are very sensitive and can easily become irritated or damaged if you are not wearing safety goggles at all times. Goggles protect your eyes from chemical splashes, harmful vapors, and sharp equipment.
2. **No food or drinks are allowed in the classroom.** Because of the nature of chemistry, you will be working with many substances that are harmful if they are ingested. It is not worth the risk to assume that what you are eating has not been contaminated in some way.
3. **Horseplay and pranks are prohibited.** Chemistry lab can be fun and exciting, but it can also be dangerous if care is not taken to ensure the safety of everyone working nearby.
4. **Unauthorized experiments are prohibited.** Even though science is based upon curiosity and seeking answers to questions, it is important to have a general idea of what to expect from an experiment so you will not accidentally cause injury to yourself or others.
5. **Locate and know how to use all safety equipment.** The fire extinguisher, fire blanket, safety shower, and eye-wash station can save your life if used properly and responsibly.

Emergency Procedures:

1. **Chemical spills.** If any chemical gets on your skin or clothing, flush the area with lots of cold water and notify the teacher immediately. Use the safety shower if the spill is extensive. If the chemical gets into your eyes, immediately irrigate the eye for 15 minutes at the eye-wash station.
2. **Burns.** Immerse the burned area in cold water and notify the teacher.
3. **Cuts and abrasions.** Immediately clean the wound with water and notify your teacher. Hold a sterile pad firmly over the wound until the bleeding stops, then apply a bandage.
4. **Fires.** A small fire at your desk can usually be put out by smothering it with a nonflammable material such as a damp rag or an inverted beaker. If your clothing catches fire, try to use your lab apron to put it out or wrap in a fire blanket. If necessary, roll on the floor. If a fire cannot be put out by the above procedures and there is little personal risk, use the fire extinguisher. Do not put yourself in danger to extinguish a fire.

Working with Chemicals:

All chemicals are potentially harmful to some degree. Avoid direct contact with any chemical. It is especially important to keep chemicals away from your hands, face, and clothing. Many substances are easily absorbed through the skin or through inhalation. Chemicals can also enter the body through the mouth or transferred to your eyes if your hands are contaminated.

1. **Never taste any chemical.**
2. **Carefully read the label twice on any bottle prior to using it.** Use chemicals only from containers that are clearly labeled.
3. **Do not carry supply bottles to your desk** as other students will need them. Bring your appropriate container to the supply table and take only what you need.
4. **Do not return unused portions of chemicals to their containers** as you could contaminate the entire bottle. See if other students in your area need the chemical or dispose of the excess as directed by your teacher.
5. **Weigh chemicals in a previously-weighed container or on weighing paper** rather than directly on the balance pan.
6. **Never smell an unknown substance directly to determine its odor.** Carefully waft the fumes toward your nose to protect yourself from harmful vapors.
7. **Pour substances from the reagent bottles holding the label side of the bottle in your hand.** This prevents dripping on the label and provides a clean side for holding the bottle.
8. **If a solution spills onto the table, dilute the spill with lots of water** and use paper towels to soak it up or to push it into the sink. Dispose of the towels. If an acid is spilled, neutralize it with sodium bicarbonate (baking soda), then clean up as before. Strong bases can be neutralized with acetic acid (vinegar).
9. **Disposal of waste chemicals:**
   (a) **Do not put any solids, paper or broken glass into the sink.** They are to be disposed of in the trash or in the waste jars provided.
   (b) **Acids, bases and water solutions may be washed down the sink** with large amounts of water, unless your instructor gives you other disposal instructions.
   (c) **Volatile or flammable liquids should not be poured down the drain**, but should be placed in specially marked containers and kept sealed.
10. **When diluting acid always add acid into the water.** The water will absorb the heat produced and also prevent the acid from splashing onto your skin.

<u>**Working with Heat:**</u>

1. **Never reach across an open flame.** It is advisable to roll up long sleeves and to tie back hair that is longer than shoulder-length.
2. **Before heating glass containers, examine them to see that they contain no cracks.** The expansion caused by heating could cause the damaged glass to break.
3. **When heating any solid or liquid in a test tube, keep the tube in constant motion and do not point the mouth of the tube at another person.** Hold the test tube with test tube clamps to avoid burning yourself.
4. **Always hold the test tube that is being heated at an angle**, and heat the sides of the tube as well as the bottom.
5. **Never look down into a tube containing a reagent or hot water**, especially if it is being heated.
6. **Never apply a direct flame to a container of volatile or flammable materials**, and never place an open flame near such containers.
7. **Hot glass looks just like cold glass, so always place hot objects on wire gauze to cool.** Hot glass can inflict severe burns.
8. **Never immerse hot glassware in cold water,** which could cause it to shatter.

Conclusion of the Lab:

1. **Clean and dry all of your glassware and your lab desk.** Return all of your equipment to its proper place.
2. **Check to see that the gas and water are turned off** before you leave your working area.
3. **Wash your hands thoroughly.**
4. **Place your goggles in the sterilizer** and your apron in its proper place.

1.6 LABORATORY EQUIPMENT

Choosing the right equipment is just as important as designing an experiment. Without the proper tools, it is difficult to achieve quality results that can be repeated by future researchers. It is important for you to be able to identify all of the equipment below and be able to use it effectively.

Measurement:
Graduated Cylinder
Buret
Volumetric Pipet
Dropper Bottle

Reaction Vessels:
Test Tube
Beaker
Erlenmeyer Flask
Florence Flask
Evaporating Dish
Crucible
Well Plate

Heating:
Bunsen Burner
Hot Plate
Wood Splint

Tools:
Scoopula
Dropper
Disposable Pipet
Beaker Tongs
Crucible Tongs
Test Tube Tongs

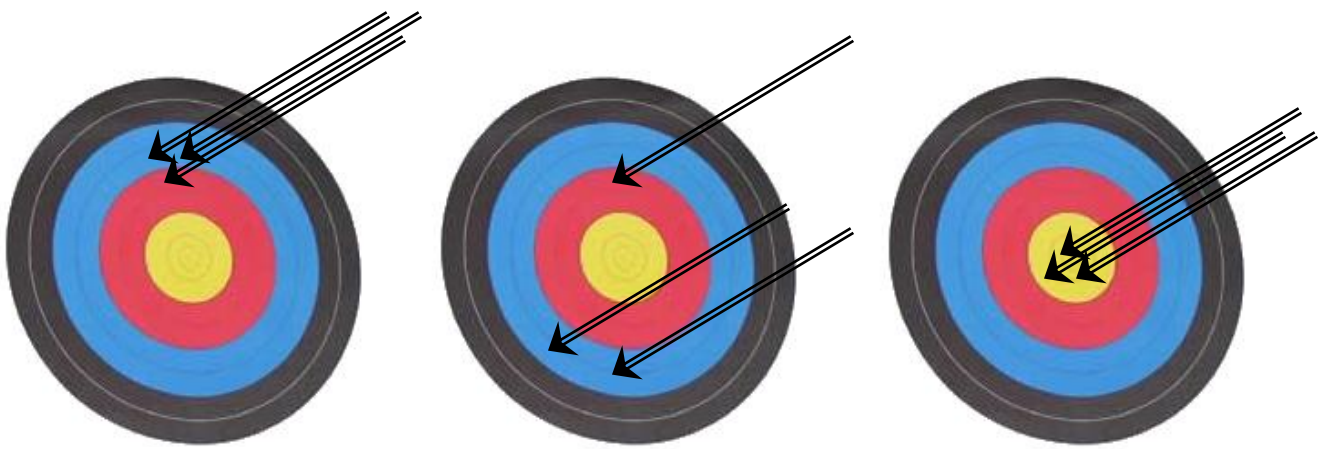
Miscellaneous:
Watch Glass
Stirring Rod
Funnel
Ring Stand
Ring Clamp
Utility Clamp
Clay Triangle
Wire Gauze

CHAPTER TWO:
MEASUREMENT AND MATTER

2.1 ACCURACY AND PRECISION

In order for valid conclusions to be drawn from quantitative measurements in science, it is essential that those measurements be both repeatable and reliable. The reliability of a measurement refers to how close a value is to the true or accepted value and is called accuracy. Careful selection and calibration of laboratory equipment can go a long way toward ensuring accurate measurements. The repeatability of a measurement is called precision, and is usually talked about in two ways – how close a series of measurements are to one another, and the "fineness" of a particular measurement, which will be discussed momentarily.

The figure below showing the archery targets is a good way to visualize the difference between accuracy and precision. In the first scenario, the arrows all strike the target near one another, but they are all far from the bull's-eye. This would represent an experiment whose data are precise but not accurate since the result is repeatable but not reliable. In the second scenario, none of the arrows hit the center, meaning the results were neither accurate nor precise. On the final target, the results are both accurate and precise as all arrows strike near the bull's-eye.



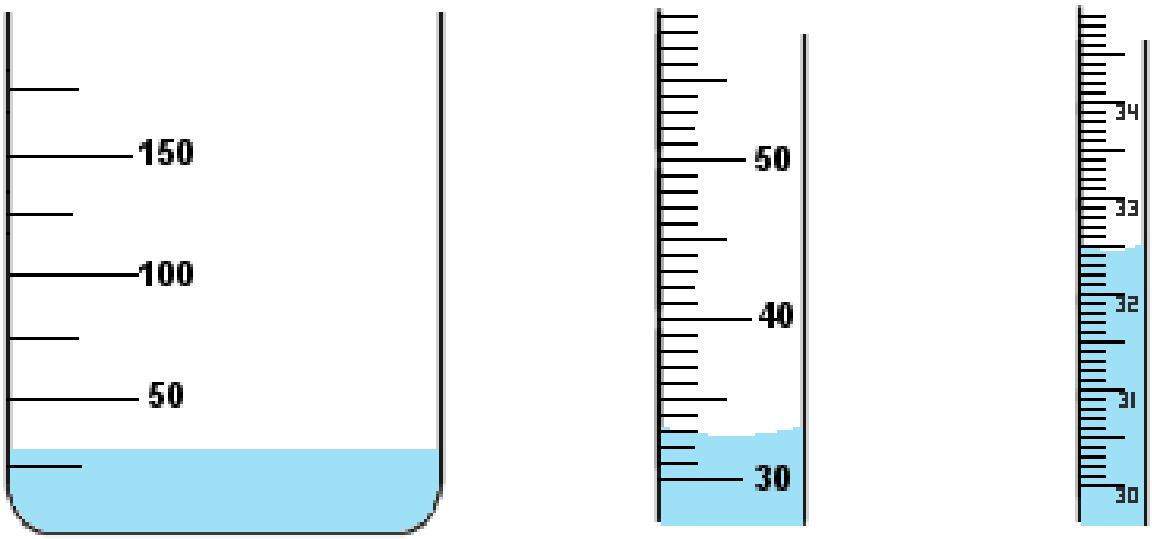
Whenever possible, you should always perform an experiment multiple times. This will eliminate much of the human error involved and produce much more accurate results. In addition, when an experiment is performed repeatedly with the same result, we have confidence that the measurement is accurate. In general, it is assumed that if the measuring instrument is in working order and is properly calibrated, precision is a good indicator of accuracy.

Suppose we set up an experiment in which three students measure the temperature of boiling water using different thermometers. Each student takes a measurement, waits one minute, takes another measurement, and so on until each thermometer has recorded four temperature values. The results shown in Table 2.1:

TABLE 2.1	Student One	Student Two	Student Three
Trial One	99.4ºC	97.3ºC	100.1ºC
Trial Two	102.3ºC	97.4ºC	100.0ºC
Trial Three	101.7ºC	97.4ºC	99.9ºC
Trial Four	101.3ºC	97.3ºC	100.0ºC

From this information we can determine which experimental data is best – both repeatable and reliable. The true boiling point of water is 100.0ºC, so accurate data would give values at or very near that number. Precise measurements are repeatable, as indicated by the results from students Two and Three. So let us analyze the data shown here: The first student's data is neither accurate nor precise. This could be due to errors made by the experimenter, or perhaps the thermometer is faulty in some way. The second student's data is precise, but is not accurate. This is likely due to an incorrectly calibrated thermometer that gives consistently low temperature readings. The third student recorded temperatures close to the true value (accurate) and reported similar temperatures repeatedly (precise).

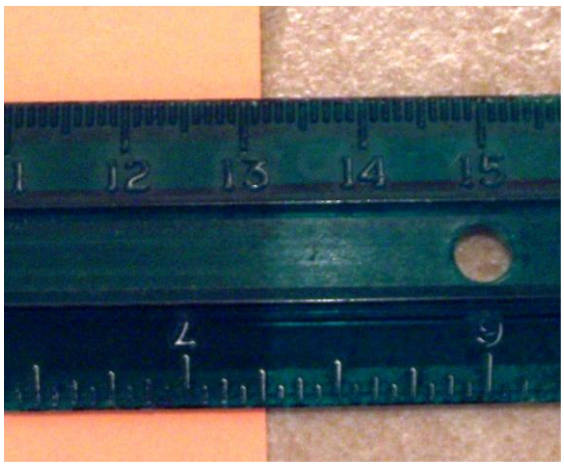
As mentioned earlier, the second use of the term precision refers to the "fineness" of a measurement. To illustrate this concept, consider a beaker, a graduated cylinder, and a buret each filled with the same quantity of water:



As we look at the markings on the beaker, we see that the smallest increments are 25 mL. When reading a measurement, we can always make a "guess" as to the very last digit in the number. We can estimate one decimal place smaller than the smallest increment on the instrument. In this case, the smallest increment is 25, so the best approximation we can make is to the tens place. Perhaps a good approximation for this volume would be 30 mL. It is impossible to know the volume in the beaker more *precisely* than this. For this reason, beakers are rarely used for measuring volume. In the graduated cylinder, the smallest increment between markings is one milliliter, so we can approximate one decimal place smaller than one, the one-tenths place. As a general rule, if the

markings are too close together to approximate ten smaller markings, then estimate by halves. Since the meniscus of the water in the cylinder rests between 32 and 33 mL, then we could report this volume as 32.5 mL. Different graduated cylinders have different increments, but these are the best tool we have for making fairly precise measurements quickily. In the buret to the left, the smallest marking is 0.1 mL, so we can estimate the one-hundredths place. Again the markings are too close together to divide each increment by ten, so we will do half. The meniscus lies between 32.4 and 32.5 mL, so the volume is 32.45 mL. Burets and volumetric pipets are used when it is important to know the volume as precisely as possible.

Keep in mind that the estimation of the last digit is an acceptable practice for every piece of laboratory equipment, including beam-balances and rulers. See if you can correctly read the measurements in the photos below:



2.2 SIGNIFICANT FIGURES

Since the instrument itself determines the degree of precision, it follows that all measurements are understood to contain a certain amount of error, either human or instrumental. It is impossible to measure exactly five and one-half grams of table salt, for example. We may be able to measure 5.5 grams, 5.50 grams, or even 5.500 grams, but none of these measurements represents *exactly* five grams. The first is accurate to only the tenths place – somewhere between 5.4 and 5.6 grams. The last is much more precise – between 5.499 and 5.501 grams – but it is still not exact. The final reported digit of any measurement is *uncertain*. It is an approximation, just like in the previous section where we estimated between graduations on an instrument. A measurement contains two kinds of *significant figures*, those that are known and one final approximated digit. The more significant figures a number has, the more precise the measurement is.

Determining the Number of Significant Figures:

This is actually a fairly simple process, as long as the rules for significant figures are understood. These rules are:

1) **All nonzero digits are considered significant**
   ex: The measurement 546 grams has three sig-figs
2) **Zeros between significant figures are considered significant**
   ex: The measurement 3.502 seconds has four sig-figs
3) **Trailing zeros after a decimal are considered significant**
   ex: The measurement 22.50 pounds has four sig-figs
4) **Zeros that serve as placeholders are not significant**
   ex: The measurements 100 mL and 0.05 ft each have only one sig-fig

There are some numbers, however, that are exact, having an infinite number of significant figures. These numbers are counting numbers (exactly 12 eggs in one dozen), defined constants (exactly 2.54 centimeters in one inch), and metric definitions (one kilometer is exactly 1000 meters).

Calculations Involving Measurements:

Significant figures indicate the degree of precision in any measurement, and when those measurements are used in calculations, the precision remains with it. This means a set of data can only be as precise as its least precise value. Let's take a simple calculation as an example: density. Density is defined as the amount of mass an object has divided by its volume. Suppose we have a 5.50 gram wooden block with a volume of 7.3 cubic centimeters. If we divide the mass by the volume, we get the value 0.753424658 g/cm^3. How can the density be known to a greater precision than either the mass or the volume? Quite simply, it can't. The mass is known to three significant figures, and the volume to two significant figures. Therefore the density can only be known to the same degree as the least precise measurement used to calculate it, two. The density of the block is 0.75 g/cm^3.

While accounting for the propagation of uncertainties is sometimes a complicated task, the following rules for significant figures generally suffice:

- **When multiplying or dividing measurements, the result must be given with the same number of significant figures as the measurement with the fewest significant figures.**
- **When adding or subtracting measurements, the result will have the same number of decimal places as the measurement with the fewest decimal places.**

Table 2. 2

From	To	Distance
McKinney	Allen	8.32 miles
Plano	Dallas	19 miles
Melissa	McKinney	7.3 miles
Allen	Plano	7.31 miles

To illustrate the second rule, suppose four students were sent out to measure distances between cities. There results are in the table to the left. If we wanted to know the distance from Melissa to Allen we would simply need to add the distance from Melissa to McKinney and the distance from McKinney to Allen. The sum of 7.3 miles and 8.32 miles is 15.62 miles, but since the least precise measurement is known only to the nearest tenth of a mile, we can only report the distance as 15.6 miles. Also, the distance from Melissa to Dallas could be obtained by adding all four measurements together, getting a total of 41.93 miles. The correct sum, however, would be 42 miles since the distance from Plano to Dallas is only known to the nearest mile. It is important to remember that rounding to the correct number of significant figures should only be done after every calculation has been performed to reduce the error introduced by excessive rounding.

2.3 LE SYSTÈME INTERNATIONAL

In any quantitative measurement, the number is only part of the value. All measurements have both a quantity and a unit, each as important as the other. If you ask someone how much pizza they ate last night and they respond "6," what does that mean? Six pizzas? Slices? Pounds? Kilograms? Ounces? Cups? Dozen? The numerical value is essentially useless without knowing the unit that is associated with it. And just as important is to know the general scale of each unit. Is six ounces of pizza a reasonable amount? Which is larger – six pounds of pizza or six kilograms of pizza?

The problem lies in the fact that there are dozens of systems of measurement used around the world, and the basis for those measurements have changed throughout history. In ancient India, length was measured with units such as the dhanus, krosa, and jojana. The ancient Mesopotamian system was based on the cubit, the distance from the elbow to the middle finger. Unfortunately this was different for every person. The Greeks and Romans inherited the foot from the Egyptians. The pace, equivalent to five Roman feet, was used to determine the mile (1000 paces). Queen Elizabeth the First of England changed the mile to 5280 feet in order to be equivalent to 8 furlongs, each furlong being 40 rods of 5.5 yards each. The meter was once defined as one ten-millionth the distance from the North Pole to the equator along a meridian. Mass was originally measured by comparing the weight of an object to the weight of a grain of wheat. Later masses were compared to standard stones with units such as the mina, sheckel, talent, and eventually the pound. And don't forget the ounce, carat, long ton, short ton, hundredweight, gram, and countless others.

When scientific discovery really began to take off and become an endeavor that had little to do with geopolitical borders, it became obvious that a standardized system of measurement was essential. In 1875 seventeen nations signed the *Convention du Mètre* and established the General Conference on

Weights and Measures (CGPM) and the International Committee for Weights and Measures (CIPM) to establish and oversee measurement standards around the world. The United States governmental body the National Institute of Standards and Technology (NIST) has a permanent seat on the CIPM which continuously updates standards of measurement.

In 1960, the 11th General Conference laid down the standards for a new system of units that would become the Système International d'Unitès, commonly referred to as SI. The SI established six base units (the seventh was added in 1971) from which all other measurements could be derived, as well as prefixes and rules for writing abbreviations. These units pertinent to chemistry are listed in Table 2.3, as well as the standard on which they are based. All seven base units are based on a universal physical constant except the kilogram, which is also the only base unit with a prefix.

Table 2. 3

The International Prototype Kilogram. There are also six official copies stored in the same vault and additional copies around the world.

While every country in the world has officially adopted the International System, a few (the United States being one of them) are reluctant to initiate change due to public resistance. Distances in inches, feet, and miles will be familiar to you since they are accepted in U.S. culture, but the scientific community and most of the world rarely use these units. The same is true for the Fahrenheit temperature scale and masses in pounds and ounces. To help you get a better understanding of the International System of Units, Table 2.4 contains some conversions between the International System and units commonly used in the United States:

Unit	Abbrev.	Quantity	Standard
Meter	m	Length (l)	The distance traveled by light in the time interval of 1/299792458 of a second
Second	s	Time (t)	The duration of 9192631770 periods of radiation corresponding to the transition between the two hyperfine levels of the ground state of the Cs-133 atom
Kilogram	kg	Mass (*m*)	The mass of the international prototype stored in a vault in Sèvres, France. The cylinder is a platinum/iridium alloy stored beneath three bell jars to limit exposure to the atmosphere.
Kelvin	K	Temperature (T)	The fraction 1/273.16 the temperature at the triple point of water. To convert degrees Celsius to Kelvin, simply add 273.15. Zero Kelvin is Absolute Zero
Mole	mol	Amount	The amount of substance in which there are as many elementary entities (atoms, molecules, ions, particles, etc) as there are in exactly 12 grams of carbon-12. This number is the constant called Avogadro's Number (N_A) and is equal to 6.022×10^{23}

Table 2. 4

LENGTH					
1 mile	=	1.61 kilometers	1 kilometer	=	0.621 miles
1 yard	=	0.914 meters	1 meter	=	1.09 yards
1 foot	=	30.5 centimeters	1 centimeter	=	0.394 inches
1 inch	=	2.54 centimeters*			
MASS					
1 carat	=	200 milligrams*	1 gram	=	0.0352 ounces
1 ounce	=	28.4 grams	1 kilogram	=	2.20 pounds
1 pound	=	0.454 kilograms			
VOLUME					
1 Tablespoon	=	14.8 milliliters	1 milliliter	=	1 cubic centimeter*
1 cup	=	237 milliliters	1 liter	=	1 cubic decimeter*
1 quart	=	0.946 liters	1 liter	=	0.264 gallons
1 gallon	=	3.79 liters			
TEMPERATURE					
0ºC = 32ºF = 273.15 K*					
100ºC = 212ºF = 373.15 K*					
ºC = (9/5)(ºF - 32)					
K = ºC + 273.15					
* denotes an exact number					

You may have noticed that some of the units in the table to the right include prefixes such as milli-, centi-, and deci-. These occur because there are many occasions on which using the standard base unit in the SI is impractical due to the magnitude of the measurement. For example, you wouldn't want to measure the mass of an electron in kilograms (approximately 0.000 000 000 000 000 000 000 000 000 0091 kg) or the distance from the sun to Jupiter in meters (778 330 000 000 m). The numbers would be either very large or very small and extremely cumbersome. There are two ways of solving this problem – using scientific notation, or altering the magnitude of the unit by changing the prefix.

Scientific Notation:

Sometimes called exponential notation, this method of writing numbers makes use of the fact that our numbering system is a *decimal* system, meaning groups of ten. If you have the number nine and add one, we don't put "10" in the ones place, but rather "1" in the tens place and start over with "0" in the ones place. The same is true for the tens place – when we reach ten "tens," we create the hundreds place and start over with zero again in the tens place. It works the same way on the other side of the decimal with the tenths and hundredths place. Other numerical systems, such as binary and hexadecimal (based on 2 and 16, respectively) are more difficult to work with because we are used to working in tens. Scientific notation simply expresses any value as a number between 1 and 10 multiplied by a factor of ten. For example, 1 000 000 would be expressed as 1×10^{6} and 0.000 000 01 would be 1×10^{-7}. When expressing a number in scientific notation, the number of significant figures should be the same as in the original number. To illustrate this, the mass of an electron from above would be 9.1×10^{-31} kg and the distance to Jupiter would be 7.7833×10^{11} m. The power of ten can be

obtained by counting the number of places the decimal must be moved to get a number between one and ten, with positive exponents representing values greater than 1 and negative exponents representing values less than 1.

Prefixes:

The alternative to scientific notation for measurements is to use a prefix that serves the same purpose as the power of ten. For example, kilo- means 1000 (10^3), so 8.3 kilometers is simply 8.3×10^3 m, or 8300 meters. Additional prefixes are in Table 2.5. Using these prefixes, we could express the distance from the sun to Jupiter in, say, terameters (Tm) or gigameters (Gm). Unfortunately we don't have a prefix small enough to represent the mass of an electron so we'll have to stick to scientific notation for that one.

Table 2.5	Prefixes		
tera-	T	Trillion	10^{12}
giga-	G	Billion	10^{9}
mega-	M	Million	10^{6}
kilo-	k	Thousand	10^{3}
hecto-	h	Hundred	10^{2}
deca-	da	Ten	10^{1}
---			
deci-	d	Tenth	10^{-1}
centi-	c	Hundredth	10^{-2}
milli-	m	Thousandth	10^{-3}
micro-	μ	Millionth	10^{-6}
nano-	n	Billionth	10^{-9}
pico-	p	Trillionth	10^{-12}

2.4 DIMENSIONAL ANALYSIS

It is useful to have a quick and straightforward way to express a measurement with different units without having to measure again and again. Once a measurement has been taken, it can be converted to any system using a process called dimensional analysis. Any unit can be converted to any other unit as long as there exists a known relationship between the two. Let's assume we measured the length of a sheet of paper and found it to be 8.5 inches. What is this value in centimeters? We can solve this or any other computational problem using dimensional analysis.

We know that a relationship exists between the unit we have (inches) and the unit we want (centimeters) – one inch is equal to exactly 2.54 centimeters. If we create a ratio, the value is fundamentally equal to one since both 1 in and 2.54 cm are the same quantity.

$$\frac{1\,\text{in}}{2.54\,\text{cm}} = 1 \qquad \frac{2.54\,\text{cm}}{1\,\text{in}} = 1$$

Multiplying the original measurement by one of these ratios will not change the physical quantity, only the units with which it is being expressed. So how do we know which ratio to use? Just like in algebraic equations, when a value or unit appears in both the numerator and the denominator, it can be cancelled out. Since we want to remove the unit "inches"

Table 2.6: Orders of Magnitude

		METERS	SECONDS	GRAMS
10^{12}	**TERA-**	**Sun to Jupiter**	**32 millennia**	**The Great Pyramid**
10^{11}		*Earth to Mars*	*Time passed since David became king of Israel*	*Fully loaded supertanker*
10^{10}		*Total length of all US roads*	*1686 to present*	*The Titanic*
10^{9}	**GIGA-**	**3x distance to the moon**	**32 years**	**The Space Shuttle**
10^{8}		*2.5x Circumference of Earth*	*Length of war with Japan 1941-1945*	*Blue Whale*
10^{7}		*Diameter of Earth*	*4 months*	*Elephant*
10^{6}	**MEGA-**	**Dallas to St. Louis**	**11 days 14 hrs**	**Passenger Car**
10^{5}		*Dallas to Oklahoma border*	*1.2 days*	*Adult male*
10^{4}		*Custer Rd. to Dallas North Tollway*	*Length of typical blockbuster movie*	*Dog*
10^{3}	**KILO-**	**5-6 city blocks**	**17 minutes**	**1 Liter soft drink**
10^{2}	*HECTO-*	*Height of the Statue of Liberty and Pedestal*	*Duration of the "Minute Waltz"*	*Human kidney*
10^{1}	*DECA-*	*Length of Killer Whale*	*World Record time in the 100m dash*	*Lethal dose of caffeine for an adult*
10^{0}		**Length of arm**	**Time for light to travel from Earth to the moon**	**Paperclip**
10^{-1}	*DECI-*	*Cell phone*	*Blink of an eye*	*One-half carat diamond*
10^{-2}	*CENTI-*	*Fingernail*	*Camera shutter speed*	*1.5 teaspoons of air*
10^{-3}	**MILLI-**	**Grain of sand**	**Duration of camera flash**	**Mosquito**
10^{-4}		*Thickness of hair*	*Sampling interval for telephone audio*	*A "Hit" of LSD*
10^{-5}		*Cell diameter*	*Sampling interval for CD audio*	*Small grain of sand*
10^{-6}	**MICRO-**	**Cellular organelles**	**Cycle time for typical AM radio signal**	**Lethal dose of botulin toxin (Botox)**
10^{-7}		*Wavelength of visible light*		*Smallest detectable amount of marijuana per mL of urine*
10^{-8}		*DNA supercoils*		*Amount of DNA needed for genetic fingerprinting*
10^{-9}	**NANO-**	**Width of DNA**	**Time for light to travel one foot**	**Human cell**
10^{-10}		*Radius of atom*		*Amount of dioxin in one hamburger*
10^{-11}		*Wavelength of gamma rays*		*Mass of 10 bacteria*
10^{-12}	**PICO-**	**100x diameter of nucleus**	**Time for light to travel 0.3 cm**	**2.5 billion Uranium nuclei**

**All measurements are approximations and are meant to provide a general size comparison*

and replace it with "centimeters," we will choose the second ratio so that inches cancels out. Then we simply multiply and round the answer to the correct number of significant figures.

$$8.5\,\cancel{\text{in}} \times \frac{2.54\,\text{cm}}{1\,\cancel{\text{in}}} = 21.59\,\text{cm} = 22\,\text{cm}$$

This method can be used with practically every problem in chemistry. It provides a standard approach to problem solving and is also easy to double check for accuracy. As long as the conversion factors are correct and the units cancel out, we can be certain of the outcome. Dimensional analysis is not limited to one-step conversions, nor is it limited to units within the same measurement scheme. Here are some additional examples to emphasize the utility of this method.

Example 1: Express the quantity 3.2 kg in milligrams.

We don't immediately know the relationship between kg and mg, but we do know that 1 kg is 1000 g and 1 g is 1000 mg. So we will first convert from kg to g, then from g to mg:

$$3.2\,\cancel{\text{kg}} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{1000\,\text{mg}}{1\,\cancel{\text{g}}} = 3\,200\,000\,\text{mg} = 3.2 \times 10^6\ \text{mg}$$

Example 2: A ream of paper weighs 5.0 kg and costs $2.50. What is the price of one gram of paper?

If we listed out the relationships we know, we would find that 1 ream = 5.0 kg, 1 ream = $2.50, and 1 kg = 1000 g. We will convert from grams to kilograms, then from kilograms to reams, then from reams to dollars:

$$1\,\cancel{\text{g}} \times \frac{1\,\cancel{\text{kg}}}{1000\,\cancel{\text{g}}} \times \frac{1\,\cancel{\text{ream}}}{1\,\cancel{\text{kg}}} \times \frac{\$2.50}{1\,\cancel{\text{ream}}} = \$0.0025$$

Example 3: How many milliseconds in 25 centuries?

There are 1000 ms in one second, 3600 s in one hour, 24 hours in a day, 365 days in a year, and 100 years per century:

$$25\,\cancel{\text{cent.}} \times \frac{100\,\cancel{\text{yr}}}{1\,\cancel{\text{cent.}}} \times \frac{365\,\cancel{\text{d}}}{1\,\cancel{\text{yr}}} \times \frac{24\,\cancel{\text{h}}}{1\,\cancel{\text{d}}} \times \frac{3600\,\cancel{\text{s}}}{1\,\cancel{\text{h}}} \times \frac{1000\,\text{ms}}{1\,\cancel{\text{s}}} = 7.884 \times 10^{13}\ \text{ms} = 7.9 \times 10^{13}\ \text{ms}$$

Scientific Notation & Significant Digits

1. Convert each of the following to proper scientific notation, retaining the same number of significant figures. Note the number of significant figures for each measurement. The first has been done for you as an example.

a) 3427 **3.427×10^3 (4 SF)**

b) 0.0056

c) 12345

d) 172

e) 0.000984

f) 0.502

g) 310.0×10^2

h) 0.011×10^4

i) 107.2

j) 0.000455

k) 2205.2

l) 30.0×10^{-2}

m) 0.982×10^{-3}

n) 0.0473

o) 650502

p) 3.03×10^{-1}

2. Convert each of the following to decimal form, retaining the same number of significant figures. Note the number of significant figures for each measurement. The first has been done for you as an example.

a) 1.56×10^4 **15,600 (3 SF)**

b) 0.56×10^{-2}

c) 3.69×10^{-2}

d) 736.9×10^5

e) 0.00259×10^5

f) 13.69×10^{-2}

3. Calculate the following. Give the answer in correct scientific notation with the correct number of significant figures.

a) $3.95 \times 10^2 / 1.5 \times 10^6 =$

b) $4.44 \times 10^7 / 2.25 \times 10^5 =$

c) $1.05 \times 10^{-26} / 4.2 \times 10^{56} =$

d) $(6.022 \times 10^{23})(3.011 \times 10^{-56}) =$

e) $(3.5 \times 10^2)(6.45 \times 10^1) =$

f) $(4.50 \times 10^{-12})(3.67 \times 10^{-12}) =$

4. Round each of the following to 3 significant figures.

a) 77.0653

b) 6300278.2

c) 0.00023350

d) 10.2030

e) 2.895×10^{21}

f) 692

g) 0.51

h) 3000

5. Calculate the following. Report each answer with the appropriate number of significant figures.

a) 97.381 + 4.2502 + 0.99195 =

b) 171.5 + 72.91 – 8.23 =

c) 1.00914 + 0.87104 + 1.2012 =

d) 21.901 – 13.21 – 4.0215 =

e) 4.184 * 100.62 * (25.27 – 24.16) =

f) 8.27 * (4.987 – 4.962) =

g) [(8.925 – 8.904) / 8.925] x 100 =

This is an example of a %-error calculation. The number "100" can be considered to be an exact number.

h) (9.5 + 4.1 + 2.8 + 3.175) / 4 =

Assume this operation is taking an average; thus the 4 in the denominator is an exact number.

i) $6.6262 \times 10^{-34} * 2.998 \times 10^{8} / 2.54 \times 10^{-9}$ =

j) $(1.00866 - 1.00728) / 6.02205 \times 10^{23}$ =

k) (9.04 – 8.23 + 21.954 + 81.0) / 3.1416 =

Dimensional Analysis #1

Show ALL work, units and proper steps on your own paper.

1. How many seconds are in 815 minutes?
2. How many centimeters are in 5.00 kilometers?
3. How many years are in 528 days?
4. How many dozens of eggs are there in 3,480 eggs?
5. How many miles are in 22,900 yards?
6. How many kilograms are in 9.00 micrograms?
7. How many yards are in 648 inches?
8. How many milliliters are in 6.30 deciliters?
9. How long, in kilometers, is a 12.0-inch ruler?
10. What is the volume, in cubic feet, of a box that measures 30.0 x 40.0 x 20.0 inches?
11. A peck (pk) is four bushels and a bushel is two gallons. Two 8oz cups make up a pint, two pints make a quart, and a gallon is four quarts. How many ounces in 0.250 pk?
12. The world record for the 800m run is 1.68 minutes. How long is this in nanoseconds?
13. How wide, in micrometers, is the Red Sea if its widest point is 306 km?

Dimensional Analysis #2

Show ALL work, units and proper steps on your own paper.

1. What is the cost of 12 onions if 3 onions weigh 1.50 lbs. and the price of onions is $0.80 per lb.?
2. How many hours will it take to drive to Los Angeles from San Francisco if an average speed of 52.0 mile/hour is maintained? The distance between the two cities is 405 miles.
3. What is the cost to drive from San Francisco to Los Angeles is the cost of the gasoline is $2.43/gal and the car has a fuel efficiency of 26.5 mpg? The distance between cities is 405 miles.
4. The price of a ream of paper is $2.50. There are exactly 500 sheets of paper in a ream. If a sheet of paper weighs 0.0690 grams, what is the price per gram of paper?
5. Sue wants to buy apples. They are on sale for $0.99 per lb. How much will she spend on 3 apples? It takes exactly 8 apples to make one pound.
6. How many oranges are in a crate if the price of a crate of oranges is $4.60 and the price of oranges is $0.70/lb.? On average there are 3.50 oranges per pound.
7. What is the cost of 15 potatoes if 3 potatoes weigh 3.25 lbs and the price is $0.88 per lb.?
8. How many 50.0 foot cars are in a long freight train traveling at 42.0 mph if it takes the entire train 60.0 seconds to pass a crossing?
9. The price of a bag of dog bones $6.85. There are exactly 12 bones in each bag. If one bone weighs 23.0 grams, what is the price per gram?
10. How many peaches are in a crate if the price of a crate of peaches is $5.30 and the price of peaches is $0.85/lb? On average there are 5.25 peaches per pound.

More Dimensional Analysis Practice Problems

Use dimensional analysis for all of the problems below. Show your work! Report your answer with the correct number of significant figures.

1. When 1.00g of gasoline burns in a car's engine, the amount of energy given off is approximately 1.03×10^4 calories. Express this quantity in joules. (1 cal = 4.184 J)

2. The pressure reading from a barometer is 742 mmHg. Express this reading in kilopascals. (760mmHg = 1.013×10^5 Pa)

3. How many megayears is equivalent to 6.02×10^{23} nanoseconds (ns)?

4. The average student is in class 330 min/day.
    a. How many hours/day (school days only) is the average student in class?

    b. How many seconds is the average student in class per week?

5. Approximately how many pounds do you weigh? _____lbs
    a. The human body is approximately 60% water by mass. Using that conversion factor, how many pounds of water are in your body? (This is only a rough estimate, since it also depends on body fat)

6. French cooks usually weigh ingredients. A French recipe uses 225g of granulated sugar. How many cups are needed if there are exactly 2 cups of sugar per pound? Round to the nearest cup.

7. Convert 9.3 $g \cdot cm^3/s^2$ to $kg \cdot m^3/hr^2$.

8. A car accelerates at 15 mi/hr·s (miles per hour per second). Write that acceleration in m/s^2. (1 mile = 1.609 km)

9. Because your 18 year-old friend never learned dimensional analysis, he started working at a fast food restaurant wrapping hamburgers. Every 3 hours he wraps 525 burgers. He works 8 hours per day, 5 days a week. He gets paid every 2 weeks with a salary of $440.34 (after taxes).
    a. Approximately how many hamburgers will he have to wrap to make his first million dollars?

    b. How much time will it take to wrap all those hamburgers? Express the time in years. Assume he works with the same efficiency every day and takes a two-week vacation every October.

    c. How old will he be when he makes his first million?

10. A child is entered into the hospital after ingesting 12 aspirin tablets. The Merck Index indicates that renal (kidney) failure can occur if as little as 3 grams in ingested, and may be fatal in as much as 10 grams is eaten. If each aspirin tablets contains 300mg of aspirin, how much aspirin (in grams) has the child ingested?

11. A patient in the hospital is given an intravenous fluid that must deliver 1000 cc (cubic centimeters) of a dextrose solution over 8 hours. The IV fluid tubing delivers 15 drops per cubic centimeter. What is the drop rate (in drops/min) that must be administered to the patient?

12. A medical doctor gives the order to administer dopamine at a rate of 3.0μg/kg·min (micrograms per kilogram per minute). The dopamine is supplied as a mixture of 400.mg dopamine in a 250.mL solution. The patient weighs 73kg. What is the infusion rate of dopamine solution into her body in mL/hr?

13. Analysis of an air sample reveals that it contains 3.5×10^{-6} g/L of carbon monoxide. Express this concentration in lb/ft^3. (1.00lb is 454g; 1in = 2.54cm)

14. A website states that the average body density is 0.001 kg/cm^3. Express this value in pounds per cubic foot.

15. The bathtub in the residential suite of the White House had to be enlarged for William H. Taft, who weighed 370.lbs. The bathtub is four feet wide, six feet long, and three feet high.
    a. How many gallons of water are needed to fill the WHT Memorial Bathtub to the rim? (1 gal = 231 in^3).

    b. That was a silly question. Why would WHT get in a tub completely filled with water? (I guess HE wouldn't have to clean it up…) You calculated average body density in question #14. How many gallons of water are needed to fill the tub so that when WHT gets in the tub, the water reaches the rim but does not flow over?

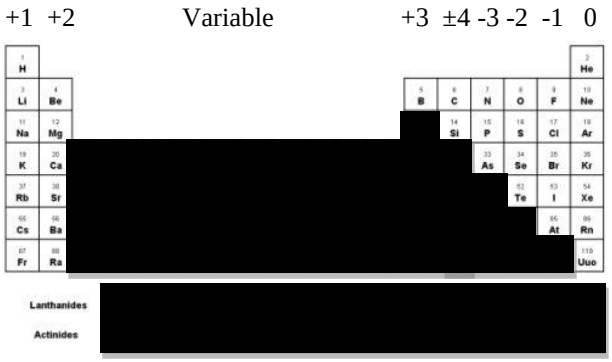
16. Albumin is a protein found in blood. If the concentration of this protein is 600. μmol/L, and its molecular mass is 68,500 g/mol, what is the concentration of this protein in mg/cm^3?

17. Remember it is never too late for the hamburger people to come get you!

Naming and writing formulas for IONIC COMPOUNDS

- Ionic compounds are made of a cation (+) and an anion (-). The cation can be a metal ion or a polyatomic ion. The anion can be a nonmetal ion or a polyatomic ion.
- They have high melting points, and the smallest unit of an ionic compound is called a formula unit.

For a neutral compound, the charges of the two ions MUST sum to zero. This allows us to determine how many of each ion are required to construct one formula unit. Many times the charge of the ion can be determined based on the element's position on the periodic table.



To name an ionic compound, follow the rules below:

1. Write the name of the cation
    a. If a polyatomic ion, write the polyatomic ion's name
    b. If a metal ion, write the name of the metal
    c. If the metal is NOT in group 1, group 2, Al^{3+}, Zn^{2+}, or Ag^{+}, write the Roman numeral designation for the charge of the ion*
2. Write the name of the anion
    a. If a polyatomic ion, write the polyatomic ion's name
    b. If a nonmetal ion, write the name of the nonmetal with the suffix "-ide."

*The system of using Roman numerals is called the Stock System and is the currently accepted nomenclature system. In the old system, metal ions with multiple possible oxidation numbers were given different names. For example, the iron(II) ion, Fe^{2+}, is called "ferrous," while the iron(III) ion, Fe^{3+}, is called "ferric." In each case, the smaller oxidation number corresponds to the "-ous" suffix and the higher oxidation number corresponds to the "-ic" suffix. Please know the following names that are still commonly used. Note that the mercurous ion is actually a polyatomic ion.

Iron(II) = Fe^{2+} = ferrous	Iron(III) = Fe^{3+} = ferric
Lead(II) = Pb^{2+} = plumbous	Lead(IV) = Pb^{4+} = plumbic
Tin(II) = Sn^{2+} = stannous	Tin(IV) = Sn^{4+} = stannic
Copper(I) = Cu^{+} = cuprous	Copper(II) = Cu^{2+} = cupric
Mercury(I) = Hg_2^{2+} = mercurous	Mercury(II) = Hg^{2+} = mercuric

Examples:

Calcium phosphide = Ca_3P_2	Lead(IV) sulfate = plumbic sulfate = $Pb(SO_4)_2$
Aluminum fluoride = AlF_3	Copper(I) oxalate = cuprous oxalate = $Cu_2C_2O_4$

Naming and writing formulas for COVALENT COMPOUNDS

- Covalent compounds are made of two nonmetals.
- They have low boiling points, and the smallest unit of a covalent compound is called a molecule.

Prefixes are used to indicate how many of each element exist in a molecule.

PREFIXES

1 = mono-	4 = tetra-	7 = hepta-	10 = deca-
2 = di-	5 = penta-	8 = octa-	
3 = tri-	6 = hexa-	9 = nona-	

To name a covalent compound, follow the rules below:

1. If there is only one of the first element, write its name
2. If there is more than one of the first element put a prefix to indicate how many, followed by the name of the element
3. Use a prefix to indicate the number of atoms of the second element, followed by the name of the element (ending changed to –ide)

Examples:

CO = carbon monoxide

N_2O_5 = dinitrogen pentoxide

P_4O_{10} = tetraphosphorus decaoxide

OF_2 = oxygen difluoride

Naming and writing formulas for ACIDS

- Binary acids are made of a hydrogen cation and a nonmetal anion.
- Oxyacids are made of a hydrogen cation and a polyatomic anion.

Like ionic compounds, the total charge of the compound is zero. This will determine how many of each ion is present in the formula.

To name a binary acid, follow the rules below:

1. Use the prefix "hydro"
2. Use the nonmetal's name with the suffix "-ic acid"

To name an oxyacid, follow the rules below:

1. Do NOT use the prefix "hydro"
2. Write the name of the polyatomic ion
    a. If the polyatomic ion ends with "-ate," change the suffix to "-ic acid"
    b. If the polyatomic ion ends with "-ite," change the suffix to "-ous acid"

Examples:

HBr = hydrobromic acid

HNO_2 = nitrous acid

$HC_2H_3O_2$ = acetic acid

H_2S = hydrosulfuric acid

Common Polyatomic Ions

+1 charge

ammonium	NH_4^+

-1 charge

hydroxide	OH^-	permanganate	MnO_4^-
nitrate	NO_3^-	chlorate	ClO_3^-
nitrite	NO_2^-	chlorite	ClO_2^-
acetate	$C_2H_3O_2^-$	perchlorate	ClO_4^-
bisulfate	HSO_4^-	hypochlorite	ClO^-
bicarbonate	HCO_3^-	cyanide	CN^-
bisulfite	HSO_3^-	thiocyanate	SCN^-

-2 charge

sulfate	SO_4^{2-}	peroxide	O_2^{2-}
sulfite	SO_3^{2-}	chromate	CrO_4^{2-}
carbonate	CO_3^{2-}	dichromate	$Cr_2O_7^{2-}$
oxalate	$C_2O_4^{2-}$		

-3 charge

phosphate	PO_4^{3-}
phosphite	PO_3^{3-}

Solubility and Strong Acids-Bases

If a compound is **NOT** listed on the solubility table below, then it must be **INSOLUBLE**. Example, $CaCO_3$ is insoluble, since Ca^{2+} and CO_3^{2-} are not listed as soluble.

SOLUBLE COMPOUNDS	EXCEPTIONS
All Group 1 salts (Li^+, Na^+, K^+, Cs^+, Rb^+)	NONE
All salts of NH_4^+	NONE
All salts of NO_3^-, ClO_3^-, ClO_4^-, and CH_3COO^-	NONE
All salts with Cl^-, I^-, Br^-	Ag^+, Hg_2^{2+}, Pb^{2+}
All salts of F^-	Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
All salts of SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

STRONG ACIDS		STRONG BASES	
Hydrochloric	HCl	Lithium hydroxide	LiOH
Hydrobromic	HBr	Sodium hydroxide	NaOH
Hydroiodic	HI	Potassium hydroxide	KOH
Nitric	HNO_3	Rubidium hydroxide	RbOH
Sulfuric	H_2SO_4 *	Cesium hydroxide	CsOH
Perchloric	$HClO_4$	Magnesium hydroxide	$Mg(OH)_2$ **
Chloric	$HClO_3$	Calcium hydroxide	$Ca(OH)_2$ ***
		Strontium hydroxide	$Sr(OH)_2$
		Barium hydroxide	$Ba(OH)_2$

* Sulfuric acid is only a strong acid for the disassociation of the first hydrogen ion, producing H^+ and HSO_4^{1-}.

** $Mg(OH)_2$ is considered a strong base since it does ionize 100%, however, it is EXTREMELY insoluble and on the AP exam it is never written in ionic form.

*** $Ca(OH)_2$ is considered a strong base, but depending on the situation, you may or may not ionize it. If the concentration is less than 0.1M, it ionizes. If the problem says "saturated" or more than 0.1M, do **NOT** ionize $Ca(OH)_2$. Usually, you will get credit for it either way on the AP Exam.

AP Chemistry WORKSHEET: Formulas and Nomenclature

I. Name the following ions. Give both the old (ous/ic) and the Stock names for those marked with an asterisk

1. S^{2-}
2. OH^{-}
3. O_2^{2-}
4. P^{3-}
*5. Fe^{3+}
6. Zn^{2+}
7. Mg^{2+}
*8. Sn^{2+}

II. Write formulas for the following ions:

1. mercurous
2. ammonium
3. chromium (III)
4. cupric
5. potassium
6. oxide
7. phosphate
8. carbide

III. Write formulas for the following ionic compounds:

1. iron(II) oxide
2. ferrous bromide
3. lead(II) hydroxide
4. lithium hydride
5. aluminum sulfide
6. calcium nitrate
7. barium peroxide
8. sodium dichromate

IV. Name the following ionic compounds:

1. CuCl
2. $KMnO_4$
3. $Al_2(SO_4)_3$
4. $(NH_4)_2CO_3$
5. $Hg(NO_3)_2$
6. $CoBr_3$
7. K_2O_2
8. $Ni(CN)_2$

V. Write formulas for the following molecular compounds and acids:

1. sulfur dioxide
2. nitrogen monoxide
3. boron tribromide
4. sulfurous acid
5. chlorous acid
6. tetraphosphorus decoxide
7. hydrocyanic acid
8. dichlorine oxide
9. hydrogen iodide
10. phosphoric acid

VI. Name the following molecular compounds and acids:

1. S_2Cl_2
2. $HI_{(aq)}$
3. H_2S
4. NO
5. $HClO_{2(aq)}$
6. XeF_4
7. $HNO_{3(aq)}$
8. NH_3

VII. Write formulas for these substances:

1. vanadium pentachloride
2. potassium phosphide
3. manganese(III) oxide
4. diiodine heptoxide
5. magnesium phosphate
6. acetic acid
7. cuprous acetate
8. perchloric acid
9. stannous bicarbonate
10. methane

VIII. Name the following substances:

1. AsF_3
2. Cl_2O
3. NH_4ClO_4
4. $H_3PO_{4(aq)}$
5. Bi_2S_3
6. FeO
7. BaH_2
8. $Pt(NO_2)_4$
9. $HBr_{(aq)}$
10. $KHSO_4$

Lots of Ionic Naming Practice Problems

Name the following ionic compounds:

1) $NaBr$ _____

2) $Sc(OH)_3$ _____

3) $V_2(SO_4)_3$ _____

4) NH_4F _____

5) $CaCO_3$ _____

6) $NiPO_4$ _____

7) Li_2SO_3 _____

8) Zn_3P_2 _____

9) $Sr(C_2H_3O_2)_2$ _____

10) Cu_2O _____

11) Ag_3PO_4 _____

12) $YClO_3$ _____

13) SnS_2 _____

14) $Ti(CN)_4$ _____

15) $KMnO_4$ _____

16) Pb_3N_2 _____

17) $CoCO_3$ _____

18) $CdSO_3$ _____

19) $Cu(NO_2)_2$ _____

20) $Fe(HCO_3)_2$ _____

Write the formulas for the following ionic compounds:

21) lithium acetate _____

22) iron (II) phosphate _____

23) titanium (II) selenide _____

24) calcium bromide _____

25) gallium chloride _____

26) sodium hydride _____

27) beryllium hydroxide _____

28) zinc carbonate _____

29) manganese (VII) arsenide _____

30) copper (II) chlorate _____

31) cobalt (III) chromate _____

32) ammonium oxide _____

33) potassium hydroxide _____

34) lead (IV) sulfate _____

35) silver cyanide _____

36) vanadium (V) nitride _____

37) strontium acetate _____

38) molybdenum sulfate _____

39) platinum (II) sulfide _____

40) ammonium sulfate _____

Naming Covalent Compounds Worksheet

Write the formulas for the following covalent compounds:

1) antimony tribromide _____

2) hexaboron silicide _____

3) chlorine dioxide _____

4) hydrogen iodide _____

5) iodine pentafluoride _____

6) dinitrogen trioxide _____

7) ammonia _____

8) phosphorus triiodide _____

Write the names for the following covalent compounds:

9) P_4S_5_____

10) O_2 _____

11) SeF_6 _____

12) Si_2Br_6 _____

13) SCl_4 _____

14) CH_4 _____

15) B_2Si _____

16) NF_3 _____

Naming Acids and Bases

Name the following acids and bases:

1) NaOH _____

2) H_2SO_3 _____

3) H_2S _____

4) H_3PO_4 _____

5) NH_3 _____

6) HCN _____

7) $Ca(OH)_2$ _____

8) $Fe(OH)_3$ _____

9) H_3P _____

Write the formulas of the following acids and bases:

10) hydrofluoric acid _____

11) hydroselenic acid _____

12) carbonic acid _____

13) lithium hydroxide _____

14) nitrous acid _____

15) cobalt (II) hydroxide _____

16) sulfuric acid _____

17) beryllium hydroxide _____

18) hydrobromic acid _____

Nomenclature - Practice sheet

$AlCl_3$		AlF_3	
$Ca(NO_3)_2$		CuI	
SO_3		$PbCl_4$	
Na_3N		$Pb(ClO_3)_2$	
$Na_2CO_3 \cdot H_2O$		$Pb(ClO)_4$	
P_4S_3		$NaHSO_4$	
Au_2O_3		H_2CO_3	
$Ca(OH)_2$		Cl_2O	
ClO_3^-		$Sn_3(PO_4)_4$	
$Al(NO_3)_3$		LiOH	
$Fe(ClO_4)_3$		$Ba(NO_2)_2$	
H_3PO_4		$Sn(ClO)_2$	
H_2Se		$Au(C_2H_3O_2)_3$	
NO		SnF_2	
$Zn(OH)_2$		HNO_3	
KOH		Mn_2O_7	
$Al(OH)_3$		$SnHPO_4$	
Cu_2CO_3		HCl	
CaO		$HClO_3$	
$(NH_4)_2SO_3$		$(NH_4)_2CO_3$	
$CuSO_4 \cdot 3H_2O$		CaSe	
FeO		$HgCl_2$	
$AuC_2H_3O_2$		Fe_2S_3	
Ca_3N_2		H_2S	
$AlPO_4$		H_2SO_4	
NO_3^-		V_2S_5	
N_2O		$MnBr_2$	
Cl_2O_7		RbH	
HBr		HNO_2	
BrF_3		Cr_2O_3	

Lead(II) hydroxide		Stannous iodide	
Sulfite ion		Cuprous chlorite	
Bromine pentafluoride		Lead(IV) perchlorate	
Aluminum sulfate		Lead(II) chlorite	
Lithium dihydrogen phosphate		Silver chloride	
Ammonium phosphate		Magnesium phosphide	
Magnesium nitrite		Tin(IV) hydroxide	
Cupric hydroxide		Chlorine dioxide	
Potassium hypochlorite		Tin(IV) hydrogen phosphate	
Sodium bicarbonate		Tin(IV) dihydrogen phosphate	
Cadmium(II) selenide		Barium hydroxide	
Carbon monoxide		Stannic nitrite	
Iodine monofluoride		Aluminum acetate	
Nitrogen dioxide		Zinc nitrate	
Ferric fluoride		Nickel(II) carbonate	
Potassium iodide		Calcium chloride hexahydrate	
Acetic acid		Manganese(VII) oxide	
Mercuric perchlorate		Chlorous acid	
Ammonium sulfide		Hypochlorous acid	
Ammonium sulfate		Perchloric acid	
Ferric nitrate		Barium sulfate	
Stannic sulfide		Barium sulfide	
Ammonium nitrite		Ferrous sulfide	
Calcium carbonate		Potassium sulfide	
Nitrogen trioxide		Hydrosulfuric acid	
Hydrogen chloride		Potassium sulfite	
Dinitrogen tetroxide		Chromium(II) acetate	
Silver nitrate		Cupric nitrate	
Iodine heptafluoride		Potassium bisulfite	
Nitrogen triiodide		Sulfurous acid	

BALANCING EQUATIONS PRACTICE

1. Aluminum hydroxide + sodium nitrate → aluminum nitrate + sodium hydroxide

   Skeletal Equation $Al(OH)_3 + 3\ NaNO_3 \rightarrow Al(NO_3)_3 + 3\ NaOH$

   Mole Ratio 1 + 3 → 1 + 3

2. Iron metal + copper (II) sulfate → iron (II) sulfate + copper metal

   Skeletal Equation _____

   Mole Ratio _____

3. Zinc metal + oxygen gas → zinc oxide

   Skeletal Equation _____

   Mole Ratio _____

4. Sodium bicarbonate → sodium carbonate + carbon dioxide + water

   Skeletal Equation _____

   Mole Ratio _____

5. Carbon tetrahydride + oxygen gas → carbon dioxide + water

   Skeletal Equation _____

   Mole Ratio _____

6. Hydrogen + Nitrogen → nitrogen trihydride

   Skeletal Equation _____

   Mole Ratio _____

7. Mercury (II) chloride + sodium metal → sodium chloride + mercury

   Skeletal Equation _____

   Mole Ratio _____

8. Ammonium bicarbonate → nitrogen trihydride + water + carbon dioxide

   Skeletal Equation _____

   Mole Ratio _____

9. Zinc sulfate + ammonium sulfide → ammonium sulfate + zinc sulfide

   Skeletal Equation _____

   Mole Ratio _____

10. Lead (II) nitrate → lead (II) oxide + nitrogen dioxide + oxygen gas

    Skeletal Equation _____

    Mole Ratio _____

11. Mercury (I) oxide + oxygen → mercury (II) oxide

Skeletal Equation _____

Mole Ratio _____

12. Aluminum oxide + carbon + chlorine → carbon monoxide + aluminum chloride

Skeletal Equation _____

Mole Ratio _____

13. Tetracarbon decahydride + oxygen → carbon dioxide + dihydrogen monoxide

Skeletal Equation _____

Mole Ratio _____

14. Barium carbonate + carbon + water → carbon monoxide + barium hydroxide

Skeletal Equation _____

Mole Ratio _____

15. Aluminum metal + water → aluminum hydroxide + hydrogen gas

Skeletal Equation _____

Mole Ratio _____

16. Zinc + chromium (III) chloride → chromium (II) chloride + zinc chloride

Skeletal Equation _____

Mole Ratio _____

17. Potassium phosphate + magnesium sulfate → magnesium phosphate + potassium sulfate

Skeletal Equation _____

Mole Ratio _____

18. Tungsten metal + tin (IV) nitrate → tungsten (II) nitrate + tin

Skeletal Equation _____

Mole Ratio _____

19. Lead (II) nitrite + potassium sulfide → lead (II) sulfide + potassium nitrite

Skeletal Equation _____

Mole Ratio _____

20. Nitrogen trihydride + oxygen gas → nitrogen gas + dihydrogen monoxide

Skeletal Equation _____

Mole Ratio _____

Balancing Equations #1

Balance the following chemical equations.

1.	___	Fe	+ ___	H_2S0_4	--> ___	$Fe_2(SO_4)_3$	+ ___	H_2
2.	___	C_2H_6	+ ___	O_2	--> ___	H_2O	+ ___	CO_2
3.	___	KOH	+ ___	H_3PO_4	--> ___	K_3PO_4	+ ___	H_2O
4.	___	SnO_2	+ ___	H_2	--> ___	Sn	+ ___	H_2O
5.	___	NH_3	+ ___	O_2	--> ___	NO	+ ___	H_2O
6.	___	KNO_3	+ ___	H_2CO_3	--> ___	K_2CO_3	+ ___	HNO_3
7.	___	B_2Br_6	+ ___	HNO_3	--> ___	$B(NO_3)_3$	+ ___	HBr
8.	___	BF_3	+ ___	Li_2SO_3	--> ___	$B_2(SO_3)_3$	+ ___	LiF
9.	___	$(NH_4)_3PO_4$	+ ___	$Pb(NO_3)_4$	--> ___	$Pb_3(PO_4)_4$	+ ___	NH_4NO_3
10.	___	$SeCl_6$	+ ___	O_2	--> ___	SeO_2	+ ___	Cl_2

Balancing Equations #2

Balance the following chemical equations.

1. ___ $SiCl_4(l)$ + ___ $H_2O(l)$ --> ___ $SiO_2(s)$ + ___ $HCl(aq)$

2. ___ As + ___ $NaOH$ --> ___ Na_3AsO_3 + ___ H_2

3. ___ Au_2S_3 + ___ H_2 --> ___ Au + ___ H_2S

4. ___ V_2O_5 + ___ HCl --> ___ $VOCl_3$ + ___ H_2O

5. ___ $Hg(OH)_2$ + ___ H_3PO_4 --> ___ $Hg_3(PO_4)_2$ + ___ H_2O

6. ___ SiO_2 + ___ HF --> ___ SiF_4 + ___ H_2O

7. ___ Zn + ___ HCl --> ___ $ZnCl_2$ + ___ H_2

8. ___ $HClO_4$ + ___ P_4O_{10} --> ___ H_3PO_4 + ___ Cl_2O_7

9. ___ $N_2(g)$ + ___ $O_2(g)$ + ___ H_2O --> ___ HNO_3

10. ___ NH_4NO_3 --> ___ N_2 + ___ O_2 + ___ H_2O

MOLAR MASS CONVERSION PRACTICE

1. **Determine the molar mass of each of the following compounds:**

   a) nitric acid

   b) ammonium nitrate

   c) zinc oxide

   d) cobalt (II) chloride

   e) Vitamin C (Ascorbic acid, $C_6H_8O_6$)

   f) Sulfuric acid

   g) Silver nitrate

   h) Saccharin ($C_7H_5NO_3S$)

2. **Calculate the mass of the following:**

   a) 2.22 mol He

   b) 0.453 mol Ni

   c) 15.0 mol $Ca(OH)_2$

   d) 3.21×10^{22} molecules oxygen gas

   e) 3.4 mol silver iodide

3. **Calculate the number of moles in each of the following:**

   a) 3.25 g Li_2O

   b) 0.345 g saccharin

   c) 5.62 g H_2O

   d) 3.25×10^{20} molecules nitrogen gas

   e) 245 g $(NH_4)_2S$

MOLARITY

(Work problems on notebook paper. Be NEAT and ORGANIZED. Show all your work and box your final answer. All answers should contain 3 S.F.)

1. What is the molarity of a solution that contains 0.800 moles sugar dissolved in 4.00 L of solution?

2. What mass of calcium chloride must be weighed out in order to prepare 3.00 L of a 0.500 M solution?

3. What volume of 1.50 M lithium nitrate would contain 0.600 g of that substance?

4. What is the molarity of a solution prepared by dissolving 6.85 g of NaCl in 0.500 L of solution?

5. How many grams of KNO_3 are in 2.00 L of a 0.400 M KNO_3 solution?

6. An experiment required that 1.00 g of cupric sulfate be used. Since the compound was available only in a 2.00 M solution, what volume of that solution must be used in order to supply the necessary 1.00 g?

7. To what volume must 35.0mL of 6.0M HCl be diluted to obtain a 1.5M solution?

8. What volume of 15.0M HNO_3 can be diluted to make 750.0mL of 3.0M nitric acid?

9. What volume of 2.50M HCl is needed to react completely with 20.0g of NaOH in a reaction that produces sodium chloride and water?

10. What mass of lead(II) chloride will be produced from the reaction of 60.0mL of 0.100M $Pb(NO_3)_2$ with 50.0mL of 0.210M KCl? (Hint: determine the limiting reactant)

Stoichiometry Practice A

Answer the following questions on a separate sheet of paper. Be neat, show your work, and report your answers with the appropriate number of significant figures.

1. $Na_2SiO_3 + 8\ HF \rightarrow H_2SiF_6 + 2\ NaF + 3\ H_2O$

   a) How many moles of HF are needed to react with 0.300mol of Na_2SiO_3?
   b) How many grams of NaF form when 0.500mol of HF react with excess Na_2SiO_3?
   c) How many grams of Na_2SiO_3 can react with 0.800g of HF?

2. $C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2$

   a) How many moles of CO_2 are produced when 0.400mol of $C_6H_{12}O_6$ react in this fashion?
   b) How many grams of $C_6H_{12}O_6$ are needed to form 7.50g of C_2H_5OH?
   c) How many grams of CO_2 form when 7.50g of C_2H_5OH are produced?

3. $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3CO_2$

   a) Calculate the number of grams of CO that can react with 0.150kg of Fe_2O_3.
   b) Calculate the number of grams of Fe and the number of grams of CO_2 formed when 0.150kg of Fe_2O_3 react.

4. $2\ NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$

   a) Which reagent is the limiting reactant when 1.85mol NaOH and 1.00mol CO_2 are allowed to react?
   b) How many moles of Na_2CO_3 can be produced?

5. $C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$

   a) What is the theoretical yield of C_6H_5Br when 30.0g of benzene reacts with 65.0g of bromine?
   b) If the actual yield of bromobenzene was 53.7g, what is the percent yield?

Stoichiometry Practice B

Answer the following questions on a separate sheet of paper. Be neat, show your work, and report your answers with the appropriate number of significant figures.

1. Calcium hydroxide, used to neutralize acid spills, reacts with hydrochloric acid according to the following equation: $Ca(OH)_2 + 2\ HCl \rightarrow ClCl_2 + 2\ H_2O$

    a. If you have spilled 6.5mol of HCl and put 3.8mol of $Ca(OH)_2$ on it, which substance is the limiting reactant?
    b. How many moles of the excess reactant remain?

2. Aluminum oxidizes according to the following equation: $4\ Al + 3O_2 \rightarrow 2\ Al_2O_3$

    a. Powdered aluminum (0.048mol) is placed into a container with 0.030mol O_2. What is the limiting reactant?
    b. How many moles of the excess reactant remain?

3. Heating zinc sulfide in the presence of O_2 yields the following: $ZnS + O_2 \rightarrow ZnO + SO_2$

    a. If 1.72mol of ZnS is heated in the presence of 3.24mol of O_2, which is the limiting reactant? (Note, equation is not balanced)
    b. How many moles of the excess reactant remain?

4. Chlorine can replace bromine in a single replacement reaction. The following is an example: $2\ KBr + Cl_2 \rightarrow 2\ KCl + Br_2$

    a. When 0.655g of Cl_2 and 3.205g of KBr are mixed in solution, which is the limiting reactant?
    b. How many grams of each product are formed?
    c. How many grams of excess reactant remain?

5. A process by which zirconium metal can be produced from the mineral zirconium (IV) orthosilicate starts by reacting it with chlorine gas: $ZrSiO_4 + 2Cl_2 \rightarrow ZrCl_4 + SiO_2 + O_2$

    a. What mass of $ZrCl_4$ can be produced if 852g of $ZrSiO_4$ and 955.g of Cl_2 are available?
    b. How many grams of excess reactant remain?

6. Aspirin is synthesized by the reaction of salicylic acid with acetic anhydride according to the equation: $2\ C_7H_6O_3 + C_4H_6O_3 \rightarrow 2\ C_9H_8O_4 + H_2O$

    a. When 20.0g of $C_7H_6O_3$ and 20.0g of $C_4H_6O_3$ react, which is the limiting reactant?
    b. What mass of aspirin is formed?
    c. How many grams of excess reactant remain?

Assigning Oxidation Numbers

The oxidation number is a bookkeeping device of the numbers of electrons gained or lost in an atom when it combines with other atoms to form a compound. The possible oxidation numbers of an atom can be derived from its ground state electron configuration. The actual oxidation number of an atom depends on the other atoms in the compound.

Rules for Assigning Oxidation Numbers

1. Any free (unattached) element with no charge has the oxidation number of zero. Diatomic gases such as O_2 and H_2 are also in this category.

$\overset{0}{Ag}$ $\overset{0}{Cu}$ $\overset{0}{Cl_2}$ $\overset{0}{Na}$

2. All compounds have a net oxidation state of zero. The oxidation number of all of the atoms add up to zero.

$\overset{+1\ -1}{NaCl}$

+1 -1 =0

3. Any ion has the oxidation state that is the charge of that ion. The ions of elements in Group I (Alkali metals), II (alkaline earth metals), and VII (halogens) and some other element have only one likely oxidation state.

$\overset{+2}{Ca}$ $\overset{-1}{F}$ $\overset{+1}{K}$

4. Polyatomic ions have an oxidation state for the whole ion which is the charge on that ion. The sum of the oxidation numbers of all the atoms that make up the ion equal the charge on the polyatomic ion.

$(\overset{+6\ -2}{SO_4})^{2-}$

+6 -8 = -2

5. Oxygen in compound has an oxidation state of minus two, except for oxygen as peroxide, which is minus one.

sodium oxide	sodium peroxide
$\overset{+1\ -2}{Na_2O}$	$\overset{+1\ -1}{Na_2O_2}$
+2 -2 =0	+1 -1 = 0

6. Hydrogen in compound has an oxidation state of plus one, except for hydrogen as hydride, which is minus one.

Hydrogen sulfide	Calcium hydride
$\overset{+1\ -2}{H_2S}$	$\overset{+2\ -1}{Ca\ H_2}$
+2 -2 =0	+2 -2 = 0

Let's Assign Oxidation Numbers!

Ex. Potassium perchlorate $KClO_4$

Potassium in a Group I ion K^{+1}

Oxygen in a compound is -2 O^{-2}

You must determine the oxidation state of Cl.

The sum of oxidation numbers in a compound equal zero, therefore

+1 x -2

$KClO_4$

+1 + 1(x) - 8 =0

The oxidation state of Cl is +7. $K^{+1}\ Cl^{+7}\ O_4^{-2}$

Assign the oxidation numbers for each individual element within the following compounds:

1. MnO_4^-
2. $Cr_2O_7^{2-}$
3. S_2O_3
4. NO
5. NO_3^-
6. N_2O_5
7. P_4O_8
8. PO_3^{3-}
9. $H_2PO_4^-$
10. LiH