CHEMISTRY 120

INTRODUCTORY CHEMISTRY

LABORATORY MANUAL

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Laboratory Experiments

“This lab manual was funded by the
Department of Education HSI STEM and Articulation Grant”

Introduction

Welcome to introductory chemistry, CHEM120. The laboratory portion of this class presents experiments which are an integral part of the class. The lab experiments will consist of a variety of activities. Most of the time, the goal will be to learn new laboratory techniques following detailed instructions. Occasionally a lecture will take place followed by practicing main concepts presented. The experiments serve to reinforce lecture concepts, as well as introduce new ideas. This part of the course will allow you to gain hands-on experience with glassware used in a chemical laboratory.

Laboratory Procedures

Prelaboratory Assignment

For all labs, you will have to come prepared to conduct the experiment. This means that you must read the introduction and procedure portion of the current experiment. Due at the beginning of lab is your prelaboratory assignment. Most will be worksheets that are turned in. **Prelaboratory assignments will be due at the <u>beginning</u> of the laboratory period.**

Locker Contents

You are responsible for the equipment and the condition of your locker. If you break or lose equipment during the semester, notify your instructor so that the equipment can be replaced.

SAFETY RULES FOR LABORATORY WORK

It is very important to the Allan Hancock College Faculty and Staff that no one is injured while working in the chemistry laboratory. The Life and Physical Science department has instituted a number of policies in order to ensure laboratory safety and efficiency. Your laboratory instructor and the stockroom staff have complete authority for enforcement of these rules and any other procedures to ensure safe practices in carrying out the laboratory work. In addition, it is essential that you prepare for each experiment by reading it carefully before entering the laboratory. Not only will this ensure that you get the maximum benefit from the experience, but it will also make a safer laboratory environment for everyone.

1. *Approved splash proof safety goggles must be worn at all times*. At no time are safety glasses of any kind acceptable in the laboratory. *Every* student in the laboratory must wear goggles until *everyone* has finished with the experimental procedure and has put away all glassware. Safety goggles are not to be modified in any manner.

2. *Shoes that completely cover your feet shall be worn in the laboratory at all times* to protect from chemical spills and broken glass. Inadequate protection often leads to injury, such as getting sharp glass shards embedded inside the shoe.

3. *Clothing that fully covers your torso (shoulders and down to the knees) is required to be worn in the laboratory at all times to protect you* from chemical burns on the skin. Lab coats are optional and may be purchased at any store that sells uniforms.

4. *Before you are allowed to work in the laboratory, you must learn the location and how to operate the nearest eyewash fountain, safety shower, fire extinguisher, and fire blanket*.

5. *In case of any chemical spill in or near your eyes* rinse your eyes with copious amounts of flowing water from the eyewash fountain for *15-20 minutes*. Ask for assistance immediately. Do not rub your eyes; keep eyes open while rinsing with water.

6. *In case of any chemical spill on the skin or clothing* rinse with copious amounts of flowing water from the faucet or safety shower for *15-20 minutes*. For minor spills, using the faucet is appropriate, however, for spills that cover larger portions of the body, use the safety shower. For heat burns on the skin also rinse with copious amounts of flowing cold water from the faucet. (Caution: Recall that *hot* equipment looks like *cold* equipment.)

7. *All accidents, injuries, explosions, or fires must be reported* at once to the laboratory instructor. You must go to the Health Center for treatment of cuts, burns, or inhalation of fumes. Transportation and an escort will be arranged. Your instructor will contact emergency services in case of serious injury.

8. *Laboratory areas must* never *be used for eating or drinking*. All food and beverages *must remain* in a purse, book bag, or backpack.

9. *Long hair must be tied back* while in the laboratory. Hair can catch on fire while using open flames or get caught in equipment.

10. *All operations in which noxious or poisonous gases are used or produced must be carried out inside a fume hood*.

11. *Proper procedures given by your instructor must be followed for specific hazardous waste and/or flammable waste*.

12. *All broken glassware must be discarded in the proper glassware disposal container*. Only the glass pieces shall be placed in these containers.

13. *Exercise great care when checking for chemical odors*. Always use your hand to waft vapors toward your nose.

14. *Do not force glass tubing into rubber stoppers*. Always lubricate with water or glycerin. Protect your hands with several paper towels when inserting tubing into stoppers. Check with your instructor for the proper procedure.

15. *Keep your work area neat and free of clutter*. If you spill water or a chemical or break a piece of glassware, clean it up immediately. If you are unsure of how to do this, consult your instructor. Clean the balance immediately after use. Before leaving the laboratory, clean off your laboratory bench.

16. *Perform no unauthorized experiments*.

17. *Containers of chemicals may not be taken out of the laboratory classroom or weighing room except to the stockroom for refilling.*

I have read the above rules and will observe them in my chemistry course.

________________________ *date*

________________________ ________________________

print your name here *sign your name here*

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LOCKER ASSIGNMENT SHEET

NAME ______________________________ Room No.______ Locker No. ______
Last (Print) First

COURSE **CHEM 120** SECTION __________ SEMESTER __________ INSTRUCTOR ____________

QUANTITY	ITEM DESCRIPTION	REMARKS (chipped, or damaged)	RECEIVED (student √)	RETURNED (instructor √)
6	**Beakers, 100, 150, 250, 400, 600, 800 (*or* 1000) mL**			
1	**Crucible with Lid**			
1	**Crucible Tongs**			
1	**Evaporating Dish**			
6	**Flasks, Erlenmeyer: 3 (125 mL) and 3 (250 mL)**			
2	**Funnels (plastic for solids/glass for liquids)**			
2	**Graduated Cylinders with guard (50 mL or 100 mL)**			
2	**Litmus Paper Vials (Red and Blue)**			
2	**Pipette Dropper (eyedropper) with bulb**			
1	**Scoopula**			
1	**Spatula, micro**			
2	**Stirring Rods (one with rubber policeman)**			
6	**Test Tubes, 16 x 150 mm, with rack**			
1	**Test Tube Brush**			
1	**Test Tube Holder**			
1	**Triangle, Clay**			
1	**Watch Glass**			
1	**Wire Gauze**			

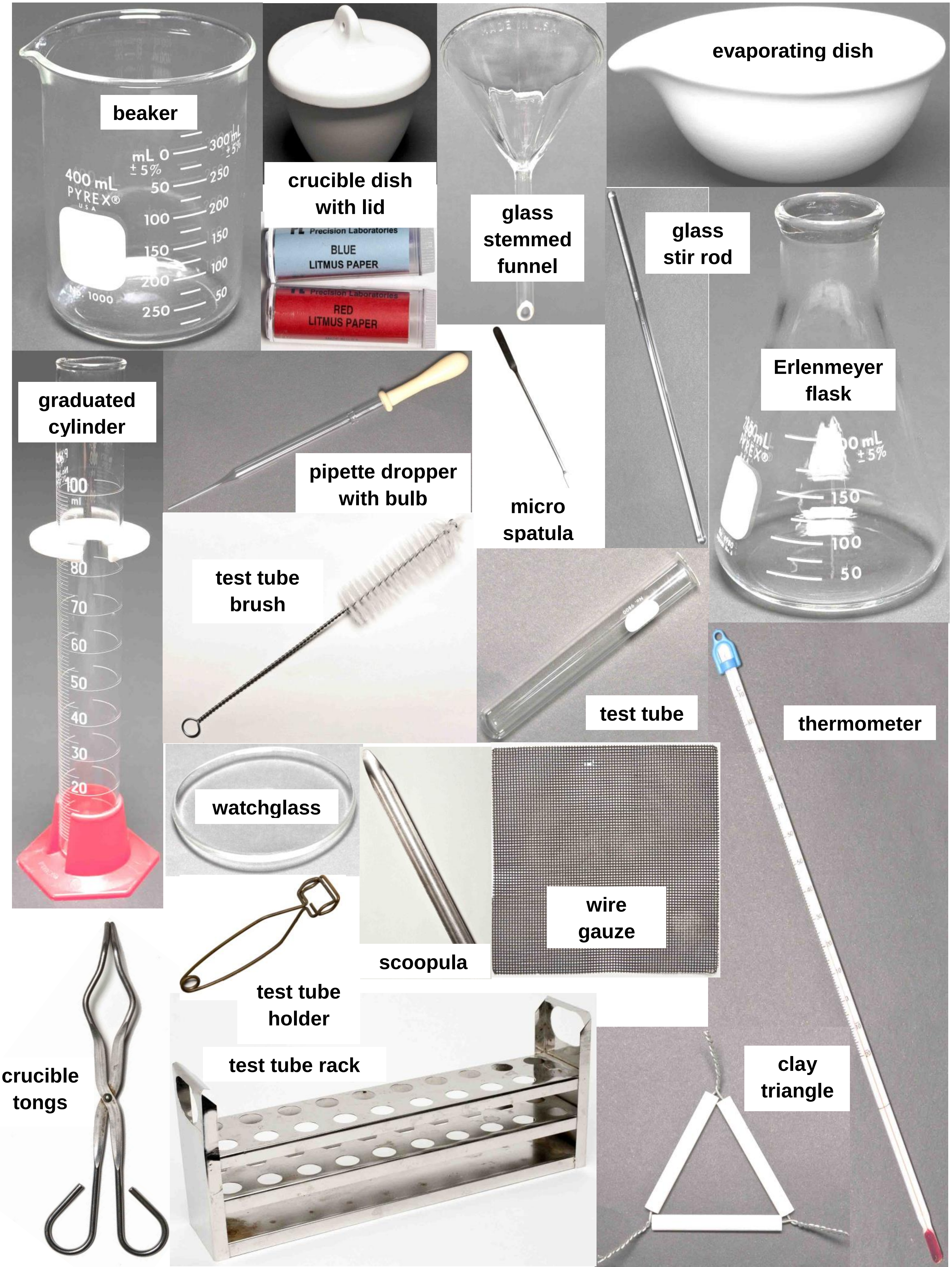
I have received the following items in good condition, as checked above:

LOCKER Checked-In By:___________________________ **DATE:**______________

....

LOCKER Checked-Out By:___________________________**DATE:**______________

(Instructor signature at end of semester.)


beaker
crucible dish with lid
glass stemmed funnel
evaporating dish
glass stir rod
Erlenmeyer flask
graduated cylinder
pipette dropper with bulb
micro spatula
test tube brush
test tube
thermometer
watchglass
scoopula
wire gauze
test tube holder
crucible tongs
test tube rack
clay triangle

Laboratory Experiments

Experiment # 1:

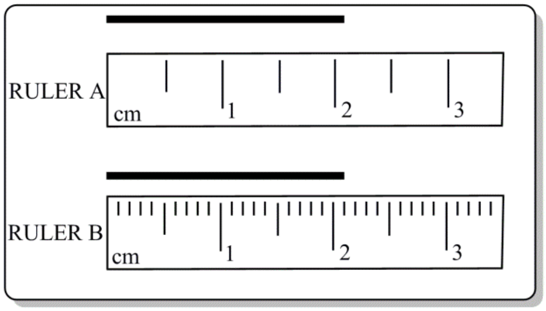
Instrumental Measurements and Calculations

Introduction:

Measurements are typically required for an experiment to obtain information that can quantitatively be compared with previous data or held to support a particular hypothesis. Instruments in the lab give us our measurements, but we need to know how to properly operate and read them to achieve the best data possible.

Instrumentation in the lab can measure energy, length, mass, temperature, time, and volume. This lab will focus on length, mass, volume, and temperature. For any instrument that we use, the markings used for measuring let us know the quality of that instrument. The more numbers that we can read from a particular instrument, typically denotes a more expensive price tag. In the following examples, when we do not have a digital device, we always want to read as many numbers from the instrument as possible. These digits are known for certain, but we always want to give a guess at the end, which communicates the limitation of that instrument.

With the first ruler, RULER A, we are only able to determine the ones' place with certainty. As we read the ruler, we always want to take one extra guess, giving us 2.1 cm. (The other marks between the numbers are just for referencing the halfway mark.) RULER B has more marks for us to confidently read 2.1 cm. With our extra guess, we get three significant figures with 2.10 cm. (Typically we see ±0.05 cm as the limitation for this device. This is because our eyes really can't resolve another 10 spaces.)



When taking mass measurements, please follow these steps:

Use Analytical Balance when possible

1) press **Tare**/**Zero**.

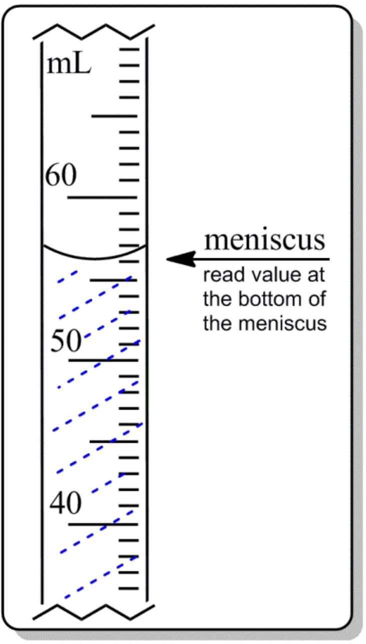
2) wait for the balance to read (**0.0000 g**).

3) place the object on the balance, shut all windows, and wait for the balance to stabilize.

4) read **ALL** digits.

5) use the same balance throughout an experiment.

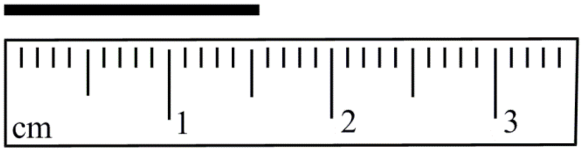
Volume measurements for liquids can be taken with specialized glassware. With this graduated cylinder we can easily read to the ones' place. We can reach 56 mL with certainty but need to make an extra guess. Exactly 56.0 mL would be the limitation for this device.

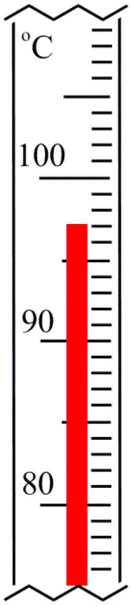


Name:____

<u>Pre-laboratory Assignment</u>:

1. Give the measurements from the following instruments:





2. Convert the following into decimal (normal) notation:

a. 5.540×10^{-4} ___________________

b. 8.002×10^{8} ___________________

3. Convert the following numbers into exponential (scientific) notation:

a. 5,000 ___________________

b. 0.00247 ___________________

c. 500.00 ___________________

d. 250,000 ___________________

4. Determine the proper number of significant figures:

a. 450 centigrams ___________________

b. 8,000,000 gallons ___________________

c. 78.30 feet ___________________

d. 1.390×10^{-34} Joules ___________________

e. 50**.** grains ___________________

f. 5.00×10^{6} seconds ___________________

g. 1.3000 grams ___________________

h. 10,000 meters ___________________

Purpose:

To perform measurements in the lab and become familiar with instruments and equipment used in taking measurements. Techniques such as significant figure rules when performing arithmetic with measurements and dimensional analysis will be utilized throughout this experiment.

Theory:

Significant figures with measurements and using dimensional analysis.

Procedure:

A. Length Measurement:

1. Measure the length of a small test tube and record your answer in centimeters (cm).
2. Measure the diameter of a watchglass and record your answer in centimeters (cm).
3. Measure the diameter of an evaporating dish (not including the spout) and record your answer in centimeters (cm).

B. Mass Measurement:

1. Using an analytical balance, determine the mass of a small test tube and record your answer in grams (g).
2. Using the same analytical balance, determine the mass of a plastic funnel and record your answer in grams (g).
3. Using the same analytical balance, determine the mass of a crucible and record your answer in grams (g).
4. Using the same analytical balance, determine the mass of a crucible cover and record your answer in grams (g).
5. Using the same analytical balance, determine the mass of a crucible AND cover and record your answer in grams (g).
6. Calculate the total mass of the crucible and crucible cover by adding #3 and #4 and record this value in your data sheet. Show your work and remember to use proper significant figure rules when doing addition calculations with measurements.
7. Calculate the difference between the observed mass and the calculated mass by subtracting #5 from #6 and record your answer in grams (g). Show your work and remember to use proper significant figure rules when doing subtraction calculations with measurements.

C. Volume Measurement:

1. Fill a 100-mL graduated cylinder with water to around the 80-mL mark. Record the volume of water in milliliters (mL) to the nearest tenths' place in your data sheet.
2. Using the water in the graduated cylinder, fill a small test tube to the brim. Record the new volume in the graduated cylinder on your data sheet.
3. Fill a different small test tube to the brim and record the new volume in the graduated cylinder on your data sheet.
4. Calculate how many milliliters of water are in the first test tube by subtracting #2 from #1 and record this value in your data sheet. Show your work and remember to use proper significant figure rules when doing subtraction calculations with measurements.
5. Calculate how many milliliters of water are in the second test tube by subtracting #3 from #2 and record this value in your data sheet. Show your work and remember to use proper significant figure rules when doing subtraction calculations with measurements.

D. Temperature Measurement:

1. Read the temperature in the laboratory by using a thermometer. Record the temperature in degrees Celsius (°C) to the nearest tenths' place in your data sheet.
2. If a cold water bath has not been prepared, fill a 100-mL beaker with ice and water to the 60-mL mark. Hold the thermometer in the ice water and record the lowest temperature observed in your data sheet.
3. If a hot water bath has not been prepared, place a 150-mL beaker filled with water to the 80-mL mark on a hot plate. Heat the water to boiling and record the hottest temperature observed in your data sheet. Remember to suspend the thermometer in the water to obtain a true reading of the boiling water. Placing on the bottom of the beaker will give the reading of the glass.

E. Mass and Volume Measurement:

1. Obtain a rectangular solid and write down the unknown number in your data sheet. Wipe off all finger prints with a napkin or Kimwipe.
2. Place the rectangular solid on an analytical balance with crucible tongs. Record the mass of the solid in your data sheet.
3. With a metric ruler, measure the length, width, and thickness of the solid in units of centimeters. Record these values in your data sheet.
4. Calculate the volume of the solid by multiplying the length, width, and thickness. Record this value in your data sheet in units of cm^3. Show your work and remember to use proper significant figure rules when doing multiplication calculations with measurements.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Length Measurement

1. Length of a small test tube (cm): ______________
2. Diameter of a watchglass (cm): ______________
3. Diameter of an evaporating dish (cm): ______________

B. Mass Measurement:

1. Mass of a small test tube (g): ______________
2. Mass of a plastic funnel (g): ______________
3. Mass of a crucible (g): ______________
4. Mass of a crucible cover (g): ______________
5. Mass of a crucible and cover (g): ______________
6. Calculated total mass of a crucible and cover (g): ______________
7. Difference between calculated and observed masses (g): ______________

C. Volume Measurement:

1. Volume in graduated cylinder (mL): ______________
2. New volume in graduated cylinder (mL): ______________
   (after adding water to first test tube)
3. New volume in graduated cylinder (mL): ______________
   (after adding water to second test tube)
4. Volume in the first test tube (mL): ______________
5. Volume in the second test tube (mL): ______________

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

D. Temperature Measurement:

1. Room temperature (°C): _______________
2. Melting/Freezing Point of Ice (°C): _______________
3. Boiling/Condensing Point of Water (°C): _______________

E. Mass and Volume Measurement:

1. Number of the unknown solid (#): _______________________________
2. Mass of the unknown solid (g): _______________________________
3. Length of the unknown solid (cm): _______________________________

   Width of the unknown solid (cm): _______________________________

   Thickness of the unknown solid (cm): _______________________________
4. Volume of the unknown solid (cm^3): _______________________________

<u>Data Analysis/Calculations</u>:

(Use this area to show all calculations performed throughout the experiment.)

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

<u>Post-laboratory exercises</u>:

Work the following problems showing all of your work through the dimensional analysis technique learned in lecture. Please **<u>box</u>** or **<u>circle</u>** your answer making sure to round to the appropriate number of significant figures and to include units. Your calculator will help you with the arithmetic.

1. A discus was thrown 190 feet. How many yards did the discus travel? (3 ft = 1 yard)

2. How many pounds are in 58 kilograms? (1 kg = 2.205 pounds)

3. Find the equivalent in Japanese yen for $376.50. Assume the rate of exchange is 105 yen equals one dollar.

4. One gallon of a certain brand of gasoline costs $4.50. How many gallons could be purchased with $75.00?

5. The distance from Santa Maria to San Francisco is 428 kilometers. Calculate the distance in miles. (1 km = 0.6214 miles)

Experiment:	Date:
Name:	Course/Section:

6. How many seconds elapse in 2.0 hours?

7. How many milligrams are in 850 nanograms?

8. In the apothecary weight system one scruple is 20 grains, and three scruples equals one dram. Determine the number of grains in 50 drams.

9. How many micrometers are in 200 decimeters?

10. Convert 55.0 miles per hour to meters per minute.
(5280 ft = 1 mile; 2.54 cm = 1 inch)

11. How many minutes are in 3.5 years?

Experiment:	Date:
Name:	Course/Section:

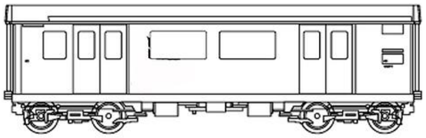
12. The gasoline in an automobile gas tank has a mass of 75.0 kg and a density of 0.752 g/cm^3. What is the volume in milliliters (mL)?

13. A medication is to be given by IV (intravenous) at the rate of 185 drops per min. The medication's concentration is 250 mg/mL. What is the weight (in grams) of the medication delivered to the patient in 5.0 days? (1 mL = 15 drops)

14. On a trip of 834 miles, an automobile consumed 35.5 gallons of gasoline costing $3.25 per gallon. What was the cost of gasoline per mile traveled?

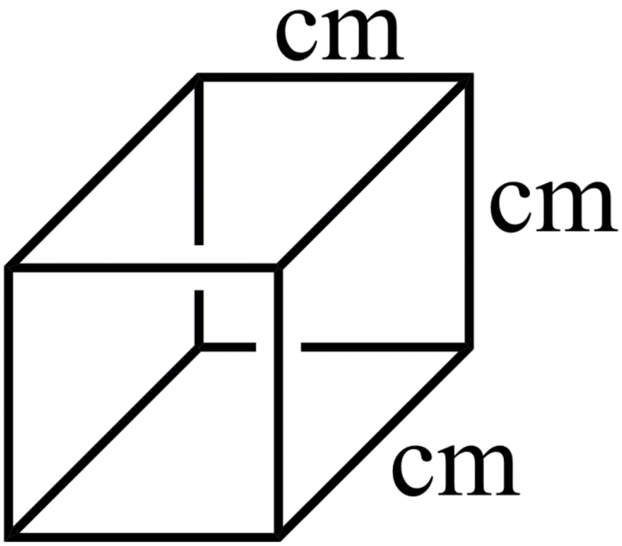
15. Find the number of cm^3 in 2.60 cubic foot (ft^3).
    (2.54 cm = 1.00 in; 12 inches = 1 ft)

16. How many boxcars are there in a long freight train if it takes the entire train 2.0 min to pass a station as it travels 40. miles per hour? The boxcars were 86 feet long.



Experiment:	Date:
Name:	Course/Section:

Experiment #2: Density



Volume = length x width x height

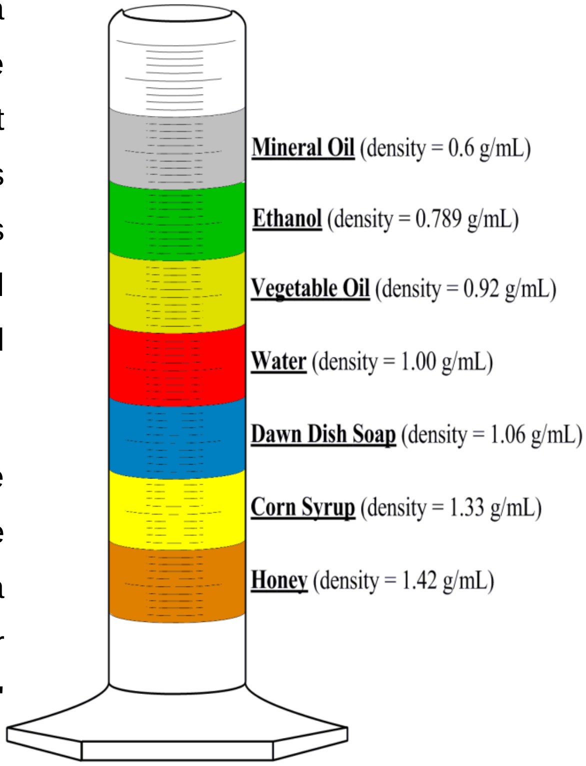
Volume = cm x cm x cm

Volume = cm^3

Introduction:

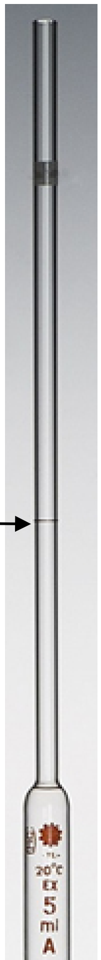
Density is a physical property that relates a substance's mass to its volume. Based on a certain volume (typically 1 mL or 1 cm^3) we learn how much mass is in that area. This gives us an idea of how tightly packed the atoms or molecules are. The larger the density, the more mass there is in a given volume. With more mass, the object will sink relative to other densities as seen in the graduated cylinder.

The equation that relates mass and volume will have units of grams for the mass; however, the units for the volume will depend on if the substance is a solid (cm^3) or a liquid (mL). This will give us g/cm^3 and g/mL. Remember that chemistry and physics will typically omit the number "1" if possible, so these density values are out of 1 cm^3 or 1 mL.



Volume can be determined by analytical equipment or by displacement. The volume by displacement technique (Archimedes' Principle) requires that we know the final and initial volumes of a liquid as we submerge an irregularly shaped object. The volume of the object comes from the difference in the volume of the liquid.

The analytical equipment that will be utilized today is known as the volumetric pipette. The volumetric pipette is considered an analytical tool because it can give a very precise volume measurement. There is a calibration line located on the top of the pipette. Using a pipetter or pipette wand(bulb), we draw up a substance into the volumetric pipette. Once the meniscus is at the **calibration line**, the substance is then drained into a pre-weighed container. The volumetric pipette is designed to have a little residual leftover in the tip. We do not want to blow or shakeout this remaining residual. The volumetric pipette is designed to deliver a specific amount from the calibration line with a little left over in the tip. The volumetric pipette can be touched to the side of the container to remove any drops hanging.



Name:____

Pre-laboratory Assignment: You must show your work in order to receive credit.

1. A wooden square block is placed on an analytical balance giving a mass of 5.1977 g. The square block has dimensions of 2.00 cm long, 2.00 cm wide, and has a height of 2.00 cm. Calculate the density of the wooden square block in units of g/cm^3 to the correct number of significant figures.

2. A marble ball is placed on an analytical balance giving a mass of 12.5640 g. This marble ball is placed in a graduated cylinder containing 40.0 mL of water. The marble ball displaces the water to a reading of 44.5 mL. Calculate the density of the marble ball in g/cm^3 to the correct number of significant figures. (1 mL = 1 cm^3)

3. An Erlenmeyer flask is placed on an analytical balance giving a reading of 124.2218 g. Exactly 25.00 mL of tap water was placed into an Erlenmeyer flask using a volumetric pipette. The Erlenmeyer flask with the tap water is now placed on the same analytical balance giving a second reading of 149.7450 g. Calculate the density of the tap water in g/mL to the correct number of significant figures.

Objectives:

Understand the concept of density and determine the densities of various liquids and solids.

Utilize density and volume to determine the thinness of an object.

Understand the volume by displacement method and weighing by difference.

Purpose: Density is defined as mass/volume. By measuring the mass and volume of solids and liquids, densities will be calculated and compared with literature values. Techniques such as volume by displacement and weighing by difference will be utilized throughout this experiment.

Theory: Weighing by difference: final mass - initial mass = total mass

Volume by displacement: final volume - initial volume = volume of object

Density = mass/volume.

Volume = length x width x height (thickness)

Procedure:

A. Density of Water:

1. Weigh a 125-mL Erlenmeyer flask with a rubber stopper that fits snuggly using the centigram balance. Record this number in your data sheet.
2. Obtain about 50 mL of deionized water in a 100-mL beaker. Using the 10-mL volumetric pipette, pipet exactly 10.00 mL of water into the 125-mL flask and record this number in your data sheet. Make sure not to pull any liquid up into the pipette wand.
3. Reweigh the 125-mL flask with the stopper using the same balance and record this number in your data sheet.
4. Determine the mass of water by subtracting #1 from #3 on your data sheet. This number is the mass of water added to the 125-mL flask and is known as **weighing by difference**. Show your work and remember to use proper significant figure rules with measurements.

5. Calculate the density of water by dividing the mass of water (#4) by the total volume (10.00 mL). Show your work and remember to use proper significant figure rules with measurements.
6. Repeat this trial two more times by repeating steps 1 through 5. It is not necessary to dry the flask between trials because we are utilizing the weighing by difference technique.
7. Determine the average density of water by adding the densities of each trial and dividing by the total number of trials (in this case, three). The number of trials were counted, so considered an exact number or part of an equation; therefore, its significant figure is NOT counted when determining the average.

B. Density of Rubbing Alcohol:

1. Weigh a 125-mL Erlenmeyer flask with a rubber stopper that fits snuggly using the centigram balance. Record this number in your data sheet.
2. Obtain about 50 mL of the alcohol solution in a 100-mL beaker. Using the 10-mL volumetric pipette, pipet exactly 10.00 mL of rubbing alcohol into the 125-mL flask and record this number in your data sheet. Make sure not to pull any liquid up into the pipette wand.
3. Reweigh the 125-mL flask with the stopper using the same balance and record this number in your data sheet.
4. Determine the mass of rubbing alcohol by subtracting #1 from #3 on your data sheet. This number is the mass of rubbing alcohol added to the 125-mL flask and is known as **weighing by difference**. Show your work and remember to use proper significant figure rules with measurements.
5. Calculate the density of rubbing alcohol by dividing the mass of rubbing alcohol (#4) by the total volume (10.00 mL). Show your work and remember to use proper significant figure rules with measurements.
6. Pour the rubbing alcohol into the appropriate waste container and rinse the volumetric pipette out with deionized water.
7. Repeat this trial two more times by repeating steps 1 through 6. It is not necessary to dry the flask between trials because we are utilizing the weighing by difference technique.
8. Determine the average density of rubbing alcohol by adding the densities of each trial and dividing by the total number of trials (in this case, three). The number of trials were counted, so considered an exact number or part of an equation; therefore, its significant figure is NOT counted when determining the average.

C. Density of a Rubber Stopper:

1. Obtain a rubber stopper (size #1 or #2) and make sure it is dry. Record the size of the rubber stopper on your data sheet. Typically, the size number is located on the top of the rubber stopper.
2. Weigh the dry rubber stopper using an analytical balance and record this number on your data sheet.
3. Obtain a 100-mL graduated cylinder and fill to the 50.0 mL mark with water. Record this number on your data sheet to the appropriate number of significant figures.
4. Gently place the rubber stopper into the 100-mL graduated cylinder without spilling any water. Tilting the cylinder to the side may help. Record the new number of the graduated cylinder to the appropriate number of significant figures.
5. Calculate the volume of the rubber stopper by subtracting #3 from #4. This is known as **volume by displacement**. Show your work and remember to use proper significant figure rules with measurements.
6. Calculate the density of the rubber stopper by dividing the mass of the rubber stopper (#2) by the volume of the rubber stopper (#5). Record this number on your data sheet. Show your work and remember to use proper significant figure rules with measurements.
7. Repeat this trial two more times by repeating steps 1 through 6 with the **same** rubber stopper and analytical balance.
8. Determine the average density of the rubber stopper by adding the densities of each trial and dividing by the total number of trials (in this case, three). The number of trials were counted, so considered an exact number or part of an equation; therefore, its significant figure is NOT counted when determining the average.

D. Density of an Unknown Solid:

1. Obtain a rectangular solid and write down the unknown number in your data sheet. Wipe off all fingerprints with a napkin or Kimwipe.
2. Place the rectangular solid on an analytical balance with crucible tongs. Record the mass of the solid in your data sheet.
3. With a metric ruler, measure the length, width, and thickness of the solid in units of centimeters. Record these values in your data sheet.
4. Calculate the volume of the solid by multiplying the length, width, and thickness. Record this value in your data sheet in units of cm^3.

5. Determine the density of the solid by dividing the mass of the solid (#2) by the calculated volume (#4). Show your work and remember to use proper significant figure rules with measurements.
6. Determine what your unknown solid is from the following known densities in grams/centimeter^3 (g/cm^3).

Acrylic	*1.16-1.19*	g/cm^3	*Nylon*	*1.13*	g/cm^3
Aluminum (Al)	*2.70*	g/cm^3	*PVC*	*1.39-1.42*	g/cm^3
Brass (Cu/Zn alloy)	*8.60*	g/cm^3	*Steel (Fe/C alloy)*	*7.87*	g/cm^3
Copper (Cu)	*8.96*	g/cm^3	*Tin (Sn) (gray)*	*5.77*	g/cm^3
Iron (Fe)	*7.87*	g/cm^3	*Tin (Sn) (white)*	*7.367*	g/cm^3
Lead (Pb)	*11.34*	g/cm^3	*Wood*	*0.35-0.90*	g/cm^3
Magnesium (Mg)	*1.74*	g/cm^3	*Zinc (Zn)*	*7.14*	g/cm^3

E. Thinness of Aluminum Foil:

1. Obtain a rectangular piece of aluminum foil. Unfold it completely if it is folded. With a metric ruler, measure the length and width of the foil in units of centimeters and record these values in your data sheet.
2. Fold it to fit in an analytical balance and write down the mass in your data sheet.
3. Using the density of aluminum (2.700 g/cm^3), determine the volume (in cm^3) of the aluminum foil. Once the volume is calculated, use the equation of a cube (Volume = length x width x thickness(height)) to determine the thinness of the aluminum foil. Show all of your work in the space provided.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Density of Water:

	TRIAL 1	***TRIAL 2***	***TRIAL 3***
1. Mass of flask and rubber stopper (g):	__________	__________	__________
2. Volume of water from 10.00 mL volumetric pipet (mL):	__________	__________	__________
3. Mass of flask, rubber stopper, and 10.00 mL of water (g):	__________	__________	__________
4. Mass of water (weighing by difference technique) (g):	__________	__________	__________
5. Density of water (**show your work for TRIAL #1**) (g/mL):	__________	__________	__________

Average density of water (g/mL): ________________

B. Density of Rubbing Alcohol:

	TRIAL 1	***TRIAL 2***	***TRIAL 3***
1. Mass of flask and rubber stopper (g):	__________	__________	__________
2. Volume of alcohol from 10.00 mL volumetric pipet (mL):	__________	__________	__________
3. Mass of flask, rubber stopper, and 10.00 mL of alcohol (g):	__________	__________	__________
4. Mass of alcohol (weighing by difference technique) (g):	__________	__________	__________
5. Density of alcohol (**show your work for TRIAL #1**) (g/mL):	__________	__________	__________

Average density of alcohol (g/mL): ________________

Experiment:	Date:
Name:	Course/Section:

C. Density of a Rubber Stopper:

	TRIAL 1	***TRIAL 2***	***TRIAL 3***
1. Size of rubber stopper (#):	__________	__________	__________
2. Mass of rubber stopper (g):	__________	__________	__________
3. Initial volume in graduated cylinder (mL):	__________	__________	__________
4. Final volume in graduate cylinder (mL):	__________	__________	__________
5. Volume of the rubber stopper (mL):	__________	__________	__________
6. Density of the rubber stopper (**show your work for TRIAL #1**) (g/mL):	__________	__________	__________

Average density of rubber stopper (g/mL): ________________

D. Density of an Unknown Solid:

1. Number of the unknown solid (#): ________________________________

2. Mass of the unknown solid (g): ________________________________

3. Length of the unknown solid (cm): ________________________________

Width of the unknown solid (cm): ________________________________

Thickness of the unknown solid (cm): ________________________________

4. Volume of the unknown solid (cm^3): ________________________________

5. Density of the unknown solid (g/cm^3): ________________________________

6. Determination of the unknown solid based on density: ________________________________

E. Thinness of Aluminum Foil:

1. Length of aluminum foil (cm): ________________________________

Width of aluminum foil (cm): ________________________________

2. Mass of aluminum foil (g): ________________________________

3. Volume of aluminum foil (cm^3): ________________________________

Thickness (or Thinness) of aluminum foil (cm): ________________________________

Experiment:	Date:
Name:	Course/Section:

Data Analysis/Calculations:

A. Density of Water:

B. Density of Rubbing Alcohol:

C. Density of a Rubber Stopper:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

D. Density of an Unknown Solid:

E. Thinness of Aluminum Foil:

Date:	Instructor Signature:	Comments:

Post-laboratory exercises:

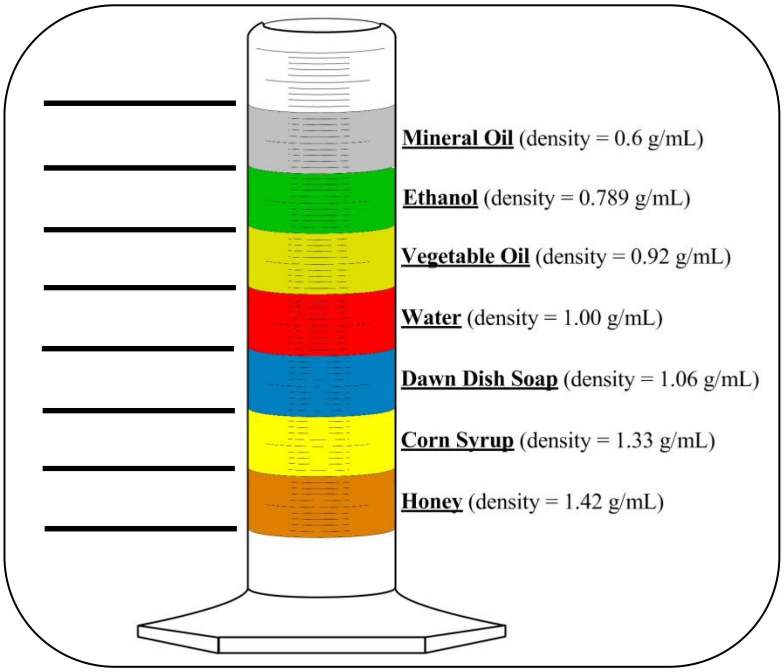
1a. Determine the volume of 1.000 g of water ($\text{Density}_{\text{water}}$ = 0.9999 g/mL).

1b. Determine the volume of 1.000 g of ice ($\text{Density}_{\text{ice}}$ = 0.9168 g/cm^3). (1 mL = 1 cm^3)

1c. From the volumes above, what can we conclude about 1.000 g of water when it freezes?

2. Show where the objects would rest once placed in the container.

substance **A**	density = 1.080 g/mL
substance **B**	density = 0.720 g/mL
substance **C**	density = 0.999 g/mL
substance **D**	density = 1.400 g/mL
substance **E**	density = 0.599 g/mL
substance **F**	density = 1.033 g/mL
substance **G**	density = 3.481 g/mL
substance **H**	density = 0.823 g/mL



3. Tequila often contains about 40 % alcohol by volume (80 proof) giving it a density of 0.812 g/mL. Calculate the mass, in grams, of a 5.0 mL shot of tequila.

4. The density of a proton is approximately 4.0 x 10^{14} g/cm^3. If the mass of the proton is 1.7 x 10^{-24} g, calculate its volume.

4 b. Using the volume of a sphere equation, determine the radius of the proton in centimeters (cm).

The volume of a sphere: $V = \frac{4}{3}\pi r^3$. (π = 3.14159)

4 c. Convert this answer to nanometers (nm).

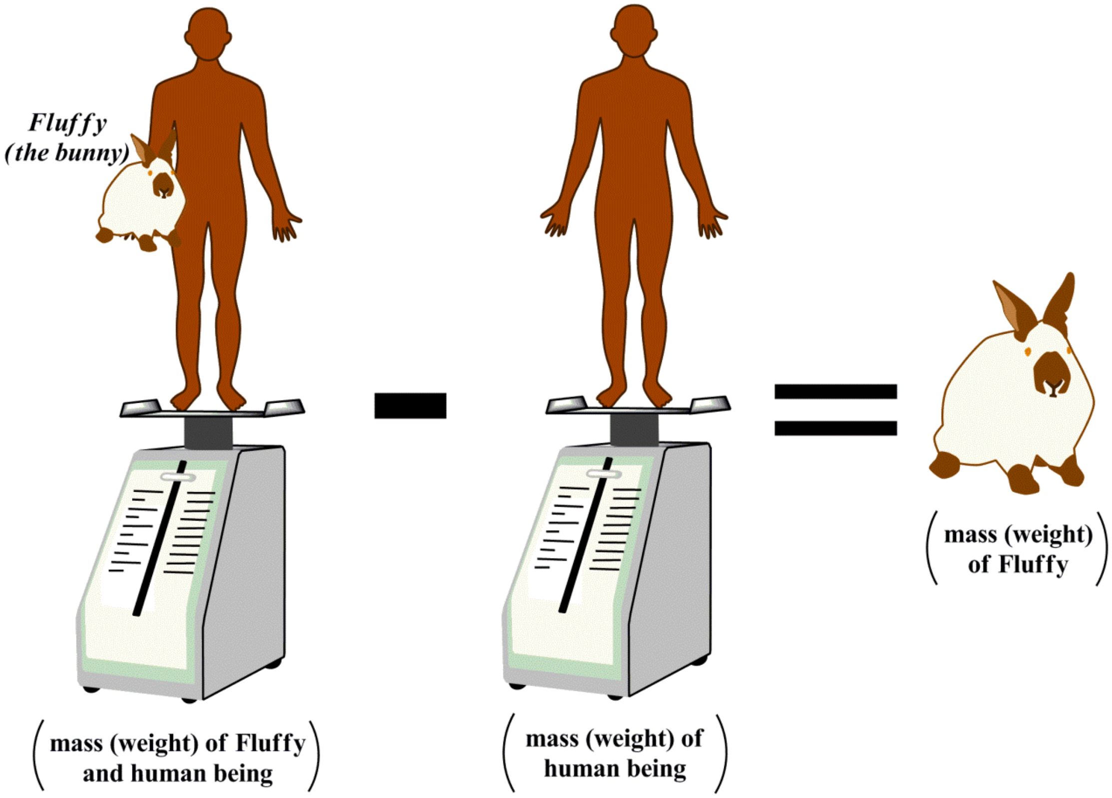
Experiment #3:

Physical Properties of Matter

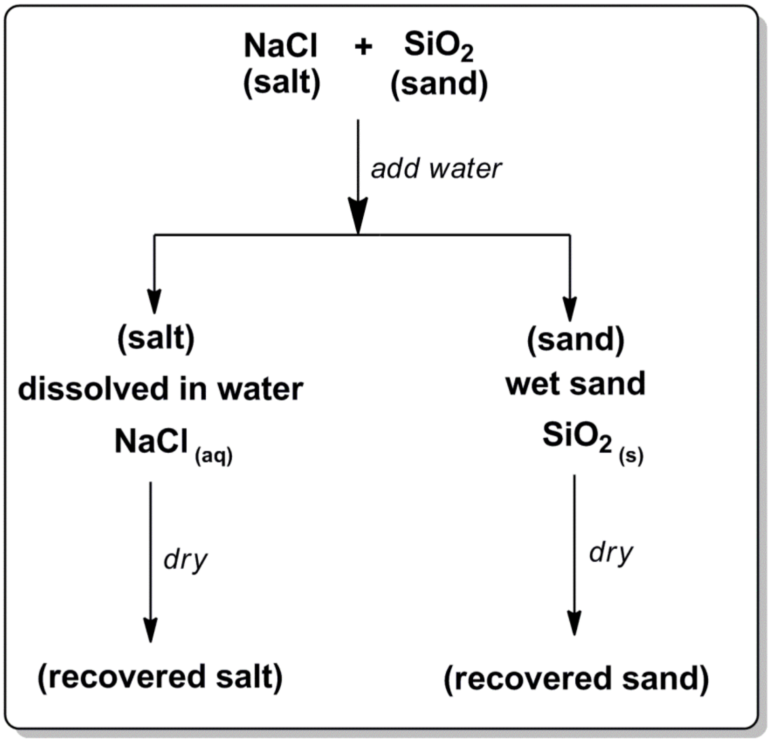
Introduction:

The weighing by difference technique is useful when certain substances are corrosive to the balance or when they like to move around (when taking the weight of animals).

Weighing By Difference Technique



By first obtaining the mass of the animal and yourself, we can then subtract out the mass of yourself to determine the mass of the animal. This process is known as the weighing by difference technique.



Physical State	Symbol
solid	**(s)**
liquid	**(l)**
gas	**(g)**
aqueous	**(aq) substance dissolved in water**

$$\text{percentage (\%)} = \frac{\text{part (sample of interest)}}{\text{whole}} \times 100 =$$

$$\text{percent salt in mixture (\%)} = \frac{\text{mass of salt recovered}}{\text{total mass of original sample}} \times 100 =$$

$$\text{percent sand in mixture (\%)} = \frac{\text{mass of sand recovered}}{\text{total mass of original sample}} \times 100 =$$

$$\text{percent recovery (\%)} = \frac{\text{total mass of recovered components}}{\text{total mass of original sample}} \times 100 =$$

Name:____

Pre-laboratory Assignment: You must show your work in order to receive credit.

1. A vial with a mixture of sand and salt is placed on an analytical balance. The mass reads 3.4155 g. Some of this mixture is poured out of the vial into a beaker. The vial with the remaining mixture is then placed on the same balance. Its mass now reads 2.8540 g. What was the mass of the sand and salt mixture transferred?

2. Give the four different types of physical states and their symbols.

3. Using the percentage equation given in the introduction, if 8.912 g of sand are recovered from an entire sand and salt mixture that weighed 10.4755 g. What was the percent of sand in the mixture?

Objectives:

Understanding the importance of weighing by difference.

Separating a mixture using physical methods.

Purpose:

To use the separation method known as filtration to separate two compounds mixed together. A mixture of salt and sand is to be separated by physical means into the individual components. Equipment and techniques such as filtration and boiling point apparatus will be utilized.

Theory:

Separation of a mixture into the individual compounds by physical means (not changing the chemical composition).

Procedure:

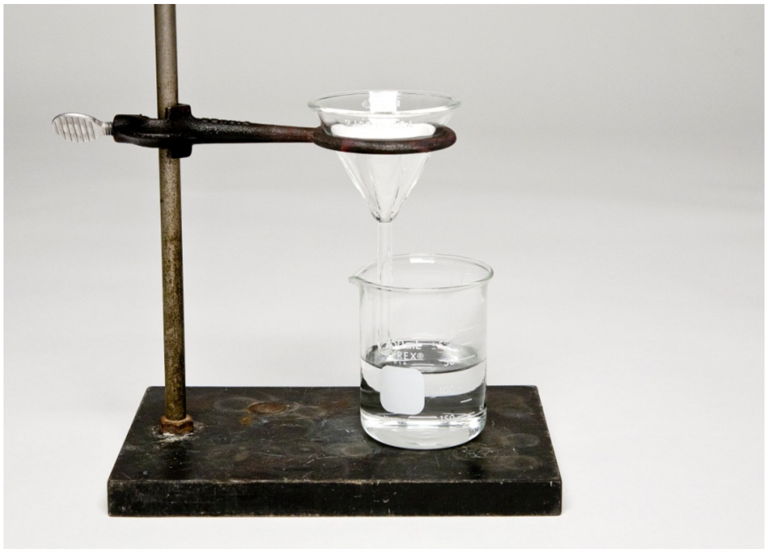
A. Physical Properties: Solubility

a. Preparing the Mixture:

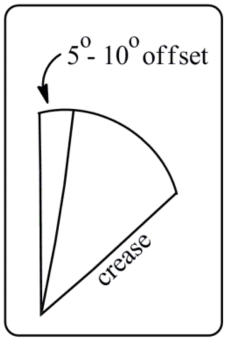
1. Obtain one vial containing the unknown sand ($SiO_{2\,(s)}$) and salt ($NaCl_{(s)}$) mixture. We will be doing two individual trials using the same vial. Write down the unknown number in your data sheet.
2. Determine the mass of the vial using the analytical balance and record this number in your data sheet.
3. We want to measure out two to three grams of the mixture of sand and salt. Do this by lightly filling the bottom of a clean, dry 150-mL breaker with the mixture. Just pour enough so that you cannot see through the bottom of the beaker. Do NOT empty the entire contents of the vial into the beaker. Reweigh the vial and record this number in your data sheet.
4. Utilizing the weighing by difference technique, determine the mass of the unknown mixture by subtracting #3 from #2 and record this value. **Notice the analytical balance was never exposed to the mixture.** Show your work and remember to use proper significant figure rules with measurements.

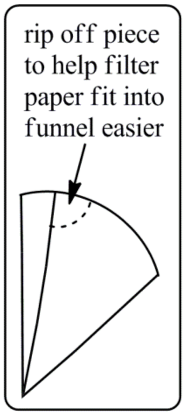
b. Separating the Mixture:

1. Obtain a piece of filter paper and place it on a clean, dry watchglass. Determine the mass of the filter paper with the watchglass and record this number in your data sheet.
2. Measure approximately 50.0 mL of deionized water in a graduated cylinder and slowly add to the 150-mL beaker containing the sand and salt mixture. Stir continuously with a glass stirring rod while adding the deionized water. Continue stirring for an additional minute.
3. Set up a small ring clamp attached to a ring stand. Place your glass-stemmed funnel inside the small ring clamp. If the ring clamp is too large for your glass funnel, use a clay-triangle to brace the funnel inside the larger ring.



4. Prepare a piece of filter paper as described below:
    - i. Fold the filter paper perfectly in half and crease.
    - ii. Fold the filter paper in half again except leave a small 5 - 10° angle between the halves before creasing.
    - iii. From the half that is in front of the back half, rip a small triangle and save.



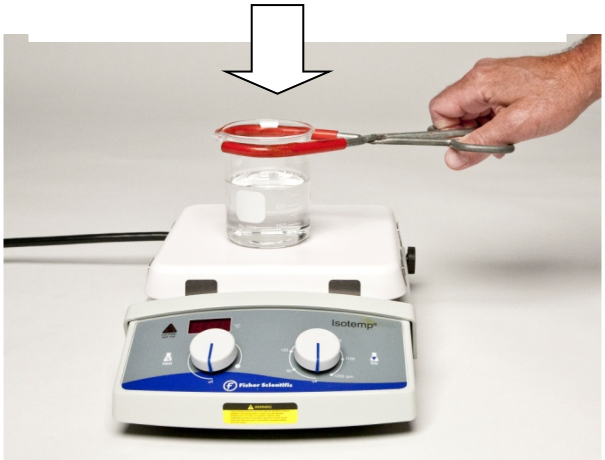


    - iv. From the back end (end away from you), open up the first leaf and form a cone. Place the cone inside the glass-stemmed funnel and place the small triangle ripped off inside. Wet with deionized water to secure paper to the glass funnel.

5. Determine the mass of another clean, dry 150-mL beaker and record this value in your data sheet. Properly label and place this beaker underneath your glass-stemmed funnel and make sure the funnel tip touches the side of the beaker.
6. Stir the sand and salt mixture for another minute and slowly decant the solution into your glass funnel using a glass stirring rod. Scrape out any remaining solid residue with a rubber policeman and rinse out the entire beaker with a wash-bottle filled with deionized water. It does not matter the amount of water added because it will be evaporated off in the next step.

decanting

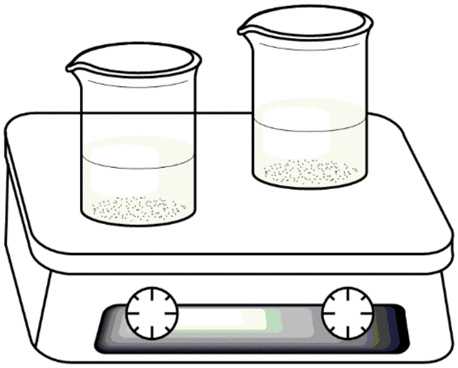
Remember that metal and glass look the same when hot or cold. Please use beaker tongs or oven mitts.



7. The liquid that flows through the glass funnel is known as the filtrate. Place the filtrate in the beaker on a hot plate and heat the solution to boiling. As the solution boils down to about 5 mL, carefully continue to heat to avoid splattering as the salt dries. *(If time is a factor, remove from the heat source and place in the fume hood area designated for this experiment).*
8. When no more water vapor is observed leaving the beaker, the salt is considered completely dry. Determine the mass of this beaker and its contents once it has cooled and record this value. Remember that the balance and the object must be at the same temperature to obtain an accurate mass. A hot object will produce convection currents which will throw off the mass reading.
9. Remove the filter paper and the residue from the glass funnel and place on the large hot plate in the fume hood to expedite drying. *(If time is a factor, place the filter paper on the same watchglass, properly label, and place this watchglass in the fume hood area designated for this experiment).*

10. Once the sand is completely dry, determine the mass of the watchglass, filter paper, and residue and record this value in your data sheet.
11. Determine the mass of sand by subtracting #1 from #10 and record this value in your data sheet. Show your work and remember to use proper significant figure rules with measurements.
12. Determine the mass of salt by subtracting #5 from #8 and record this value in your data sheet. Show your work and remember to use proper significant figure rules with measurements.
13. Calculate the percent sand in your mixture by dividing the mass of sand recovered (#11) by the total mass of the sample (#4 from part a) and multiply this number by 100. Record this value. Show your work and remember to use proper significant figure rules with measurements.
14. Calculate the percent salt in your mixture by dividing the mass of salt recovered (#12) by the total mass of the sample (#4 from part a) and multiply this number by 100. Record this value. Show your work and remember to use proper significant figure rules with measurements.
15. Calculate the total mass of sand and salt recovered by adding #11 and #12. Record this value.
16. Calculate the percent recovery by dividing #15 by (#4 from part a) and multiply by 100. Record this value. Show your work and remember to use proper significant figure rules with measurements.
17. Before you throw the sand and filter paper in the chemical waste area, check your calculations to make sure you do not still have water weight. If your percent recovery is over 100%, continue to dry the sand and salt for an additional ten minutes or until no more water vapor is observed leaving the system.
18. If time permits, repeat all the steps above with the <u>**same**</u> vial.

HEAT
Temperature readings only work when a connected probe is inserted directly into a solution.



STIR
Stir feature only works when a magnetic spin-bar or spinvane is placed directly into the solution.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Physical Properties: Solubility

a. Preparing the Mixture:

	TRIAL 1:	TRIAL 2:
1. Unknown number on vial (#):	________________	________________
2. Mass of vial and mixture before pour (g):	________________	________________
3. Mass of vial and mixture after pour (g):	________________	________________
4. Mass of original sample mixture (g):	________________	________________

b. Separating the Mixture:

	TRIAL 1:	TRIAL 2:
1. Mass of filter paper and watchglass (g):	________________	________________
2. Amount of water (mL):	________________	________________
3. Setting up equipment (no data):		
4. Preparing filter paper (no data):		
5. Mass of 150 mL beaker:	________________	________________
6. Filtering (no data):		
7. Evaporating (no data):		
8. Mass of 150 mL beaker and salt (g):	________________	________________
9. Drying (no data):		
10. Mass of watchglass, filter paper, and residue (g):	________________	________________
11. Mass of sand (SiO_2) (g):	________________	________________
12. Mass of salt (NaCl) (g):	________________	________________
13. Percentage of sand in mixture (%):	________________	________________
14. Percentage of salt in mixture (%):	________________	________________
15. Total mass recovered (g):	________________	________________
16. Percentage recovered (%):	________________	________________

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Data Analysis/Calculations:

Date:	Instructor Signature:	Comments:

Post-laboratory exercises:

1a. Throughout the experiment, what are some possible sources of error that could have led to someone recovering a significantly smaller amount of table salt (NaCl) than was originally present in the sample mixture.

1b. What are some possible sources of error that could have led to someone to appear to have recovered a significantly larger amount of sand (SiO_2) than was originally present in the sample mixture.

2. Use the table of substances and their solubilities in water to answer the following questions:

substance	*soluble (dissolves) in cold water*	*soluble (dissolves) in hot water*
cotunnite ($PbCl_2$)	**no**	**yes**
milk of magnesia ($Mg(OH)_2$)	**no**	**no**
table salt (NaCl)	**yes**	**yes**

a. Describe how you would separate a mixture of cotunnite and table salt and recover the two substances.

b. Could you separate a mixture of cotunnite and milk of magnesia using only cold water? Why or why not?

c. A sample mixture contains two of the three compounds listed in the table. Both substances will not dissolve (are insoluble) in cold water. The sample mixture containing substances A and B was placed into hot water, dissolving one of the substances and allowing us to extract it from the mixture. After filtration, substance A flowed through the funnel and into the receiving flask. This filtrate was evaporated to dryness. Identify which is which **AND** explain why.

Experiment #4:

Nomenclature of Inorganic Compounds and Common Acids

Introduction:

Purpose: To learn the language of chemistry by learning how atoms come together to make ionic compounds. Polyatomic ions are memorized. Techniques such as the “cross-over” rule and common mnemonics will be utilized.

Theory: International Union of Pure and Applied Chemistry (IUPAC) nomenclature rules on ionic compounds.

Objectives:

Learn the difference between an ionic and molecular compound.

Be able to name ionic compounds given compounds and determine empirical formulas from names.

Be able to name molecular compounds given molecules and determine molecular formulas from names.

Memorize the six common acids (acetic, carbonic, hydrochloric, nitric, phosphoric, and sulfuric acid).

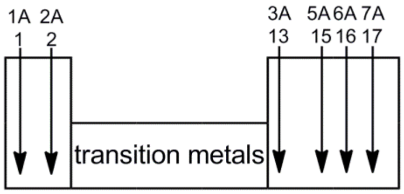
Learn the seven diatomic molecules (hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine).

H — SYMBOL
1 — ATOMIC NUMBER
1.008 — ATOMIC WEIGHT
Hydrogen — NAME

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	**H** 1 1.008 Hydrogen																	**He** 2 4.00 Helium
2	**Li** 3 6.94 Lithium	**Be** 4 9.01 Beryllium											**B** 5 10.81 Boron	**C** 6 12.01 Carbon	**N** 7 14.01 Nitrogen	**O** 8 16.00 Oxygen	**F** 9 19.00 Fluorine	**Ne** 10 20.18 Neon
3	**Na** 11 22.99 Sodium	**Mg** 12 24.31 Magnesium											**Al** 13 26.98 Aluminum	**Si** 14 28.09 Silicon	**P** 15 30.97 Phosphorus	**S** 16 32.07 Sulfur	**Cl** 17 35.45 Chlorine	**Ar** 18 39.95 Argon
4	**K** 19 39.10 Potassium	**Ca** 20 40.08 Calcium	**Sc** 21 44.96 Scandium	**Ti** 22 47.88 Titanium	**V** 23 50.94 Vanadium	**Cr** 24 52.00 Chromium	**Mn** 25 54.94 Manganese	**Fe** 26 55.85 Iron	**Co** 27 58.93 Cobalt	**Ni** 28 58.69 Nickel	**Cu** 29 63.55 Copper	**Zn** 30 65.39 Zinc	**Ga** 31 69.72 Gallium	**Ge** 32 72.61 Germanium	**As** 33 74.92 Arsenic	**Se** 34 78.96 Selenium	**Br** 35 79.90 Bromine	**Kr** 36 83.80 Krypton
5	**Rb** 37 85.47 Rubidium	**Sr** 38 87.62 Strontium	**Y** 39 88.91 Yttrium	**Zr** 40 91.22 Zirconium	**Nb** 41 92.91 Niobium	**Mo** 42 95.94 Molybdenum	**Tc** 43 (97.9) Technetium	**Ru** 44 101.07 Ruthenium	**Rh** 45 102.91 Rhodium	**Pd** 46 106.42 Palladium	**Ag** 47 107.87 Silver	**Cd** 48 112.41 Cadmium	**In** 49 114.82 Indium	**Sn** 50 118.71 Tin	**Sb** 51 121.76 Antimony	**Te** 52 127.60 Tellurium	**I** 53 126.90 Iodine	**Xe** 54 131.29 Xenon
6	**Cs** 55 132.91 Cesium	**Ba** 56 137.33 Barium	**La** 57 138.91 Lanthanum	**Hf** 72 178.49 Hafnium	**Ta** 73 180.95 Tantalum	**W** 74 183.85 Tungsten	**Re** 75 186.21 Rhenium	**Os** 76 190.2 Osmium	**Ir** 77 192.22 Iridium	**Pt** 78 195.08 Platinum	**Au** 79 196.97 Gold	**Hg** 80 200.59 Mercury	**Tl** 81 204.38 Thallium	**Pb** 82 207.2 Lead	**Bi** 83 208.98 Bismuth	**Po** 84 (209) Polonium	**At** 85 (210) Astatine	**Rn** 86 (222) Radon
7	**Fr** 87 223.02 Francium	**Ra** 88 226.03 Radium	**Ac** 89 227.03 Actinium	**Rf** 104 (261) Rutherfordium	**Db** 105 (262) Dubnium	**Sg** 106 (263) Seaborgium	**Bh** 107 (262) Bohrium	**Hs** 108 (265) Hassium	**Mt** 109 (266) Meitnerium									

Experimental Procedure:

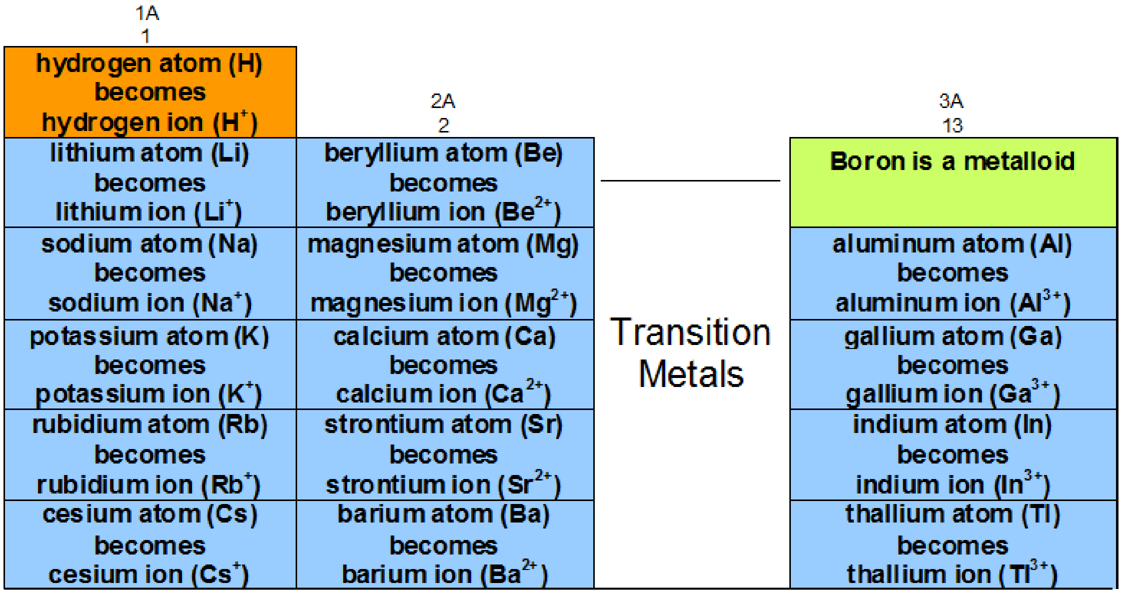
A. Ionic Compounds: Metals and Nonmetals



a. Ionic compounds with main-group metals (Groups 1, 2, and 13):

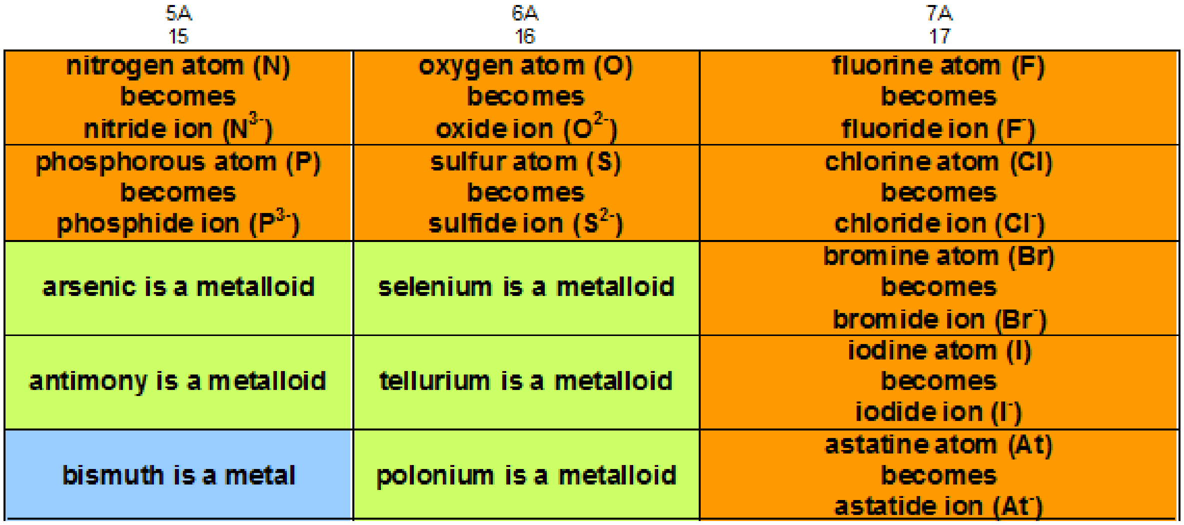
1. When dealing with ionic compounds, we have a metal that will form a 1+, 2+, or 3+ charge depending on what group they are in. This positively charged ion (cation) will get with a nonmetal or a polyatomic ion that is negatively charged (anions). The nonmetals will have a 1-, 2-, or 3- charge depending on what group they are in. The polyatomic ions have charges that are memorized.

2. The metals form specific charged species depending on what group they are in. If the metal is in group 1/1A, then the metal will have a 1+ charge. Group 2/2A metals form 2+ charges and group 3/3A metals form 3+ charges as shown below. Notice the name changes from atom to ion.



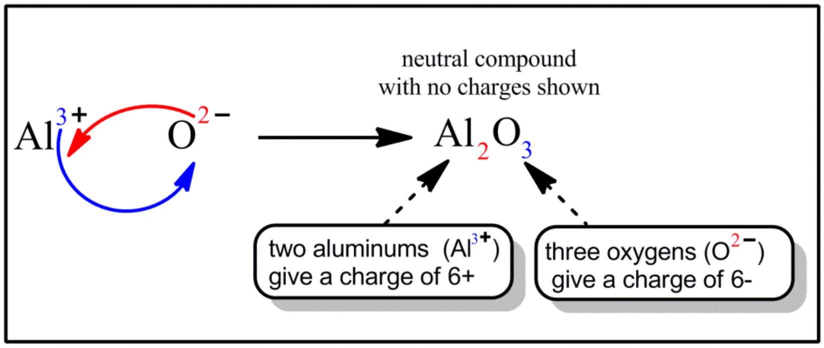
1A 1	2A 2		3A 13
hydrogen atom (H) becomes hydrogen ion (H^+)			
lithium atom (Li) becomes lithium ion (Li^+)	beryllium atom (Be) becomes beryllium ion (Be^{2+})	Transition Metals	Boron is a metalloid
sodium atom (Na) becomes sodium ion (Na^+)	magnesium atom (Mg) becomes magnesium ion (Mg^{2+})		aluminum atom (Al) becomes aluminum ion (Al^{3+})
potassium atom (K) becomes potassium ion (K^+)	calcium atom (Ca) becomes calcium ion (Ca^{2+})		gallium atom (Ga) becomes gallium ion (Ga^{3+})
rubidium atom (Rb) becomes rubidium ion (Rb^+)	strontium atom (Sr) becomes strontium ion (Sr^{2+})		indium atom (In) becomes indium ion (In^{3+})
cesium atom (Cs) becomes cesium ion (Cs^+)	barium atom (Ba) becomes barium ion (Ba^{2+})		thallium atom (Tl) becomes thallium ion (Tl^{3+})

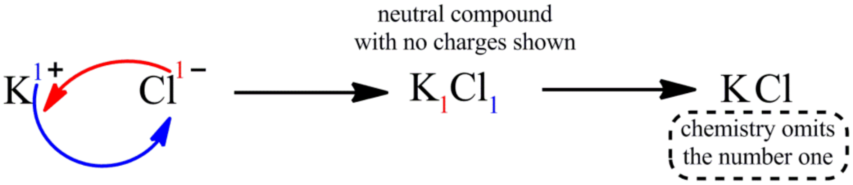
3. The nonmetals form specific charged species as well depending on what group they are in. If the nonmetal is in group 15/5A, then the nonmetal will have a 3- charge. Group 16/6A nonmetals form 2- charges and group 17/7A nonmetals form 1- charges as shown below. Notice the name ending always ends in "-ide". Nitrogen, oxygen, phosphorous and sulfur have unique names.

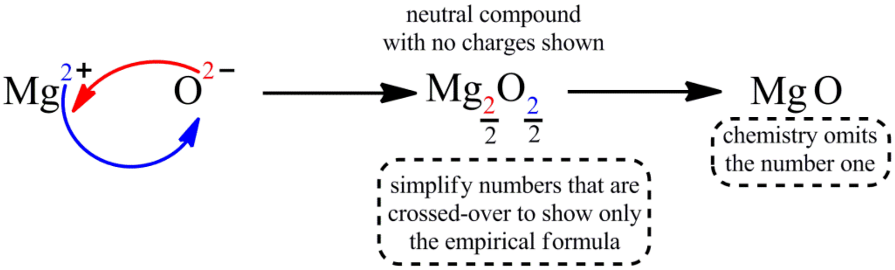


5A 15	6A 16	7A 17
nitrogen atom (N) becomes nitride ion (N^{3-})	oxygen atom (O) becomes oxide ion (O^{2-})	fluorine atom (F) becomes fluoride ion (F^-)
phosphorous atom (P) becomes phosphide ion (P^{3-})	sulfur atom (S) becomes sulfide ion (S^{2-})	chlorine atom (Cl) becomes chloride ion (Cl^-)
arsenic is a metalloid	selenium is a metalloid	bromine atom (Br) becomes bromide ion (Br^-)
antimony is a metalloid	tellurium is a metalloid	iodine atom (I) becomes iodide ion (I^-)
bismuth is a metal	polonium is a metalloid	astatine atom (At) becomes astatide ion (At^-)

4. As metals and nonmetals form their respective charged species, Nature forces them together to balance out their charge differences. This results in the formation of ionic compounds with very high melting and boiling points due to the electrostatic forces between the cation and anion. To make an ionic compound, we need to make sure that the charges balance out to give a neutral compound. We can think about this based on simple charges or with a technique known as the cross-over rule. Metals are **always** listed first!







5. Utilizing the cross-over rule, fill in the following ionic compound formulas:

potassium ion (K^{1+}) and the chloride ion (Cl^{1-})

magnesium ion (Mg^{2+}) and the bromide ion (Br^{1-})

______________________________ ______________________________

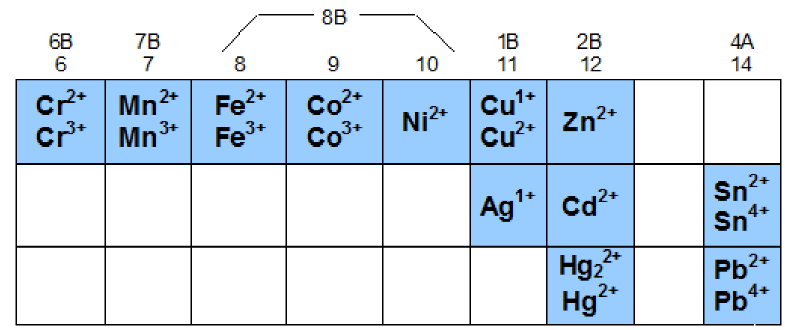
6. Once we have placed the metal and nonmetal together, the charges on each are now cancelled and we have a neutral ionic compound. To name this compound, simply place the metal name in front of the nonmetal name and drop the name ion on each. (E.g. KCl came from the potassium ion and chloride ion forming potassium chloride.) Name the following ionic compounds given the formulas.

$MgBr_2$ ____________________

CaF_2 ____________________

b. Ionic compounds with transition metals:

(Groups 3 through 12 including lead and tin from group 14):



6B	7B	8B			1B	2B		4A
6	7	8	9	10	11	12		14
Cr^{2+} **Cr^{3+}**	**Mn^{2+}** **Mn^{3+}**	**Fe^{2+}** **Fe^{3+}**	**Co^{2+}** **Co^{3+}**	**Ni^{2+}**	**Cu^{1+}** **Cu^{2+}**	**Zn^{2+}**		
					Ag^{1+}	**Cd^{2+}**		**Sn^{2+}** **Sn^{4+}**
						Hg_2^{2+} **Hg^{2+}**		**Pb^{2+}** **Pb^{4+}**

Stock name	**Ion**	**Classical**
chromium(II) ion	Cr^{2+}	chromous ion
chromium(III) ion	Cr^{3+}	chromic ion
cobalt(II) ion	Co^{2+}	cobaltous ion
cobalt(III) ion	Co^{3+}	cobaltic ion
copper(I) ion	Cu^{1+}	cuprous ion
copper(II) ion	Cu^{2+}	cupric ion
iron(II) ion	Fe^{2+}	ferrous ion
iron(III) ion	Fe^{3+}	ferric ion
lead(II) ion	Pb^{2+}	plumbous ion
lead(IV) ion	Pb^{4+}	plumbic ion
manganese(II) ion	Mn^{2+}	manganous ion
manganese(III) ion	Mn^{3+}	manganic ion
mercury(I) ion	Hg_2^{2+}	mercurous ion
mercury(II) ion	Hg^{2+}	mercuric ion
tin(II) ion	Sn^{2+}	stannous ion
tin(IV) ion	Sn^{4+}	stannic ion

1. When forming ionic compounds in which the metal is a transition metal (group 3 through 12), we have a simple rule we have to remember. Transition metals can form various stable positive charges; therefore, the charge on the metal is shown through a roman numeral in parentheses. Ti^{4+} would be the titanium(IV) ion and chromium(II) ion is shown as Cr^{2+}. The cross-over rule still works because we are putting a metal and a nonmetal together. Remember, metals are **<u>always</u>** listed first.

iron(III) ion (Fe^{3+}) and the oxide ion (O^{2-})

____

copper(I) ion (Cu^{1+}) and the oxide ion (O^{2-})

____

2. When naming these ionic compounds, remember to include the Roman numeral in parentheses. $CrBr_3$ comes from the chromium(III) ion and the bromide ion. The name would then be chromium(III) bromide since it is a neutral compound and the ionic charges are now balanced. The numbers brought down can also simplify, but the roman numeral still communicates the charge on the atom, not how many atoms there are.

Fe_2O_3 ____________________

$TiCl_4$ ____________________

CrO ____________________

$CuCl_2$ ____________________

SnF_2 ____________________

FeS ____________________

c. Ionic compounds with polyatomic ions:

1. A lot of the times, when forming ionic compounds, the nonmetals can be grouped together into species we call polyatomic ions. The only way to know these are to memorize the formulas and their charges. With this information, we are able to utilize the cross-over rule as long as we are placing a positively charged species with a negatively charged species.

2. When utilizing the cross-over rule, if the charge number being brought down on a polyatomic ion is greater than one, parentheses must be used to conserve the polyatomic ion. The calcium ion (Ca^{2+}) combining with the hydroxide ion (OH^{1-}) will result in parentheses used to create $Ca(OH)_2$. The polyatomic ions are treated as a single unit when performing the cross-over rule. Convert the following ions into ionic compounds:

magnesium ion (Mg^{2+}) and the hydroxide ion (OH^{1-}) ____________________

copper(I) ion (Cu^{1+}) and the sulfate ion (SO_4^{2-}) ____________________

calcium ion (Ca^{2+}) and the sulfate ion (SO_4^{2-}) ____________________

ammonium ion (NH_4^{1+}) and the phosphate ion (PO_4^{3-}) ____________________

3. When naming ionic compounds containing polyatomic ions, the compounds formed are once again neutral. By treating the polyatomic ions as individual units, they help neutralize the charge on the metal.

$CaCO_3$ ____________________

$Ti(SO_3)_2$ ____________________

$NaHCO_3$ ____________________

$KMnO_4$ ____________________

H_2O_2 ____________________

$NH_4C_2H_3O_2$ ____________________

B. Molecular Compounds: Nonmetal and nonmetal

1. Different from the ionic compounds in section A, nonmetals and nonmetals are forced to shared electrons instead of forming discrete charged species like cations and anions. Since our nonmetals are forming negatively charged species, we are unable to place them together using the cross-over rule. Instead we must communicate through the use of molecular formulas and prefixes. The following prefixes are utilized when showing the number of each nonmetal type.

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

2. When naming molecular compounds, use a prefix in front of each nonmetal. Do not use mono- on the first nonmetal. We call (CO) carbon monoxide, we do <u>**NOT**</u> call it monocarbon monoxide. Using the prefix table above, fill in the following molecular formulas and molecular names. When using oxygen, the -a on the prefix is dropped. BrO_5 would be bromine pentoxide, NOT bromine pentaoxide. There are no other rules that state to drop a vowel when two are placed together. I_2O_3 would be diiodine trioxide. Notice the second nonmetal always ends in "-ide" just like with ionic compounds.

prefix for 1st nonmetal **(do not use mono- prefix)**	**1st nonmetal**	**prefix for 2nd nonmetal**	**2nd nonmetal** **(ending in "-ide")**

CO_2 ______________________

I_2F_5 ______________________

CBr_4 ______________________

NO ______________________

diphosphorous pentoxide ____________

sulfur dioxide ___________

carbon monoxide ____________

phosphorous pentachloride ___________

C. Six common Acids: Hydrogen and Nonmetal or Polyatomic ion

1) **$HC_2H_3O_2$ (aq)** **acetic acid** from hydrogen ion (H^{1+}) and the acetate ion ($C_2H_3O_2^{1-}$)

2) **H_2CO_3 (aq)** **carbonic acid** from hydrogen ion (H^{1+}) and the carbonate ion (CO_3^{2-})

3) **HCl (aq)** **hydrochloric acid** from hydrogen ion (H^{1+}) and the chloride ion (Cl^{1-})

4) **HNO_3 (aq)** **nitric acid** from hydrogen ion (H^{1+}) and the nitrate ion (NO_3^{1-})

5) **H_3PO_4 (aq)** **phosphoric acid** from hydrogen ion (H^{1+}) and the phosphate ion (PO_4^{3-})

6) **H_2SO_4 (aq)** **sulfuric acid** from hydrogen ion (H^{1+}) and the sulfate ion (SO_4^{2-})

D. Seven diatomic molecular elements:

1. The following atoms always exist as diatomic elements. Because they like to pair up with themselves, they are known as molecular elements. Hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine are always diatomic existing as H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. They make a number seven on the periodic table, or you can remember them by the following mnemonics:

(Horses **N**eed **O**ats **F**or **Cl**ear **Br**own **I**'s**)** or **(H**ave **N**o **F**ear **O**f **I**ce **C**old **B**eer**.)**

Horses → **H_2**

Need → **N_2**

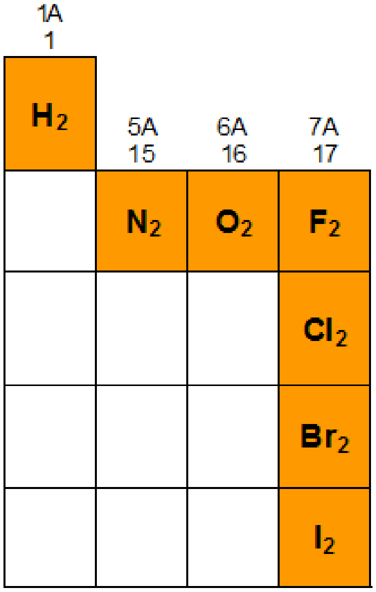
Oats → **O_2**

For → **F_2**

Clear → **Cl_2**

Brown → **Br_2**

I's → **I_2**



Pre-laboratory Assignment:

Memorize the following polyatomic ions by making flashcards. The flashcards can be 3"x5" index cards or pieces of paper cut up into squares. On one side, have the formula with the charge and on the other, have the polyatomic ion name. These flashcards will be checked at the beginning of the lab period.

polyatomic ion	**formula**
ammonium ion	NH_4^{1+}
acetate ion	$C_2H_3O_2^{1-}$ or $CH_3CO_2^{1-}$
hypochlorite ion	ClO^{1-}
chlorite ion	ClO_2^{1-}
chlorate ion	ClO_3^{1-}
perchlorate ion	ClO_4^{1-}
cyanide ion	CN^{1-}
hydrogen carbonate ion OR bicarbonate ion	HCO_3^{1-}
hydrogen sulfite OR bisulfite ion	HSO_3^{1-}
hydrogen sulfate OR bisulfate ion	HSO_4^{1-}
permanganate ion	MnO_4^{1-}
hydroxide ion	OH^{1-}
nitrite ion	NO_2^{1-}
nitrate ion	NO_3^{1-}
peroxide ion	O_2^{2-}
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen phosphate ion	HPO_4^{2-}
phosphate ion	PO_4^{3-}
sulfite ion	SO_3^{2-}
sulfate ion	SO_4^{2-}

Nomenclature Lab

Name:____

Give the formulas of the following compounds:

1. carbon dioxide ________________
2. triphosphorous octaoxide ___________
3. water ________________
4. dinitrogen tetroxide ________________
5. bromine pentafluoride ________________
6. sulfur trioxide ________________
7. strontium oxide ________________
8. lithium iodide ________________
9. calcium fluoride ________________
10. aluminum chloride ________________
11. sodium sulfide ________________
12. aluminum nitride ________________
13. calcium bromide ________________
14. sodium perchlorate ________________
15. potassium cyanide ________________
16. barium hydroxide ________________
17. cesium phosphate ________________
18. ammonium sulfate ________________
19. magnesium acetate ________________
20. calcium oxide ________________
21. sodium hydrogen carbonate ___________
22. ammonium chloride ___________
23. iron(II) sulfate ______________________
24. nickel(II) sulfite ______________________
25. chromium(II) sulfate ________________
26. bismuth(III) nitrate ________________
27. tin(II) chromate ________________
28. lead(II) hypochlorite ________________
29. scandium(III) iodide ___________
30. platinum(IV) cyanide ________________
31. copper(II) carbonate ________________
32. barium chlorate ______________________
33. lithium hydroxide ________________
34. hydrogen peroxide ________________

Give the names of the following compounds:

1. CCl_4 ______________________
2. SO_2 ______________________
3. IF_3 ______________________
4. P_2O_3 ______________________
5. NO_2 ______________________
6. CO ______________________
7. KI ______________________
8. $MgCl_2$ ______________________
9. Na_2O ______________________
10. $SrBr_2$ ______________________
11. Al_2S_3 ______________________
12. $AgCl$ ______________________
13. P_4S_3 ______________________
14. $(NH_4)_2SO_3$ ______________________
15. $NaCN$ ______________________
16. K_3PO_4 ______________________
17. $CaCO_3$ ______________________
18. $Ba(OH)_2$ ______________________
19. $Al(ClO_2)_3$ ______________________
20. Na_3N ______________________
21. $Ca(C_2H_3O_2)_2$ ______________________
22. $HNO_{3\,(aq)}$ ______________________
23. $H_3PO_{4\,(aq)}$ ______________________
24. $CrBr_3$ ______________________
25. PbO ______________________
26. Fe_2O_3 ______________________
27. $CuSO_4$ ______________________
28. $SnCl_2$ ______________________
29. CoS ______________________
30. $Pb(HCO_3)_2$ ______________________
31. $CuClO_2$ ______________________
32. $Fe(C_2H_3O_2)_2$ ______________________

Experiment #5:

Determining the Mass Percentage of Water in an Unknown Hydrate

Introduction:

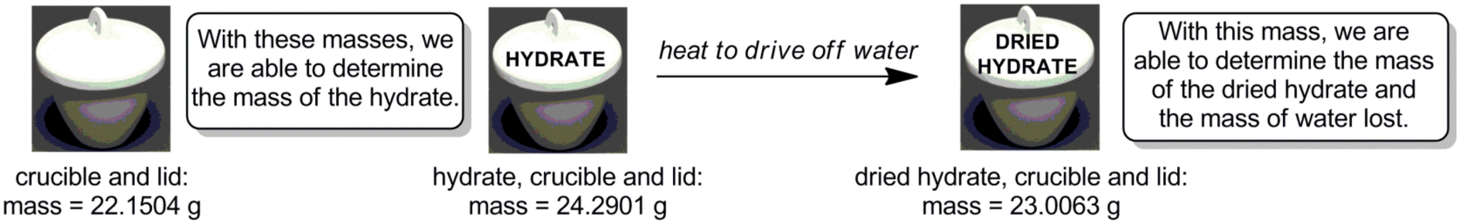
Hydrates:

Hydrates are compounds that have water molecules attached (chelated) to themselves. There are only so many locations water can adhere to a particular compound. This gives us hydrates such as $LiNO_3 \cdot 3H_2O$ and $MgSO_4 \cdot 7H_2O$. Even though they are solids, they still contain water molecules. By heating these substances, the water from the hydrates can be evaporated off and turned into their anhydrous (dried) form. Heating hydrates does not destroy the substance. By simply adding water to the anhydrous compound, the substance can be rehydrated back to its hydrated form.

Mass Percent Composition:

We can determine the amount of water each hydrate contains by driving off the water through heating. A crucible dish is preferred for several reasons. Ceramic (porcelain) can be heated to high temperatures and also heats and cools very rapidly; however, it does not have a very large thermal expansion coefficient. This means it does not take extreme temperature changes well and tends to crack.

Heating the hydrates to drive off the water allows us to measure the mass of the anhydrous form and the mass of water lost.



Comparing this mass to the mass of the hydrate will allow us to determine the mass percentage of water in that particular hydrate.

$$\begin{array}{c}\text{percent (\%)}\\ \text{composition}\\ \text{of water}\end{array} = \frac{\text{mass of water lost}}{\text{mass of entire hydrate}} \times 100$$

Name:____

<u>Pre-laboratory Assignment</u>: You must show your work in order to receive credit.

1. Explain the differences between **$MgSO_4$** $_{(s)}$, **$MgSO_4$** $_{(aq)}$, and **$MgSO_4 \bullet 7\ H_2O$** $_{(s)}$.

2. The picture in the introduction shows that the mass of a crucible and lid was determined to be 22.1504 g. The hydrate was placed into the crucible, and now the crucible with lid and hydrate weighed 24.2901 g. The hydrate in the crucible was then heated and all water was removed. Once cooled, the crucible with lid and dried hydrate weighed 23.0063 g.

    a) Calculate the mass of the hydrate.

    b) Calculated the mass of water lost.

    c) Calculate the mass percentage of water in the hydrate.

$$\begin{array}{c}\text{percent (\%)}\\ \text{composition}\\ \text{of water}\end{array} = \frac{\text{mass of water lost}}{\text{mass of entire hydrate}} \times 100$$

Purpose:

The purpose of this experiment is to determine an unknown hydrate by desiccation. By completely removing all water molecules associated with a hydrate compound, analysis will enable comparisons made through calculation of some possible hydrates. Common lab equipment such as the Bunsen burner and crucible tongs will be mastered as well as the weighing by difference technique.

Theory:

Mass Percent Composition, hydrates (hygroscopic compounds)

% water = (mass of water lost/mass of entire hydrate) x 100

Weighing by difference: final mass - initial mass = total mass

Procedure:

A. Mass Percent of an Unknown Hydrate

1. Obtain a vial containing a pure unknown hydrate. We will be doing two individual trials using the contents in the vial. Write down the unknown number or letter in your data sheet.
2. Set up the apparatus as shown:

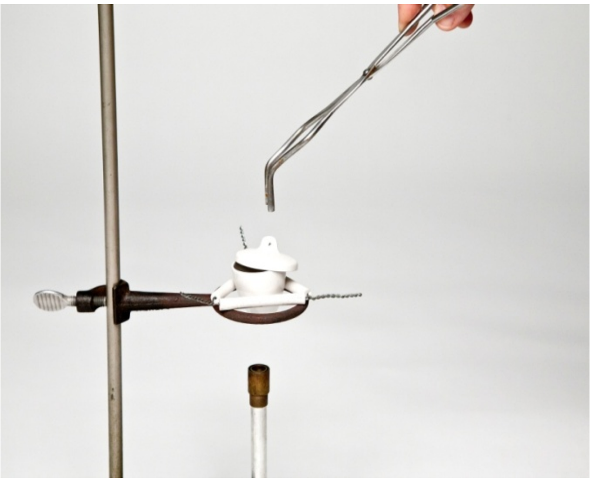
3. Heat the crucible dish to make sure it is completely dry. To keep the crucible dish from cracking, remember to slowly heat the crucible for three to four minutes before placing the Bunsen burner directly under the crucible. Make sure the inner-cone of the flame touches the dish.
4. Get the crucible dish red hot for two minutes.

   Wait for it to cool until you cannot feel any heat radiating from the surface of the dish (place your finger 1 - 2 cm away from the dish to detect any heat).

   Using crucible tongs, determine the mass of the crucible dish **with** a crucible lid using the analytical balance and record this number in your data sheet.

5. Place about two grams of your hydrate into the clean, dry crucible that was just weighed (Fill the crucible only about 1/8 full). Reweigh the crucible with lid and hydrate and record this number in your data sheet.
6. Determine the mass of the hydrate by subtracting #4 from #5 and record this value.
7. Place the crucible dish back onto the apparatus in the middle of the clay triangle. Place the crucible lid on the crucible dish at a slight angle as shown. We want to allow the water to escape as we heat the hydrate, but we do not want to lose any hydrate if it splatters when heated.
8. Slowly heat the bottom of the crucible dish by slowly moving the Bunsen burner back and forth for five minutes. By slowly heating, we are avoiding the water evaporating too quickly and causing the hydrate to splatter outside of the crucible dish.

   Once it appears that the hydrate will not splatter any, strongly heat the crucible dish for ten minutes. Make sure the inner-cone of the flame almost touches the bottom of the crucible dish making it red hot.



9. Remove the heat and allow the crucible dish to completely cool for ten minutes either in the clay triangle or on a wire gauze screen. Using crucible tongs, reweigh the crucible dish with lid and dried hydrate on the same analytical balance used in step 4. Record this number in your data sheet.
10. Determine the mass of dried hydrate by subtracting #4 from #9 and record this value.
11. Calculate the mass of water lost from the unknown hydrate by subtracting #10 from #6 and record this value. Show your work and remember to use proper significant figure rules with measurements.
12. Calculate the percent of water in the unknown hydrate by dividing #11 by #6 and multiply this number by 100. Show your work and record this value to the tenths' place.
13. Repeat all the steps above with the same vial and about 2 grams. Calculate the average and record this value in the data sheet.
14. Place the dried hydrate in the appropriate waste container.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

Mass Percent of an Unknown Hydrate

	TRIAL 1:	TRIAL 2:
1. Unknown number on vials (#):	________________	________________
2. Setting up equipment (no data):		
3. Heating (no data):		
4. Mass of dry crucible dish with lid (g):	________________	________________
5. Mass of dry crucible dish with lid and hydrate (g):	_______________	________________
6. Mass of unknown hydrate (g):	________________	________________
7. Heating (no data):		
8. Heating (no data):		
9. Mass of crucible dish with lid and dried hydrate (g):	______________	________________
10. Mass of dried hydrate (g):	________________	________________
11. Mass of water lost (g):	________________	________________
12. Percent of water in unknown hydrate (%):	________________	________________

13. Average percent of water in hydrate (%): ________________

Data Analysis/Calculations:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

<u>Post-laboratory exercises</u>:

In this exercise, we are going to calculate the percent composition of water in a hydrate and compare it to your experimental values you just obtained in lab. To calculate the percent composition of water in a hydrate, we need to look at our equation.

$$\begin{array}{c}\text{percent (\%)}\\ \text{composition}\\ \text{of water}\end{array} = \frac{\text{mass of water}}{\text{mass of entire hydrate}} \text{ x } 100$$

For $LiNO_3 \cdot 3H_2O$ (lithium nitrate trihydrate) we have to calculate the mass of each element. Notice the three in front of the water will give us a total of six hydrogens and three oxygens.

Li: 1 x 6.94 g/mol
N: 1 x 14.01 g/mol
<u>O: 3 x 16.00 g/mol</u>
68.95 g/mol

H: 6 x 1.01 g/mol
<u>O: 3 x 16.00 g/mol</u>
54.06 g/mol

$$\begin{array}{c}\text{percent (\%)}\\ \text{composition}\\ \text{of water}\end{array} = \frac{3\ H_2O}{LiNO_3 \cdot 3H_2O} \text{ x } 100$$

$$\begin{array}{c}\text{percent (\%)}\\ \text{composition}\\ \text{of water}\end{array} = \frac{54.06 \text{ g/mol}}{(68.95 \text{ g/mol} + 54.06 \text{ g/mol})} \text{ x } 100$$

percent (%) composition of water in $LiNO_3 \bullet 3\ H_2O$ = 43.9 %

1. Calculate the percent composition of water in the following hydrates: (show all work)

$BaCl_2 \cdot 2H_2O$	**$MgSO_4 \cdot 7H_2O$**	**$ZnSO_4 \cdot 7H_2O$**

There are additional questions on the other side.

2. Based on your percent composition obtained in lab, was there a trial that matched up with one of the hydrates calculated on the previous page? Fill in the following table.

percent (%) composition from **Trial #1**.	percent (%) composition from **Trial #2**.	calculated percent (%) composition (from question 1).

If your trial did not match up, what are some possible sources of error that could have led to your experimental numbers being off from the calculated percent composition of the hydrate?

3a. Another student performed the same percent composition of a hydrate experiment, but kept the crucible completely covered throughout the entire experiment. What effect would this have on the experiment?

3b. Would the student's calculated percent water in the hydrate be high, low, or unaffected? Explain why or why not.

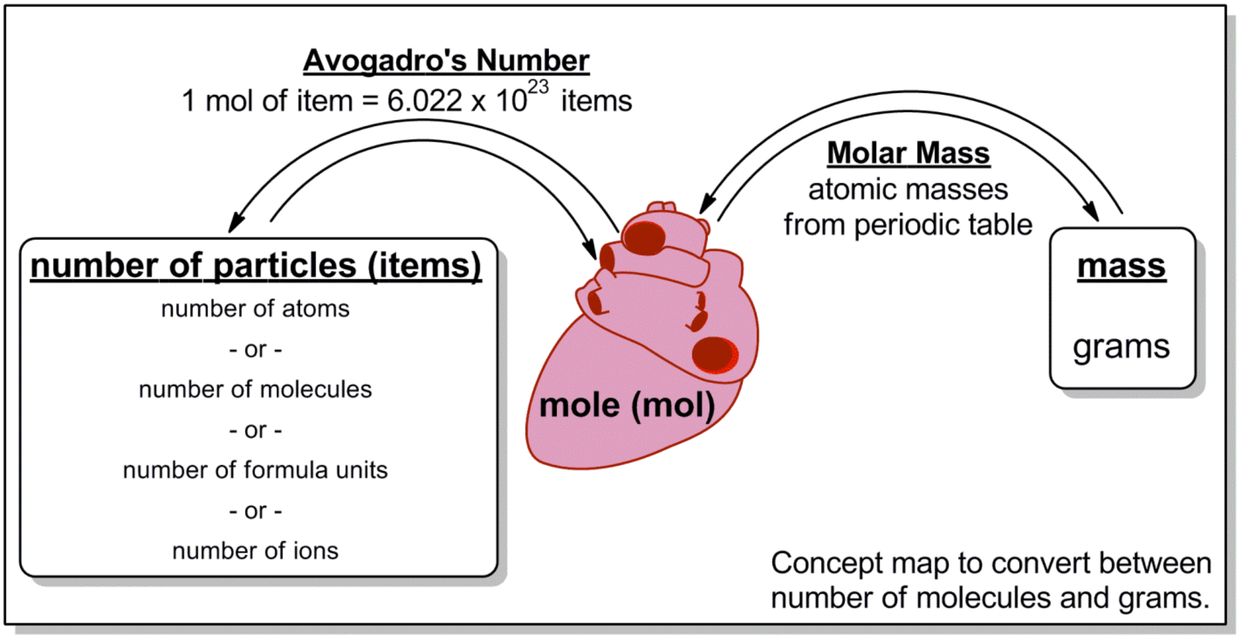
Experiment #6:

Determination of an Empirical Formula

Title: Determining an empirical (experimental) formula.

Introduction:

The masses of substances we measure out for reactions allow us to determine how many molecules are available for the reaction. The chemist's unit, the mole (mol), is at the heart of our concept map seen below. To convert from the mass, we utilize the periodic table and the atomic masses for each atom to calculate the molar mass. This value in (grams/mole) is used to calculate our number of moles. The molecular formula needs to be known to correctly determine the molar mass. This experiment shows how these formulas can be determined.



Concept map to convert between number of molecules and grams.

Use this concept map to help solve the conversions asked in the pre-laboratory assignment.

An example has been shown below:

A large square of aluminum foil weighs 4.508 g. How many aluminum atoms are in the piece of foil?

$$\frac{4.508\ \cancel{\text{g Al}}}{}\ \Bigg|\ \frac{\cancel{\text{mole}}\text{ Al}}{26.98\ \cancel{\text{g Al}}}\ \Bigg|\ \frac{6.022 \times 10^{23}\text{ aluminum atoms}}{1\ \cancel{\text{mole}}\text{ Al}} = 1.006 \times 10^{23}\text{ aluminum atoms}$$

Name:____

Pre-laboratory Assignment: You must show your work in order to receive credit.

1. How many molecules of oxygen are in 50.0 g of oxygen?

2. What is the mass (in grams) of 7.11 X 10^{22} molecules of tetraphosphorous trisulfide?

3. How many water molecules are in 2.00 liters of water? The density of water is 1.00 g/mL.

Purpose:

The purpose of this experiment is to determine an empirical formula by experimentation. A metal will be reacted with air and weighing by difference will be used to determine the oxide created. Common lab equipment such as the Bunsen burner and crucible tongs will be mastered as well as the weighing by difference technique.

Theory:

Experimental Formula, Weighing by difference: final mass - initial mass = total mass

Procedure:

Determination of an Empirical Formula

1. Set up the apparatus as shown:
2. Obtain a crucible dish with lid and place it into the clay triangle at a slight angle. Set the lid slight ajar on the crucible dish. To keep the crucible dish from cracking, remember to slowly heat the crucible for three to four minutes before placing the Bunsen burner under the dish and heat strongly for five minutes. Make sure the inner-cone of the flame almost touches the bottom of the crucible dish making it red hot. Remove the heat and allow the crucible dish to completely cool for ten minutes. Using crucible tongs, weigh the crucible dish and lid on an analytical balance. Record this number in your data sheet.
3. Obtain approximately 0.3 -- 0.5 grams of magnesium ribbon and clean with steel wool or sand paper.
   Roll the magnesium ribbon into a loosely rolled ball and place this into the crucible dish once it has completely cooled. Using crucible tongs, weigh the crucible dish with the magnesium ribbon inside using the analytical balance and record this number in your data sheet.
4. Determine the mass of magnesium ribbon in the crucible dish by subtracting #2 from #3 and record this value.

5. Place the lid slightly ajar on the crucible dish and slowly heat the bottom. If white smoke begins to escape, stop heating and place the crucible lid completely on top of the crucible dish. The white smoke may contain some of the oxide we are trying to trap. After 30 seconds, open the crucible lid and begin slowly heating again. Continue doing this until no white smoke appears. After five minutes of gently warming, heat strongly for ten minutes. Make sure the inner-cone of the flame almost touches the bottom of the crucible dish making it red hot for at least ten minutes. Continue heating until all of the metal can be broken up with a spatula. Most should be turning a white powder.
6. Allow the crucible dish to completely cool by removing the crucible lid and placing it on a wire gauze screen (the crucible dish is cool enough when you cannot feel heat 1 – 2 cm away). Once cooled, slowly add deionized water with a medicine dropper until all of the solid material is covered. This will convert any magnesium nitride ($Mg_3N_{2\,(s)}$) back into magnesium oxide ($MgO_{(s)}$) and ammonia gas ($NH_{3\,(g)}$).
7. Place the crucible dish on a hotplate located in the fumehood. Boil the water off with the crucible lid removed to allow the ammonia gas and water to evaporate. Since the Mg also reacted with H_2O and CO_2 in the atmosphere, this also created $Mg(OH)_2$ and $MgCO_3$. This process also converts any $Mg(OH)_2$ and $MgCO_3$ back to MgO along with H_2O or CO_2 respectively.
8. Using your crucible tongs, remove the crucible dish from the hot plate and continue to heat with the Bunsen burner again with the lid slightly ajar. Remember to gently heat for five minutes followed by strongly heating for another five minutes. Make sure the inner-cone of the flame almost touches the bottom of the crucible dish making it red hot.
9. Allow the crucible dish and lid to completely cool for ten minutes. Using crucible tongs, weight the crucible dish and lid using the same analytical balance as in step #3. Record this number in your data sheet.
10. Determine the mass of magnesium oxide by subtracting #2 from #9 and record this value.
11. Calculate the mass of oxygen reacted by subtracting #4 from #10 and record this value.
12. Calculate the number of moles of magnesium reacted from the amount of magnesium in step #4 using the atomic weight of magnesium (24.31 g/mol).
13. Calculate the number of moles of oxygen reacted from the amount of oxygen in step #11 using the atomic weight of oxygen (16.00 g/mol).
14. Calculate the ratio of magnesium by dividing step #12 by the smaller mole number.
15. Calculate the ratio of oxygen by dividing step #13 by the smaller mole number.
16. Give the empirical formula of magnesium oxide determined experimentally.
17. Place the crucible contents in the appropriate waste container.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

Determination of an Empirical Formula

1. Setting up equipment (no data):
2. Mass of dry crucible dish with lid (g): ________________
3. Mass of dry crucible dish with lid and magnesium (g): ________________
4. Mass of magnesium (g): ________________
5. Heating (no data):
6. Cooling (no data):
7. Heating (no data):
8. Heating (no data):
9. Mass of crucible dish with lid and magnesium oxide (g): ________________
10. Mass of magnesium oxide (g): ________________
11. Mass of oxygen (g): ________________
12. Moles of magnesium (mol): ________________
13. Moles of oxygen (mol): ________________
14. Ratio of magnesium: ________________
15. Ratio of oxygen: ________________
16. Empirical Formula of magnesium oxide: ________________

Data Analysis/Calculations:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

Post-laboratory exercises:

1. Answer the following questions based on the determination of an empirical formula experiment just performed in lab.

   a. During the experiment, water was added to make sure all of the magnesium nitride (Mg_3N_2) had converted to magnesium oxide (MgO). A student performing this lab forgot to add the water to the crucible. Describe, in detail, the effect of failing to add the water on the empirical formula of magnesium oxide. Would this make the magnesium or the oxygen ratio higher, lower, or unaffected? Explain.

   b. A student weighed the crucible dish while it was still warm. Discuss the problem with weighing objects when they are not at the same temperature as the balance.

2. What is the empirical formula for a compound containing 26.57% potassium, 35.36% chromium, and 38.07% oxygen?

3. An unknown compound has a molar mass of 118.10 g/mol. Given the following mass percent composition: 40.7% C, 5.13% H, and 54.2% O, calculate the empirical and molecular formula, respectively.

4. Determine the empirical formula of a compound with the following composition by mass: 48.0 % C, 8.0 % H, 28.0 % N and 16.0 % O. If this compound has a molar mass of 200 g/mol, what is its molecular formula?

Experiment:	Date:
Name:	Course/Section:

Experiment #7:

Observations of Chemical Reactions

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Objectives:

Identify a chemical reaction through the four main types of observations:

- formation of a precipitate (solid forms or solution becomes cloudy)
- formation of a gas (odor, fizzing, bubbles, or bubbles adhering to the side of container)
- permanent color change (lasts longer than one minute)
- release or absorption of energy (feel if the test tube gets hot or cold)

Purpose:

To identify the four main observations associated with a chemical reaction. Evidence of a precipitate (solid) or gas forming, a permanent color change, or release or absorption of energy are signs that the matter is changing composition. When matter changes, we have a chemical reaction that can be expressed by a balanced chemical equation. When matter changes physical states, we have a physical change like melting, freezing, vaporization, condensation, deposition, or sublimation, and the matter is not changing composition. Deposition and sublimation physical changes will be directly observed.

Theory:

Observations of a chemical reaction can communicate when matter is changing composition. Physical changes do not change the composition of the matter.

Procedure:

A. Chemical Changes

a. Reaction A:

1. Label a clean, dry test tube with the letter A and record this letter in your data sheet.
2. Transfer a small amount (tip of a spatula) of $CoCl_2 \bullet 6\ H_2O$ to test tube A.
3. Add approximately 2 mL of 95 % ethanol (C_2H_5OH) to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction A.
6. Rinse and clean your test tube at your work station.

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

b. Reaction B:

1. Label a clean, dry test tube with the letter B and record this letter in your data sheet.
2. Add approximately 1 mL of $CoCl_2 \cdot 6\ H_2O$ in 95 % ethanol (C_2H_5OH) to test tube B. Remember: (1 mL = 15 drops).
3. Using a 10-mL graduated cylinder, add 1 mL of deionized water to test tube B and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction B.
6. Rinse and clean your test tube at your work station.

c. Reaction C:

1. Label a clean, dry test tube with the letter C and record this letter in your data sheet.
2. Transfer a small amount (tip of a spatula) of zinc metal to test tube C.
3. Add approximately 2 mL of 3 M $HCl_{(aq)}$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction C.
6. Rinse and clean your test tube at your work station.

d. Reaction D:

1. Label a clean, dry test tube with the letter D and record this letter in your data sheet.
2. Transfer a small amount (tip of a spatula) of zinc metal to test tube D.
3. Add approximately 2 mL of 0.1 M $CuSO_4$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction D.
6. Rinse and clean your test tube at your work station.

e. Reaction E:

1. Label a clean, dry test tube with the letter E and record this letter in your data sheet.
2. Add approximately 1 mL of 0.1 M $BaCl_2$ to test tube E. Remember: (1 mL = 15 drops).
3. Add approximately 1 mL of 0.1 M $AgNO_3$ to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction E.
6. Rinse and clean your test tube at your work station.

Date:	Instructor Signature:	Comments:

f. Reaction F:

1. Label a clean, dry test tube with the letter F and record this letter in your data sheet.
2. Add approximately 1 mL of 0.1 M $CuSO_4$ to test tube F. Remember: (1 mL = 15 drops).
3. Add approximately 1 mL of 0.5 M $NaHCO_3$ to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction F.
6. Rinse and clean your test tube at your work station.

g. Reaction G:

1. Label a clean, dry test tube with the letter G and record this letter in your data sheet.
2. Add approximately 1 mL of 0.1 M $CuSO_4$ to test tube G. Remember: (1 mL = 15 drops).
3. Add approximately 1mL of 1 M NH_4OH to the test tube and mix well by lightly flicking the bottom of the test tube. NH_4OH is gaseous ammonia ($NH_{3\,(g)}$) dissolved in water. ($NH_{3\,(g)} + H_2O_{(l)} \rightarrow NH_4OH_{(aq)}$)
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction G.
6. Rinse and clean your test tube at your work station.

h. Reaction H:

1. Label a clean, dry test tube with the letter H and record this letter in your data sheet.
2. Transfer a small amount (1 cm) of magnesium metal to test tube H.
3. Add approximately 2 mL of 0.1 M $CuSO_4$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction H.
6. Rinse and clean your test tube at your work station.

i. Reaction I:

1. Label a clean, dry test tube with the letter I and record this letter in your data sheet.
2. Transfer a small amount (1 cm) of magnesium metal to test tube I.
3. Add approximately 2 mL of 3 M $HCl_{(aq)}$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Make sure to carefully feel the test tube. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction I.
6. Rinse and clean your test tube at your work station.

Experiment:	Date:
Name:	Course/Section:

j. Reaction J:

1. Label a clean, dry test tube with the letter J and record this letter in your data sheet.
2. Add approximately 1 mL of 0.1 M $Cu(NO_3)_2$ to test tube J.
   Remember: (1 mL = 15 drops).
3. Add approximately 1 mL of 0.1 M NaOH to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Make sure to carefully feel the test tube. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction J.
6. Rinse and clean your test tube at your work station.

k. Reaction K:

1. Label a clean, dry test tube with the letter K and record this letter in your data sheet.
2. Add approximately 1 mL of 3 M $HCl_{(aq)}$ to test tube K. Remember: (1 mL = 15 drops).
3. Add approximately 1 mL of 0.5 M $NaHCO_3$ to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction K.
6. Rinse and clean your test tube at your work station.

l. Reaction L:

1. Label a clean, dry test tube with the letter L and record this letter in your data sheet.
2. Transfer a small amount (tip of a spatula) of NH_4Cl to test tube L.
3. Using a 10-mL graduated cylinder, add 2 mL of deionized water to test tube L and mix well by lightly flicking the bottom of the test tube.
4. After a minute, feel the test tube. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction L.
6. Rinse and clean your test tube at your work station.

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Chemical Changes

a. Reaction A: $CoCl_2 \cdot 6\ H_2O$ and C_2H_5OH

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

b. Reaction B: $CoCl_2 \cdot 6\ H_2O$ in C_2H_5OH

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

c. Reaction C: Zn and HCl

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

d. Reaction D: Zn and $CuSO_4$

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

e. Reaction E: $BaCl_2$ and $AgNO_3$

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

f. Reaction F: $CuSO_4$ and $NaHCO_3$

1. Letter on test tube: _______________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

g. Reaction G: $CuSO_4$ and NH_4OH

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

h. Reaction H: Mg and $CuSO_4$

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

i. Reaction I: Mg and HCl

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

j. Reaction J: $Cu(NO_3)_2$ and NaOH

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

k. Reaction K: HCl and $NaHCO_3$

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

l. Reaction L: NH_4Cl and H_2O

1. Letter on test tube: ________________
2. Transferring (no data):
3. Mixing (no data):
4. Observation: ____

Date:	Instructor Signature:	Comments:

<u>Post-laboratory exercises</u>: (both sides)

1. Answer the following questions based on your observations from today’s experiment.
    a. List the reactions that produced a precipitate (solid):
    b. List the reactions that produced a gas (bubbles or odor):
    c. List the reactions that resulted in permanent color changes:
    d. List the reactions that resulted in energy changes:
    e. List the reactions that did not react at all:

2. Based on your observations from today’s experiment, can we conclude that precipitates only form when two solutions are mixed? Why or why not?

3. Based on your observations from today’s experiment, can we conclude that the gases released were the same gas each time? Why or why not?

4. Based on your observations from today’s experiment, are all chemical reactions that exchange energies always releasing heat (exothermic). Why or why not?

Post-laboratory exercises:

5. Balance the following equations:

a. $AgI + Fe_3(PO_4)_2 \rightarrow Ag_3PO_4 + FeI_2$

b. $C_2H_6 + O_2 \rightarrow H_2O + CO_2$

c. $KOH + H_2SO_4 \rightarrow K_2SO_4 + H_2O$

d. $(NH_4)_3PO_4 + Pb(NO_3)_4 \rightarrow Pb_3(PO_4)_4 + NH_4NO_3$

Experiment #8:

Chemical reactions and Their Equations

Introduction:

Calcium carbonate (limestone) is manufactured from the shells of marine organisms. Companies such as Tums and Rolaids purchase limestone and sell it as an antacid. Calcium carbonate was also decomposed to calcium oxide and carbon dioxide. The calcium oxide (lime) was burned at plays/theaters to produce a spotlight, hence the phrase (being in the lime light). We represent these chemical reactions through molecular equations. This lab focuses on predicting the products produced as reactants are mixed. Being able to break down the molecular equation to the complete ionic equation allows us to eventually see the main participants through the net ionic equation once all of the spectator ions have been cancelled. Below is an example when aqueous calcium chloride and silver(I) nitrate are mixed.

Molecular Equation: Break apart reactants and switch partners (cations and anions). Remember to balance the equation and determine the physical states by using the solubility rules given.

$$\mathbf{CaCl_{2\,(aq)} + 2\ AgNO_{3\,(aq)} \longrightarrow Ca(NO_3)_{2\,(aq)} + 2\ AgCl_{(s)}}$$

Complete Ionic Equation: Only break apart aqueous (aq) systems.
Leave solids (s), liquids (l), and gases (g) together.

$$\mathbf{Ca^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)} + 2\ Ag^{+}_{(aq)} + 2\ NO_3^{-}{}_{(aq)} \longrightarrow Ca^{2+}_{(aq)} + 2\ NO_3^{-}{}_{(aq)} + 2\ AgCl_{(s)}}$$

spectator ions

Net Ionic Equation: Only bring down the remaining species that are not considered spectator ions.

$$\mathbf{2\ Cl^{-}_{(aq)} + 2\ Ag^{+}_{(aq)} \longrightarrow 2\ AgCl_{(s)}}$$

Compounds containing the following ions are generally soluble	**Exceptions (when combined with ions on the left the compound is insoluble)**		**Compounds containing the following ions are generally insoluble**	**Exceptions (when combined with ions on the left the compound is soluble or slightly soluble)**
Li^+, Na^+, K^+, NH_4^+	none		$OH^–$	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
$NO_3^–$, $C_2H_3O_2^–$	none		$S^{2–}$	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
$Cl^–$, $Br^–$, $I^–$	Ag^+, Hg_2^{2+}, Pb^{2+}		$CO_3^{2–}$, $PO_4^{3–}$	Li^+, Na^+, K^+, NH_4^+
$SO_4^{2–}$	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}			

Objectives:

Identify a chemical reaction through the four main types of observations:

- formation of a precipitate (solid forms or solution becomes cloudy)
- formation of a gas (odor, fizzing, bubbles, or bubbles adhering to the side of container)
- permanent color change (lasts longer than one minute)
- release or absorption of energy (feel if the test tube gets hot or cold)

Be able to identify what molecule is responsible for the chemical observation. This requires knowledge of synthesis, decomposition, single-replacement, and double-replacement reactions.

Purpose:

To make observations of mixing reactants and predicting the products formed through decomposition, double-displacement, and neutralization reactions. Chemical reactions will show reactants forming into products and what observations were made during the experiment. The molecular equations will be balanced, and if a chemical reaction is observed, the complete and net ion equations will also be given.

Theory:

Balanced chemical equations communicate what matter is changing into. Combination, decomposition, double-displacement, neutralization, oxidation-reduction, and single-displacement reactions.

Experimental Procedure:

A. Chemical Reactions

a. Reaction A:

1. Label a clean, dry test tube with the letter A and record this letter in your data sheet.
2. Add approximately 2 mL of 0.3 M Na_2CO_3 to test tube A. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 3 M HCl to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction A.
6. Rinse and clean your test tube at your work station.

b. Reaction B:

1. Label a clean, dry test tube with the letter B and record this letter in your data sheet.
2. Add approximately 2 mL of 3 M KOH to test tube B. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M $Fe(NO_3)_3$ to the test tube and mix well.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction B.
6. Rinse and clean your test tube at your work station.

c. Reaction C:

1. Label a clean, dry test tube with the letter C and record this letter in your data sheet.
2. Add approximately 2 mL of 0.1 M NaCl to test tube C. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.3 M Na_2CO_3 to the test tube and mix well.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction C.
6. Rinse and clean your test tube at your work station.

d. Reaction D:

1. Label a clean, dry test tube with the letter D and record this letter in your data sheet.
2. Add approximately 2 mL of 0.1 M $Pb(NO_3)_2$ to test tube D.
   Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M NaI to the test tube and mix well.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction D.
6. Rinse and clean your test tube at your work station.

e. Reaction E:

1. Label a clean, dry test tube with the letter E and record this letter in your data sheet.
2. Add approximately 2 mL of 0.1 M Na_2SO_4 to test tube E. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M $BaCl_2$ to the test tube and mix well.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction E.
6. Rinse and clean your test tube at your work station.

f. Reaction F:

1. Label a clean, dry test tube with the letter F and record this letter in your data sheet.
2. Add approximately 2 mL of 0.3 M Na_2CO_3 to test tube F. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M $CaCl_2$ to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction F.
6. Rinse and clean your test tube at your work station.

g. Reaction G:

1. Label a clean, dry test tube with the letter G and record this letter in your data sheet.
2. Add approximately 2 mL of 3 M KOH to test tube G. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.3 M $Mg(NO_3)_2$ to the test tube and mix well by lightly flicking the bottom of the test tube.
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction G.
6. Rinse and clean your test tube at your work station.

h. Reaction H:

1. Label a clean, dry test tube with the letter H and record this letter in your data sheet.
2. Add approximately 2 mL of 0.1 M NaOH to test tube H. Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M $Ni(NO_3)_2$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction H.
6. Rinse and clean your test tube at your work station.

i. Reaction I:

1. Label a clean, dry test tube with the letter I and record this letter in your data sheet.
2. Add approximately 2 mL of 3 M HCl to test tube I.
3. Add approximately 2 mL of 1 M $KHCO_3$ to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction I.
6. Rinse and clean your test tube at your work station.

j. Reaction J:

1. Label a clean, dry test tube with the letter J and record this letter in your data sheet.
2. Add approximately 2 mL of 0.1 M $Co(NO_3)_2$ to test tube J.
   Remember: (1 mL = 15 drops).
3. Add approximately 2 mL of 0.1 M NaOH to the test tube and mix well by lightly flicking the bottom of the test tube. Remember: (1 mL = 15 drops).
4. Wait five minutes and record your observation in your data sheet.
5. Place the contents of the test tube in the Waste Container labeled Reaction J.
6. Rinse and clean your test tube at your work station.

Name:____

Letter on test tube: _____A_____

Observations when sodium carbonate and hydrochloric acid were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{Na_2CO_3\ (aq) \quad + \quad HCl\ (aq) \quad \rightarrow}$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____B_____

Observations when potassium hydroxide and iron(III) nitrate were mixed:______________________________________

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{KOH\ (aq) \quad + \quad Fe(NO_3)_3\ (aq) \rightarrow}$$

complete ionic equation:____

net ionic equation:____

Name:____

Letter on test tube: _____C_____

Observations when sodium chloride and sodium carbonate were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{NaCl}_{(aq)} + \mathbf{Na_2CO_3}_{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____D_____

Observations when lead(II) nitrate and sodium iodide were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{Pb(NO_3)_2}_{(aq)} + \mathbf{NaI}_{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____E_____

Observations when sodium sulfate and barium chloride were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{Na_2SO_4\ (aq)\ +\ BaCl_2\ (aq) \rightarrow}$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____F_____

Observations when sodium carbonate and calcium chloride were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{Na_2CO_3\ (aq)\ +\ CaCl_2\ (aq) \rightarrow}$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: ______G______

Observations when potassium hydroxide and magnesium nitrate were mixed:____

Did a reaction occur?____________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{KOH}\ _{(aq)} + \mathbf{Mg(NO_3)_2}\ _{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: ______H______

Observations when sodium hydroxide and nickel(II) nitrate were mixed:____

Did a reaction occur?____________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{NaOH}\ _{(aq)} + \mathbf{Ni(NO_3)_2}\ _{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____I_____

Observations when hydrochloric acid and potassium hydrogen carbonate were mixed:____________________________________

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{HCl}_{(aq)} + \mathbf{KHCO_3}_{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Letter on test tube: _____J_____

Observations when cobalt(II) nitrate and sodium hydroxide were mixed:____

Did a reaction occur?___________ If so, complete the remainder of the data sheet. If not, write **NO REACTION.**

$$\mathbf{Co(NO_3)_2}_{(aq)} + \mathbf{NaOH}_{(aq)} \rightarrow$$

complete ionic equation:____

net ionic equation:____

Mass Stoichiometry Handout

Name____

1. For the reaction shown, find the limiting reactant for each:

$$4\ Al_{(s)} + 3\ O_{2\,(g)} \rightarrow 2\ Al_2O_{3\,(s)}$$

   a. 1.0 g Al and 1.0 g of O_2

   b. 2.2 g Al and 1.8 g of O_2

   c. 0.353 g Al and 0.482 g of O_2

2. For the reaction shown, find the theoretical yield for each:

$$Ti_{(s)} + 2\ F_{2\,(g)} \rightarrow TiF_{4\,(s)}$$

   a. 1.0 g Ti and 1.0 g of F_2

   b. 4.8 g Ti and 3.2 g of F_2

   c. 0.388 g Ti and 0.341 g of F_2

3. If the theoretical yield of a reaction is 0.118 g and the actual yield is 0.104 g, what is the percent yield?

4. Consider the reaction between sulfur trioxide and water:

$$\mathbf{SO_3\,_{(g)} + H_2O\,_{(l)} \rightarrow H_2SO_4\,_{(aq)}}$$

A chemist allows 61.5 g of SO_3 and 11.2 g of H_2O to react. When the reaction is finished, the chemist collects 54.9 g H_2SO_4. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

5. Consider the reaction between HCl and O_2:

$$\mathbf{4\ HCl_{(g)} + O_2\,_{(g)} \rightarrow 2\ H_2O\,_{(l)} + 2\ Cl_2\,_{(g)}}$$

When 63.1 g of HCl react with 17.2 g of O_2, 49.3 g of Cl_2 are collected. Determine the limiting reactant, theoretical yield of Cl_2, and percent yield for the reaction.

6. Magnesium oxide can be produced by heating magnesium metal in the presence of oxygen. The balanced chemical equation for the reaction is:

$$\mathbf{2\ Mg_{(s)} + O_{2\,(g)} \rightarrow 2\ MgO_{(s)}}$$

When 10.1 g of Mg react with 10.5 g of O_2, 11.9 g of MgO are collected. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

7. Classify each process as exothermic or endothermic and indicate the sign of ΔH^o_{rxn}.
    a. ice melting
    b. a sparkler burning
    c. acetone evaporating from skin

8. The equation for the combustion of methane, CH_4 (the main component in natural gas) is shown below. How much heat is produced by the complete combustion of 237 g of CH_4?

$$\mathbf{CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)} \qquad \Delta H^o_{rxn} = -802.3\ kJ/mol}$$

9. The evaporation of water is endothermic:

$$\mathbf{H_2O_{(l)} \rightarrow H_2O_{(g)} \qquad \Delta H^o_{rxn} = 44.01\ kJ/mol}$$

What minimum mass of water (in grams) has to evaporate to absorb 175 kJ of heat?

10. Toilet bowl cleaners often contain hydrochloric acid to dissolve the calcium carbonate deposits that accumulate within a toilet bowl. How much calcium carbonate in grams can be dissolved by 5.8 g of $HCl_{(aq)}$? Begin by writing a balanced chemical equation for calcium carbonate reacting with hydrochloric acid.

11. Many home barbecues are fueled with propane gas ($C_3H_{8\,(g)}$). How much carbon dioxide in kilograms is produced upon the complete combustion of 18.9 L of propane. Assume the density of liquid propane in the tank is 0.621 g/mL. Begin by writing a balanced chemical equation of propane reacting with molecular oxygen. Remember a complete combustion reaction always produces gaseous carbon dioxide and steam.

Name: ____

Experiment #9:

Atomic Spectra and Electron Configuration Handout

Purpose:

The Bohr model will be presented to understand the atomic emission spectrum of hydrogen. Other atom emission spectra will be observed and compared. Emission spectra tubes and spectroscopes will be utilized to determine the atoms excited in a fluorescent tube.

Theory:

Bohr Model, Quantum Mechanics, Emission Spectra, and Electron Configurations

Procedure:

A. Emission Spectra

a. Continuous Spectrum -- White Light

Using the hand-held spectroscope, go outside and observe the continuous spectrum of natural light. Do not point the spectroscope directly at the sun. Pointing towards the sky allows enough light to enter the slit of the spectroscope. A rainbow pattern should be observed. Using colored pencils, color in this pattern with respect to the corresponding wavelengths. Use the wavelength scale on your data sheet.

b. Emission Spectra -- Argon, Helium, Hydrogen, Neon, Krypton, Mercury, and Xenon.

Using the hand-held spectroscope, observe the emission spectrum of argon, helium, hydrogen, neon, krypton, mercury, and xenon. Aim the spectroscope's slit directly at the gas discharge tubes. Pointing towards the light allows enough to enter the spectroscope. Using colored pencils, color in the emission spectrum with respect to the corresponding wavelengths. Use the wavelength scale on your data sheet.

c. Identifying Elements in Fluorescent Light

Using the hand-held spectroscope, observe the emission spectrum of fluorescent light. Aim the spectroscope's slit directly at the fluorescent lights overhead. Pointing towards the light allows enough to enter the spectroscope. Using colored pencils, color in the emission spectrum with respect to the corresponding wavelengths. Use the wavelength scale on your data sheet. Compare this spectrum with those in part B to determine the two main elements in fluorescent light.

B. Electron Configurations

Follow the directions of the electron configuration handout filling out the necessary parts.

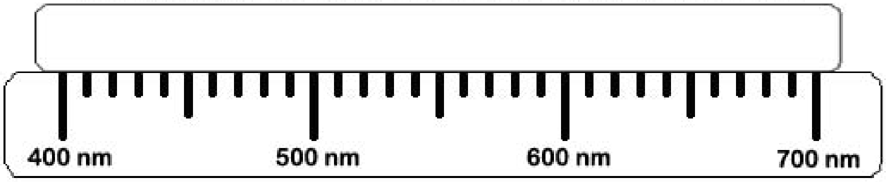
Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Emission Spectra

a. Continuous Spectrum -- White Light

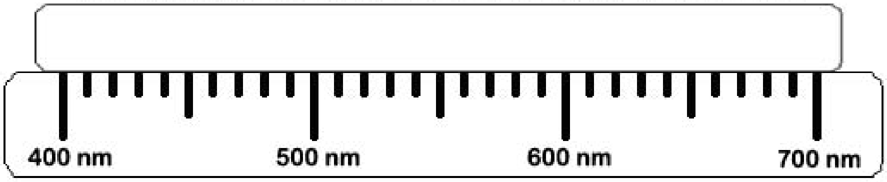
Use colored pencils to fill in the emission spectrum observed.



b. Emission Spectra -- Argon, Helium, Hydrogen, Neon, Krypton, Mercury, and Xenon.

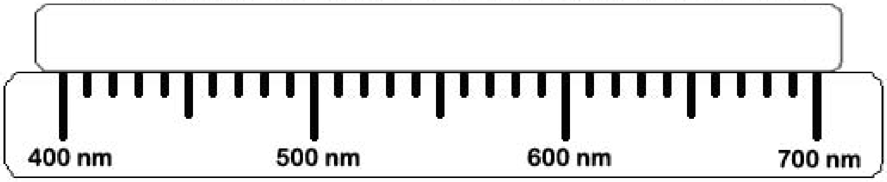
i. **Argon**

Use colored pencils to fill in the emission spectrum observed.



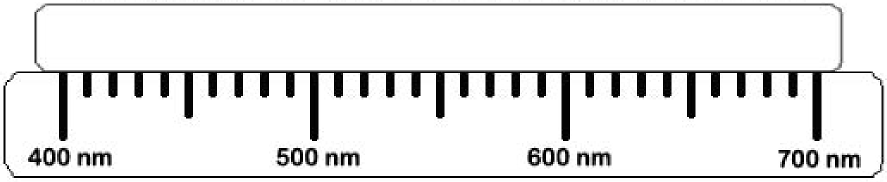
ii. **Helium**

Use colored pencils to fill in the emission spectrum observed.



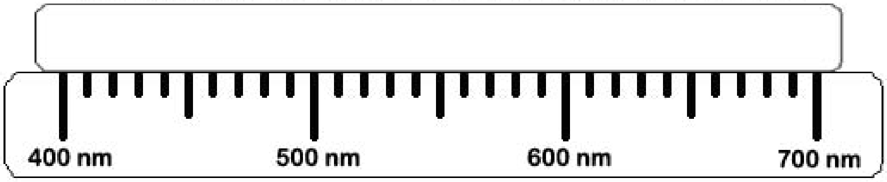
iii. **Hydrogen**

Use colored pencils to fill in the emission spectrum observed.



iv. **Neon**

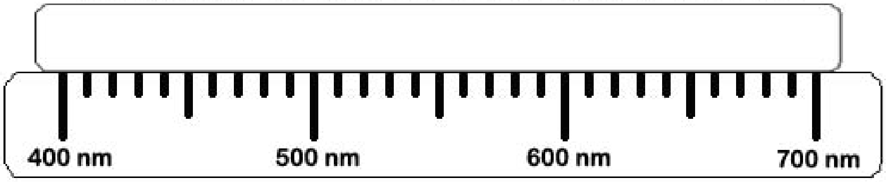
Use colored pencils to fill in the emission spectrum observed.



Experiment:	Date:
Name:	Course/Section:

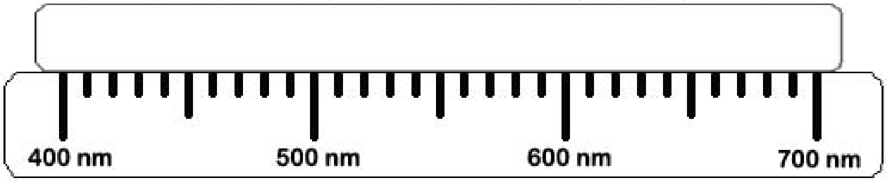
v. **Krypton**

Use colored pencils to fill in the emission spectrum observed.



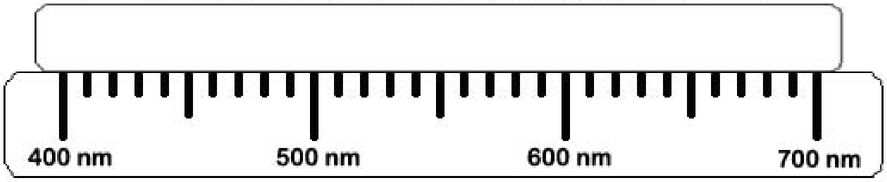
vi. **Mercury**

Use colored pencils to fill in the emission spectrum observed.



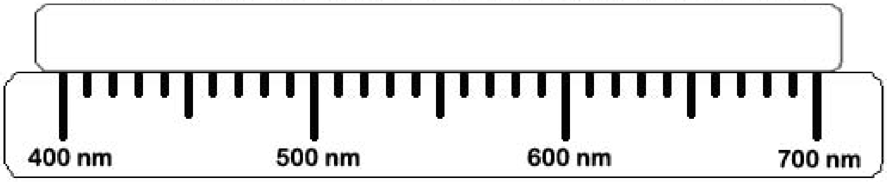
vii. **Xenon**

Use colored pencils to fill in the emission spectrum observed.



c. Identifying Elements in Fluorescent Light

Use colored pencils to fill in the emission spectrum observed.



Compare each individual emission spectra to that of the fluorescent light emission spectrum. Based on your observations in part B, identify the two elements (argon, helium, hydrogen, neon, krypton, mercury, or xenon) found in fluorescent lights.

<u>Introduction</u>:

Certain experiments prove that electrons in atoms occupy specific energy levels. These energy levels form an orderly pattern, and allow us to determine what regions the electrons occupy for any element. Once we know this electron arrangement, also called the electron configuration, we can predict a great deal about how these electrons will behave in chemical reactions.

Electron Configuration energy levels:

Principal Energy Level	The Energy Level has as many sublevels
5	5 sublevels = 5s, 5p, 5d, 5f, 5g
4	4 sublevels = 4s, 4p, 4d, 4f
3	3 sublevels = 3s, 3p, 3d
2	2 sublevels = 2s, 2p
1	1 sublevel = 1s

Electron Configuration sublevels (subshells):

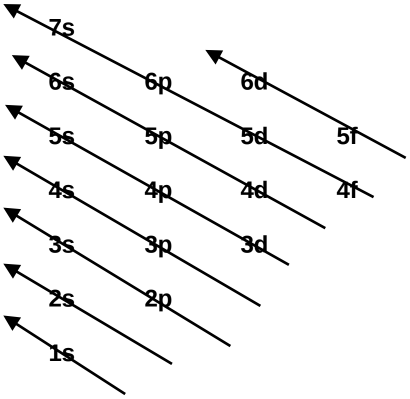
Principle energy levels give how many subshells are allowed and their designation is based on historical and experimental reasons:

1 ...	1s	s = sharp	(*simply*)
2 ...	2s, 2p	p = principle	(*pure*)
3 ...	3s, 3p, 3d	d = diffuse lines	(*darn*)
4 ...	4s, 4p, 4d, 4f	f = fine lines	(*foolishness*)

Question 1: For energy level 6, list the sublevels: ____

Question 2: What sort of experiment(s) have you seen in lecture that indicate that energy levels are discrete instead of continuous? Please explain:

To determine the order of subshells from lowest energy to highest energy (Aufbau Principle), we construct the following table. Construct your own to the right and connect the arrows properly. This will give the correct order of sequence from lowest to highest energy.



give order of sequence:____

subshell have orbitals: s = 1, p = 3, d = 5, f = 7

each orbital can hold two electrons (Pauli Exclusion Principle): s = $2e^-$, p = $6e^-$, d = $10e^-$, f = $14e^-$

energy level	subshell	number of electrons	maximum number of electrons
1	1s	2	2
2	2s	2	2 + 6 = 8
	2p	6	
3	3s	2	2 + 6 + 10 = 18
	3p	6	
	3d	10	
4	4s, 4p, 4d, and 4f	2, 6, 10, and 14	2 + 6 + 10 + 14 = 32

Question 3: Show how many electrons the following subshells can hold. The first has been done for you:

3p ↑↓ ↑↓ ↑↓ _6_ electrons
6s ___electrons
4p ___electrons
3d ___electrons

5f ___electrons
1s ___electrons
4d ___electrons
8p ___electrons

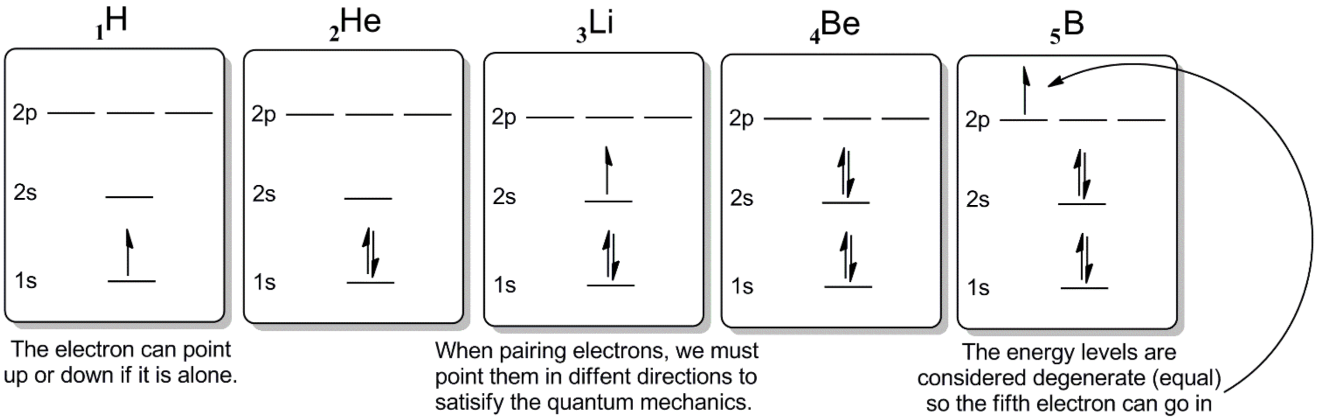
Question 4: What is the maximum number of electrons in energy level 1? ________________________

Question 5: What is the maximum number of electrons in energy level 2? ________________________

Question 6: What is the maximum number of electrons in energy level 3? ________________________

Guidelines for orbital diagrams:

Pauli's Exclusion Principle relates to the quantum numbers that describe an electron's location. As quantum mechanics determines the whereabouts of the electrons in a system, the address cannot be the same for two different electrons. The spin quantum numbers relate to the orientation of the single-headed electrons. We must draw one up and the other down if they have very similar addresses.



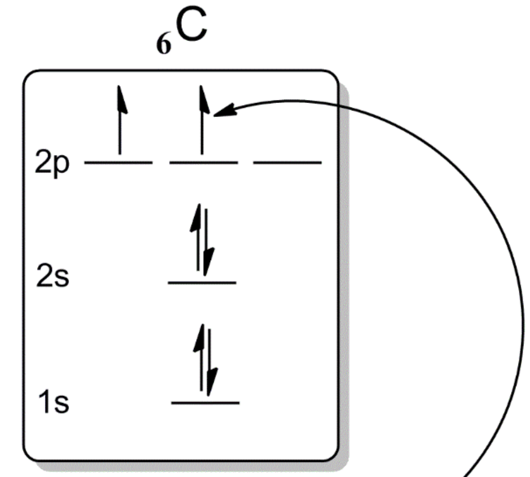
The electron can point up or down if it is alone.

When pairing electrons, we must point them in diffent directions to satisify the quantum mechanics.

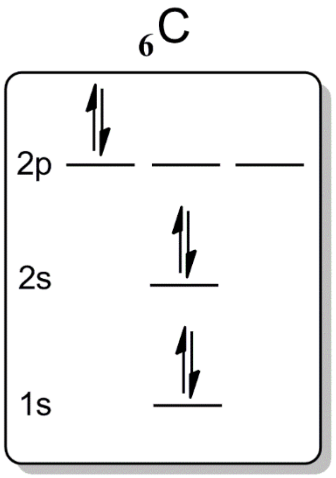
The energy levels are considered degenerate (equal) so the fifth electron can go in any of the three spots.

Hund's Rule:

Hund's Rule dictates that electrons are negatively charged, and we know they tend to spread out; therefore, when we have degenerate energy levels, remember to space the electrons out before pairing them. It does not matter the orientation you draw the arrows if they occupy an orbital by themselves.

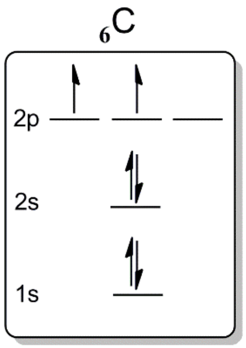


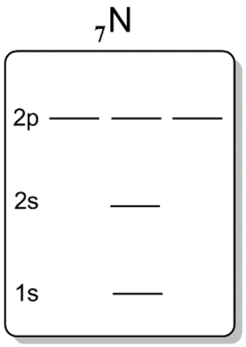
When placing the sixth electron, remember to place into a separate location before pairing.

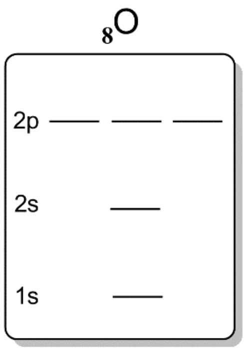


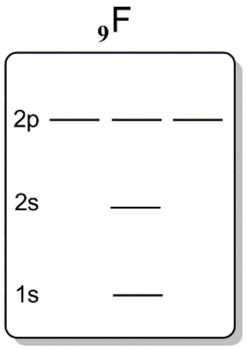
Pairing the electrons is higher in energy and is considered **INCORRECT!**

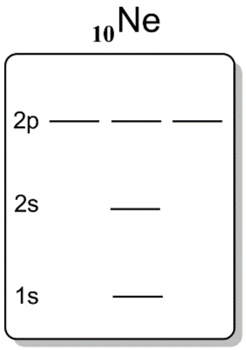
Fill in the remaining diagrams with the corresponding number of electrons following the Pauli Exclusion Principle and Hund's Rule.



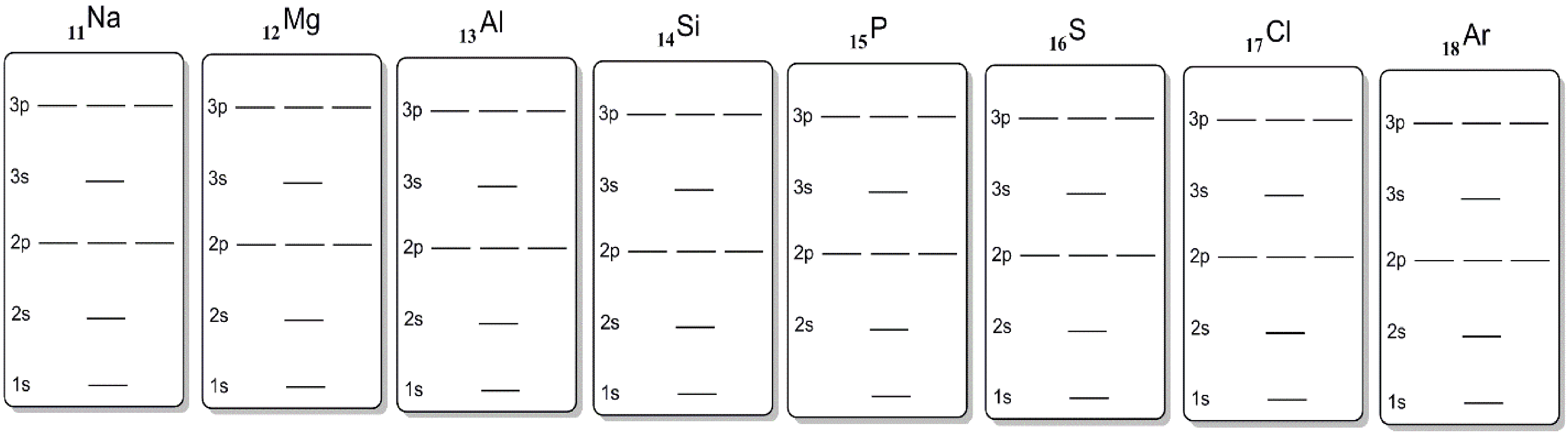








Fill in the remaining diagrams with the corresponding number of electrons following the Pauli Exclusion Principle and Hund's Rule.



Complete the following electron configurations #1-18: (The first five have been done as examples)

$_{1}$H _____$1s^1$

$_{2}$He _____$1s^2$

$_{3}$Li _____$1s^2 2s^1$

$_{4}$Be _____$1s^2 2s^2$

$_{5}$B _____$1s^2 2s^2 2p^1$

$_{6}$C________________

$_{7}$N________________

$_{8}$O________________

$_{9}$F ________________

$_{10}$Ne________________

$_{11}$Na ________________

$_{12}$Mg________________

$_{13}$Al ________________

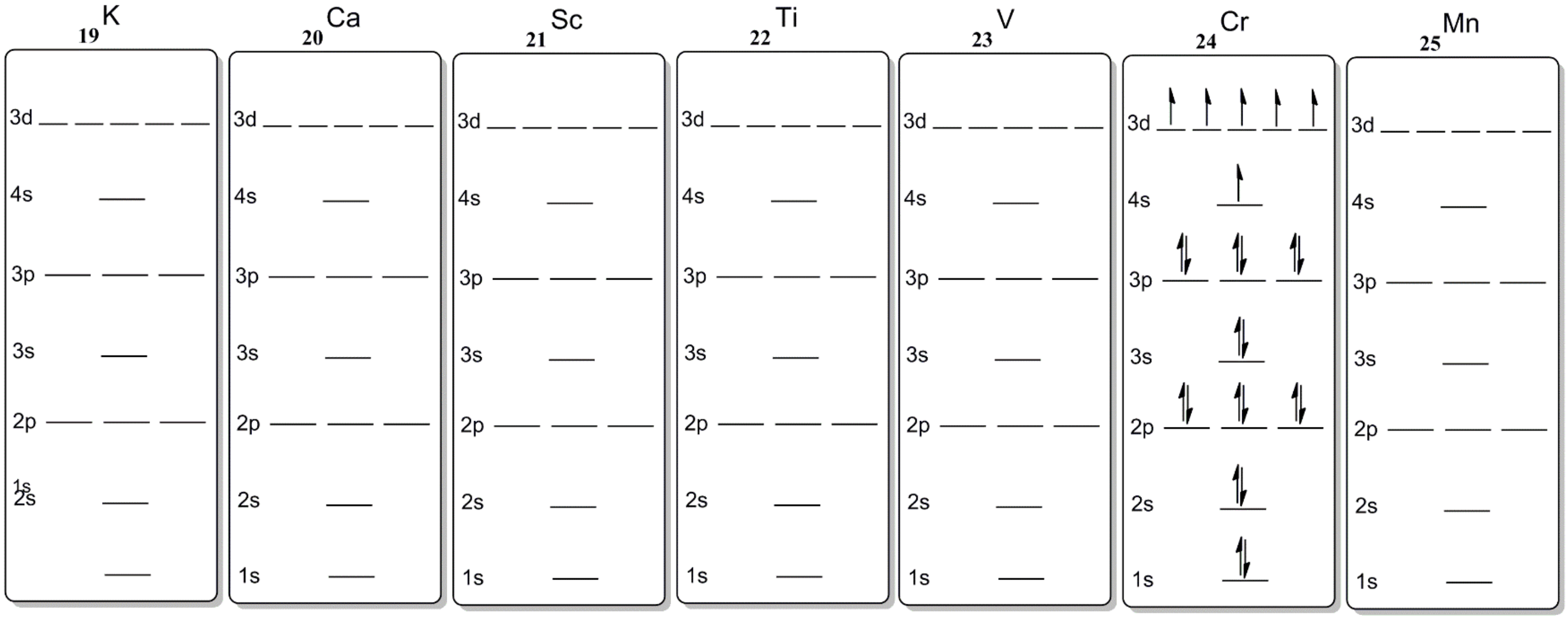
$_{14}$Si ________________

$_{15}$P ________________

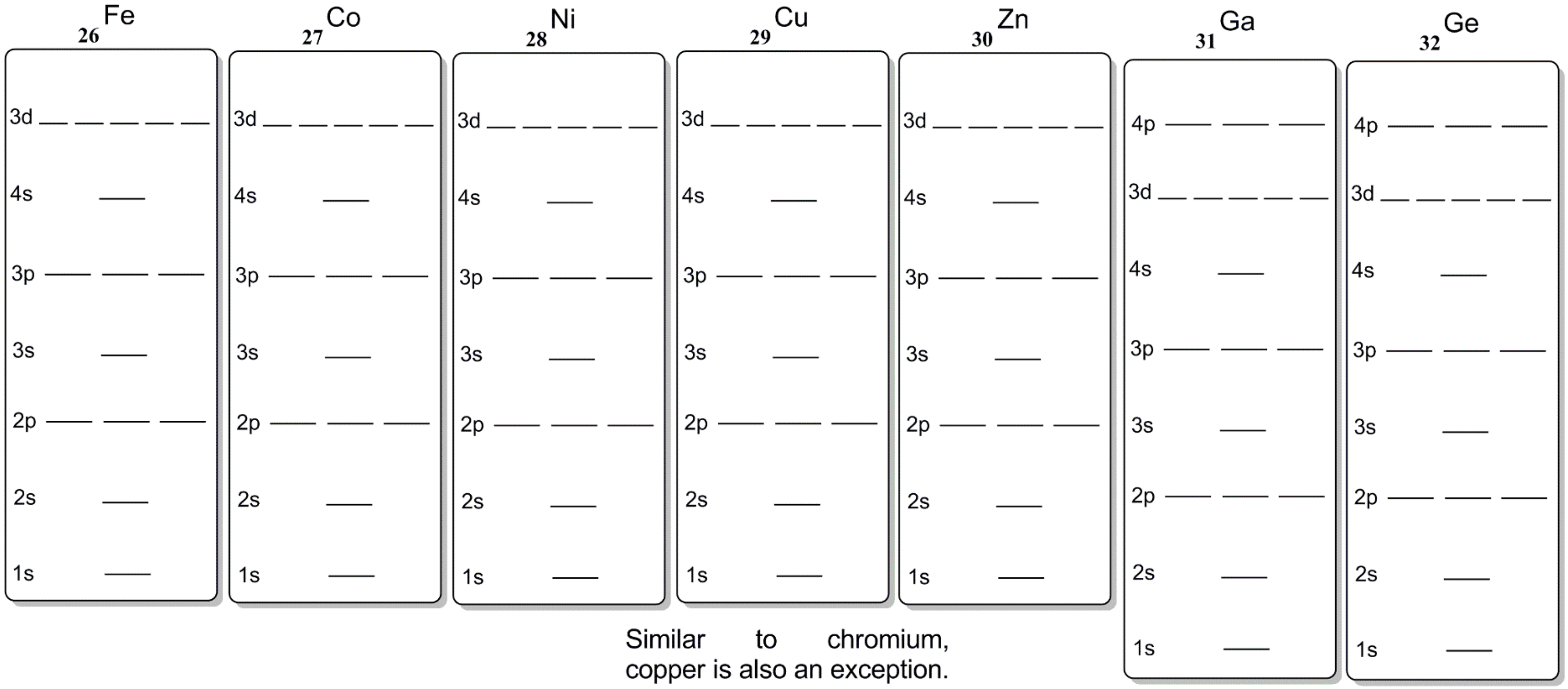
$_{16}$S ________________

$_{17}$Cl ________________

$_{18}$Ar________________



There is extra stability in having a half-filled shell as shown in chromium's 3d energy level. Because of this, an electron will jump from the 4s to the 3d energy level since they are so close in energy. This same phenomenon is observed for half-filled and completely-filled energy levels.



Similar to chromium, copper is also an exception.

From the pages just completed, use the orbital diagrams to help complete the following periodic table. In each empty box, write the symbol of the element and the electron configuration of the outermost sublevel only. A few examples have been done.

1A	2A		3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Na $3s^1$																		
					V $3d^3$													
																	I $5p^5$	
		*	Lu										Tl $6p^1$					
		* *	Lr															

*** Lanthanide series**	La													
****Actinide series**	Ac													

Utilizing the periodic table just completed, see if you observe a pattern in the arrangement of electron configurations. Answer the following questions using the periodic table on the previous page.

1. Which sublevels (s, p, d, or f) are filling up as one goes across a row of elements?

2. Which sublevels (s, p, d, or f) are filling up as one goes across a row of transition elements?

3. Which sublevel is being filled by the lanthanides?

4. Which sublevel is being filled by the actinides?

5. Which sublevel is being filled by the elements of atomic number 104 and 105?

6. Give the electron configurations ($1s^2 2s^2 2p$…) for:

Al:______________________________

Ge:______________________________

S:______________________________

Au:______________________________

Diamagnetism vs. Paramagnetism:

Definitions: **Diamagnetic – No unpaired electrons**

Paramagnetic – One or more unpaired electrons

State whether the following atoms are diamagnetic or paramagnetic:

He B C Ne Ti Fe Zn S Sn

Post-laboratory exercises:

1. What is the electron configuration ($1s^2$, etc.) for sodium (Na)? Identify the valence electrons by placing a box around them.

2. What is the electron configuration ($1s^2$, etc.) for phosphorous (P)? Identify the valence electrons by placing a box around them.

3. What is the electron configuration ($1s^2$, etc.) for Silver (Ag)? Identify the valence electrons by placing a box around them.

4. What is the electron configuration ($1s^2$, etc.) for Barium (Ba)? Identify the valence electrons by placing a box around them.

The noble gas electron configuration uses noble gases in column 8A or 18 to stand for a certain number of electrons. [Ne] = 10 electrons, [Kr] = 36 electrons, and so on. Magnesium, with twelve electrons, is $1s^2 2s^2 2p^6 3s^2$ or we can use noble gas configuration: [Ne] $3s^2$. The [Ne] stands for $1s^2 2s^2 2p^6$ (ten electrons).

5. What is the noble gas electron configuration for Bromine (Br)? Identify the valence electrons by placing a box around them.

6. What is the noble gas electron configuration for Barium (Ba)? Identify the valence electrons by placing a box around them.

7. What is the noble gas electron configuration for Radon (Rn)? Identify the valence electrons by placing a box around them.

Experiment:	Date:
Name:	Course/Section:

Experiment #10: Lewis Dot Structures and Molecular Geometry

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Rules for writing Lewis Dot Structures for covalent molecules:

By Dr. Sylvia Esjornson, Southwestern Oklahoma State University

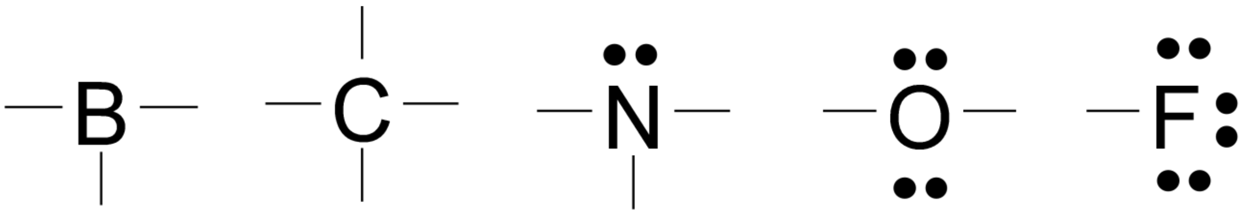
1. Attach the atoms together in a skeletal structure.
    - ✓ Most metallic element is generally central.
        - ➢ In PCl_3, the P is central because it is further left on the periodic table and therefore more metallic.
    - ✓ Halogens and hydrogen are generally terminal.
        - ➢ In C_2Cl_4, the Cs are attached together in the center and the Cls are surrounding them.
    - ✓ Many molecules tend to be symmetrical.
        - ➢ Though there are many exceptions to this, chemical formulas are often written to indicate the order of atom attachment.
        - ➢ In C_2Cl_4, there are two Cls on each C.

2. Calculate the total number of valence electrons available for bonding.
    - ✓ Use group number of periodic table to find number of valence electrons for each atom.
    - ✓ If you have a cation, subtract 1 electron for each + charge.
    - ✓ If you have an anion, add 1 electron for each − charge.

3. Attach atoms with pairs of electrons and subtract electrons used from total.
    - ✓ Bonding electrons.

4. Add remaining electrons in pairs to complete the octets of all the atoms.
    - ✓ Remember H only wants 2 electrons.
    - ✓ Don’t forget to keep subtracting from the total.
    - ✓ Complete octets on the terminal atoms first, then work toward central atoms.

5. If there are not enough electrons to complete the octet of the central atom, bring pairs of electrons from an attached atom in to share with the central atom until it has an octet.
    - ✓ Try to follow common bonding patterns.
    - ✓ C = 4 bonds & 0 lone pairs.
        - ✓ 4 bonds = 4 single, or 2 double, or single + triple, or 2 single + double.
    - ✓ N = 3 bonds & 1 lone pair.
    - ✓ O = 2 bonds & 2 lone pairs.
    - ✓ H and halogen = 1 bond.



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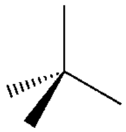
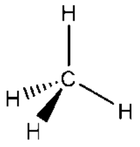
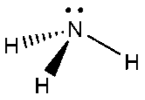
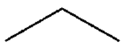
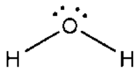
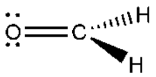
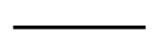
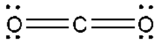
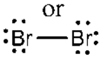
Geometry:

A valid electron-dot structure is a two-dimensional representation of how the atoms are bonded together. This is helpful, but does not give us the complete picture of the molecule's true structure. Geometries of molecules communicate the orientation of the atoms relative to one another in three-dimensions. Knowing the number of bonded and non-bonded groups gives us the geometry around a particular atom. We will focus on molecular geometries shown below. Molecular geometries help to minimize the repulsive forces between the different electron groups. All geometric structures in this experiment fall into the following five patterns:

1. **Tetrahedral**: Four groups of shared electrons (four single bonds) and no nonbonding pairs (lone pairs) around a central atom
2. **Pyramidal**: Three groups of shared electrons (three single bonds) and one lone pair of unshared electrons around a central atom.
3. **Bent**: Two groups of shared electrons (two single bonds) and two lone pairs around a central atom *OR* two groups of shared electrons (one single bond and one double bond) and a lone pair around a central atom.
4. **Trigonal planar**: Three groups of shared electrons (two single bonds and one double bond) and no lone pairs around a central atom.
5. **Linear**: Two groups of shared electrons (two double bonds) and no lone pairs around a central atom *OR* two groups of shared electrons (one single and one triple bond) and no lone pairs around a central atom. When there are only two atoms in a molecule (CO for example), the geometry is also linear. Any triple bond coming off of an atom is always considered linear.

	Molecular Geometries	Bonding Groups	Lone Pairs (nonbonding electrons)	Examples of geometries around central atom
1.	 tetrahedral	$\boxed{4}$ single bonds	$\boxed{0}$ lone pairs	
2.	 pyramidal	$\boxed{3}$ single bonds	$\boxed{1}$ lone pair	
3.	 bent	$\boxed{2}$ single bonds	$\boxed{2}$ lone pairs	
4.	 trigonal planar	2 single bonds 1 double bond	0 lone pairs	
5.	 linear	2 double bonds or a diatomic molecule	N/A	 or 

Date:	Instructor Signature:	Comments:

Experiment:	Date:
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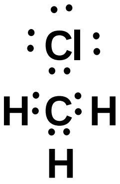
Fill in the following chart using the rules for writing Lewis Dot Structures.

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around central atom(s).
$\mathbf{C}\mathrm{H_3Cl}$	C: 4 x 1 = 4 H: 1 x 3 = 3 Cl: 7 x 1 = 7 Total: 14		tetrahedral
H_2	H:		
Cl_2	Cl:		
HCl	H: Cl:		
HBr	H: Br:		
ICl	I: Cl:		
$\mathbf{C}\mathrm{H_4}$	C: H:		

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use table 10.1)	Electron dot formula	Geometry around central atom(s).
$\mathbf{C}H_2Br_2$	C: H: Br:		
$H\mathbf{O}Br$	H: O: Br:		
$H_2\mathbf{O}_2$	H: O:		O: O:
$\mathbf{N}H_3$	N: H:		
$\mathbf{N}_2H_4$	N: H:		N: N:
$\mathbf{N}H_2CH_3$	N: H: C:		N: C:
O_2	O:		

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use table 10.1)	Electron dot formula	Geometry around central atom(s).
$\mathbf{C}_2H_4$	C: H:		C: C:
$\mathbf{S}O_4^{2-}$	(Add 2 valence electrons for the -2 charge) S: O:		
$\mathbf{N}H_4^+$	(Subtract 1 valence electron for the +1 charge) H: N:		
CN^-	(Add 1 valence electron for the -1 charge) C: N:		
$\mathbf{C}_2H_2$	C: H:		C: C:
$H_3C\mathbf{C}N$	H: C: N:		C: C:
$\mathbf{Si}O_2$	Si: O:		
$\mathbf{C}_3H_8$	C: H:		C: C: C:

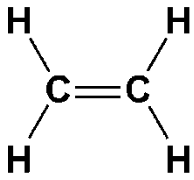
Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

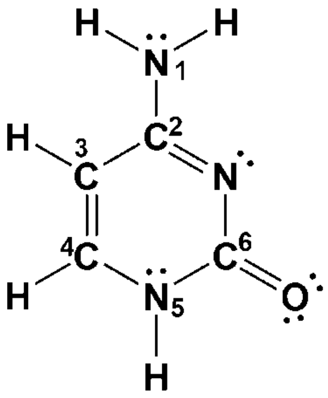
Date:	Instructor Signature:	Comments:

<u>Post-laboratory exercises</u>: (both sides)

1. Ethene is a molecule released by ripening fruit. Identify the geometry around each of the central atoms below:

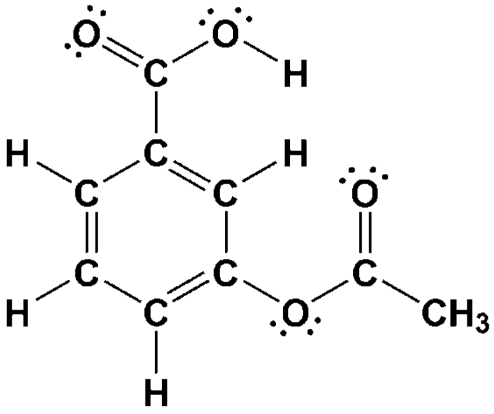


2. Cytosine is a pyrimidine base found in both RNA and DNA. Identify the geometry around each of the central atoms numbered below:



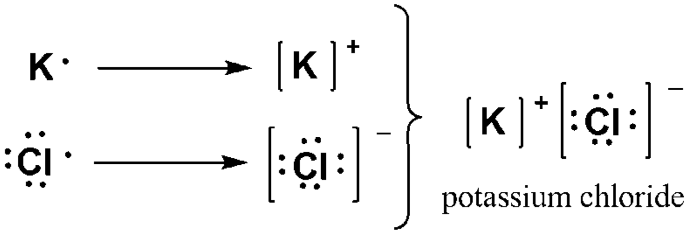
Molecular geometry:
1. _________________
2. _________________
3. _________________
4. _________________
5. _________________
6. _________________

3. Aspirin is a very popular analgesic (pain reliever). Identify the geometry around each of the central atoms below. There are a total of eleven atoms that are central:



Molecular geometry:
1. _________________
2. _________________
3. _________________
4. _________________
5. _________________
6. _________________
7. _________________
8. _________________
9. _________________
10. _________________
11. _________________

Ionic bonds are the interactions between a cation and an anion. Metals tend to form cations giving a discrete positive charge while nonmetals tend to form anions having a negative charge. These charges were developed as the atoms were attempting to satisfy the octet rule; however, Nature prefers matter to be electronically neutral. This results in ionic compounds between metals and nonmetals which have literally transferred their electrons and are now forced together simply based on the electrostatic interaction between positive and negative charges. This electrostatic interaction is so strong in an ionic bond, that melting and boiling points are very high.



potassium chloride

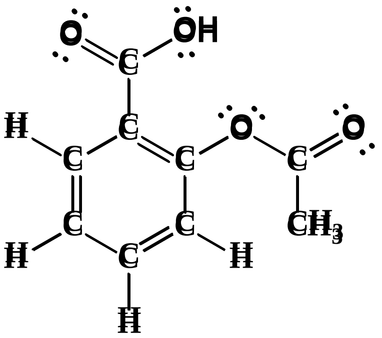
4. Fill in the following table. An example has been given:

Metal	Cation Formed	Nonmetal	Anion Formed	Ionic Compound	Lewis Structure
$\cdot\mathbf{Mg}\cdot$	$[\mathbf{Mg}]^{2+}$	$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\cdot$	$[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:]^{-}$	$MgCl_2$ melting point = 714 °C	$[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:]^{-}[\mathbf{Mg}]^{2+}[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:]^{-}$
$\mathbf{Na}\cdot$		$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}\cdot$		melting point = 801 °C	
$\cdot\mathbf{Ba}\cdot$		$\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{O}}}\cdot$		melting point = 1923 °C	
$\cdot\mathbf{Ca}\cdot$		$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{F}}}\cdot$		melting point = 1380 °C	

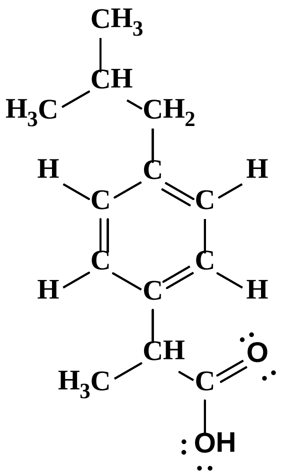
Information for using the Molecular Model Kits:

Element	Color			Number of holes
Carbon	**Black**	**Black**	**Black**	4
Hydrogen	**White**		**White**	1
Oxygen	**Red**	**Red**	**Red**	2
Halogens (F, Cl, Br, or I)	**Green**	**Green**	**Green**	1
Nitrogen	**Blue**	**Blue**	**Blue**	4
Sulfur	**Yellow**	**Yellow**	**Yellow**	4

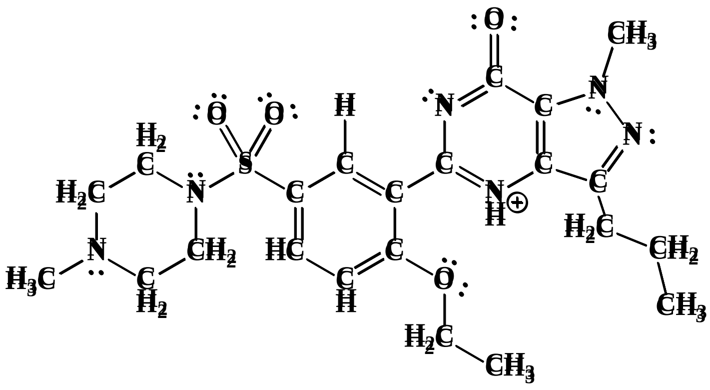
Rigid bonds	Used for making single bonds.
Flexible bonds	Used for making double and triple bonds.



i. aspirin



ii. ibuprofen



iii. viagra

Experiment #11: Molecular Geometry and Electronegativity

Introduction:

Molecules represent nonmetallic atoms sharing electrons which result in covalent bonds. Unlike ionic bonds between a metal and a nonmetal where the electrons are transferred, nonmetals bonded to other nonmetals must share their electrons in order to achieve extra stability. This stability is usually found in molecules that help each atom achieve two or eight electrons to help satisfy the duet and octet rule respectively. Hydrogen will be the most common atom satisfied by the duet rule (two valence electrons). All other nonmetals will follow the octet rule in which eight valence electrons surround the atom. With the octet rule, valid electron-dot structures will be constructed. With these structures we will predict geometry which allows us to make certain predictions about a molecule's polarity.

Valence electrons:

All atoms have a nucleus containing protons and neutrons, and surrounding this nucleus is a region of electron probability. This region contains enough electrons to exactly balance out the positive charge from the protons giving a neutral atom. The electrons that are on the outermost energy level are known as valence electrons. Valence electrons are **transferred** between a metal and a nonmetal when making an ionic compound; however, valence electrons are **shared** between two nonmetals when making a molecule. When constructing valid electron-dot structures, the total number of valence electrons needs to be known for each atom of the molecule. The periodic table is set up to easily determine the number of valence electrons that a particular atom has. The column numbers above the groups correspond to the number of valence electrons (see Table 11.1). Valence electrons are represented as dots surrounding the atom in the 12-, 3-, 6-, and 9-o'clock fashion. Notice the noble gases have eight valence electrons, satisfying the octet rule, and lending explanation for their unique chemical reactivity.

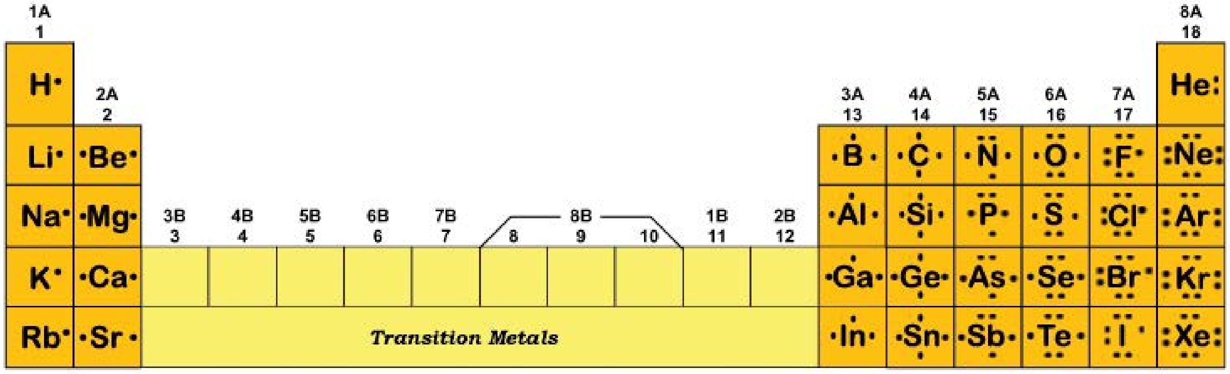


Table 11.1: Valence electron determination from column numbers in periodic table.

Lewis Structures (electron-dot structures):

Electron-dot structures are representations of nonmetal atoms and their attachment to other nonmetal atoms. Valence electrons from each nonmetal atom in the molecule are totaled and spread out over the molecule to help satisfy the duet and octet rules. Since the electrons are shared by the molecule as a whole, it is legal to move valence electrons until a valid electron-dot structure is achieved. Valence electrons between two nonmetal atoms give rise to a covalent bond which represents sharing of electrons. Any valence electrons surrounding an atom and are nonbonding are known as lone pairs as shown in Figure 11.1.

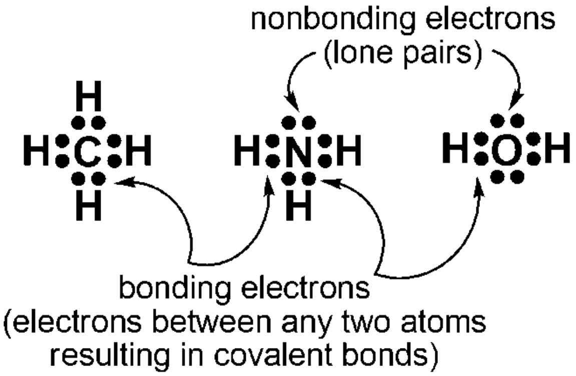


Figure 11.1: **Valence electrons in electron-dot structures representing bonding and nonbonding electrons.**

Line-bond Structures:

The valence electron dots can be tedious to draw and sometimes lead to errors. Once a valid electron-dot structure is found, the bonding electrons can be represented by solid lines. These solid lines now represent a two-electron covalent bond. These lines can be drawn multiple times to give double and triple bonds respectively as shown in Figure 11.2.

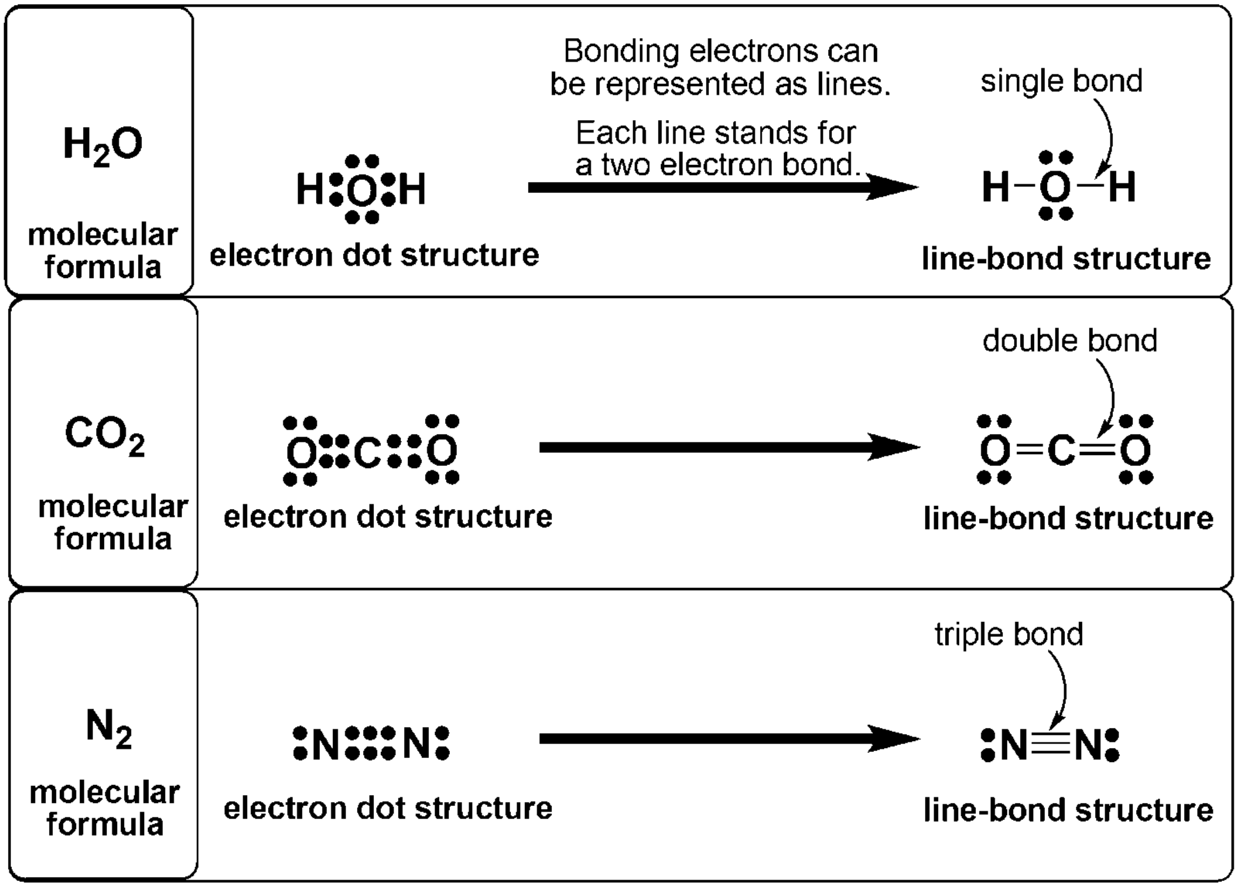


Figure 11.2: Molecules with single, double, and triple covalent bonds. Each line represents two-electron covalent bonds.

Geometry:

A valid electron-dot structure is a two-dimensional representation of how the atoms are bonded together. This is helpful, but does not give us the complete picture of the molecule's true structure. Geometries of molecules communicate the orientation of the atoms relative to one another in three-dimensions. Knowing the number of bonded and non-bonded groups gives us the geometry around a particular atom. We will focus on molecular geometries shown in Table 11.2. Molecular geometries help to minimize the repulsive forces between the different electron groups. All geometric structures in this experiment fall into the following five patterns:

1. **Tetrahedral**: <u>Four</u> groups of shared electrons (four single bonds) and <u>no</u> nonbonding pairs (lone pairs) around a central atom
2. **Trigonal pyramidal**: <u>Three</u> groups of shared electrons (three single bonds) and <u>one</u> lone pair of unshared electrons around a central atom.
3. **Bent**: <u>Two</u> groups of shared electrons (two single bonds) and <u>two</u> lone pairs around a central atom *-OR-* two groups of shared electrons (one single bond and one double bond) and a lone pair around a central atom.
4. **Trigonal planar**: Three groups of shared electrons (two single bonds and one double bond) and no lone pairs around a central atom.
5. **Linear**: Two groups of shared electrons (two double bonds) and no lone pairs around a central atom *OR* two groups of shared electrons (one single and one triple bond) and no lone pairs around a central atom. When there are only two atoms in a molecule (CO for example), the geometry is also linear.

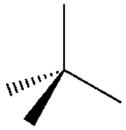
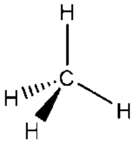
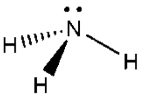
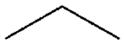
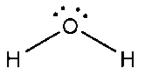
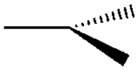
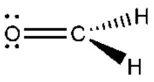
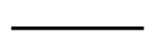
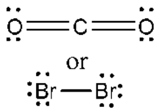
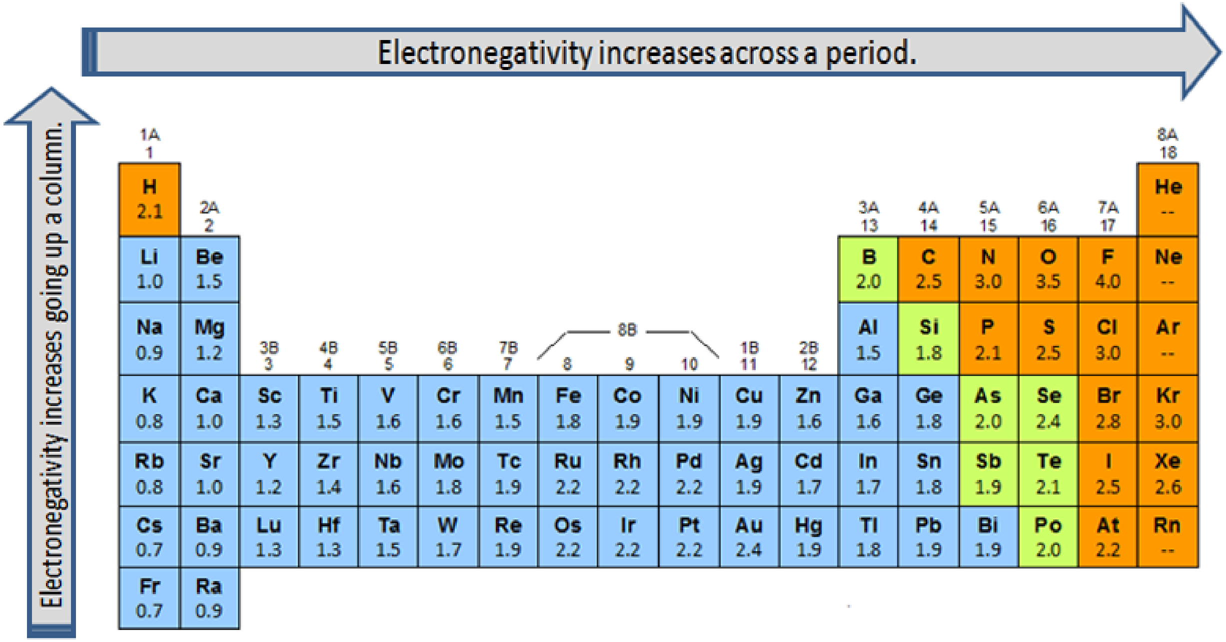
	Molecular Geometries	Bonding Groups	Lone Pairs (nonbonding electrons)	Examples of geometries around central atom
1.	 tetrahedral	$\boxed{4}$ single bonds	$\boxed{0}$ lone pairs	
2.	 pyramidal	$\boxed{3}$ single bonds	$\boxed{1}$ lone pair	
3.	 bent	$\boxed{2}$ single bonds	$\boxed{2}$ lone pairs	
4.	 trigonal planar	2 single bonds 1 double bond	0 lone pairs	
5.	 linear	2 double bonds or a diatomic molecule	N/A	

Table 11.2: Molecular geometries determined by bonding and nonbonding groups around the central atom.

Electronegativity:

Even though covalent bonds are valence electrons being shared between atoms, the electrons may be shared unequally. Since each atom has a nucleus containing positively charged protons, there is desire for the shared electrons. Depending on the atom, some are able to pull electron density towards itself more than others. This ability to attract electrons is known as electronegativity, lending to periodic trends observed on the periodic table. Electronegativity increases as we go up a column and to the right on a period making fluorine the most electronegative atom. The scale shown in Table 11.3 is known as the Linus Pauling's electronegativity scale.



1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 2.1																	He --
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	Lu 1.3	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9																

Table 11.3: Linus Pauling electronegativity numbers.

Bond Polarity:

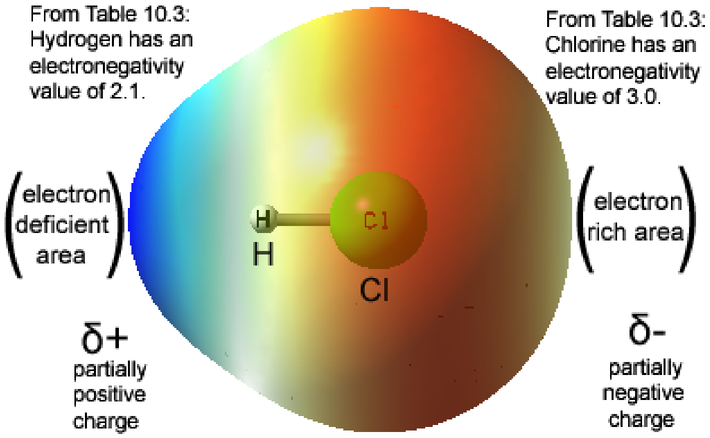
Atoms having different attractive forces for the shared electrons in a covalent bond give rise to a clarification as to what type of covalent bond exists. If the atoms have similar electronegativity, the electrons are shared equally resulting in a nonpolar covalent bond. By subtracting the electronegativity numbers of each atom, we find the electronegativity difference (ΔEN). If that difference is less than 0.4, we consider the type of bond to be a nonpolar covalent bond. If the atoms have a great enough electronegativity difference (0.5 - 1.9), the electrons in the covalent bond will be pulled towards the more electronegative atom. This results in an unequal sharing of electrons giving a polar covalent bond. Table 11.4 shows the electronegativity differences that result in different types of bonds.

Electronegativity difference $\Delta EN_1 - EN_2$	type of bond	notation shown depending on bond type (electronegativity numbers are given below each atom)	
0.0 - 0.4	nonpolar covalent bond	none	H—H 2.1 2.1
0.5 - 1.9	polar covalent bond	delta notation	$\delta+$ $\delta-$ H—Cl 2.1 3.0
> 1.9	ionic bond	explicit charges	$[K]^+[Cl]^-$ 0.8 3.0

Notice the larger electronegative number results in the atom having a partial negative charge (δ^-).

Table 11.4: Electronegativity differences resulting in different types of bonds.

The fact that the electron density of the covalent bond is unevenly shared results in charges on the molecule. These charges are not explicit as in ionic bonds and so are represented as partial charges. The delta symbol (δ) implies partial and each atom of the polar covalent bond will either be partially positively charged (δ^+) or partially negativity charged (δ^-) depending on the electronegativity numbers. The more electronegative an atom, the more attraction it has for electrons; therefore, the atom with the greater electronegativity will be labeled as partially negativity charged (δ^-). An electrostatic potential is mapped onto a molecule to represent electron rich and poor regions as shown in Figure 11.3.



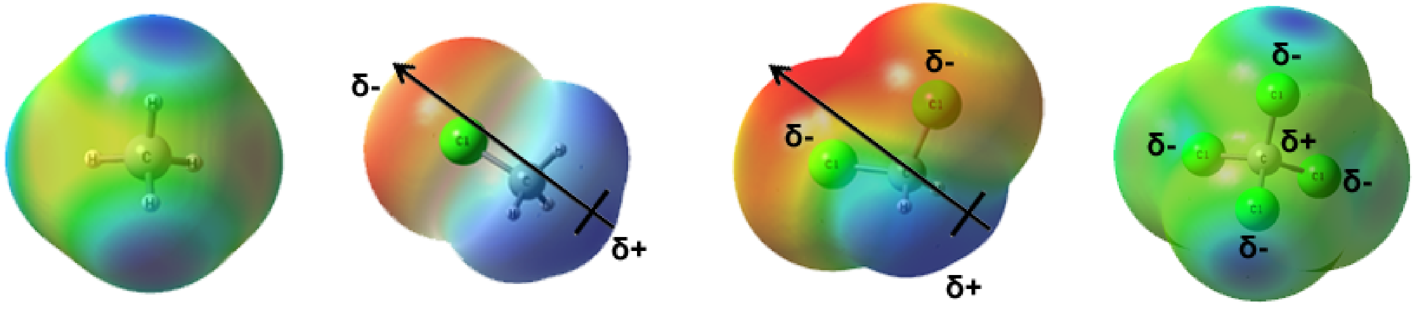
From Table 10.3: Hydrogen has an electronegativity value of 2.1.

From Table 10.3: Chlorine has an electronegativity value of 3.0.

Figure 11.3: Electrostatic potential map of hydrogen chloride (HCl) with red representing electron rich area and blue representing electron deficient region.

Polar versus Nonpolar Molecules:

The bond polarity determines the overall electron distribution between two atoms. Most molecules are larger than two atoms; therefore, each bond must be taken into consideration. The polarity of each individual bond add together like vectors to give the overall polarity of the molecule. If a molecule has an overall greater electron density on one area, a dipole moment is observed and can be represented by a dipole arrow ($+\!\!\!\longrightarrow$). The arrow points to the more negatively charged region of the molecule where the tail represents a positively charged area as shown in Figure 11.4. Most molecules with polar covalent bonds are polar molecules; however, it is possible for dipoles to add together (net dipole) and cancel out resulting in a nonpolar molecule.



All bonds are nonpolar covalent bonds resulting in no net dipole moment so CH$_4$is a nonpolar molecule.

A polar covalent bond pulling in one direction resulting in a net dipole moment so CH$_3$Cl is a polar molecule.

Two polar covalent bonds pulling in unequal directions resutling in a net dipole moment so CH$_2$Cl$_2$ is a polar molecule.

Four polar covalent bonds pulling in equal and opposite directions resulting in no net dipole moment so CCl$_4$ is a nonpolar molecule.

Figure 11.4: Examples of dipoles. Dipoles can be added like vectors and may cancel out.

Objectives:

- Understand electronegativity and its periodic trend.

- Learn how to construct valid electron-dot (Lewis) structures.

- Predict the molecular geometry of a molecule.

- Determine whether a molecule is polar or nonpolar.

Experimental Procedure:

Fill in the table on the data sheet with the following column information. The column begins with a molecule with the central atom bolded.

A. **Number of valence electrons:**
Using Table 11.1, determine the number of valence electrons for each atom and give the total.

B. **Electron-dot formula:**
With the bolded atom in the center of the molecule, distribute the valence electrons as discussed in class and the textbook.

C. **Geometry:**
Build a molecular model and use Table 11.2 to determine the electron groups around the central atom.

1. Tetrahedral 2. Pyramidal 3. Trigonal planar 4. Bent 5. Linear

If there is more than one central atom then identify the geometry around each central atom. Each model that is constructed must be checked by the instructor in groups of four to five.

D. **Line-bond structure and Polarity:**
Bond polarity can be determined by looking up the difference in electronegativity values for both atoms. Use Table 11.3 to see if the electronegativity difference falls into the polar covalent bond range (ΔEN = 0.5 - 1.9). If so, label the bonds with δ^- on the more electronegative atom and δ^+ on the less electronegative atom. If the bond is nonpolar covalent (ΔEN = 0.0 - 0.4), do NOT use delta notation.

E. **Overall Molecular Polarity:**
To determine a molecule's overall polarity, we need to consider the type of bonds and symmetry. If a molecule only has nonpolar covalent bonds, then the molecule as a whole is considered nonpolar. If a molecule contains one or more polar covalent bonds, typically the molecule is polar overall; however, symmetry can cause the polar covalent bonds to cancel out resulting in an overall nonpolar molecule as shown in Figure 11.4.
Model Building:
Use the balls and sticks from the kit provided to build a three-dimensional model of the molecule for each line bond structure shown in the report form.

a. Use the different colored balls in your model kit to represent atoms such as C, H, N, or O. The exact color for each atom is noted in the cover of the box.
b. Use the small, rigid sticks for single bonds
c. Use the flexible connectors for double or triple bonds.

Leave the model together until checked by the instructor. Start building the next model on the list. Your instructor will check the models in groups of four.

NAME:____

Pre-laboratory Assignment:

1. What type of bond is formed when a metal bonds with a nonmetal?

2. What type of bond is formed when a nonmetal bonds with a nonmetal?

3. Refer to Table 11.1 in order to predict the number of valence electrons for each of the following elements:

oxygen: __________ **nitrogen:** ___________ **carbon:** _________

4. What numerical difference in electronegativity results in a nonpolar covalent bond?

5. What numerical difference in electronegativity results in an ionic bond?

6. What numerical difference in electronegativity results in a polar covalent bond?

7. Find the electronegativity (EN) of each atom in the following compounds. Find the difference in electronegativity (ΔEN) between the atoms in each pair. Predict whether the bond between the two atoms is ionic, polar covalent, or nonpolar covalent. An example is provided.

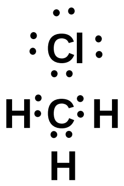
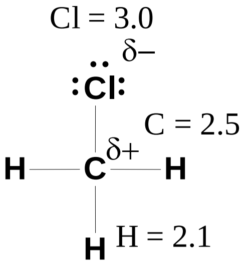
	EN_1 (atom 1)	**EN_2 (atom 2)**	**$\Delta EN = EN_2 - EN_1$**	**Bond type**
K-F				
H-Br				
Cl-F	3.0	4.0	4.0 - 3.0 = 1.0	polar covalent
Na-Cl				
Cl-Cl				

8. What generalizations (increase/decrease) can you make about the trend in electronegativity as you go across a period?

9. What generalization (increase/decrease) can you make about the trend in electronegativity as you go down a column?

Experiment:	Date:
Name:	Course/Section:

Read carefully the procedure for building molecular models and complete the following table. Build the models and have your instructor check it before you disassemble the molecules. Label only polar bonds with partially positive charges (δ^+) and partially negative charges (δ^-).

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use table 11.1)	Electron dot formula	Geometry around central atom(s). (use table 11.2)	Structural formula with bond polarity shown. (use table 11.3)	Is the molecule polar or nonpolar?
CH_3Cl	C: 4 x 1 = 4 H: 1 x 3 = 3 Cl: 7 x 1 = 7 Total: 14		tetrahedral	 ΔEN=3.0 - 2.5 = 0.5 ΔEN=2.5 - 2.1 = 0.4	polar
H_2	H:				
Cl_2	Cl:				
HCl	H: Cl:				
HBr	H: Br:				
ICl	I: Cl:				
CH_4	C: H:				

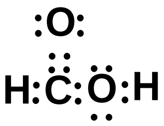
Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use table 11.1)	Electron dot formula	Geometry around central atom(s). (use table 11.2)	Structural formula with bond polarity shown. (use table 11.3)	Is the molecule polar or nonpolar?
$\mathbf{C}H_2Cl_2$	C: H: Cl:				
$H\mathbf{O}Cl$	H: O: Cl:				
$H_2\mathbf{O}_2$	H: O:		O: O:		
$\mathbf{N}H_3$	N: H:				
$\mathbf{N}_2H_4$	N: H:		N: N:		
$\mathbf{N}H_2SH$	N: H: S:		N: S:		
O_2	O:				

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use table 11.1)	Electron dot formula	Geometry around central atom(s). (use table 11.2)	Structural formula with bond polarity shown. (use table 11.3)	Is the molecule polar or nonpolar?
$\mathbf{C}_2H_4$	C: H:		C: C:		
$\mathbf{C}O_3{}^{2-}$	(Add 2 valence electrons for the -2 charge) C: O:				
H**C**OOH	H: C: O:	 $\mathrm{H{:}\overset{:O:}{\overset{..}{C}}{:}\underset{..}{\overset{..}{O}}{:}H}$	C: O:		
N_2	N:				
$\mathbf{C}_2H_2$	C: H:		C: C:		
HO**C**N	H: O: C: N:		O: C:		
$\mathbf{C}O_2$	C: O:				
$\mathbf{C}_3H_4$	C: H:		C: C: C:		

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

Experiment #12:

Calculating Molar Volume at Standard Temperature and Pressure (STP)

Introduction:

The gaseous physical state has infinite volume and shape which is considered an uncondensed state. We are able to apply the Kinetic Molecular Theory of gases which allows us to simplify how we treat gaseous states when performing calculations.

Kinetic Molecular Theory of Gases

1. Gases are in constant motion.
2. Collisions are considered elastic, which means there is no net loss of energy as gaseous particles collide with themselves or the sides of the container.
3. The gaseous state has infinite shape and volume; therefore, there is a lot of empty space between gaseous particles. This also means that gases have low densities.
4. As the temperature increases, so does the speed of the gaseous molecules.

Another set of conditions that are utilized is Standard Temperature and Pressure (STP). This allows us to work at typical laboratory conditions, yet compare our values with those around the world. We can use the combined gas law to correct our experimental data to those of STP conditions (Pressure = 1atm; Temperature = 0 ºC.)

As we plug these values into the Ideal Gas Law (PV = nRT), we get what is known as the Molar Volume at STP.

$$\boxed{PV = nRT}$$

$P = 1$ atm
$V = ?$
$n = 1$ mole
$R = 0.0821$ atm*L/mol*K
$T = 273$ K

Molar Volume

$$V = \frac{22.4\ \text{L}}{1\ \text{mole}}$$

No matter the gas, at these conditions, 1 mole of a gas will take up 22.4 Liters of volume (space).

Name:____

<u>Pre-laboratory Assignment</u>: You must show your work in order to receive credit.

1. The density of air is given as 0.00118 g/mL, and the total volume of air is found to be 105.89 mL. What is the mass of this air in grams?

2. The mass of an "empty" flask with its rubber stopper (the mass of air has already been subtracted) was 104.6164 g when placed on an analytical balance. A piece of dry ice was placed inside the flask and allowed to sublime. The flask, rubber stopper, and carbon dioxide gas were placed on the same balance and obtained a mass of 104.8959 g.

a. What was the mass of the piece of dry ice?

b. Convert this mass of dry ice (carbon dioxide) into moles of carbon dioxide.

c. The volume of the flask was found to be 0.139 L. Calculate the molar volume using the moles determined in part 2b.

(molar volume = L/moles of carbon dioxide)

Objectives:

Use the combined gas law to determine the volume of a gas at Standard Temperature and Pressure (STP).

Determine the molar volume of carbon dioxide by subliming dry ice.

Utilize the laboratory barometer and thermometer and the Handbook of Chemistry and Physics for experimental data.

Purpose:

The molar volume of carbon dioxide will be calculated by subliming dry ice in a fixed volume. The number of molecules will be determined and the volume will be corrected to STP by the combined gas law. The laboratory barometer and thermometer will be used to give the temperature and pressure of the carbon dioxide once it has fully sublimed.

Theory:

Molar volume of (22.4 L of gas/mol of gas) at standard temperature and pressure.

Combined Gas Law: $\frac{P_1 V_1}{T_1} = \frac{P_{STP} V_{STP}}{T_{STP}}$ **percent (%) error** = $\frac{\text{actual} - \text{theoretical}}{\text{theoretical}} \times 100$

Procedure:

A. Molar Volume of Carbon Dioxide (CO_2):

1. Using the laboratory barometer, record the atmospheric pressure on your data sheet.
2. Convert the laboratory pressure to atmospheres (atm).
3. Using the laboratory thermometer, record the temperature on your data sheet.
4. Convert the laboratory temperature to Kelvin (K).
5. Weigh a 125-mL Erlenmeyer flask with a rubber stopper that fits snuggly. Record this number in your data sheet.
6. Obtain a small piece (about the size of a pea) of solid carbon dioxide (dry ice). The size should be less than 0.5 grams, or 10 mm on each side. Place this piece of dry ice into the Erlenmeyer flask just weighed. Do **NOT** stopper the flask. The solid carbon dioxide will sublime into gaseous carbon dioxide displacing the air inside the flask. The carbon dioxide will remain on the bottom of the flask due to its density being 1.977 g/L.
7. Once all of the carbon dioxide has sublimed completely, stopper the flask tightly, and reweigh the 125-mL flask with the stopper and record this number in your data sheet.

8. To determine the total amount of volume in the flask, we will use water. Remove the stopper and completely fill the flask with water to the bottom of where the rubber stopper was. Place the stopper back into the flask and displace some of the water in case you added too much. Pour the water from the flask into a graduated cylinder to determine the total amount of water in milliliters. Depending on the size of the graduated cylinder, multiple pours may be needed. Record the total volume on your data sheet in mL.
9. Record the density of air at the laboratory pressure and temperature. *(CRC Handbook)*

Temperature	**Density of Air**
16 - 17 °C	*0.00122 g/mL*
18 - 19 °C	*0.00121 g/mL*
20 - 22 °C	*0.00120 g/mL*
23 - 24 °C	*0.00119 g/mL*
25 - 27 °C	*0.00118 g/mL*

10. Determine the mass of air in the flask using the volume in #8 and density in #9. Record this value on your data sheet.
11. Determine the mass of the "empty" glass flask with stopper by subtracting the mass of air (#10) from the mass of the flask, air, and stopper (#5).
12. Determine the mass of carbon dioxide by subtracting the mass of the empty flask and stopper (#11) from the mass of the flask, stopper, and CO_2 (#7) on your data sheet.
13. Determine the total moles of carbon dioxide by using the molar mass of carbon dioxide to convert the mass of carbon dioxide in #12 to moles of carbon dioxide. Record this value on your data sheet.
14. To compare our experimental values to conditions at STP, we can use the combined gas law to calculate what the volume would be at those conditions (P_2 = 1 atm; T_2 = 0 °C). To correct the volume of the flask to STP, use the combined gas law and standard temperature (273 K) and standard pressure (1.0 atm) to solve for V_2. Show your work making sure all of your units are consistent, and record this value in mL.

P_1 = laboratory pressure
V_1 = volume of flask
T_1 = laboratory temperature

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

P_2 = standard pressure
V_2 = corrected volume at STP
T_2 = standard temperature

15. Convert the corrected volume to Liters.
16. Calculate the molar volume of carbon dioxide at STP by dividing #15 by #13.
17. Repeat this trial two more times by repeating steps 1 through 15.
18. Determine the average molar volume of carbon dioxide by adding the molar volumes of each trial and dividing by the total number of trials (in this case, three).
19. The molar volume of any gas at STP is 22.4 L/mol. Calculate the percent error for the average molar volume by subtracting 22.4 L/mol from #18. Divide this number by 22.4 L/mol and multiply by 100.

$$\textbf{percent (\%) error} = \frac{\textit{actual} - \textit{theoretical}}{\textit{theoretical}} \times 100$$

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Molar Volume of Carbon Dioxide (CO_2):

	TRIAL 1	*TRIAL 2*	*TRIAL 3*
1. Laboratory barometric pressure (mm Hg or in. Hg):	__________	__________	__________
2. Laboratory barometric pressure (atm):	__________	__________	__________
3. Laboratory temperature (°C):	__________	__________	__________
4. Laboratory temperature (K):	__________	__________	__________
5. Mass of flask, air, and rubber stopper (g):	__________	__________	__________
6. Sublimation (no data):			
7. Mass of flask, rubber stopper, and carbon dioxide (g):	__________	__________	__________
8. Volume of flask (total volume from water measured in graduated cylinder) (mL):	__________	__________	__________
9. Density of air at laboratory temperature and pressure (g/mL):	__________	__________	__________
10. Mass of air (g):	__________	__________	__________
11. Mass of the empty flask with stopper (g):	__________	__________	__________
12. Mass of carbon dioxide (g):	__________	__________	__________
13. Moles of carbon dioxide (mole):	__________	__________	__________
14. Corrected volume of flask to STP (mL):	__________	__________	__________
15. Corrected volume of flask to STP (L):	__________	__________	__________
16. Molar volume of carbon dioxide at STP (L/mol):	__________	__________	__________
17. Repeat steps for Trials 2 and 3:			
18. Average molar volume (L/mol):		__________	
19. Percent Error (%):		__________	

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Data Analysis/Calculations:

Date:	Instructor Signature:	Comments:

Post-laboratory exercises:

Substance	Density of gas at STP
air (78 % N_2, 21 % O_2)	1.279 g/L
$CO_{2\ (g)}$	1.977 g/L

1. The table above gives the density of CO_2 gas and air at STP. Explain the assumption that the $CO_{2\ (g)}$ does not need to be capped while subliming and that air in the flask is being displaced as the $CO_{2\ (g)}$ concentration increases.

2. Performing this same experiment, what would happen if a larger amount of dry ice was used? Would this produce a larger, smaller, or have no effect on the calculated molar volume of CO_2? Please explain.

3. In the middle of the experiment (after obtaining the original mass of the flask, stopper, and air) a stopper with a smaller mass was used to obtain the new mass with all of the dry ice sublimed. What would happen to the calculated molar volume of CO_2?

4. A patient at a nearby hospital at sea level (1.00 atm) with a resting lung volume of 1.20 L is asked to take a deep breath. As the intercostal muscles pull up and the diaphragm lowers, their lungs open to a capacity of 4.85 L. What is the pressure in the lungs at this exact moment?

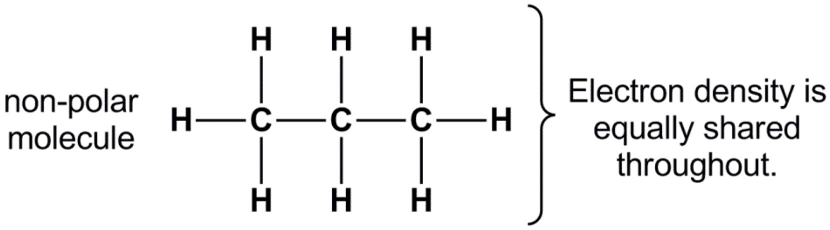
5. The initial volume of a balloon is 118 mL. If the pressure of the gas inside the balloon changes from 755 mm Hg to 709 mm Hg, what is the final volume of the balloon?

Experiment #13:

Solubility

Introduction:

As our atoms come together to make larger molecules, they take on a certain molecular geometry (shape). As the valence electrons are placed onto our Lewis structure, these electrons can be influenced by what atoms make up the molecule. If the electron density is equally shared we refer to that molecule as non-polar. Electronegativity determines if an atom pulls electrons towards itself. We determine this by measuring the electronegativity difference (ΔEN) between two atoms. As we subtract electronegativity values from a carbon to a hydrogen bond, we obtain an ΔEN = 0.4. Having an electronegativity difference value between (0 - 0.4) means that neither atom is electronegative relative to the other, and the electron density is equal throughout the entire bond giving us a non-polar bond. If all of our bonds are considered non-polar, then the molecule as a whole is considered non-polar.



The intermolecular force that is present between two non-polar molecules are called London or dispersion forces. Even though the electrons are not explicitly pulled to one direction by electronegative atoms, the electron density is able to move around the molecule. This results in a temporary dipole. If a molecule has an overall greater electron density on one area, a dipole moment is observed and can be represented by a dipole arrow (+→). The arrow points to the more electron rich area of the molecule where the tail represents the electron deficient area. Since the electron density resides on one side of the molecule this induces dipoles in nearby molecules allowing there to be some small electrostatic interaction between the molecules. This electrostatic interaction is the "glue" holding these molecules near and is referred to as London or dispersion forces. These forces increase with size and surface area depending on how much the molecules are able to interact (touch).

Permanent dipoles result from a molecule having a polar covalent bond. An electronegativity difference (ΔEN = 0.5 - 1.9) communicates that the electron density is unequally shared across that bond. Most molecules with polar covalent bonds are polar molecules; however, it is possible for dipoles to cancel out based on their molecular geometry resulting in a non-polar molecule. With a polar molecule, the electron density is unequally shared, and can be represented with a dipole arrow (+→). Simple electrostatics states that positive charges are attracted to negative charges; therefore, Dipole-Dipole forces are the intermolecular force resulting from permanent dipoles. These dipoles become more extreme when one atom is a hydrogen and the other is something very electronegative (Oxygen, Nitrogen, or Fluorine). These bonds are referred to as a super-dipoles and result in molecules having an interaction known as Hydrogen-Bonding. Both Dipole-Dipole and Hydrogen-Bonding forces exist in polar molecules and are both able to interact through electrostatic forces.

This brings us to an important concept known as the "LIKE dissolves LIKE" rule. Because non-polar molecules exhibit London forces and polar molecules exhibit Dipole-Dipole forces, these types of forces are actually unable to interact. The electrostatic differences between the molecules doesn't provide enough positive and negative interactions and they will not be able to mix into a homogeneous solution.

So, **<u>POLAR dissolves POLAR</u>**: molecules with Dipole-Dipole and Hydrogen Bonding are both polar and are able to interact and mix together.

<u>NONPOLAR dissolves NONPOLAR</u>: molecules with London forces are both non-polar and are able to interact and mix together.

If our substances are able to interact and form a homogeneous solution, they are referred to as miscible. If the substances are unable to interact like olive oil (non-polar) and water (polar), then the substances are unable to mix and are said to be immiscible. We will be able to see a separation of the layers.

Name:____

Pre-laboratory Assignment: You must show your work in order to receive credit.

1. Answer the following questions about the following molecules:
   $\mathbf{CH_3CH_2CH_2CH_2CH_2CH_3}$ and $\mathbf{Br_2}$

   a. What is the strongest intermolecular force present in $\mathbf{CH_3CH_2CH_2CH_2CH_2CH_3}$?

   b. Is $\mathbf{CH_3CH_2CH_2CH_2CH_2CH_3}$ polar or nonpolar? (EN: C = 2.5, H = 2.1)

   c. What is the strongest intermolecular force present in $\mathbf{Br_2}$?

   d. Is $\mathbf{Br_2}$ polar or nonpolar? (EN: Br = 2.8)

   e. Are $\mathbf{CH_3CH_2CH_2CH_2CH_2CH_3}$ and $\mathbf{Br_2}$ miscible? Why or why not?

2. Answer the following questions about the following molecules:
   $\mathbf{CH_3OH}$ and $\mathbf{H_2O}$

   a. What is the strongest intermolecular force present in $\mathbf{CH_3OH}$?

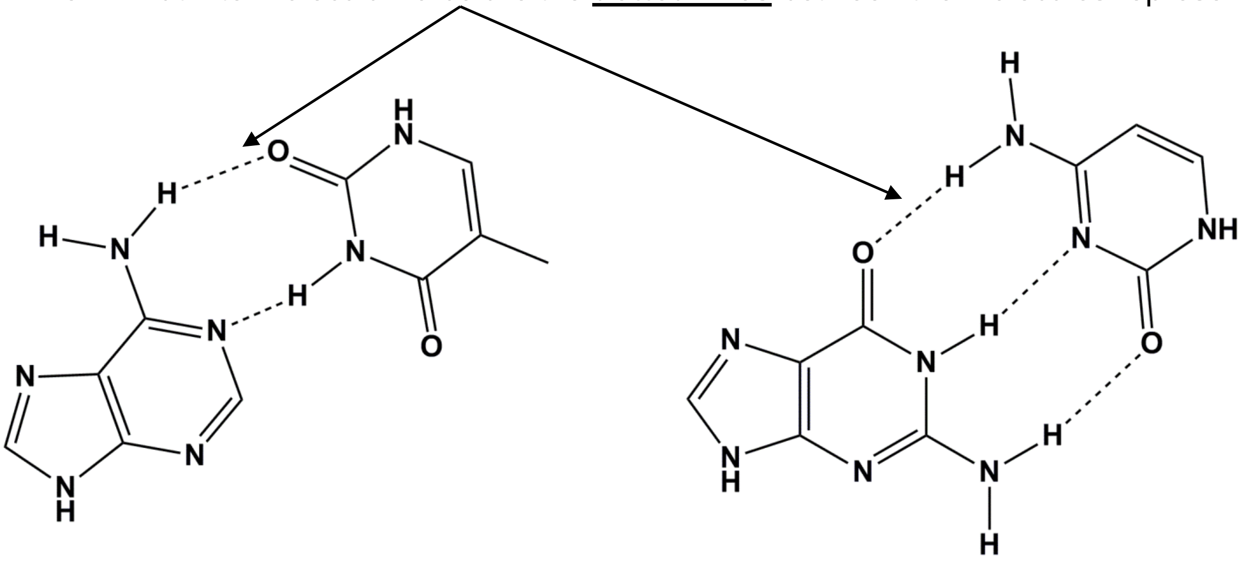
   b. Is $\mathbf{CH_3OH}$ polar or nonpolar? (EN: C = 2.5, H = 2.1, O = 3.5)

   c. What is the strongest intermolecular force present in $\mathbf{H_2O}$?

   d. Is $\mathbf{H_2O}$ polar or nonpolar? (EN: O = 3.5, H = 2.1)

   e. Are $\mathbf{CH_3OH}$ and $\mathbf{H_2O}$ miscible? Why or why not?

3. What intermolecular force are the **dotted lines** between the molecules representing?



Objectives:

Understanding the solubility of solutes and solvents and be able to discern between miscible and immiscible mixtures.

Understand mass percent concentrations and determine the amount of sodium chloride in salt water.

Purpose:

To test chemical solutions for polarity and understand the "LIKE DISSOLVES LIKE" rule. Solubility and concentrations of solutions will be analyzed for their molar concentration. Techniques such as volume delivery by 10-mL and 25-mL volumetric pipets will be utilized throughout this experiment.

Theory:

Polarity, solubility, and concentration of solutions.

Procedure:

A. Supersaturated Solution

1. Set up the boiling point apparatus as shown. You will need a ring stand, a ring clamp, 400-mL beaker, wire gauze, a utility clamp, and a Bunsen burner.
2. Bring about 300 mL of tap water to a rolling boil and add a boiling chip.
3. Obtain a supersaturated solution mixture from your instructor. If none are available, place enough sodium acetate trihydrate ($NaC_2H_3O_2 \cdot 3H_2O$) to fill one-third of a large test tube. Add just enough deionized water to just cover the crystals.
4. Using a utility clamp, support the test tube with sodium acetate trihydrate in the boiling water bath. Once the crystals have all completely dissolved, turn off the gas, and remove the test tube from the boiling water, but leave supported by the utility clamp on the ring stand.
5. While waiting for the supersaturated solution to cool, move on to the next section. Once the solution has completely cooled, place a seed crystal of sodium acetate trihydrate into the supersaturated solution. Record your observations in your data sheet.

B. Solutes and Solvents: Solubility

a. Solubility: (WEAR GLOVES)

1. Set up three clean, dry test tubes supported by your test tube rack. Place a small crystal of potassium permanganate ($KMnO_4$) into each of the test tubes.
2. Add 20 drops of deionized water to the first test tube and record whether water and $KMnO_4$ is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container.
3. Add 20 drops of hexane (C_6H_{14}) to the second test tube and record whether hexane and $KMnO_4$ is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container. Be careful, hexane is flammable.
4. Add 20 drops of methanol (CH_3OH) to the third test tube and record whether methanol and $KMnO_4$ is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container. Be careful, methanol is flammable.
5. Set up three more clean, dry test tubes supported by your test tube rack. Place a small crystal of iodine (I_2) into each of the test tubes.
6. Add 20 drops of deionized water to the first test tube and record whether water and I_2 is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container.
7. Add 20 drops of hexane (C_6H_{14}) to the second test tube and record whether C_6H_{14} and I_2 is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container. Be careful, hexane is flammable.
8. Add 20 drops of methanol (CH_3OH) to the third test tube and record whether CH_3OH and I_2 is soluble or insoluble in your data sheet. Empty the test tube into the appropriate waste container. Be careful, methanol is flammable.

Caution: Iodine (I_2) causes severe burns to the skin. If you come into contact with the iodine wash it off with soap and water. If necessary rinse the affected area with a dilute solution of sodium thiosulfate (1 M $Na_2S_2O_3$). Be sure to wipe up any iodine that you spill so that others will not come in contact with it. Do not breathe excessive iodine vapors.

b. Miscibility:

1. Set up three clean, dry test tubes supported by your test tube rack. Place 20 drops of deionized water into each of the test tubes.
2. Add 20 drops of acetone (C_3H_6O) to the first test tube and stir by lightly flicking the bottom the test tube. Record whether acetone and water are miscible or immiscible in your data sheet. Empty the test tube into the appropriate waste container.
3. Add 20 drops of ethanol (C_2H_5OH) to the second test tube and stir by lightly flicking the bottom the test tube. Record whether ethanol and water are miscible or immiscible in your data sheet. Empty the test tube into the appropriate waste container.
4. Add 20 drops of hexane (C_6H_{14}) to the third test tube and stir by lightly flicking the bottom the test tube. Record whether hexane and water are miscible or immiscible in your data sheet. Empty the test tube into the appropriate waste container.

C. Mass Percent Concentration of Saltwater

1. Set up a ring stand, a ring clamp, 250-mL beaker, wire gauze, and a Bunsen burner.
2. Bring about 200 mL of tap water to a rolling boil and make sure to add a boiling chip. Make sure that you or the instructor has extra hot water boiling to replace evaporated water. Multiple beakers can be placed on hot plates.
3. Determine the mass of a clean, dry evaporating dish using an analytical balance. Record this value in your data sheet.
4. Obtain a 10-mL volumetric pipette and clean with dilute soap solution (1% Alconox solution) by drawing up a little soap solution into the pipette. Completely rinse with tap water by using a pipette wand to draw up the water into the pipette until no more suds exist. Make sure not to suck up any soap or water into the pipette wand itself. If you do, the wand is contaminated and should be replaced with a dry one. Notify your instructor so it can be placed aside to properly dry.

   Rinse three more times with deionized water. Do NOT pipet directly from the deionized water bottles. Pour the deionized water into a clean, dry beaker and pipet from the beaker.
5. Obtain about 75 mL of unknown saltwater solution in a clean, dry 150-mL beaker. Do not pipet directly out of the stock solution. Record the unknown number in your data sheet.
6. Condition the 10-mL volumetric pipette by pulling a little bit of the saltwater solution into the pipette with the pipette wand. Tilt the volumetric pipette and roll the saltwater solution around in the pipette. Drain this solution into the sink.

7. Fill the 10-mL volumetric pipette with the saltwater solution making sure not to suck up any liquid into the pipette wand. Fill the pipette up to the calibration mark. If using a pipette wand, press the release button to dispense. {If using a pipette bulb, remove the pipette bulb and place your thumb over the top of the pipette so that the solution remains above the calibration mark. With a wet thumb, slowly lower the solution down to the calibration mark by rotating the pipette.} Once the meniscus is at the calibration mark, allow the solution to pour in a clean, dry evaporating dish. Allow the solution to drain while touching the tip to the side of the evaporating dish. A little will remain in the pipette. Volumetric pipettes are specially calibrated, so do not blow or shake out the remainder in the tip. Record this value (10.00 mL) on your data sheet.
8. Determine the mass of a clean, dry evaporating dish with the 10.00 mL saltwater solution using an analytical balance. Record this value in your data sheet.
9. Calculate the mass of saltwater solution by subtracting #3 from #8 and record this value.
10. Carefully place the evaporating dish containing the saltwater solution onto the 250-mL beaker boiling water bath. The energy from the boiling water will slowly evaporate off the water solvent. Do NOT allow the beaker to boil dry. Replace the boiling water as needed.
11. Once all of the water has evaporated, use crucible tongs (not beaker tongs) to remove the evaporating dish. Dry the bottom by lightly heat the evaporating dish with a Bunsen burner to completely dry. Make sure to gently heat to avoid splattering of the salt solute.
12. Allow the dish to cool on a wire gauze and obtain the mass of the evaporating dish and residue with an analytical balance. Record this value in your data sheet.
13. Calculate the mass of salt solute by subtracting #3 from #12 and record this in your data sheet.
14. Determine the mass percent concentration of salt in saltwater by dividing #13 by #9 and multiplying by 100 and record this value.
15. Calculate the number of moles of solute in the saltwater by converting the mass of the salt (NaCl) to moles by using the molar mass of sodium chloride (NaCl). Record this value in your data sheet.
16. Convert the saltwater solution volume from mL to liters and record this value in your data sheet.
17. Calculate the molarity of the salt water by dividing #15 by #16 and record this value.
18. Repeat this trial one more time with the **same** unknown number by repeating steps 7 through 17.

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Supersaturated Solution

1. Setting up boiling apparatus (no data):
2. Boiling water (no data):
3. Heating supersaturated solution (no data):
4. Cooling supersaturated solution (no data):
5. Observation from supersaturated solution: ______________________________

____

____

B. Solutes and Solvents: Solubility

a. Solubility:

	$KMnO_4$	**I_2**
1. Setting up test tube rack (no data):		
2. The solvent water is **soluble/insoluble**:	____________	____________
3. The solvent hexane is **soluble/insoluble**:	____________	____________
4. The solvent methanol is **soluble/insoluble**:	____________	____________

b. Miscibility:

	water (H_2O)
1. Setting up test tube rack (no data):	
2. The solvent acetone is **miscible/immiscible** with water:	____________
3. The solvent ethanol is **miscible/immiscible** with water:	____________
4. The solvent hexane is **miscible/immiscible** with water:	____________

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

C. Mass Percent Concentration of Saltwater

	TRIAL 1:	TRIAL 2:
1. Apparatus setup (no data):		
2. Apparatus setup (no data):		
3. Mass of evaporating dish (g):	________	________
4. Cleaning volumetric pipet (no data):		
5. Unknown saltwater solution:	________	________
6. Conditioning volumetric pipet (no data):		
7. Volume from 10-mL volumetric pipet (mL):	________	________
8. Mass of evaporating dish and solution (g):	________	________
9. Mass of saltwater solution (g):	________	________
10. Evaporating solution (no data):		
11. Heating solution (no data):		
12. Mass of evaporating dish and residue (g):	________	________
13. Mass of salt (g):	________	________
14. Mass percent of salt (%):	________	________
15. Moles of salt (NaCl solute) (mole):	________	________
16. Volume from 10-mL volumetric pipet (L):	________	________
17. Molarity (M):	________	________

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Data Analysis/Calculations:

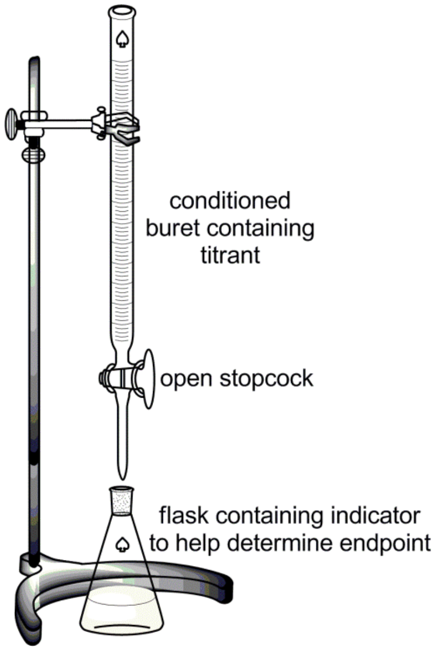
Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:

Experiment #14:

Titrations



Introduction:

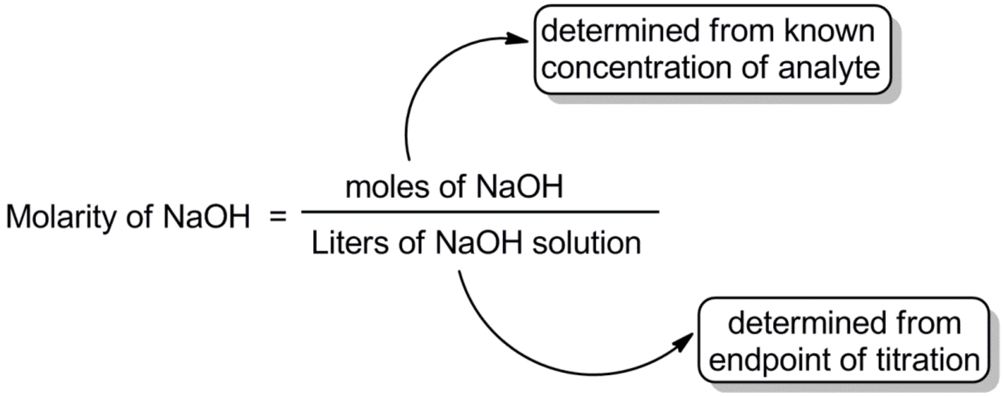
Titrations utilize analytical glassware to determine unknown concentrations. By knowing a certain concentration, we can utilize a balanced chemical reaction to determine what the concentration is of the unknown substance. The unknown substance that goes into a buret is referred to as the **titrant**. This unknown substance is then transferred from the buret into a flask containing a known concentration referred to as the **analyte**. An indicator that changes color based on acidity or basicity is utilized to determine the endpoint of a reaction.

A typical acid-base reaction known as a neutralization reaction:

Acid + Base → Salt + Water

This process is extremely useful when standardizing solutions. A standardized solution is a solution with a known concentration to at least the thousandths' place. Using analytical glassware such as burets and volumetric pipettes, we are able to achieve up to four significant figures.

Molarity concentration is moles of the solute divided by Liters of solution. When we attempt to calculate the concentration of an unknown substance, we begin with a known amount of a known concentration, and perform a titration. The known concentration and amount of the analyte is used to calculate the number of moles of that substance. A balanced chemical reaction can then determine the number of moles of our unknown NaOH solution.

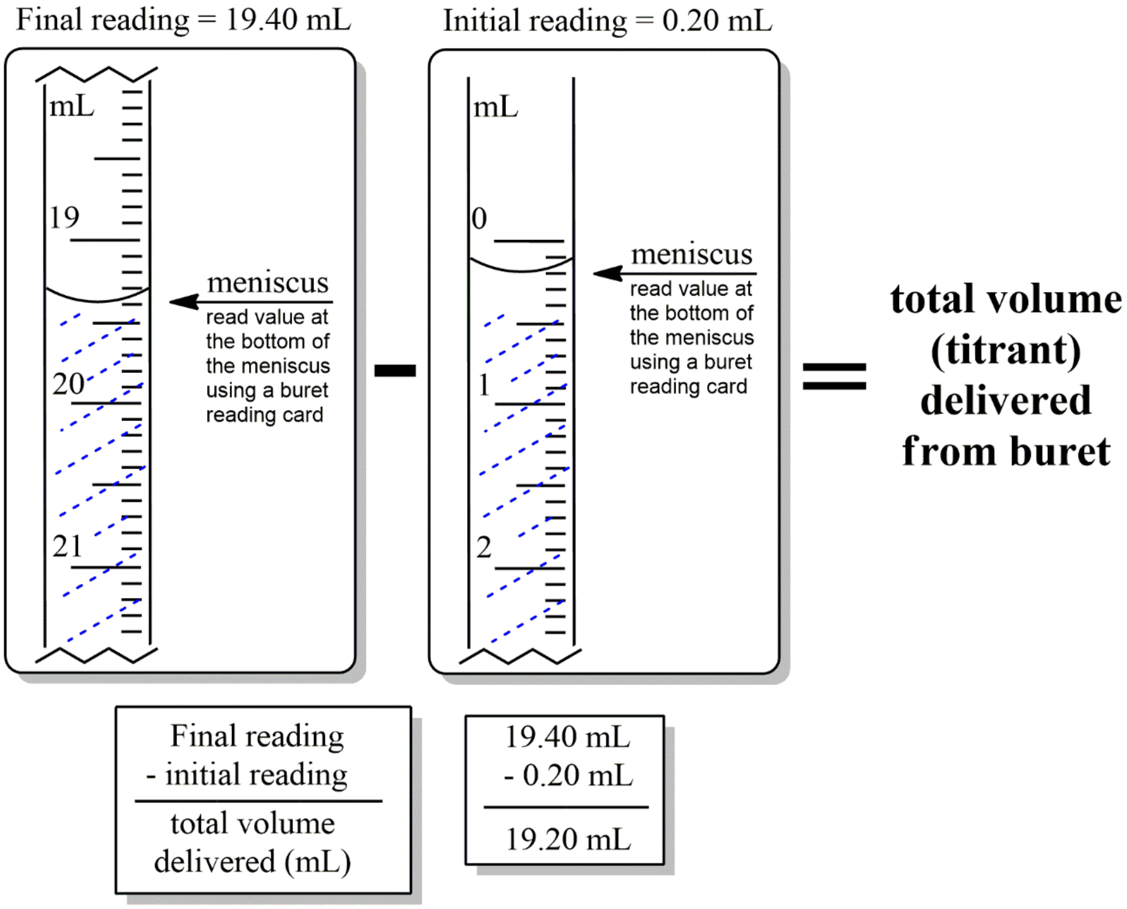


$$\text{Molarity of NaOH} = \frac{\text{moles of NaOH}}{\text{Liters of NaOH solution}}$$

The amount of unknown solution it took to reach the end point is the amount of solution in the flask. When we calculate the number of moles of our unknown NaOH substance and divide it by the amount of Liters of NaOH solution we poured in, we have successfully determined the Molarity of our unknown concentration.

To successfully read a buret, make sure to use a buret reading card. The buret reading card allows us to see the true meniscus. By slowly bringing the buret reading card behind the layer of liquid, you will observe the true meniscus. Do not place the reading card directly behind the liquid layer, it must be slightly below it. By moving it up and down, you should see the meniscus appear and disappear.

The buret is set up having a 0.00 mL reading at the top and a 50.00 mL reading at the bottom. We must keep the liquid layers between these numbers or else we are unable to get an accurate reading. We always begin with our initial reading near the top to the hundredths' place. We then perform our titration until the endpoint is reached. We then take our final reading to the hundredth's place. To determine the amount of solution delivered, we subtract the initial reading from the final reading (Final - Initial = amount delivered) as shown below:



Name:____

Pre-laboratory Assignment:

1. Calculate the Molarity of 0.524 mol of glucose in 895 mL of solution.

2. Calculate the Molarity of 0.917 mole of NaCl in 0.500 L of solution.

3. How many moles of NaCl are in 1.85 L of a 0.458 M NaCl solution?

4. How many liters are required to obtain 0.182 moles of $NaNO_3$ from a 0.824 M $NaNO_3$ solution?

Objectives:

To learn how to standardize a NaOH solution and to then determine the unknown concentration in vinegar using that solution.

Purpose:

To understand and develop proper titrating technique. A standardized hydrochloric acid solution will be used to standardize an unknown sodium hydroxide solution. This standardized sodium hydroxide solution will be utilized to analyze the concentration of acetic acid in household vinegar. Techniques such as titrating, and proper cleaning, conditioning, and reading of burets will be practiced throughout this experiment.

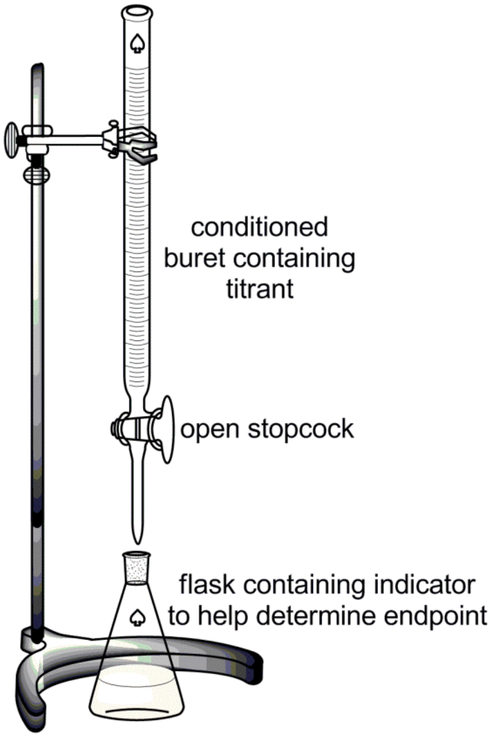
Theory:

Concentration of Solutions:

(Molarity = moles of solute/ Liters of solution)

Standardizing a solution by determining its molarity to the thousandths' place.

Titrating solutions to determine their unknown molar concentration.



Procedure:

Experimental Procedure:

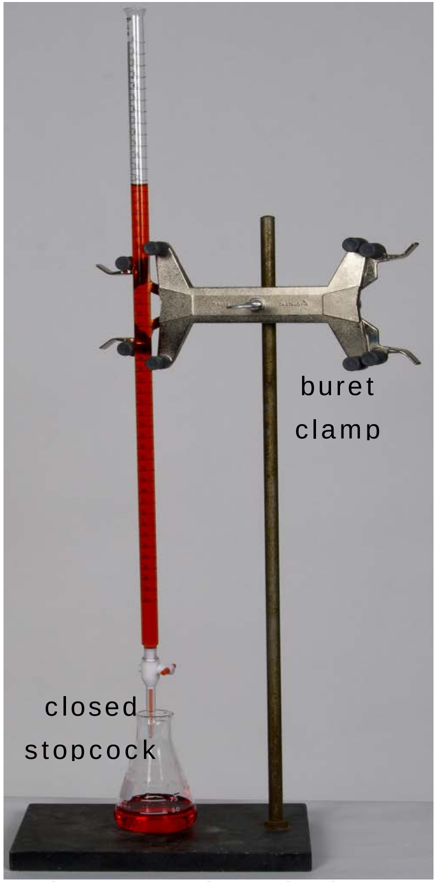
A. Standardization of a NaOH solution

1. Obtain a 25-mL volumetric pipette and clean with dilute soap solution by drawing up a little soap solution into the pipette. Rinse with tap water and deionized water by using a pipette wand to draw up the water into the pipette. Make sure not to suck up any soap or water into the pipette wand itself. If you do, the wand is contaminated and should be replaced with a dry one.
2. Record the concentration of the HCl solution in your data sheet and condition the 25-mL volumetric pipette by pulling a little bit of HCl solution into the pipette with the pipette wand. Tilt the volumetric pipette and roll the HCl solution around in the pipette. Drain this solution into a discard beaker.

3. Obtain about 50 mL of the HCl solution in a 150-mL beaker. From the beaker, fill the 25-mL volumetric pipette with the HCl solution making sure not to suck up any liquid into the pipette wand. **Do NOT pipet directly from the stock solution.** Once the meniscus is at the calibration mark, remove your thumb and allow the solution to pour in a clean, dry 125-mL Erlenmeyer flask. Allow the solution to drain while touching the tip to the side of the flask. A little will remain in the pipette. Volumetric pipettes are specially calibrated, so do not blow or shake out the remainder in the tip. Record this value on your data sheet. (If using a pipette bulb, fill the pipette past the calibration mark. Remove the pipette bulb and place your thumb over the top of the pipette so that the solution remains above the calibration mark. With a wet thumb, slowly lower the solution down to the calibration mark by rotating the pipette.)
4. Assemble the buret apparatus as shown below:

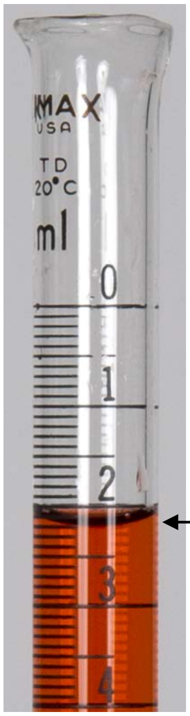
The buret is closed if the stopcock is perpendicular to the buret.

The buret is open if the stopcock is parallel to the buret.

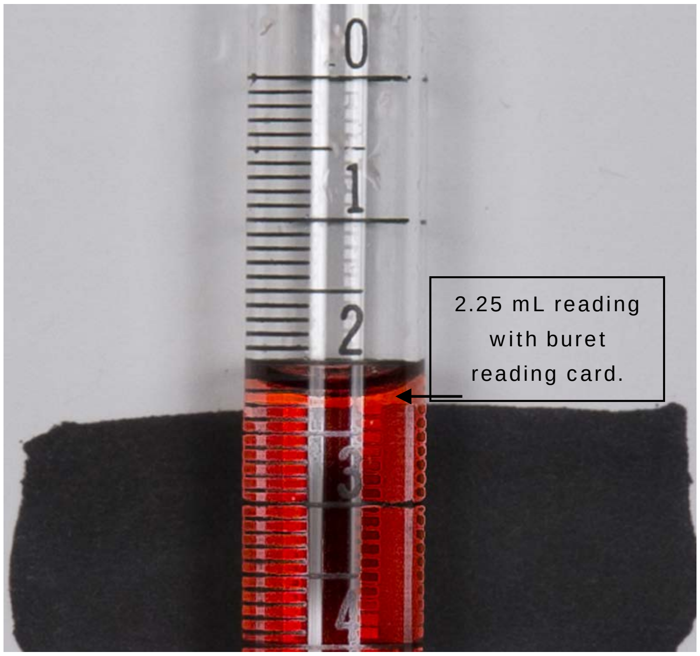


5. Determine if the buret needs to be cleaned. Close the stopcock on the buret and add deionized water until the buret is full. Drain the buret through the buret tip into the sink. If any drops adhere to the inside of the buret, the buret needs to be cleaned. Spray or dip the buret brush with a dilute soap solution (1% Alconox solution). Gently brush the interior of the buret making sure not to scrub below the 50.00-mL mark. Rinse the buret by running tap water through the buret tip until all of the soap solution is removed. Once all suds have been removed, rinse twice more with deionized water. Remember, one rinse gets 90%, two rinses gets 99%, and three rinses gets 100% of all contaminants.

6. Each time glassware such as a buret is going to be used, the glassware must be conditioned with the solution that is going to be dispensed. This makes sure that all deionized water is removed so that the solution is not diluted any. To condition the buret, pour 20 mL of the titrant (the substance that will be dispensed) into the buret. Make sure the stopcock is closed to trap the titrant in the buret. Tilt the buret to allow the titrant to coat the inside of the buret without spilling any out of the top. Once the buret has been coated, drain the titrant into a discard beaker. We cannot use this titrant any longer because it has been contaminated and diluted. Do this again with 20 mL of fresh titrant to ensure the buret is completely conditioned.
7. Now that the buret is conditioned, close the stopcock and fill the buret with the titrant to the 0.00-mL calibration mark. Open up the buret tip and drain the titrant out into the discard beaker until all air bubbles are removed from the buret tip. Fill the buret back up to around the 0.00-mL mark if a lot of titrant was lost. Determine the initial buret reading by using a buret reading card making sure to read to the hundredths' place. Record this value in your data sheet. An example of a buret reading with and without a reading card is shown:



2.20 mL reading without buret reading card.



2.25 mL reading with buret reading card.

8. Add two to three drops of the phenolphthalein indicator to the 125-mL Erlenmeyer flask containing the 25.00 mL HCl solution prepared earlier and gently swirl to mix the contents. Lower the buret tip into the 125-mL Erlenmeyer flask containing the HCl solution and indicator. Add the sodium hydroxide solution and continuously swirl until a slight pink color is observed and persists for at least one minute. Placing a white sheet of paper underneath the flask can help to watch for the faint pink that comes from the phenolphthalein indicator.

9. The faint pink color marks the endpoint of the titration. Read the final buret reading using a buret reading card making sure to read to the hundredths' place. Record this value.
10. Determine the total volume of sodium hydroxide solution added by subtracting the initial buret reading #7 from the final buret reading #9 and record this value.
11. Convert this volume into liters and record this value.
12. Calculate the number of moles of HCl that were in the 125-mL Erlenmeyer flask by multiplying the volume of HCl dispensed from the volumetric pipette #3 by the known concentration of HCl #2.
13. Determine the number of moles of sodium hydroxide using a balanced chemical equation (hydrochloric acid reacting with sodium hydroxide) and record this value in your data sheet.
14. Calculate the molarity of sodium hydroxide solution by dividing the number of moles #13 by the volume #11 and record this value in your data sheet.
15. Perform Trial #2 by repeating steps 1 through 14. You do not need to clean or recondition the glassware for additional trials in Part A.
16. Determine the average molarity of sodium hydroxide by adding the molarities of each trial and dividing by the total number of trials (in this case, two).

B. Volume-volume stoichiometry: Determination of the concentration of acetic acid in vinegar

1. Place 50 mL of vinegar in to a clean, dry 100-mL beaker.
2. Obtain a 5-mL volumetric pipette and clean with dilute soap solution by drawing up a little soap solution into the pipette. Rinse with tap water and deionized water by using a pipette wand to draw up the water into the pipette. Make sure not to suck up any soap or water into the pipette wand itself. If you do, the bulb is contaminated and should be replaced with a dry one.
3. Condition the 5-mL volumetric pipette by pulling a little bit of the vinegar into the pipette with the pipette wand. Tilt the volumetric pipette and roll the vinegar around in the pipette. Drain this solution into a discard beaker.
4. Fill the 5-mL volumetric pipette with vinegar making sure not to suck up any liquid into the pipette wand. Once the meniscus is at the calibration mark, remove your thumb and allow the solution to pour in a clean, dry 125-mL Erlenmeyer flask. Allow the solution to drain while touching the tip to the side of the flask. A little will remain in the pipette. Volumetric pipettes are specially calibrated, so do not blow or shake out the remainder in the tip. Record this value on your data sheet. (If using a pipette bulb, fill the pipette past the calibration mark. Remove the pipette bulb and place your thumb over the top of the pipette so that the solution

remains above the calibration mark. With a wet thumb, slowly lower the solution down to the calibration mark by rotating the pipette.)

5. Convert this volume into liters and record this value.
6. Using a graduated cylinder, add 25.0 mL of deionized water into the 125-mL Erlenmeyer flask that contains the 5.00 mL of vinegar. Make sure to stopper the flask with the appropriate rubber stopper size as shown on the side of the flask.
7. Fill the buret with the titrant (standardized sodium hydroxide from Part A) above the 0.00-mL calibration mark. Open up the buret tip and drain the titrant out into the discard beaker until all air bubbles are removed from the buret tip. Fill the buret up to around the 0.00-mL mark if a lot of titrant was lost. Determine the initial buret reading by using a buret reading card making sure to read to the hundredths' place. Record this value in your data sheet
8. Lower the buret tip into the 125-mL Erlenmeyer flask containing the vinegar solution prepared earlier using the 5-mL volumetric pipette and don't forget to add two to three drops of the phenolphthalein indicator. Add the sodium hydroxide solution until a slight pink color is observed and persists for at least 30 seconds. Placing a white sheet of paper can help to watch for the faint pink that comes from the phenolphthalein indicator.
9. The faint pink color marks the endpoint of the titration. Read the final buret reading using a buret reading card making sure to read to the hundredths' place. Record this value.
10. Determine the total volume of sodium hydroxide solution added by subtracting the final buret reading #9 from the initial buret reading #7 and record this value.
11. Convert this volume into liters and record this value.
12. Calculate the number of moles of sodium hydroxide that were added to the 125-mL Erlenmeyer flask by multiplying the volume of sodium hydroxide dispensed from the buret #10 by the average known concentration of NaOH (#16 from part A).
13. Determine the number of moles of vinegar using the balanced chemical equation shown and record this value in your data sheet.

$$\mathbf{HC_2H_3O_2}\ _{(aq)} + \mathbf{NaOH}\ _{(aq)} \longrightarrow \mathbf{NaC_2H_3O_2}\ _{(aq)} + \mathbf{H_2O}\ _{(l)}$$

14. Calculate the molarity of vinegar by dividing the number of moles #12 by the volume delivered by the 10-mL volumetric pipette #5 and record this value in your data sheet.
15. Perform Trial #2 by repeating steps 1 through 14. You do not need to clean or recondition the glassware for additional trials in Part B.
16. Determine the average molarity of vinegar by adding the molarities of each trial and dividing by the total number of trials (in this case, two).

Experiment:	Date:
Name:	Course/Section:

Data/Observations:

A. Standardization of a NaOH solution

	TRIAL 1	TRIAL 2
1. Setting up equipment (no data):		
2. Molarity of HCl solution (M):	_______	________
3. Volume delivered from 25.00 mL volumetric pipette (mL):	_______	________
4. Setting up equipment (no data):		
5. Cleaning the buret (no data):		
6. Conditioning the buret (no data):		
7. Initial buret reading (mL):	_______	________
8. Titrating (no data):		
9. Final buret reading (mL):	_______	________
10. Volume delivered from buret (mL):	_______	________
11. Volume converted to liters (L):	_______	________
12. Moles of hydrochloric acid (mol):	_______	________
13. Moles of sodium hydroxide (mol):	_______	________
14. Molarity of NaOH solution (M):	_______	________
15. Obtaining data for Trial #2:		
16. Average Molarity of NaOH solution (M):	_______________	

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

B. Volume-volume stoichiometry: Determination of the concentration of acetic acid in vinegar

	TRIAL 1	TRIAL 2
1. Obtaining vinegar (no data):		
2. Cleaning the pipette (no data):		
3. Conditioning the pipette (no data):		
4. Volume of vinegar from 5.00 mL volumetric pipette (mL):	_______	________
5. Volume converted to liters (L):	_______	________
6. Adding deionized water (no data):		
7. Initial buret reading (mL):	_______	________
8. Titrating (no data):		
9. Final buret reading (mL):	_______	________
10. Volume delivered from buret (mL):	_______	________
11. Volume converted to liters (L):	_______	________
12. Moles of sodium hydroxide (mol):	_______	________
13. Moles of vinegar (mol):	_______	________
14. Molarity of vinegar (M):	_______	________
15. Obtaining data for Trial #2:		
16. Average Molarity of Vinegar (M):	_______________	

Date:	Instructor Signature:	Comments:

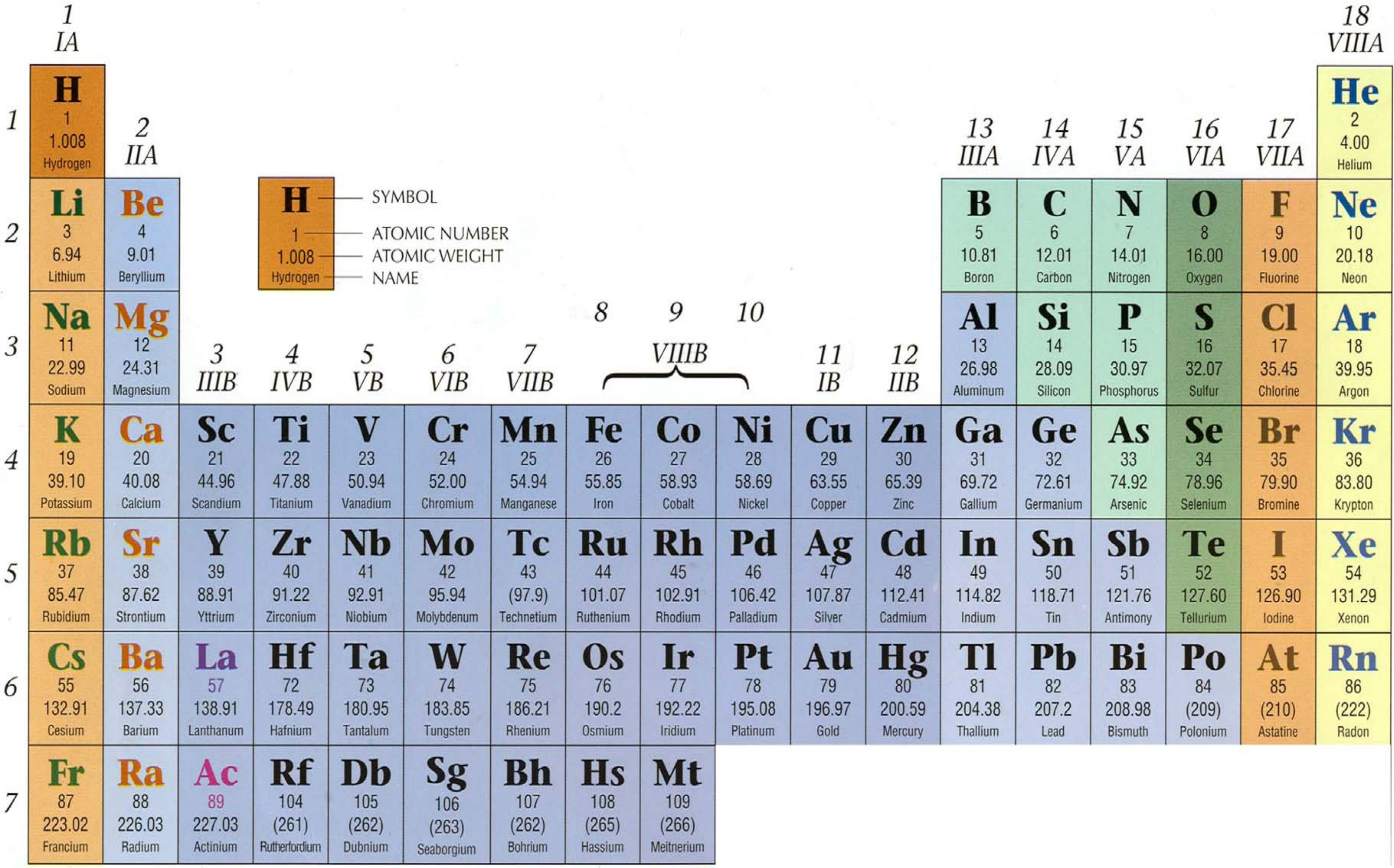
Experiment:	Date:
Name:	Course/Section:

Data Analysis/Calculations:

Date:	Instructor Signature:	Comments:

Experiment:	Date:
Name:	Course/Section:

Date:	Instructor Signature:	Comments:



H — SYMBOL
1 — ATOMIC NUMBER
1.008 — ATOMIC WEIGHT
Hydrogen — NAME

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	H 1 1.008 Hydrogen																	He 2 4.00 Helium
2	Li 3 6.94 Lithium	Be 4 9.01 Beryllium											B 5 10.81 Boron	C 6 12.01 Carbon	N 7 14.01 Nitrogen	O 8 16.00 Oxygen	F 9 19.00 Fluorine	Ne 10 20.18 Neon
3	Na 11 22.99 Sodium	Mg 12 24.31 Magnesium											Al 13 26.98 Aluminum	Si 14 28.09 Silicon	P 15 30.97 Phosphorus	S 16 32.07 Sulfur	Cl 17 35.45 Chlorine	Ar 18 39.95 Argon
4	K 19 39.10 Potassium	Ca 20 40.08 Calcium	Sc 21 44.96 Scandium	Ti 22 47.88 Titanium	V 23 50.94 Vanadium	Cr 24 52.00 Chromium	Mn 25 54.94 Manganese	Fe 26 55.85 Iron	Co 27 58.93 Cobalt	Ni 28 58.69 Nickel	Cu 29 63.55 Copper	Zn 30 65.39 Zinc	Ga 31 69.72 Gallium	Ge 32 72.61 Germanium	As 33 74.92 Arsenic	Se 34 78.96 Selenium	Br 35 79.90 Bromine	Kr 36 83.80 Krypton
5	Rb 37 85.47 Rubidium	Sr 38 87.62 Strontium	Y 39 88.91 Yttrium	Zr 40 91.22 Zirconium	Nb 41 92.91 Niobium	Mo 42 95.94 Molybdenum	Tc 43 (97.9) Technetium	Ru 44 101.07 Ruthenium	Rh 45 102.91 Rhodium	Pd 46 106.42 Palladium	Ag 47 107.87 Silver	Cd 48 112.41 Cadmium	In 49 114.82 Indium	Sn 50 118.71 Tin	Sb 51 121.76 Antimony	Te 52 127.60 Tellurium	I 53 126.90 Iodine	Xe 54 131.29 Xenon
6	Cs 55 132.91 Cesium	Ba 56 137.33 Barium	La 57 138.91 Lanthanum	Hf 72 178.49 Hafnium	Ta 73 180.95 Tantalum	W 74 183.85 Tungsten	Re 75 186.21 Rhenium	Os 76 190.2 Osmium	Ir 77 192.22 Iridium	Pt 78 195.08 Platinum	Au 79 196.97 Gold	Hg 80 200.59 Mercury	Tl 81 204.38 Thallium	Pb 82 207.2 Lead	Bi 83 208.98 Bismuth	Po 84 (209) Polonium	At 85 (210) Astatine	Rn 86 (222) Radon
7	Fr 87 223.02 Francium	Ra 88 226.03 Radium	Ac 89 227.03 Actinium	Rf 104 (261) Rutherfordium	Db 105 (262) Dubnium	Sg 106 (263) Seaborgium	Bh 107 (262) Bohrium	Hs 108 (265) Hassium	Mt 109 (266) Meitnerium									