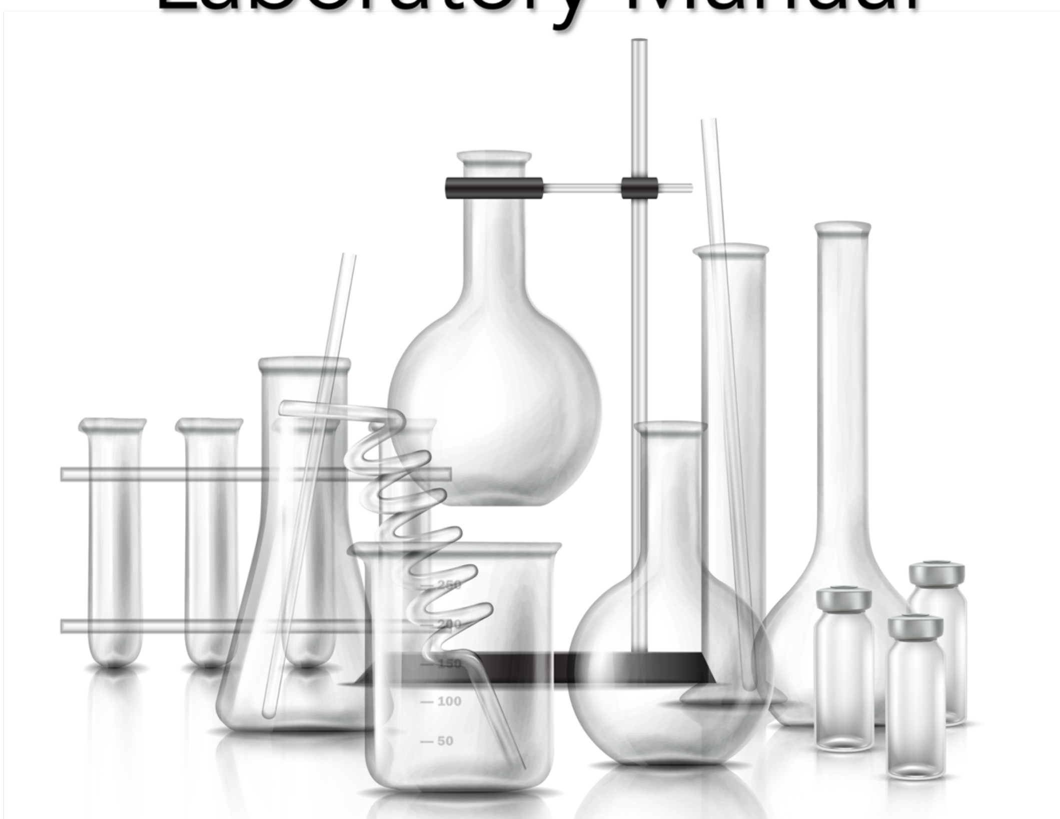
Chemistry 120

Introductory Chemistry

Laboratory Manual



ALLAN HANCOCK COLLEGE

CHANGING *the* ODDS

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This edition of the lab manual was funded by Allan Hancock College and the Culturally Responsive Curriculum Grant

Cover image provided by Freepik.com

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Safety First!

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.

Safety First!

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. **All hazardous waste and/or flammable waste must be disposed of in a labeled waste container** as indicated by your instructor.

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**. 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.

____

Printed name

____ ________________

Signature *Date*

About using this lab manual

This lab manual is designed using an active engagement curriculum (AEC) model for learning. Student participation and engagement in the weekly pre-laboratory discussion will be necessary for the completion of each lab. The AEC teaching philosophy places the responsibility for understanding how to execute the experiment onto the student, and encourages this with clear supplementary material. The expectation is that this approach will result in students having a clearer understanding of the purpose, implementation and expected outcomes of each experiment. Improved engagement is expected to also foster a safer laboratory environment and improve connections between laboratory and lecture curriculum. Each section of the lab manual is separated into the following categories:

Introduction: Gives theory related to the lab that should be read through before your lab period.

Purpose: Relays the learning objectives of the experiment. This section is expanded upon to have a list of learned lab technique skilled and a list of learned knowledge related to the theory discussed in lecture.

What we will do, briefly: Gives some (but not all!) of the information students will need to complete the lab. Think of this section as the "big picture". The crucial details needed to conduct the lab will need to be gleaned from your instructor and notated in the following section.

Pre-lab Assignment: An assignment related to the concepts in the experiment, is due at the start of the laboratory session

Image Guide and Procedure: This section is to be completed <u>during class</u> when the instructor is giving their <u>pre-laboratory lecture.</u> At this time, the instructor will give you all the information needed to fill in each black, and it is important that you listen and ask questions when something is unclear—this guide will serve as your procedure for the day's lab and will be checked for completion by your instructor upon starting the lab.

Data Sheets: Where all of your measured values and observations for each lab are recorded

Post-lab Assignment: An assignment which reiterates the concepts of the experiment.

Locker Check-in Sheet

Name: ____ Room # ________ Locker # ________

Course: ________________ Section: _______ Semester: _______ Instructor ______________

Quantity	Item description	Remarks (chipped or damaged)	Checked-In (start of term)	Checked-out (end of term)
6	Beakers (100, 150, 250, 400, 600, and 800/1000 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 and 100 mL)			
2	Litmus paper vials (red and blue)			
2	Pipette dropper 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	Tringle, Clay			
1	Watch Glass			
1	Wire Gauze			

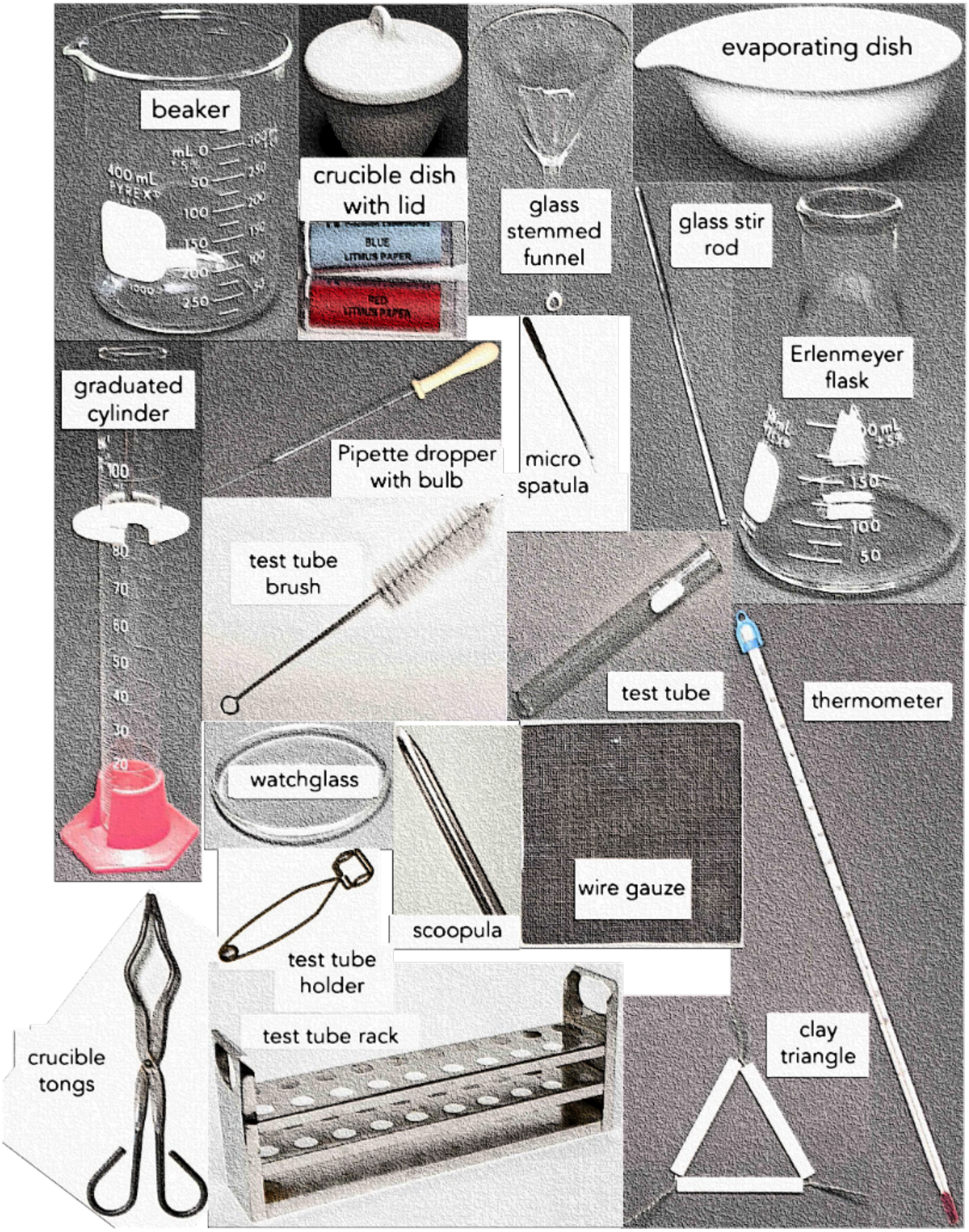
Check-in: I have received the following items in good condition, as checked above:

Student signature, start of term: __________________________________ Date: _________

Check-out: I have returned the following items in good condition, as checked above:

Instructor signature, end of term: _________________________________ Date: _________

Glassware


beaker
crucible dish with lid
BLUE LITMUS PAPER
RED LITMUS PAPER
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
test tube rack
crucible tongs
clay triangle

Lab 1:

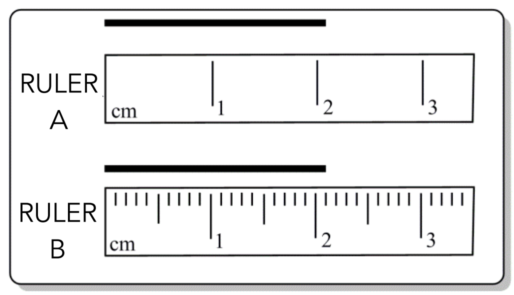
Instrumental Measurements and Calculations

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. Digits that come from the specific markings on the instrument are considered **certain digits**, but scientific measurements always include an extra **estimated digit** at the end, which communicates the limitation of that instrument.

For example, the number of **significant figures** we can achieve from a **length measurement** varies with the markings on our ruler. 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 provide one extra estimated digit, giving us a reading of 2.1 cm. **RULER B** has more marks for us to confidently read 2.1 cm, but adding our estimated digit, gives 2.10 cm, a value with three significant figures. (Typically, we see ±0.05 cm as the limitation for this device. This is because our eyes really can't resolve another 10 spaces.)

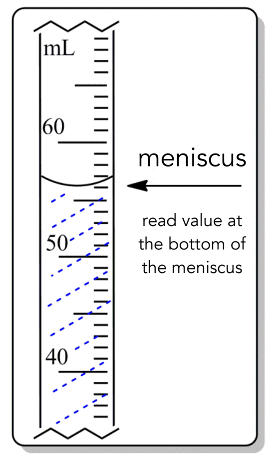


When taking **mass measurements,** the follow steps should be taken:

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 and record ALL digits.
5) use the same balance throughout an experiment.

Volume measurements for liquids can be taken with specialized glassware. With the graduated cylinder on the right, we can easily read to the ones' place. We can reach 56 mL with certainty but need to add a final estimated digit. 56.0 mL would be the correct way to record this value.



Purpose

Students will learn skills and apply knowledge necessary to **accurately and precisely measure length, mass and volume using appropriate laboratory equipment.**

Skills:

1. Correct use of a ruler, analytical balance, and graduated cylinder. These will be necessary skills for most future laboratory experiments.
2. Calculate volume from length measurements

Knowledge:

1. Familiarity with names of glassware and lab equipment
2. Application of significant figures
3. Demonstrate dimensional analysis

What you will do, briefly:

A. Use the **centimeter** side of a **ruler** to measure the **length** of a **small test tube** and **diameter** of a **watchglass** and **evaporating dish** to **correct significance** considering the uncertainty of the ruler.

B. Use an **analytical balance** to individually measure the **mass** of a **small test tube**, **plastic funnel**, **crucible**, and **crucible cover**. Then measure the mass of the **crucible and crucible cover together** and **determine the mass difference** between the individually weighed crucible + individually weighed crucible and the crucible + cover measured together.

C. Use a **100-mL graduated cylinder** to measure the **volume of a small test tube by difference**. You will first fill the graduated cylinder up to about the 80 mL mark, record its exact volume, then pour that water into a small test tube until the test tube is filled to the brim. Taking the difference between the initial volume in the graduated cylinder and the volume in the graduated cylinder after filling up the small test tube will give you the volume of the test tube. Do this twice and calculate the volume difference between the two test tubes.

D. Use a **thermometer** to measure the **temperature** of the **laboratory room**, a **hot water bath** (prepared in a small beaker on a hot plate) and a **cold-water bath** (prepared in a beaker with ice and water) in **degrees Celsius.**

E. Use a **ruler** to determine the **volume** of a **rectangular solid** by measuring the solid's **length, width and thickness in centimeters**. V = length x width x height

Criteria for Success:

1. The correct number of **significant figures** must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware you are using.

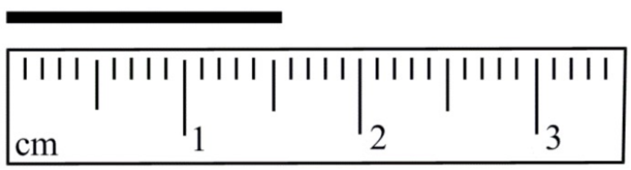
2. The **correct units** must be used for each measurement, following the metric system.

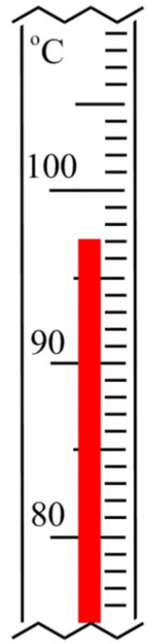
3. Your answers should make **logical sense**, and calculations/measurements should be repeated otherwise.

Name: ______________________ Experiment: ____

Pre-laboratory Assignment:

1. Give the measurement from the following instruments:





2. Convert the following into decimal notation:

   a. 5.540×10^{-4} ______________________

   b. 8.002×10^{8} ______________________

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

   a. 5,000 ______________________

   b. .00247 ______________________

   c. 500.00 ______________________

   d. 250,000 ______________________

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

   a. 5,000

   ____

   a. .00247

   ____

   a. 500.00

   ____

   a. 250,000

   ____

   e. 50. grains

   ____

   f. 5.00×10^{6} seconds

   ____

   g. 1.3000 grams

   ____

   h. 10,000 meters

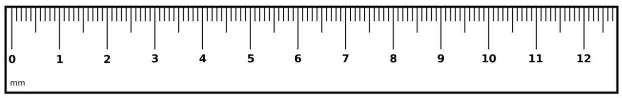
   ____

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

A

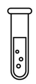
Uncertainty: ______________ Units: ______________ Measures: ______________



The minor hashmarks are millimeters

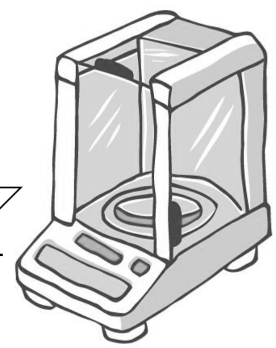
If the measurement is on the line, the estimated digit will be: ______________

If the measurement is between the lines, the estimated digit will be: ______________

With the ruler, you will measure: ______________  ______________  ______________ 

B

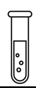
True or False: You estimate digits on the analytical balance. **True** **False**



Uncertainty: ______________ Units: ______________

Measures: ______________

Never put anything hot on the analytical balance

With the balance, you will measure: ______________  ______________ 


____

____

____

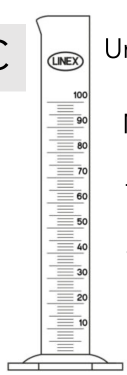
How will you determine the difference between the calculated and observed mass? ______________________________________

____

____

Always keep the balance free of powder and debris

C



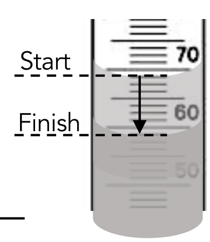
Uncertainty: ______________ Units: ______________

Measures: ______________

The starting volume: ___________

The final volume: ___________

Amount added to test tube: ___________



Measure from the **TOP** / **BOTTOM** of the meniscus (circle one)

Additional Notes:

____

Safety Considerations:

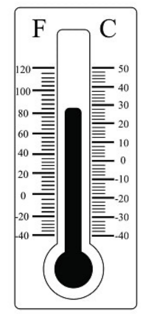
____

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

D

Uncertainty: ______________ Units: ______________ Measures: ______________



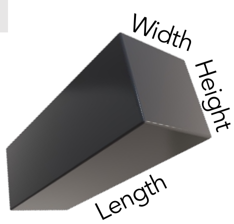
If the measurement is on the line, the estimated digit will be: __________

If the measurement is between the lines, the estimated digit will be: __________

With the thermometer, you will measure: __________ __________

____

E



The mass will be measured with the: ________________________

The number of decimal points in mass will be: __________________

The volume will be measured with the: ________________________

The number of decimal points in volume will be: __________________

The equation for volume will be: ____

Additional Notes:

____

____

____

____

____

____

____

Safety Considerations:

____

____

____

Name: ____________________ Experiment: ______________________________________

Data / Observations

A. Length Measurements

	Object	Measurement (cm)
1.	____________________	____________________
2.	____________________	____________________
3.	____________________	____________________

B. Mass Measurements

	Object	Measurement (g)
1.	____________________	____________________
2.	____________________	____________________
3.	____________________	____________________
4.	____________________	____________________
5.	____________________	____________________
6.	Crucible and cover together	____________________

7. Difference between calculated and observed masses:________________

C. Volume Measurements

1. Volume in graduated cylinder (mL) ____________________
2. Volume in graduated cylinder after first test tube (mL) ____________________
3. Volume in graduated cylinder after first test tube (mL) ____________________
4. Volume in first test tube (mL) ____________________
5. Volume in second test tube (mL)____________________

Name: ____________________ Experiment: ______________________________________

Data / Observations

D. Temperature Measurements

Measurement (°C)

1. Room temperature ____________________
2. Ice water ____________________
3. Boiling water ____________________

E. Mass and Volume Measurements

1. Unknown number ____________________
2. Mass of unknown solid (g) ____________________
3. Length of unknown solid (cm) ____________________
4. Width of unknown solid (cm) ____________________
5. Thickness of unknown solid (cm) ____________________
6. Volume of unknown solid (cm^3) ____________________

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

Use dimensional analysis to complete the following exercises. Use the unit conversion guide (Appendix A) for assistance. Round to the appropriate number of significant figures and include units.

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 may 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)

Name: ______________________ Experiment: ____

6. How many seconds elapse in 2.0 hours?

7. How many milligrams are in 850 nm?

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 hours to meters per minute. (5280 ft = 1 mile; 2.54 cm = 1 inch)

11. How many minutes are in 3.5 years?

Name: ______________________ Experiment: ____

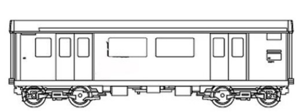
12. The gasoline in an automobile gas tank has a mass of 75.0 kg and a density of 0.752 g/cm^3. What is the volume in mL?

13. A medicine is to be given by IV (intravenous) at the rate of 185 drops per min. The medicine's concentration is 250 mg/mL. What is the weight (in grams) of the medicine 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 feet (ft^3) (2.54 cm = 1 inch; 12 inches = 1 foot)

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 are 86 feet long.

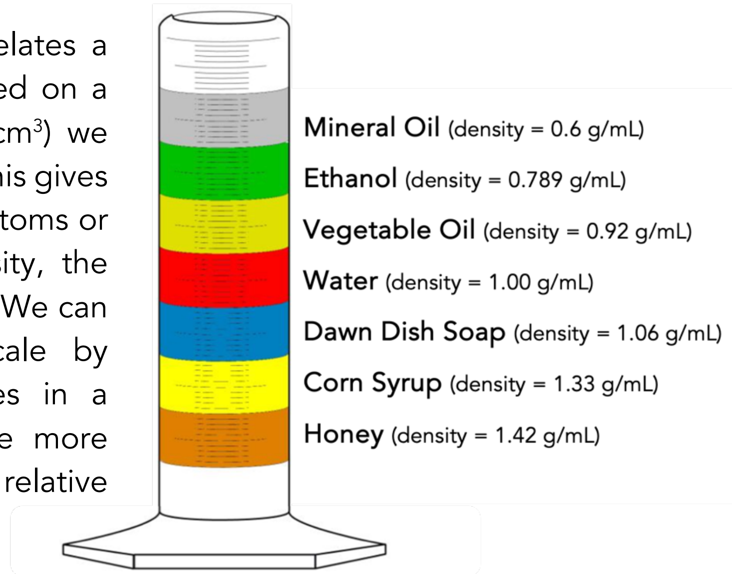


Lab 2:

Density

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. We can observe this on a macroscopic scale by layering liquids with varied densities in a graduated cylinder: liquids that have more mass per unit area (i.e. density) will sink relative to substances with lower density.



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 units of g/cm^3 or 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 pipettor 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 deliver a specific amount from the calibration line with a little left over in the tip. We do not want to blow or shakeout this remaining residual. The volumetric pipette can be touched to the side of the container to remove any drops hanging.

Purpose

Students will learn skills and apply knowledge necessary to **calculate the density of liquids and solids using mass and volume measurements and utilize density and volume to determine the thinness of a piece of aluminum foil.**

Skills:

1. Determine volume by displacement
2. Correct use of volumetric pipette
3. Determine mass by weighing by difference

Knowledge:

1. Demonstrate the relationships between mass, volume and density (d = m/v)
2. Solve for an unknown dimension given length, width, mass and density

What you will do, briefly:

A. Use a 10 mL volumetric pipette to transfer 10.00 mL of water into a pre-weighed 125 mL flask, then obtain the mass of the 125 mL flask with the water inside. Calculate the mass of the water and determine its density (mass/volume) to appropriate significance. Repeat for 3 total trials and calculate the average density.

B. Repeat Part A using rubbing alcohol as your liquid instead of water. Pour your waste into the appropriate waste container.

C. Determine the mass of a dry rubber stopper (size 1 or 2) and then carefully place the stopper into a 100 mL graduated cylinder that has at least 50 mL of water in it to begin with (you must record the initial volume before adding the stopper). Record the final volume of the graduated cylinder with the stopper inside. Calculate the volume of the stopper (final – initial volume) and then calculate the density of the stopper (mass/volume). Repeat for 3 total trials and calculate the average density of the stopper.

D. Use a ruler to determine the volume of an unknown solid by measuring the solid's length, width and thickness in centimeters. (V = length x width x height). Determine the mass of the unknown solid using the analytical balance and then calculate its density (mass/volume) Using the given density tables, identify your known solid.

E. Determine the mass of a square piece of aluminum foil. Calculate its volume using the density of aluminum (2.700 g/cm^3). Determine the thinness of the piece of aluminum. Remember: V = length x width x thinness (height)

Criteria for Success:

1. The correct number of significant figures must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware you are using.

2. The correct units must be used for each measurement, following the metric system

3. Your answers should make logical sense, and calculations/measurements should be repeated otherwise

Name: ______________________ Experiment: ____

Pre-laboratory Assignment:

Please show your work, round using proper significant figures and include units

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 2.00 cm high. Calculate the density of the wooden square block in units of g/cm^3.

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 (1 mL = 1 cm^3)

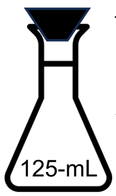
3. An Erlenmeyer flask is placed on an analytical balance giving a reading of 124.2218 g. Exactly 25.00 mL of tap water is placed into the flask using a volumetric pipette. The Erlenmeyer flask with the tap water now show a reading of 149.7450 g when measured on the same analytical balance. Calculate the density of tap water in g/mL.

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

A

True or False: You can round the value given from the analytical balance. **True** **False**



Make sure you collect:

____

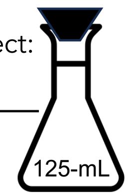
Collect about 50. mL of:

____

Mark the fill line to collect: ______________

Make sure you collect:

____



To determine mass: ____________________ To determine density: ____________________

B

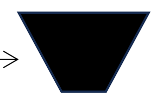
Part B will use the same protocol as part A with the only change being:

____

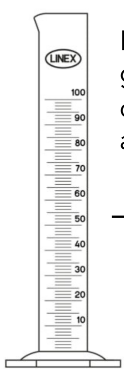
C

Make sure you collect: ____________________

The size stopper you should use for part C:______________



Location of stopper size is on the (circle one): **Bottom / Top/ Side/ Not On** the stopper



Fill the graduated cylinders to about:

____

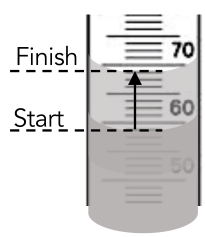
Record the:

____

and

____

volumes



To determine volume of stopper: ____________________________

To determine density of stopper: ____________________________________

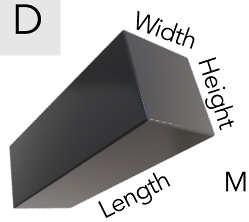
Additional Notes:

Safety Considerations:

Image Guide and Procedure

Name: ______________________ Experiment: ____

D



Before you begin, record the: ____________________________

First, make sure you record: ____________________________

Measure the length, width and height with the: ___________________________

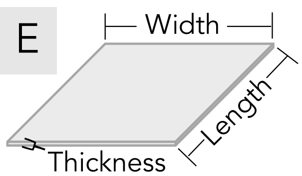
Uncertainty: _____________ meaning the measurement will have _______ decimal points

which will end in ______ or ______

To determine volume of unknown solid: ____

To determine density of unknown solid: ____

E



Make sure you record: ____________________________

Measure the length, width with the: ___________________________

Uncertainty: _____________

Meaning the measurement will have _______ decimal points which will end in ______ or ______

The density of aluminum: _________________ Density equation: __________________________

Rearranged density equation to solve for thickness: _______________________________________
(remember volume (cm^3) = length x width x thickness)

Additional Notes:

____

____

____

____

____

____

____

Safety Considerations:

____

____

____

Name: ____________________ Experiment: ____

Data / Observations

A. Density of water

	Trial 1	Trial 2	Trial 3
1. Mass of flask and rubber stopper (g)	________	________	________
2. Mass of flask, rubber stopper, and 10.00 mL of water (g)	________	________	________
3. Mass of 10.00 mL of water (g)	________	________	________
4. Density of water (g/mL)	________	________	________
5. 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. Mass of flask, rubber stopper, and 10.00 mL of alcohol (g)	________	________	________
3. Mass of 10.00 mL of alcohol (g)	________	________	________
4. Density of alcohol (g/mL)	________	________	________
5. Average density of alcohol (g/mL)		____________	

Name: ____________________ Experiment: ____

C. Density of a rubber stopper	Trial 1	Trial 2	Trial 3
1. Size of the rubber stopper (#)		____________	
2. Mass of the rubber stopper (g)	________	________	________
3. Initial vol. of graduated cylinder (mL)	________	________	________
4. Final vol. of graduated cylinder (mL)	________	________	________
5. Volume of rubber stopper (mL)	________	________	________
6. Density of rubber stopper (g/mL)	________	________	________
7. Average density of rubber stopper (g/mL)		____________	

D. Density of Unknown solid

1. Unknown number ____________________
2. Mass of unknown solid (g) ____________________
3. Length of unknown solid (cm) ____________________
4. Width of unknown solid (cm) ____________________
5. Thickness of unknown solid (cm) ____________________
6. Volume of unknown solid (cm^3) ____________________
7. Density of the unknow metal (g/cm^3) ____________________
8. Identity of metal based on density ____________________

E. Thickness of aluminum foil

1. Mass of aluminum foil (g) ____________________
2. Length of aluminum foil (cm) ____________________
3. Width of aluminum foil (cm) ____________________
4. Volume of aluminum foil (cm^3) ____________________
5. Thickness of aluminum foil (cm) ____________________

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

1. Determine the volume of 1.000g of water ($Density_{water}$ = 0.9999 g/mL)

2. Determine the volume of 1.000g of ice ($Density_{ice}$ = 0.9168 g/mL)

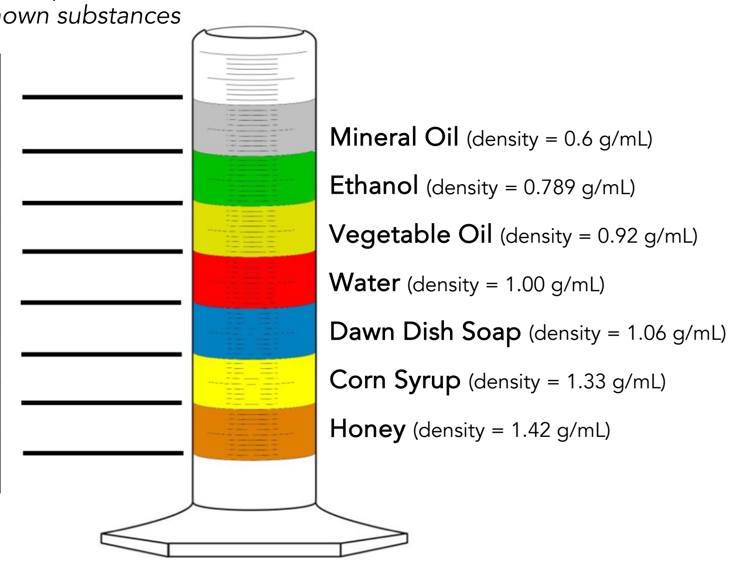
3. Considering the answers to questions 1 and 2, what can we conclude about 1.000 g of water when it freezes?

4. Show where the objects would rest when placed in the container.

Determine placement of unknown substances

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

Unknown substances to order by density



5. 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.

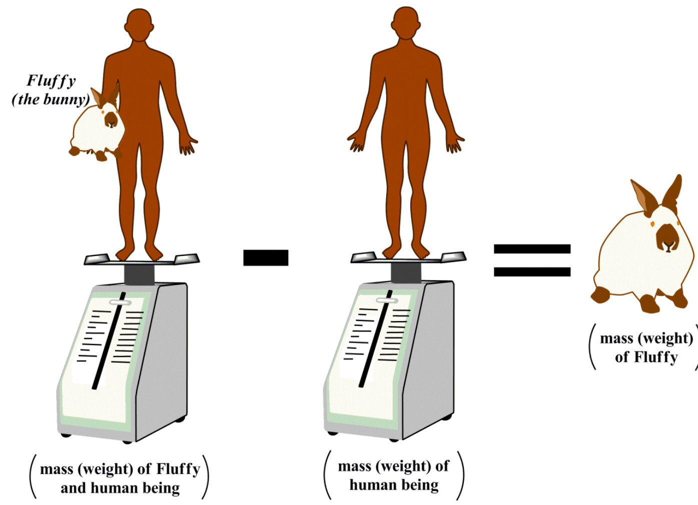
Lab 3:

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).

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.



NaCl (salt) + SiO_2 (sand)

↓ *add water*

(salt) dissolved in water $NaCl_{(aq)}$ → *dry* → **(recovered salt)**

(sand) wet sand $SiO_{2\,(s)}$ → *dry* → **(recovered sand)**

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

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

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

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

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

Purpose

Students will learn skills and apply knowledge necessary to **use physical separation techniques (filtration and boiling) to separate a mixture of salt and sand into its components and calculate the percent composition of each present in the original mixture.**

Skills:

1. Use of a graduated cylinder, ring stand/clamps
2. Determine mass via weigh by difference
3. Separate a mixture via boiling and via filtration
4. Proper handling of hotplate and hot glassware

Knowledge:

1. Understand how physical separation techniques differ from chemical reactions
2. Calculate percent composition of a mixture using the recovered mass of each component

What you will do, briefly:

A. Determine the mass of a sample of a mixture of salt and sand

i. Determine the mass of a vial containing an unknown mixture of sand and salt, then pour some of the mixture out to lightly fill the bottom of a clean, dry 150 mL beaker (Beaker #1) with the sand/salt mixture. Reweigh the original vial (which will still contain some sand/salt). The difference between the vial before and the vial after gives the mass of the sand/salt mixture which was transferred to Beaker #1.

B. Dissolve the salt in water, separate the sand using filter paper and evaporate the water

i. Obtain the mass of a watchglass + filter paper. Use a graduated cylinder to measure out about 50.0 mL of DI water, transfer the water into Beaker #1 and stir for 1-2 minutes.

ii. Following the diagrams in your lab manual, assemble a filtration apparatus. Obtain a second dry 150 mL beaker (Beaker #2) and record its mass, then slowly decant the water/salt/sand mixture onto the filter paper and into Beaker #2 until there is no remaining residue in Beaker #1. When finished, place the filter paper + watch glass on the large hotplate in the fumehood to allow the sand to dry.

iii. Heat Beaker #2 and its contents (i.e. the filtrate) on a hotplate and until the salt is dry. Once Beaker #2 has cooled back down to room temperature, record the mass of Beaker #2 plus its contents (now dried salt).
iv. Using your data, calculate: the mass of sand in the mixture, the mass of salt in the mixture, the percent sand in the mixture, the percent salt in the mixture and the total percent recovery.

Criteria for Success:

1. The correct number of significant figures must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware you are using.

2. The correct units must be used for each measurement, following the metric system

3. Your answers should make logical sense, and calculations/measurements should be repeated otherwise

Name: ______________________ Experiment: ____

Pre-laboratory Assignment:

Please show your work, round using proper significant figures and include units

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?

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

Before beginning, make sure you record the: ______________________________________

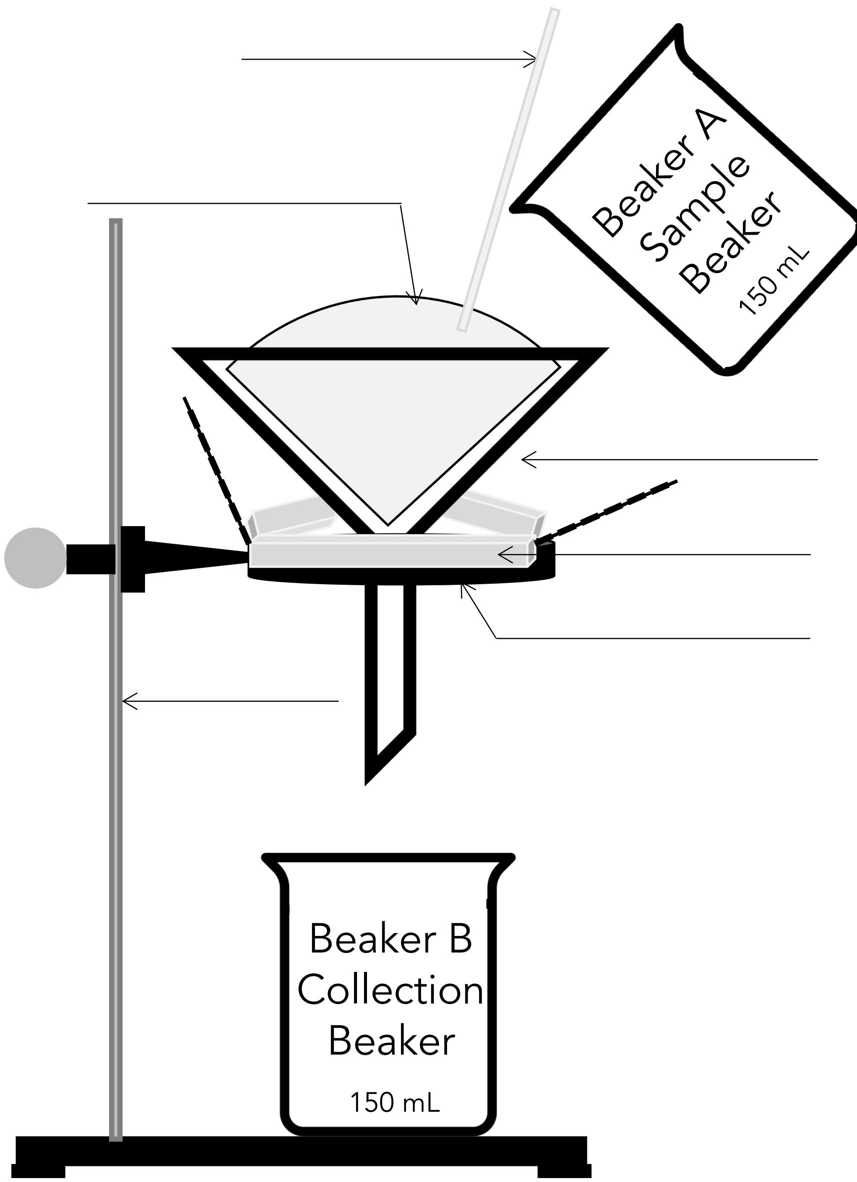
After recording your sample number, determine: ______________________

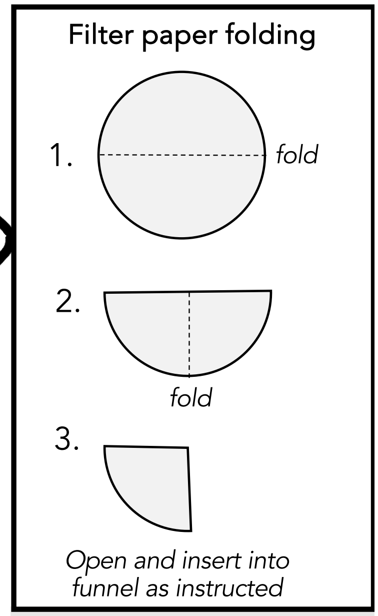
Then add approximately ____________ of ____________ to ____________

Finally determine: ______________________________________

To calculate the mass of sample added: ____________________________

Add approximately ____________ of ____________ to ____________





Additional Notes:

____

____

____

____

____

____

____

Safety Considerations:

____

____

____

Image Guide and Procedure

Name: ______________________ Experiment: ____

Once the sample is ready and setup complete:

Make sure you record:____

Also, you must record:____

1.____

Note: When the sample sticks to the beaker you should:____________________________

____

Note: The liquid (filtrate) that flows to the collection beaker contains: **NaCl (salt)** **SiO_2 (sand)**

Note: The solid collected in the filter paper contains: **NaCl (salt)** **SiO_2 (sand)**

2.____

Note: To avoid splattering, when the volume is about 5 mL you should:________________

3.____

Determine the mass of:______________________________and_________________________

To determine the mass of salt (NaCl):____

To determine the mass of sand (SiO_2):____

To determine mass percent of salt (NaCl):____

To determine mass percent of sand (SiO_2):____

To determine total mass of salt and sand recovered:__________________________________

To determine percent recovery:____

Additional Notes:____

____

____

Name: ____________________ Experiment: ______________________________________

Data / Observations

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 mixture (g):	________	________

B. Separating the Mixture:

		Trial 1	Trial 2
1.	Mass of filter paper and watchglass (g):	________	________
2.	Mass of 150 mL collection beaker (g):	________	________
3.	Mass of 150 mL collection beaker and salt (g):	________	________
4.	Mass of filter paper and watchglass and sand (g):	________	________
5.	Mass of sand (SiO_2) (g):	________	________
6.	Mass of salt (NaCl) (g):	________	________
7.	Percentage of sand in mixture (%):	________	________
8.	Percentage of salt in mixture (%):	________	________
9.	Total mass recovered (g):	________	________
10.	Percentage recovered (%):	________	________

Name: ____________________ Experiment: ______________________________________

Post-laboratory Exercises:

1. 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.

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

3. 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 table salt and recover the two substances.

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.

Lab 4:

Nomenclature of Inorganic Compounds and Common Acids

Introduction

Being able to understand and "speak" the language of chemistry is one of the most crucial skills to master as a beginning chemist. This requires first learning how to identify what type of compound you are dealing with (molecular, ionic fixed charge, ionic variable charge, oxoacid or binary acid), and then learning the rules for naming each type of compound.

Naming category	Identifying factor	Example Formula	Example Name
Molecular	Contain only nonmetals (no ions)	NO_2	Nitrogen dioxide
Ionic (fixed charge)	Contain a metal which forms fixed charge ions (or polyatomic cation) paired with an anion	$AlCl_3$	Aluminum chloride
Ionic (variable charge)	Contains a metal which forms variable charge ions paired with an anion	CuO	Copper (II) oxide
Oxoacid	Are a subcategory of ionic and have H^+ as their cation, are in the aqueous state and their anion contains oxygen	H_2SO_4 (aq)	Sulfuric acid
Binary Acid	Are a subcategory of ionic and have H^+ as their cation, are in the aqueous state and their anion does not contain oxygen	HCl (aq)	Hydrochloric acid

Another skill that is necessary at this juncture is learning how to appropriately balance ionic compounds. In short, a balanced ionic compound needs to have equal magnitudes of positive and negative charges. That means if we are to make an ionic compound between sodium (which ionizes to have a +1 charge, Na^+) and sulfur (which ionizes to have a 2- charge, S^{2-}) our balanced ionic compound must have two sodium ions for every one sulfide ion, which we indicate as subscripts, resulting in an ionic formula of Na_2S.

Purpose

Students will learn skills and apply knowledge necessary to **identity whether a given compound is molecular, ionic (fixed charge), ionic (variable charge), an oxoacid or a binary acid, give the correct name for the compound following IUPAC nomenclature rules and determine the correct chemical formula for an ionic or molecular compound given its IUPAC name.**

What you will do, briefly:

Following the nomenclature guide from **Appendix 2,** complete the worksheet that follows, remembering to take into consideration that the first step is to identify which type of compound you are looking at. When determining the chemical formula for an ionic compound, remember that the total positive ionic charge must equal the total negative ionic charge.

Criteria for Success

1. Compounds will only nonmetals are named with molecular naming conventions.

2. Compounds with fixed cation charge are named with fixed-charge ionic conventions.

3. Compounds with variable cation charge are named with variable-charge ionic conventions.

4. Compounds with H^+ as their cation, are in the aqueous state and contain oxygen are named with oxoacid conventions

5. Compounds with H^+ as their cation, are in the aqueous state and do not contain oxygen are named with oxoacid conventions

---- Pre-laboratory Assignment ----

This week's prelaboratory assignment is to **make flashcards of the polyatomic ions** listed on this page. **For the remainder of the class, knowledge of the name, chemical formula and charge of each polyatomic ion will be crucial.** You will need to bring the completed flashcards to class the day of this lab and show them to your instructor to receive credit.

Polyatomic ion	Formula	Polyatomic ion	Formula
Ammonium ion	NH_4^{1+}	Acetate ion	$C_2H_3O_2^{1-}$
Hydronium ion	H_3O^{1+}	Permanganate ion	MnO_4^{1-}
Hypochlorite ion	ClO^{1-}	Hydroxide ion	OH^{1-}
Chlorite ion	ClO_2^{1-}	Nitrite ion	NO_2^{1-}
Chlorate ion	ClO_3^{1-}	Nitrate ion	NO_3^{1-}
Perchlorate ion	ClO_4^{1-}	Chromate ion	CrO_4^{2-}
Carbonate ion	CO_3^{2-}	Dichromate ion	$Cr_2O_7^{2-}$
Hydrogen carbonate or bicarbonate ion	HCO_3^{1-}	Hydrogen sulfite or bisulfite ion	HSO_3^{1-}
Hydrogen phosphate or biphosphate ion	HPO_4^{2-}	Hydrogen sulfate or bisulfate ion	HSO_4^{1-}
Phosphate ion	PO_4^{3-}	Sulfite ion	SO_3^{2-}
Cyanide ion	CN^{1-}	Sulfate ion	SO_4^{2-}

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

To begin nomenclature analysis, start with the question:

Is there a metal?

Yes — *Ionic species*

No — *Molecular species**

Note: Greek prefixes are **only** used when no metals or polyatomic ions are present

Do we know it's charge?

Yes — Name the metal then use the anion name with the suffix replaced with –ide i.e. sodium fluoride (NaF)

No — Name the metal, place parenthesis with the metal's charge in roman numerals, then use the anion name with the suffix replaced with –ide i.e. iron (II) chloride ($FeCl_2$)

Name using Greek prefixes to indicate the number of each element present. The only exception is when the first element has one, the mono- can be dropped. The last element uses the suffix –ide.
i.e. dinitrogen pentoxide (N_2O_5)
carbon monoxide (CO)

Use the nomenclature rules to name the following species:

1. CsCl ______________________________
2. SO_3 ______________________________
3. PbI_2 ______________________________
4. $KClO_3$ ______________________________
5. Fe_2O_3 ______________________________

Use the nomenclature rules to write the formula for the following:

6. sodium sulfate ______________________
7. potassium hydroxide __________________
8. xenon tetraoxide ____________________
9. lead (II) chromate ___________________
10. diphosphorus pentasulfide _____________

Name: ____________________ Experiment: ____

A Give the formula for the following compounds:

1. carbon dioxide______________________
2. triphosphorous octaoxide____________
3. water ______________________________
4. dinitrogen pentafluoride____________
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__________________

Name: ____________________ Experiment: ____

B Give the names for 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. $Fe(C_2H_3O_2)_2$ ______________________________

There is no post-laboratory assignment this week

Lab 5:

Analysis of Hydrates

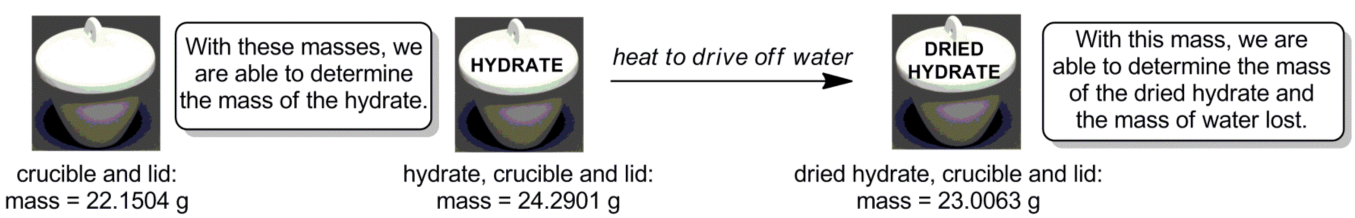
Introduction

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:

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

Purpose

Students will learn skills and apply knowledge necessary to **determine the mass percent of water in an unknown hydrate sample and use it to correctly identify the chemical formula of the hydrate.**

Skills

1. Safe and correct use of Bunsen burners for both slow and strong heating
2. Determine mass via weigh by difference
3. Proper use of crucible/crucible tongs for evaporation

Knowledge

1. Relate mass of a sample to the mass of its components (I.e. mass percent)
2. Describe why a hydrate can be dried using physical separation techniques
3. Determine chemical formula of a hydrate from its mass percent

What you will do, briefly:

A. Obtain the mass of a clean, dry crucible then fill it with about 2.0 grams of unknown hydrate. Reweigh the crucible (now containing hydrated salt) and determine the mass of the hydrated salt by difference. Heat your crucible, with the lid ajar over a Bunsen burner, slowly for 5 minutes and strongly for 10 minutes. Allow the crucible to cool completely, then obtain the mass of the crucible, lid and dehydrated salt.

B. Use your data to determine the mass of the unknown hydrate, mass of the dehydrated salt, mass of water lost and the percent water in the unknown hydrate.

Criteria for Success:

1. Correct significant figures and units must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware used.
2. The mass percent of water calculated should closely match (within 5%) one of the options for your unknown in the post-laboratory assignment
3. One or more errors are identified which feasibly explain any variation from calculated mass percent

Name: ____________________ Experiment: ______________________________________

Pre-laboratory Assignment:

Please show your work, round using proper significant figures and include units

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

2. 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.

   a. Calculate the mass of the water lost.

   a. Calculate the mass percent of the water in the hydrate

$$\text{Precent composition of water (\%)} = \frac{\text{mass of water lost}}{\text{mass of entire hydrate}} \times 100$$

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

Before you collect your sample, you must: ______________________________________

____

____

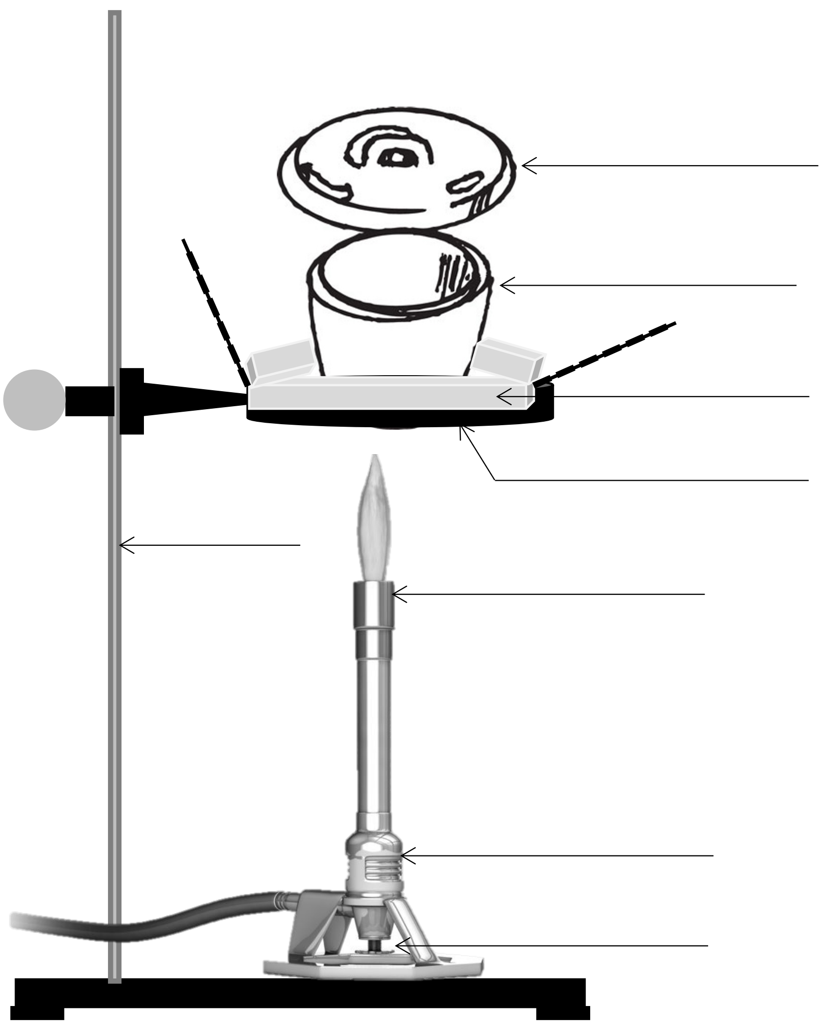
Then, make sure you record the: ______________________________________

After recording your sample number, determine: ______________________

Then add approximately __________ of ______________________________

Finally determine: ____

To calculate the mass of sample added: ____________________________



Additional Notes:

Safety Considerations:

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

Once the sample is ready and setup complete:

1. ____
2. ____
3. ____

To determine the mass of the dehydrated salt: ______________________________

To determine the mass of the water lost by the salt: ___________________________

To determine the percent of water in hydrated salt: ___________________________

Data:

Unknown #: ________	Trial 1	Trial 2
Mass of dry crucible dish with lid (g)	________	________
Mass of dry crucible dish with lid and hydrated salt (g)	________	________
Mass of unknown hydrate (g)	________	________
Mass of crucible dish with lid and dehydrated salt (g)	________	________
Mass of dehydrated salt (g)	________	________
Mass of water lost (g)	________	________
Percent of water in unknown hydrate (%)	________	________
Average percent of water in hydrate (%)	________	

Additional Notes: ____

Name: ____________________ Experiment: ______________________________________

Post-laboratory Exercises:

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.

$$\text{Percent composition of water (\%)} = \frac{\text{mass of water}}{\text{mass of entire hydrate}} \times 100$$

For $LiNO_3 \cdot 3H_2O$ (lithium nitrate trihydrate) we must 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	H: 6 x 1.01 g/mol
O:3 x 16.00 g/ mol	O: 3 x 16.00 g/mol
68.95 g/mol	54.06 g/mol

$$\text{Percent composition of water (\%)} = \frac{3\ H_2O}{LiNO_3 \cdot 3H_2O} \times 100$$

$$\text{Percent composition of water (\%)} = \frac{54.06\ \text{g/mol}}{(68.95\ \text{g/mol} + 54.06\ \text{g/mol})} \times 100$$

Percent composition of water in $LiNO_3 \cdot 3H_2O$ = 43.9%

1. Calculate the percent composition of water in the following hydrates. Please show all work. One of these values will correspond with your experimental results.

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

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

2. Based on your results from question 1, did your experimental results match one of the three calculated percent compositions? Fill in the following table

Percent composition from Trial 1	Percent composition from Trial 2	Calculated percent composition

If you 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?

3. 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?

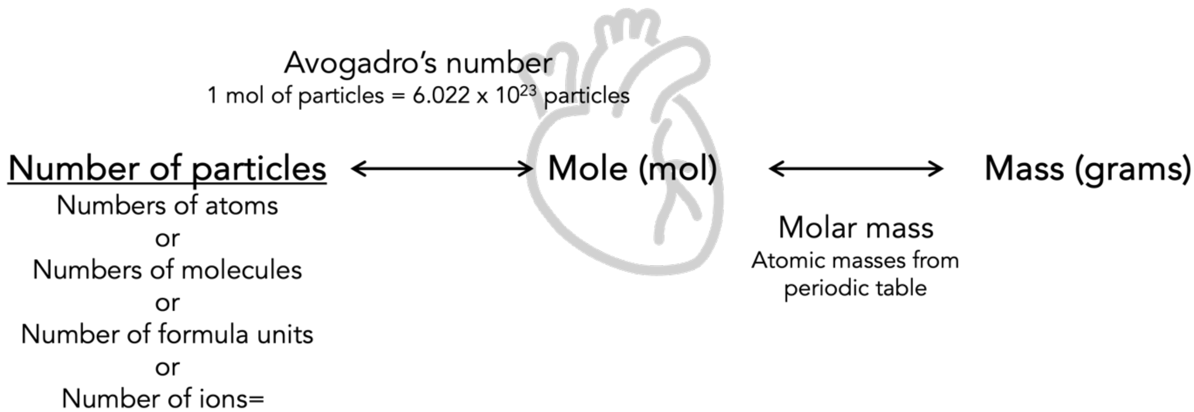
4. Would the student from question 3 calculate the percent water in the hydrate to be high, low, or unaffected? Explain why or why not.

Lab 6:

Analysis of Compounds

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.



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?

$$\frac{4.508\text{ g Al}}{} \left| \frac{1\text{ mol Al}}{26.98\text{ g Al}} \right| \frac{6.022 \times 10^{23}\text{ Al atoms}}{1\text{ mol Al}} = 1.006 \times 10^{23}\text{ Al atoms}$$

Purpose

Students will learn skills and apply knowledge necessary to **experimentally determine the empirical formula of magnesium oxide, using techniques such as weighing by difference and careful heating of a solid with a Bunsen burner, using a crucible and crucible tongs.**

Skills:

1. Safe and correct use of Bunsen burners for both slow and strong heating
2. Determine mass via weigh by difference
3. Proper use of crucible/crucible tongs
4. Proper waste disposal

Knowledge:

1. Convert mass to moles using molar mass
2. Determine the mole ratio of two elements in a compound to one another

What you will do, briefly:

A. Add a small, rolled up strip of magnesium metal to a clean, dry pre-weighed crucible, and determine its mass by difference. Assemble the apparatus shown on the following page, ensuring the crucible lid is slightly ajar to allow for oxygen gas to enter the crucible.

B. Heat the crucible slowly for 5 minutes, then heat strongly for 10 minutes. Ensure the magnesium has become mostly white powder, then cool the crucible until it no longer radiates heat.

C. To the crucible, use a dropper to add just enough DI water to cover the solid, then place the crucible on the hotplate in the fumehood to allow the water to evaporate. This converts undesired side products (like Mg_3N_2, $Mg(OH)_2$ and $MgCO_3$) back into magnesium oxide, NH_3 (g) CO_2 (g) and H_2O (g).

D. Repeat the Bunsen burner heating process again to ensure only magnesium oxide remains, then allow your crucible to cool completely. Once it is cooled, obtain the mass of the crucible, lid and the magnesium oxide.

E. Use your data to calculate: the mass of magnesium metal used, the mass of magnesium oxide produced, the mass of oxygen gas that reacted, the moles of magnesium and moles of oxygen, ratio of magnesium, ratio of oxygen and the empirical formula.

Criteria for Success:

1. Correct significant figures and units must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware used.
2. The experimentally determined empirical formula for magnesium oxide should correlate with the expected chemical formula for magnesium oxide, considering the ions made from Mg & O.
3. One or more errors are identified which feasibly explain any variation from the expected empirical formula.

Name: ____________________ Experiment: ______________________________________

Pre-laboratory Assignment:

Please show your work, round using proper significant figures and include units

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

2. What is the mass (in grams) of 7.11 X 1022 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.

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________

Before you collect your sample, you must: ______________________________

____

____

Before beginning, record: ______________________________________

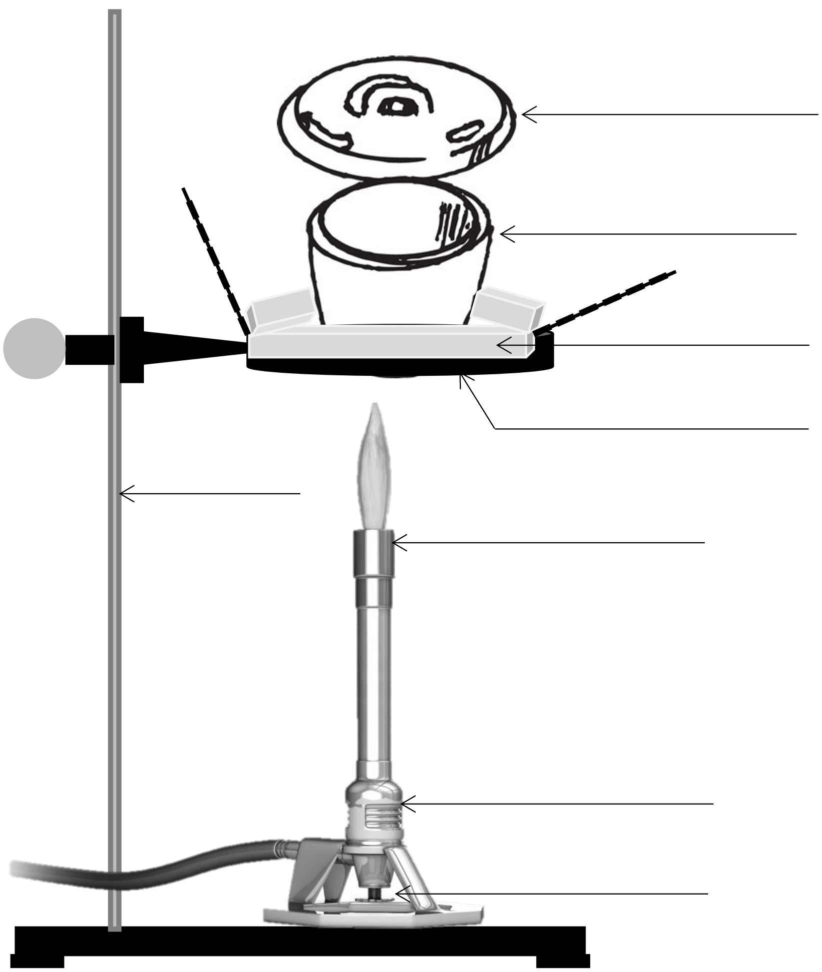
Obtain approximately ________________ of ______________________

To your sample make sure you: ________________________________

Before putting in crucible, you must: ____________________________

Finally, record: ____

To determine the mass of Mg: ________________________________



Additional Notes:

Safety Considerations:

Image Guide and Procedure

Name: ______________________ Experiment: ____

Once the sample is ready and setup complete:

1. ____

Note: If you observe white smoke: ____

Note: The reaction is complete when the product is: _____________________________________

2. ____

Note: Step 2 ensures the removal of __________________ from the product.

3. ____

Note: Step 3 ensures the removal of ______________ and ______________ from the product.

4. ____

5. ____

To determine the mass of magnesium: ____

To determine the mass of oxygen: ____

To determine the moles of magnesium: ____

To determine the moles of oxygen: ____

$$\text{Empirical Formula} = Mg_{\frac{\text{moles}Mg}{\text{smallest moles}}}O_{\frac{\text{moles}O}{\text{smallest moles}}}$$

Additional Notes: ____

____

____

Name: ____________________ Experiment: ____

Data / Observations

Determination of an Empirical Formula

1. Mass of dry crucible dish with lid (g): ____________
2. Mass of dry crucible dish with lid and magnesium (g): ____________
3. Mass of magnesium (g): ____________
4. Mass of crucible dish with lid and magnesium oxide (g): ____________
5. Mass of magnesium oxide (g): ____________
6. Mass of oxygen (g): ____________
7. Moles of magnesium (mol): ____________
8. Moles of oxygen (mol): ____________
9. Ratio of magnesium : oxygen ____________
10. Ratio of oxygen : magnesium ____________
11. Empirical formula of magnesium oxide: ____________

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

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

    a. During the experiment, water was added to make sure all 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 while the dish was still warm. Discuss the problem with weighing objects when they are not the same temperature as the balance.

2. What is the empirical formula for a compound containing 26.5% potassium, 35.36% chromium, and 38.07% oxygen

Name: ____________________ Experiment: ______________________________________

Post-laboratory Exercises:

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?

Lab 7:

Observations of Chemical Reactions

Introduction

Chemistry is the study of matter and the energy it exchanges when it rearranges. When two substances are combined together in a mixture, the resulting mixture tends to be an averaging of the physical properties of the components in the mixture. For example, when salt is mixed with water, we end up with salty water. On the other hand, when two substances are combined and result in a chemical reaction, the properties of the product are often nothing like the properties of the individual reactants. For example, the formation of NaCl (table salt) from its elements involves sodium metal (a reactive, silver metal) reacting with elemental chlorine (a green gas). In today's lab you will be learning what types of observations indicate a chemical reaction (or a rearrangement of matter) has taken place.

Most commonly, these observations can be binned into one of four categories:

- Formation of a precipitate (solid forms or a solution becomes cloudy)
- Formation of a gas (bubbles, fizzing, new odor)
- Permanent unexpected color change
- Release or absorption of energy (the test tube becomes hotter or colder after the substances are combined)

It is worth noting that a change of phase (like melting, freezing, vaporization, condensation, deposition or sublimation) is a physical change, not a chemical reaction, since the composition of the matter is not changed. In other words, ice, water and steam are all comprised of H_2O, and these molecules remain intact when shifting from one physical phase to another (it is the interaction between water molecules that changes).

Purpose

Students will learn skills and apply knowledge necessary to **determine if a chemical reaction has taken place based on observations of solid formation, gas formation, unexpected color change and release/absorption of energy.** This lab will prepare students for next week's lab on writing and balancing chemical equations to express the reactions they observe.

Skills

1. Learn how to efficiently conduct a qualitative experiment
2. Learn to identify a chemical reaction has taken place
3. Learn to properly record a scientific observation

Knowledge

1. Identify the four types of observations which indicate a chemical reaction has taken place
2. Differentiate between a chemical and physical change

What you will do, briefly:

For reactions A – L, add the two chemicals specified in approximate amounts in a test tube, mix well and record your observations.

Criteria for Success:

1. Recorded observations should be descriptive and fall into one or more of the four main observation types
2. Waste should be disposed of in the correct container for each experiment.

There is no pre-laboratory assignment this week

It will be useful to review:

1. Reaction types
2. Evidence of chemical reactions

Image Guide, Procedure and Data Sheet

Name: ____________________ Experiment: ______________________________________

Four ways to identify a chemical reaction:

1. ______________________________ 3. ______________________________

2. ______________________________ 4. ______________________________

True or False: A cloudy or opaque product solution from two clear reactants can be considered a precipitate. **True** **False**

A

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

B

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

C

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

D

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

Safety Considerations: ____

____

____

____

Image Guide, Procedure and Data Sheet

Name: ____________________ Experiment: ______________________________________

E

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

F

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

G

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

H

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

I

Add __________ of ____________________ ( )

Add __________ of ____________________ ( )
phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

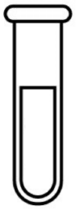
Observations: ____

Safety Considerations: ____

Image Guide, Procedure and Data Sheet

Name: ____________________ Experiment: ______________________________________

E



Add ____________ of ______________________ ( )

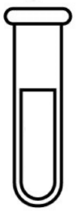
Add ____________ of ______________________ ( )

phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

F



Add ____________ of ______________________ ( )

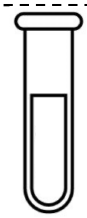
Add ____________ of ______________________ ( )

phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

G



Add ____________ of ______________________ ( )

Add ____________ of ______________________ ( )

phase

Mix well and make observations. Make sure to feel the temperature of the test tube and dispose of waste correctly

Observations: ____

Safety Considerations: ____

Name: ____________________ Experiment: ______________________________________

Post-laboratory Exercises:

1. Answer the following questions based on the 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?

2. 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?

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

Name: ____________________ Experiment: ______________________________________

Post-laboratory Exercises:

5. Balance the following equations:

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

   b. C_2H_6 + O_2 → H_2O + CO_2

   c. KOH + H_2SO_4 → K_2SO_4 + H_2O

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

Lab 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 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.

$$CaCl_2(aq) + 2\,AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\,AgCl(s)$$

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

$$\boxed{Ca^{2+}(aq)} + 2\,Cl^-(aq) + 2\,Ag^+(aq) + \boxed{2\,NO_3^-(aq)} \rightarrow \boxed{Ca^{2+}(aq)} + \boxed{2\,NO_3^-(aq)} + 2\,AgCl(s)$$

Spectator ions appear on both the products and reactants side in the same form

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

$$2\,Cl^-(aq) + 2\,Ag^+(aq) \rightarrow 2\,AgCl(s)$$

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

Compounds containing the following ions are generally **insoluble**	**Exceptions** (when combined with ions on the left the compound is **soluble**)
OH^-	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
S^{2-}	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
CO_3^{2-}, PO_4^{3-}	Li^+, Na^+, K^+, NH_4^+

For a more complete solubility table, reference Appendix 3

Purpose

Students will learn skills and apply knowledge necessary to **determine if a chemical reaction has taken place based on observations of solid formation, gas formation, unexpected color change and release/absorption of energy.** Students will then **write and balance a chemical reaction to describe what they have observed.**

Skills:

1. Learn how to efficiently conduct a qualitative experiment, identify a chemical reaction has taken place and properly record a scientific observation
2. Predicting a chemical reaction based on the reactants and identify which product is responsible for a given observation

Knowledge:

1. Identify the four types of observations which indicate a chemical reaction has taken place
2. Understand synthesis, decomposition, single-replacement and double displacement reaction patterns
3. Correctly balance a chemical reaction
4. Write a balanced complete and net ionic equation for a reaction

What you will do, briefly:

For reactions A – J, add the two chemicals specified in approximate amounts in a test tube, mix well and record your observations. For any combinations in which a chemical reaction was observed, write the balanced chemical equation, the balanced complete ionic equation and the balanced net ionic equation.

Criteria for Success:

1. Recorded observations should be descriptive and fall into one or more of the four main observation types.

2. Waste should be disposed of in the correct container for each experiment.

3. Your written chemical reactions should be properly balanced and should support your observations.

There is no pre-laboratory assignment this week

It will be useful to review:

1. Double displacement reactions
2. Balancing ionic compounds
3. Balancing molecular, complete and net ionic equations

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________

A

Add __________ of ____________________

Add __________ of ____________________

We expect the reaction to form (circle one):

Precipitate / Gas / Water + Salt / No Reaction Expected

Beaker diagram

Expected chemical reaction:

____

Total Ionic reaction: ____

Net Ionic reaction: ____

B

Add __________ of ____________________

Add __________ of ____________________

We expect the reaction to form (circle one):

Precipitate / Gas / Water + Salt / No Reaction Expected

Beaker diagram

Expected chemical reaction:

____

Total Ionic reaction: ____

Net Ionic reaction: ____

C

Add __________ of ____________________

Add __________ of ____________________

D

Add __________ of ____________________

Add __________ of ____________________

Additional Notes:

Safety Considerations:

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

E

Add ____________ of ____________________

Add ____________ of ____________________

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

---

F

Add ____________ of ____________________

Add ____________ of ____________________

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

---

G

Add ____________ of ____________________

Add ____________ of ____________________

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

---

H

Add ____________ of ____________________

Add ____________ of ____________________

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

---

I

Add ____________ of ____________________

Add ____________ of ____________________

J

Add ____________ of ____________________

Add ____________ of ____________________

Additional Notes:

____

____

____

____

Safety Considerations:

____

____

____

Mix well and make observations. Then complete the molecular, total ionic and net ionic equations

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

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

3. Consider the reaction between sulfur trioxide and water:

$$SO_3\ (g) + H_2O\ (l) \rightarrow H_2SO_4\ (aq)$$

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.

Name: ______________________ Experiment: ____

Post-laboratory Exercises:

4. 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?

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)\;\Delta H^{\circ}_{rxn} = -802.3\text{ kJ/mol}$$

5. 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.

Lab 9:

Atomic Emission Spectra and Electron Configurations

Introduction

When you observe a rainbow after a rainstorm, this is an example of a **continuous light spectrum**. Raindrops lingering in the air act as a prism, separating white light from the sun by wavelength. Our eyes perceive a portion of the electromagnetic spectrum called **visible light.** Within the visible light spectrum, red light has the longest wavelength and violet light has the shortest wavelength. A prism will bend light of shorter wavelengths more than light of longer wavelengths, which is why we always observe rainbows with red at the top and violet at the bottom!

Elements also emit light when sufficient energy is added to them. However, unlike sunlight, when elements are "excited" (meaning we have added sufficient electrical or heat energy to allow their electrons to "jump" from a lower energy state to a higher energy state), they do not produce a continuous spectrum, but instead produce what is called **atomic line spectra**. Every element has a distinct line spectrum, and these lines can be observed using an instrument called a **spectroscope**, which acts as a prism.

These emission lines are evidence of **quantum mechanics**, which is the idea that electrons occupy orbitals of distinct energy levels. Like rungs on a ladder, orbitals only exist in specific locations. So, when electrons "relax" from a higher energy state to a lower energy state, a "packet" of light, called a **photon**, is emitted which has the exact energy as the energetic distance between the higher and lower states. Since the energy of a photon is directly related to its wavelength, we can now see that the colored lines we observe from excited elements each correspond to a specific transition between two energy states in that element.

We can determine **the ground state electron configuration** (or the lowest energetic state of its electrons) for an element based on the following three key principles:

1. **Aufbau Principle:** Electrons always fill from the lowest energy orbital to the highest energy orbital. This is analogous to filling up a glass with water: The water must always fill from bottom to top.

2. **Hund's Rule:** When there are isoenergetic orbitals (multiple orbitals with the same energy), electrons will always occupy an orbital singly before pairing with another electron.

3. **Pauli exclusion principle:** An orbital can hold a maximum of two electrons, and the two electrons must have opposing spins. We indicate this by drawing upward pointing arrow for "spin up" and a downward pointing arrow for "spin down".

Each **principal energy level** is made up of one or more **subshells**, and each subshell is made up of **orbitals.** S subshells have 1 orbital, p subshells have 3 orbitals, d orbitals have 5 orbitals, and f orbitals have 7 orbitals. The following table can help us see how many electrons can exist in a given principal energy level:

Principal energy level	Subshell	# of orbitals	# of electrons	Maximum # of electrons
1	1s	1	2	2
2	2s	1	2	2 + 6 = 8
	2p	3	6	
3	3s	1	2	2 + 6 + 10 = 18
	3p	3	6	
	3d	5	10	
4	4s, 4p, 4d, 4f	1, 3, 5, 7	2, 6, 10, 14	2 + 6 + 10 + 14 = 32

Purpose

Students will learn skills and apply knowledge necessary to **differentiate between and observe continuous and atomic emission spectra using a spectroscope, and to determine the electron configuration for elements by following Hund's Rule, the Aufbau Principle and the Pauli Exclusion principle.**

Skills:

1. Learn how to use a spectroscope to approximate the wavelength of emission spectral lines
2. Learn how to determine the electron configuration for a given element based on its location on the periodic table

Knowledge:

1. Relate the line spectra of an element to its valence electronic transitions
2. Differentiate between energy absorption and emission
3. Apply Hund's Rule, the Aufbau Principle and the Pauli Exclusion Principle to show the ground state electron configuration for any element

What you will do, briefly:

A. Using the spectroscope and colored pencils, observe and color in the emission spectra from (1) sunlight (2) elements lamps and (3) a fluorescent light.

B. Complete a "paper lab" assignment to complete orbital diagrams and write electron configurations for elements.

Criteria for Success:

1. You should observe/draw a continuous spectrum for sunlight and line spectra for the element lamps and fluorescent light
2. Your orbital diagrams and electron configurations for each element should be correct, following Hund's Rule, the Aufbau Principle, and the Pauli Exclusion Principle

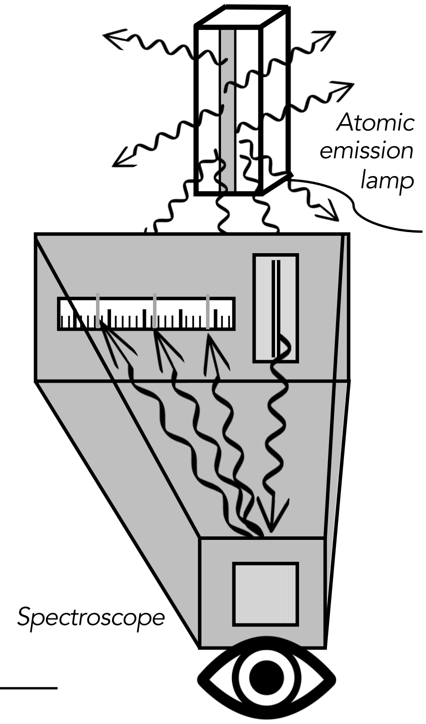
There is no pre-laboratory assignment this week

It will be useful to review:

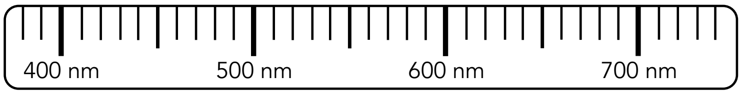
Chapter 9 of the textbook, discussing the Bohr model of the atom and the electromagnetic spectrum

Image Guide and Procedure

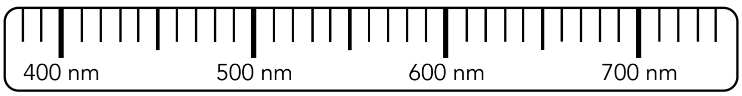
Name: ____________________ Experiment: ______________________________________



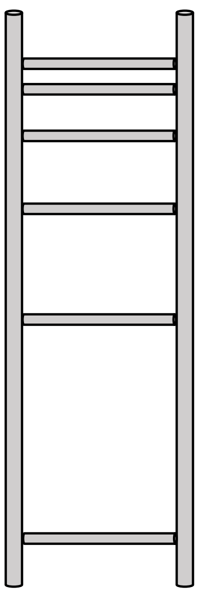
Estimated continuous spectrum



Estimated atomic spectrum



Mark three atomic lines and label what color you expect based on the continuous spectrum



n = _____
n = _____
n = _____
n = _____
n = _____
n = _____
(also called the _________ state)

Aufbau principle: ______________________________

Pauli exclusion principle: ______________________________

Hund's rule: ______________________________

7s			
6s	6p	6d	
5s	5p	5d	5f
4s	4p	4d	4f
3s	3p	3d	
2s	2p		
1s			

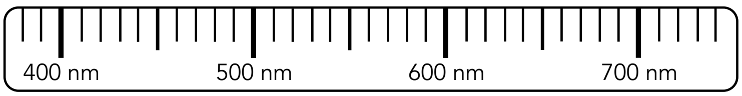
Complete the atomic filling order: ______________________________________

Additional Notes and Safety Considerations:

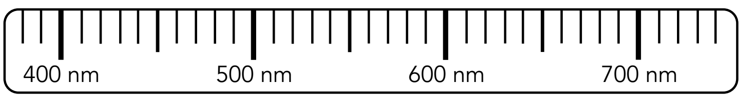
Name: ____________________ Experiment: ______________________________________

A Emission Spectrum

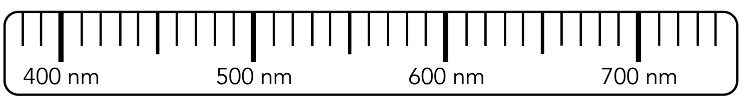
Draw the spectrum as seen through the spectroscope, taking care to apply an appropriate color and location for an observed line



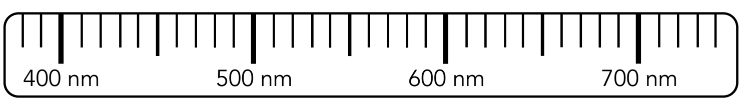
Continuous spectrum (white light)



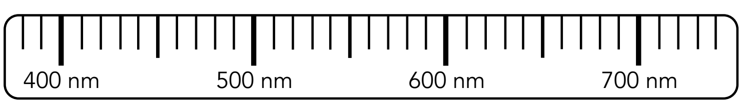
Argon



Helium



Hydrogen

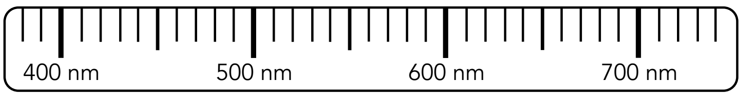


Neon

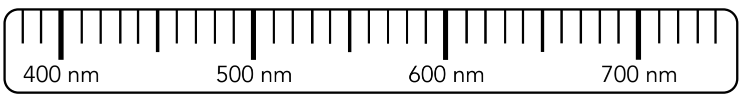
Name: ____________________ Experiment: ______________________________________

A Emission Spectrum

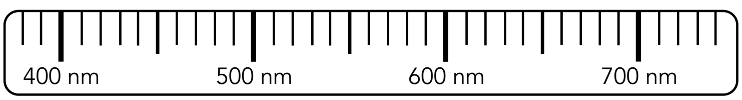
Draw the spectrum as seen through the spectroscope, taking care to apply an appropriate color and location for an observed line



Krypton



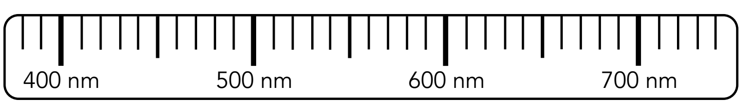
Mercury



Xenon

Identifying Elements in Fluorescent Light

Compare each individual emission spectra to that of the fluorescent light emission spectrum. Identify two elements found in fluorescent lights based on your observations.

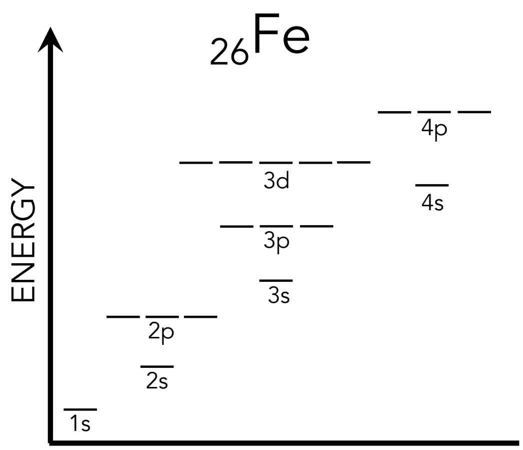


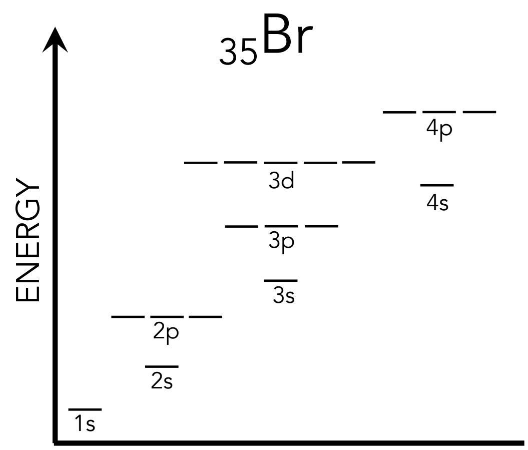
Fluorescent Light
(overhead lights)

Name: ____________________ Experiment: ______________________________________

B Orbital Diagrams & Electron Configuration

Fill in the electrons for the given elements considering the Aufbau principle, Pauli exclusion principle and Hund's rule





Give the electron configurations for the following and **circle** the valence electrons

Electron configuration for V=____

Electron configuration for As=____

Give the electron configuration, number of valence electron(s) and group number for the following elements:

	Electron Configuration	# of valence e^-	Group #
$_{6}C$			
$_{7}N$			
$_{8}O$			
$_{38}Sr$			
$_{16}Sr$			
$_{16}S$			
$_{55}Cs$			
$_{17}Cl$			
$_{13}Al$			

Name: ____________________ Experiment: ______________________________________

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, or condensed electron configuration uses noble gases in column 8A or 18 to stand for a certain number of core electrons. [Ne] = 10 electrons, [Kr] = 36 electrons, and so on. Magnesium, with twelve electrons, is $1s^22s^22p^63s^2$ or we can use condensed electron configuration: [Ne] $3s^2$. The [Ne] stands for $1s^22s^22p^6$ (ten electrons).

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

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

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

Lab 10:

Lewis Structures and Molecular Geometry

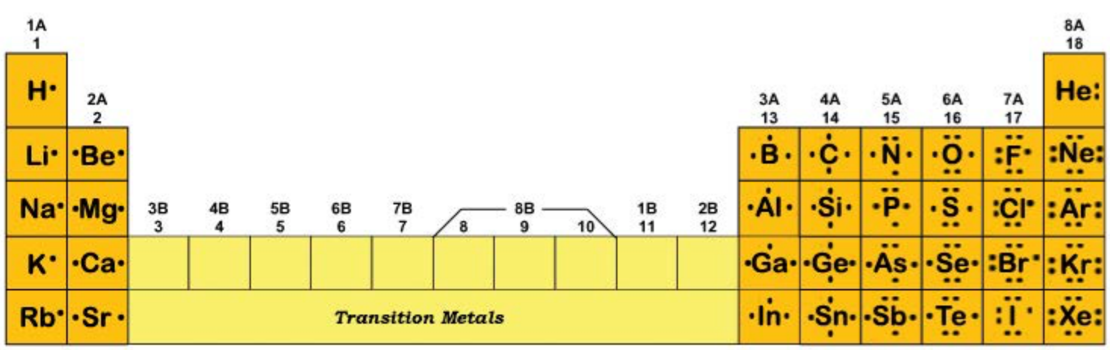
Introduction

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.

Determining 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 10.1).** 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 10.1: 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.

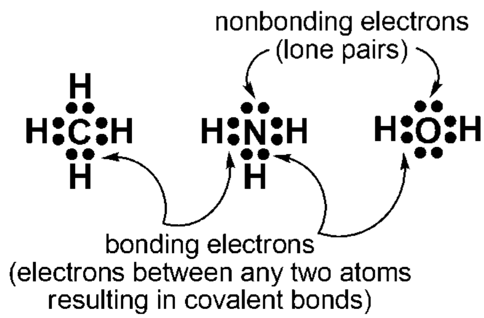


Figure 10.1: Valence electrons in electron-dot structures representing bonding and nonbonding electrons.

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.

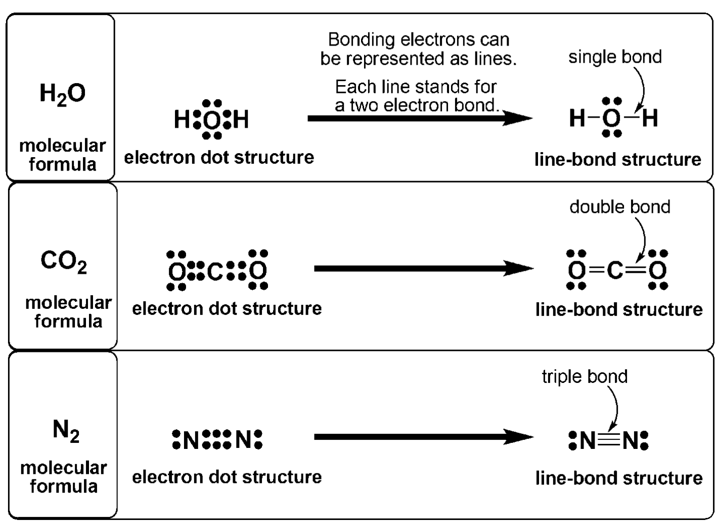


Figure 10.2: Molecules with single, double, and triple covalent bonds. Each line represents two- electron covalent bonds.

When determining the **Lewis Structure** for a covalent molecule, the following steps should be taken:

1. Place the atoms relative to each other, creating a **skeletal structure** for the molecule. Place atoms with the lower group number in the center, because they need more electrons to fill their octet. Hydrogen is an important exception to this rule, and like the halogens, will always be a terminal atom.

2. Determine the **total number of valence electrons** in the molecule. Add up the number of valence electrons from each individual atom in the molecule. This is the total number of electrons which can contribute to forming covalent bonds or lone pairs.

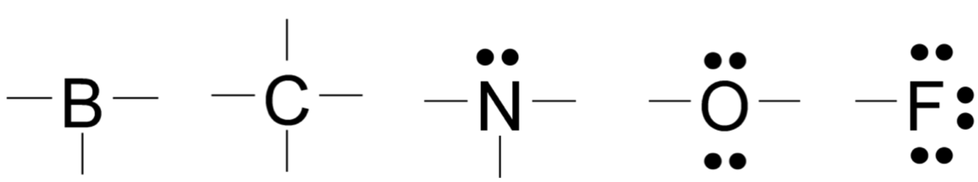
3. Draw **single bonds** from each surrounding atom to the central atom and subtract two electrons from the valence total for each bond. These are the electrons that are used for forming single bonds.

4. Distribute the remaining electrons as **lone pairs** around each atom, first placing lone pairs on the surrounding atoms, then the central atom last.

5. If the central atom does not end up with a full octet, move a lone pair to make a **double or triple bond**.

There are some common **bonding patterns** that are useful to keep in mind when you are getting used to drawing Lewis structures:

1. Hydrogen will only ever found one bond, and so must be **terminal** (bonded to only one other atom)
2. Halogens usually form one bond and have three lone pairs (i.e. F)
3. Carbon usually form four bonds (either single or multiple) and has no lone pairs, so is usually **central** (bonded to two or more atoms)
4. Nitrogen usually forms three bonds (either single or multiple) and has one lone pair
5. Oxygen usually forms two bonds (either single or multiple) and has two lone pairs



Sometimes (there are many exceptions to this rule) molecular formulas are written in a way which indicates the way the molecule is bonded. One example is CH_3OH, methanol, which is a carbon bonded to three hydrogen atoms and one oxygen, and that oxygen is bonded to a hydrogen.

One other consideration is that molecules tend to be symmetrical. So the molecule with formula C_2Cl_4 is drawn as two carbons double bonded to each other, and each carbon single bonded to two chlorines.

Molecular Geometry

A valid Lewis 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

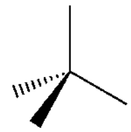
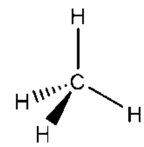
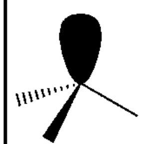
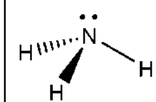
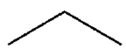
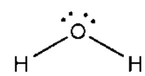
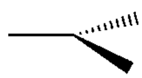
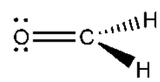
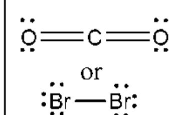
2. **Pyramidal**: Three groups of shared electrons (three single bonds) and one lone pair of unshared electrons around a central atom.

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 an atom is always considered linear.

Table 10.2: Molecular geometries determined by bonding and nonbonding groups around the central atom

	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	

Purpose

Students will learn skills and apply knowledge necessary to **determine the total number of valence electrons in a compound, draw the Lewis structure and determine the geometry around each central atom based on its electron groups.**

There is no pre-laboratory assignment this week

It will be useful to review:

Chapter 10 of the textbook: Lewis Structures and molecular geometry

Image Guide and Procedure

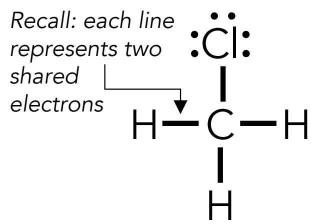
Name: ____________________ Experiment: ______________________________________

Rules for Lewis structures

1. Place the atoms relative to each other
   Place the atoms with a lower group number in the center because it needs more electrons to fill its octet
2. Determine the total number of valence electrons
   For polyatomic ions, add an electron for every negative charge on the ion, subtract an electron for every positive charge on the ion
3. Draw a single bond from each surrounding atom to the central atom and count two electrons from the valence total for each bond
4. Distribute the remaining electrons in pairs around each atom so that each end up with a full octet
5. If a central atom does not end up with an octet, form one or more multiple bonds

*Molecule (central atom in **bold**)*	*Number of valence electrons (use periodic table)*	*Lewis structure*	*Molecular shape around central atom*
CH_3F			
$\mathbf{C_2}H_4$			
CN^-			

Name: ____________________ Experiment: ______________________________________

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)
CH_3Cl	C: 4 x 1 = 4 e$^-$ H: 1 x 3 = 3 e$^-$ Cl: 7 x 1 = 7 e$^-$ Total: 14 e$^-$		tetrahedral
H_2	H:		
Cl_2	Cl:		
HCl	H: Cl:		
HBr	H: Br:		
ICl	I: Cl:		
CH_4	C: H:		

Name: ____________________ Experiment: ______________________________________

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)
$\mathbf{C}H_3Br_2$	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:		

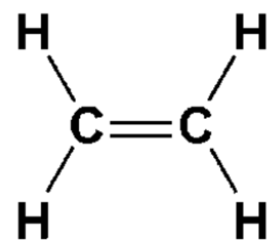
Name: ____________________ Experiment: ______________________________

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)
$\mathbf{C}_2H_4$	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:		
$\mathbf{C}_2H_2$	N: H:		C: C:
$H_3\mathbf{CCN}$	H: C: N:		C: C:
$\mathbf{Si}O_2$	Si: O:		
$\mathbf{C}_3H_8$	C: H:		C: C: C:

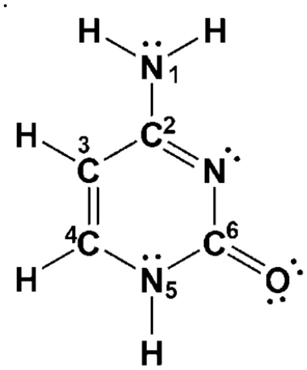
Name: ______________________ Experiment: ____

Post-laboratory Exercises:

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



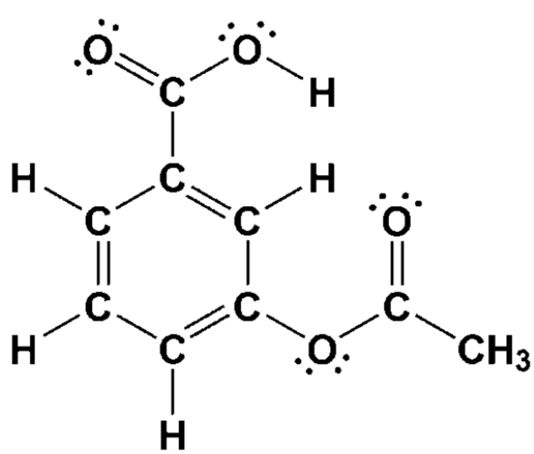
2. Cytosine is a pyrimidine base found in both RNA and DNA. Identify the geometry around each of the central atoms below:



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 central atoms.



Molecular geometry

1.________________
2.________________
3.________________
4.________________
5.________________
6.________________
7.________________
8.________________
9.________________
10.________________
11.________________

Lab 11:

Electronegativity and Molecular Polarity

Introduction

This week will expand upon the Lewis Structures we drew last week, and it may be useful to re-read Lab 10's Introduction section. We will now consider electronegativity differences between atoms within a molecule and determine the polarity of each bond and ultimately the polarity of the molecule.

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 11.1 is known as the Linus Pauling's electronegativity scale.

Table 11.1: Linus Pauling electronegativity numbers

Electronegativity increases across a period.

Electronegativity increases going up a column.

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																

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 11.2 shows the electronegativity differences that result in different types of bonds.

Table 11.2: Electronegativity differences resulting 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 → $\delta+$ H—Cl $\delta-$ (2.1, 3.0)
> 1.9	ionic bond	explicit charges $[K]^+[Cl]^-$ (0.8, 3.0)

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 ($\boldsymbol{\delta}^+$) or partially negativity charged ($\boldsymbol{\delta}^-$) 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 negatively charged ($\boldsymbol{\delta}^-$). An electrostatic potential is mapped onto a molecule to represent electron rich and poor regions as shown in Figure 11.1.

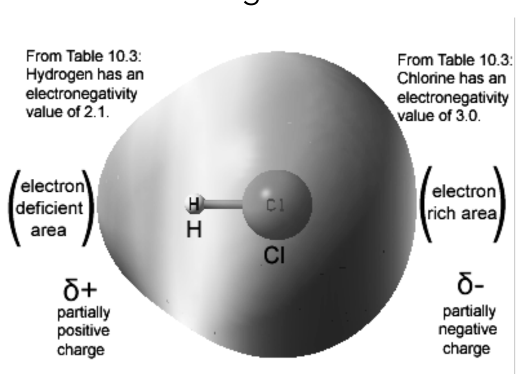


Figure 11.1: Electrostatic potential map of hydrogen chloride (HCl) with darker gray representing electron rich area and lighter gray representing electron deficient region.

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. 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**.

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 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 11.2.

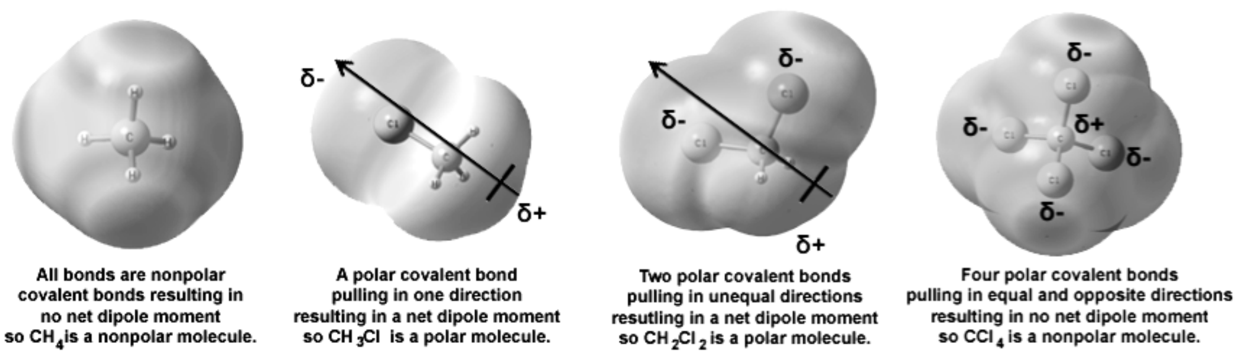


Figure 11.2: Examples of dipoles. Dipoles can be added like vectors and may cancel out.

In general, **with geometries up to tetrahedral, one or more lone pairs on the central atom leads to asymmetry, and thus indicates a polar molecule**. You can see this in H_2O in Figure 11.3, where the lone pairs repel the bonding electron groups and cause this molecule to have a "bent" shape, with dipoles that do not cancel.

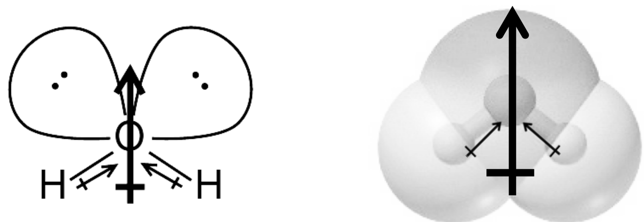


Figure 11.3: Examples of a polar molecule with dipoles that do not cancel.

Purpose

Students will learn skills and apply knowledge necessary to **use molecular models to show 3-D representations of molecules, determine the polarity of bonds within a molecule and determine the polarity of molecules.**

What you will do, briefly

A. Use molecular model kits to build the molecules you determined Lewis Structures for last week and use it to determine the number of electron groups around each central atom.
   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

B. Determine the polarity of each bond (ΔEN) using Tables 11.1 and 11.2 and redraw the Lewis structure to include dipole arrows for each polar bond.

C. Determine whether each molecule is polar by considering the bond polarity and the symmetry of the overall molecule.

Name: ______________________ Experiment: ____

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 another nonmetal?

3. Using the periodic table, predict the number of valence electrons in 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 generalizations (increase/decrease) can you make about the trend in electronegativity as you go down a group?

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

*Molecule (central atom in **bold**)*	*Number of valence electrons (use periodic table)*	*Lewis structure*	*Molecular shape around central atom*
CH_3Br			
C_2 H_4			
CN^-			

*Molecule (central atom in **bold**)*	*Structure with bond polarity shown*	*Is the molecule symmetrical?*	*Is the molecule polar?*
CH_3Br			
C_2 H_4			
CN^-			

Electronegativity:

H: 2.1

C: 2.5

N: 3.0

O: 3.5

F: 4.0

P: 2.1

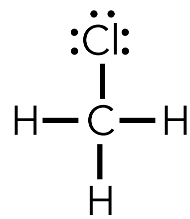
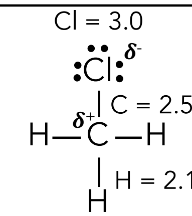
S: 2.5

Cl: 3.0

Br: 2.8

I: 2.5

Name: ______________________ Experiment: ____

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)	Structure with bond polarity shown	Is the molecule polar or nonpolar?
$\mathbf{C}H_3Cl$	C: 4 x 1 = 4 e^- H: 1 x 3 = 3 e^- <u>Cl: 7 x 1 = 7 e^-</u> Total: 14 e^-		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:				
$\mathbf{C}H_4$	C: H:				

Name: ____________________ Experiment: ______________________________________

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)	Structure with bond polarity shown	Is the molecule polar or nonpolar?
CH_3Br_2	C: H: Br:				
H**O**Br	H: O: Br:				
H_2**O_2**	H: O:		O: O:		
NH_3	N: H:				
N_2 H_4	N: H:		N: N:		
NH_2**C**H_3	N: H: C:		N: C:		
O_2	O:				

Name: ____________________ Experiment: ______________________________________

Molecule (Central atom is in **BOLD**)	Number of valence electrons (use periodic table)	Electron dot structure	Geometry around the central atom(s) (molecular shape)	Structure with bond polarity shown	Is the molecule polar or nonpolar?
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 electrons for the -1 charge)* C: N:				
C_2H_2	N: H:		C: C:		
$H_3\mathbf{CCN}$	H: C: N:		C: C:		
$\mathbf{Si}O_2$	Si: O:				
C_3H_8	C: H:		C: C: C:		

There is no post-laboratory assignment this week

Lab 12:

Molar Volume of a Gas

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 gas 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** (below) to correct our experimental data to those of STP conditions (Pressure = 1atm; Temperature = 0 °C.)

$$\frac{P_{lab}V_{lab}}{T_{1ab}} = \frac{P_{STP}V_{STP}}{T_{STP}}$$

Plugging STP values into the Ideal Gas Law (PV = nRT) as shown below, reveals an interesting truth: One mole of any ideal gas occupies exactly the same volume at STP conditions, 22.4 L. This is known as the Molar Volume at STP, and is a useful conversion factor for relating quantity (moles or grams) of a gas to volume of a gas.

For one mole of gas at STP Conditions:

(1 atm)(V) = (1 mol)(0.0821 atm*L/mol*K)(273 K)

V = 22.4 L/mol (Molar Volume)

Purpose

Students will learn skills and apply knowledge necessary to **experimentally measure the volume, pressure and temperature of a sample of CO_2 gas, use the combined gas law to convert CO_2 volume to STP conditions and use STP volume and the measured mass of CO_2 to confirm the Molar Volume of an ideal gas is 22.4 L/mol.**

Skills:

1. Accurately read a laboratory barometer and thermometer
2. Use weigh by difference and the density of air to calculate the mass of a gas in a stoppered flask.

Knowledge:

1. Use the combined gas law to convert between gas volume measured in laboratory conditions to gas volume at STP conditions
2. Dividing STP volume by the moles of gas measured yields the molar volume of a gas (22.4 L/mol)

What you will do, briefly:

A. Record the pressure (P_{lab}) and temperature (T_{lab}) of the laboratory, weigh a dry 125-mL flask with a rubber stopper. Place a small piece of dry ice (solid CO_2) into 125-mL flask to sublimate. Once the CO_2 has sublimated completely, tightly stopper the flask and reweigh, now filled with CO_2 gas. Then determine the volume of the CO_2 gas (V_{lab}) using water, and then use the combined gas law and known values of P_{STP} and T_{STP} to convert V_{lab} to V_{STP}.

B. Determine the mass of CO_2 gas by difference and convert to moles of CO_2.

C. Confirm molar volume of a gas at STP by dividing V_{STP} by moles of CO_2 gas and calculate the percent error from known STP molar volume (22.4 L/mol)

Criteria for Success:

1. Correct significant figures and units must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware used.
2. The calculations leading to determining the molar volume were carried out correctly.
3. One or more specific errors are identified which feasibly explain significant variation from the expected molar volume.

Name: ____________________ Experiment: ______________________________________

Pre-laboratory Assignment:

1. The density of air is given as 0.00118 g/mL, and the total volume of a sample 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 = liters / moles of carbon dioxide)

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

Before we begin:

Today's lab will determine the number of ______________ of the CO_2 (g)

and the ______________ at STP. Together we can observe the molar volume at STP which is

expected to be: ______________ and can be calculated by: ______________________________

Note: At STP, the pressure is: ______________ and the temperature is: ______________________

Note: When calculating gas laws your temperature should be calculated in: __________________

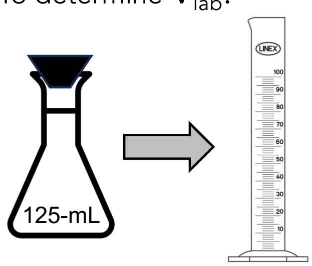
A Calculating the volume at STP

Obtain the laboratory ______________ and ____________ and covert units to match STP units.

$$\frac{P_{lab}V_{lab}}{T_{lab}} = \frac{P_{STP}V_{STP}}{T_{STP}} \quad \therefore \quad V_{STP} = \frac{P_{lab}V_{lab}T_{STP}}{T_{lab}P_{STP}}$$

To solve for volume at STP, we need to determine the volume at laboratory conditions

To determine V_{lab}:



1. ____
2. ____
3. ____
4. Covert from mL to L

Note: After determining the volume, you must dry your flask and stopper COMPLETELY.

B Calculating the moles of CO_2

1. Determine the mass of the flask, rubber stopper and the air trapped inside

$$\text{Volume} \times \frac{\text{Mass}}{\text{Volume}} = \text{Mass of air}$$

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

CRC Handbook

2. Using the volume of the flask found in part A and the density of air at the laboratory temperature, determine the mass of air in the flask

Additional Notes:

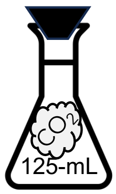
Safety Considerations:

Image Guide and Procedure

Name: ____________________ Experiment: ______________________________________

3. Calculate the mass of the flask and stopper with no air by: ______________________________

4. Collect a piece of dry ice about __________ size and put it in dry 125 mL flask with stopper OFF



5. Once the sublimation is complete, stopper the flask and get the mass of the flask, rubber stopper and CO_2

6. The mass of CO_2 can be determined by: ______________________________

____

7. The moles of CO_2 can be determined by: ______________________________________

8. The molar volume of carbon dioxide at STP (L/mol) can be calculated by:

____

Note: This is considered the "actual" value. Found in experiment, this is expected to contain error.

9. Calculate the percent error using your determined "actual" value and the theoretical value of 22.4 L/mol

$$\text{Percent error (\%)} = \frac{\text{actual - theoretical}}{\text{theoretical}} \text{ X } 100$$

Additional Notes:

Safety Considerations:

Name: ____________________ Experiment: ______________________________________

Data / Observations

A Calculating the volume at STP

	Trial 1	Trial 2	Trial 3
1. Laboratory barometric pressure (mmHg or inHg):	________	________	________
2. Laboratory barometric pressure (atm):	________	________	________
3. Laboratory temperature (°C):	________	________	________
4. Laboratory temperature (K):	________	________	________
5. Volume of flask from water (mL):	________	________	________
6. Corrected volume to STP (mL):	________	________	________
7. Corrected volume to STP (L):	________	________	________

B Calculating the moles of CO_2

1. Mass of flask, stopper and air (g):	________	________	________
2. Density of air at lab temperature (g/ml):	________	________	________
3. Mass of air in flask (g):	________	________	________
4. Mass of flask and stopper without air (g):	________	________	________
5. Mass of the flask, stopper and CO_2 (g):	________	________	________
6. Mass of the CO2 (g):	________	________	________
7. Moles of CO2 (mol):	________	________	________

C Calculating the molar volume of CO_2

1. Molar volume of CO2 at STP (L/mol):	________	________	________
2. Average molar volume (L/mol)		________	
3. Percent error (%):		________	

Name: ______________________ Experiment: ____

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 CO2? 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?

Lab 13:

Solubility

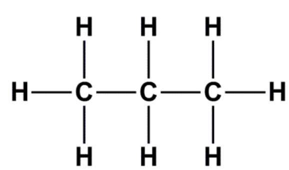
Introduction

As 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.

Atom	H	C	N	O	F	Cl	Br	I
EN value	2.1	2.5	3.0	3.5	4.0	3.0	2.8	2.5

As we subtract electronegativity values from a carbon to a hydrogen bond, for example, 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.

Each bond has
$\Delta EN \leq 0.4$
$\therefore$ each bond is nonpolar
$\therefore$ the molecule is nonpolar



electron density is shared equally throughout the molecule

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** ($+\!\!\!\longrightarrow$). 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 also 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.

When a solute forms a homogenous mixture with a solute, they are referred to as **soluble**. If the substance is unable to dissolve fully in the solvent, we call it **insoluble.** Similarly, if two substances are able to interact and form a homogeneous solution in any proportion, they are referred to as **miscible.** If two 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,** which is observed as a separation into layers.

Purpose

Students will learn skills and apply knowledge necessary to **understand solubility limits of solutions, relate the “like dissolves like” rule to polarity and discern between miscible and immiscible mixtures, and determine the concentration of sodium chloride in a saltwater solution.**

Skills:

1. Properly dispose of chemical waste
2. Identify if a mixture is soluble/insoluble or miscible/immiscible
3. Use volumetric glassware for accurate volume measurements
4. Boil off a solvent and use weigh by difference to determine solvent mass

Knowledge:

1. Differentiate between saturated, unsaturated and supersaturated solutions
2. Relate the solubility or miscibility of a mixture to the intermolecular forces of the two substances mixed
3. Mass percent can be calculated by: mass of solute / mass of solution
4. Molarity can be calculated by: mols of solute / L of solution

What you will do, briefly:

A. Dissolve a test tube of supersaturated solution in a boiling water bath, then once cooled, add a seed crystal to the test tube and observe.

B. Combine one solute and one solvent into six total test tubes to determine whether two solutes ($KMnO_4$ and I_2) are soluble or insoluble when paired with three different solvents (DI H_2O, hexane and methanol). Then in three test tubes, combine three different liquids (acetone, ethanol and hexane) with water to determine the miscibility of the resulting solutions.

C. Calculate the concentration of sodium chloride (in mass percent and molarity) of a provided saltwater solution by obtaining the mass of a volumetrically-measured saltwater sample, boiling off the water, measuring the mass of the resulting dried salt and employing weigh by difference.

Criteria for Success:

1. Correct significant figures and units must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware used.

2. Your observations and conclusions for the solubility and miscibility experiments coincide with the "like dissolves like" rule.

3. The calculations leading to determining the mass percent and molarity were carried out correctly.

4. One or more specific errors are identified which feasibly explain significant variation from the expected saltwater concentration.

Name: ______________________ Experiment: ____

Pre-laboratory Assignment:

1. Answer the following questions about the following molecules:

$CH_3CH_2CH_2CH_2CH_2CH_3$ and Br_2

   a. What is the strongest intermolecular force present in $CH_3CH_2CH_2CH_2CH_2CH_3$?

   b. Is $CH_3CH_2CH_2CH_2CH_2CH_3$ polar or nonpolar? (EN: C = 2.5; H = 2.1)

   c. What is the strongest intermolecular force present in Br_2?

   d. Is Br_2 polar or nonpolar? (EN: Br = 2.8)

   e. Are $CH_3CH_2CH_2CH_2CH_2CH_3$ and Br_2 miscible? Why or why not?

2. Answer the following questions about the following molecules:

CH_3OH and H_2O

   a. What is the strongest intermolecular force present in CH_3OH?

   b. Is CH_3OH polar or nonpolar? (EN: C = 2.5; H = 2.1; O = 3.5)

   c. What is the strongest intermolecular force present in H_2O?

   d. Is H_2O polar or nonpolar? (EN: H = 2.1; O = 3.5)

   e. Are CH_3OH and H_2O miscible? Why or why not?

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

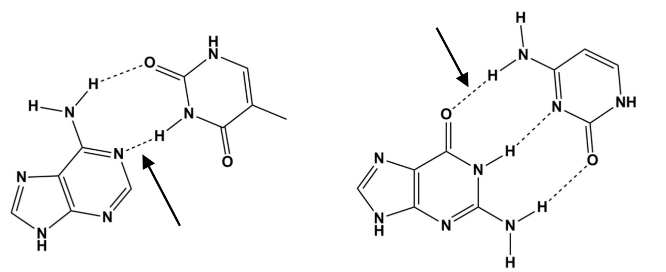


Image Guide and Procedure

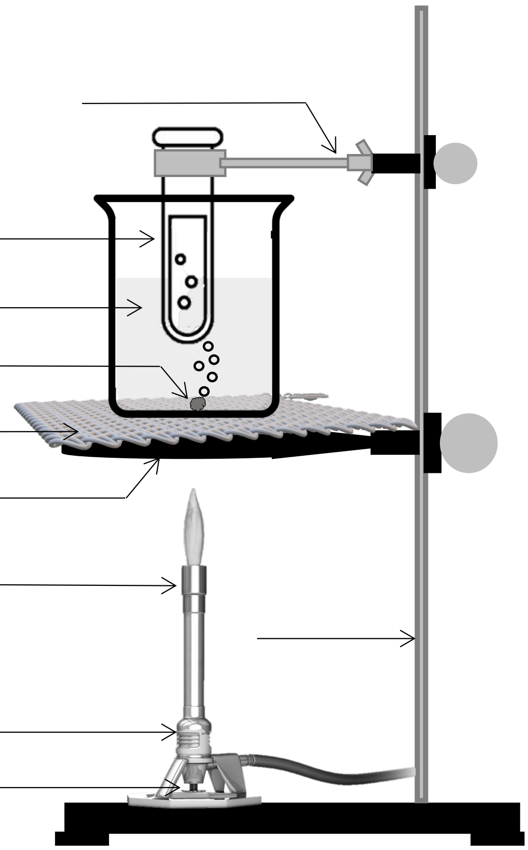
Name: ____________________ Experiment: ______________________________________

A Supersaturated solution

1. Bring __________ mL of H2O to a boil
2. Suspend supersaturated solution in water bath until: ________________

   ____
3. Allow solution to cool and then:

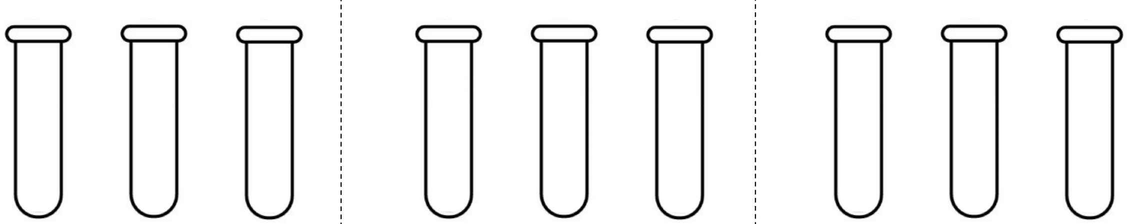
   ____
4. Record your observations



B Solubility and Miscibility

To each test tube, mix the two listed chemicals and determine if the solution is soluble or insoluble

Mix a pea size amount of $KMnO_4(s)$ Which is: **Polar / Non Polar**	Mix a pea size amount of $I_2(s)$ Which is: **Polar / Non Polar**	Mix 20 drops of $H_2O(l)$ Which is: **Polar / Non Polar**



Add to soln:									
	20 drops DI H_2O	20 drops Hexane	20 drops Methanol	20 drops DI H_2O	20 drops Hexane	20 drops Methanol	20 drops Acetone	20 drops Ethanol	20 drops Hexane

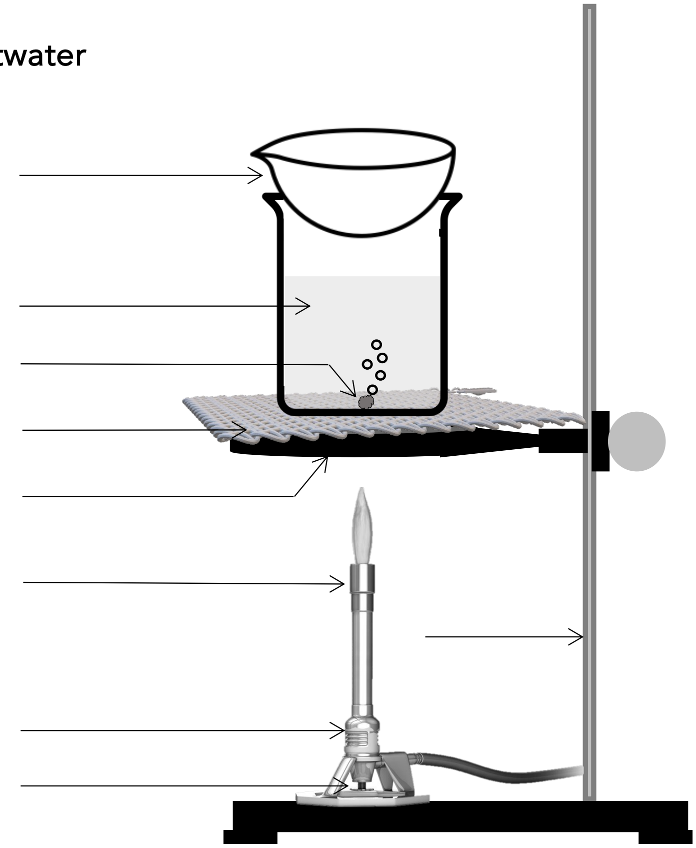
Safety Considerations: ____

____

Image Guide and Procedure

Name: ______________________ Experiment: ____

C Concentration analysis of Saltwater



1. Determine the mass of a clean, dry evaporating dish.
2. Collect__________of unknown saltwater solution in beaker.
3. Condition 10.00 mL volumetric pipette as shown by instructor.
4. Pipette exactly 10.00 mL of solution into evaporating dish.
5. Determine the mass of evaporating dish with 10.00 mL of saltwater.
6. The mass of the saltwater can be calculated by:____
7. Place the evaporating dish on the boiling water bath and evaporate the solvent in the saltwater solution. Do not let the boiling water bath run dry.
8. Once cool, determine the mass of the evaporating dish and salt. This can be done:____
9. The mass percent of the salt in the saltwater solution can be calculated:____
10. Calculate moles of NaCl:____________________________________
11. Convert the volume of solution to L and calculate the molarity of the solution:____

Additional Notes:

Safety Considerations:

Name: ______________________ Experiment: ____

Data / Observations

A Supersaturated solution

Observations from supersaturated solution:______________________________________

____

____

B Solubility and Miscibility

Circle whether the solution is soluble or insoluble; miscible or immiscible

Mix a pea size amount of $KMnO_4(s)$

With:

20 drops DI water H_2O	20 drops Hexane C_6H_{14}	20 drops Methanol CH_3OH
Soluble	**Soluble**	**Soluble**
Insoluble	**Insoluble**	**Insoluble**

Mix a pea size amount of $I_2(s)$

With:

20 drops DI water H_2O	20 drops Hexane C_6H_{14}	20 drops Methanol CH_3OH
Soluble	**Soluble**	**Soluble**
Insoluble	**Insoluble**	**Insoluble**

Mix 20 drops of $H_2O(l)$

With:

20 drops Acetone	20 drops Ethanol	20 drops Hexane
Miscible	**Miscible**	**Miscible**
Immiscible	**Immiscible**	**Immiscible**

Name: ____________________ Experiment: ______________________________________

Data / Observations

C Concentration analysis of Saltwater

	Trial 1	Trial 2
1. Unknown saltwater solution number:	________	________
2. Mass of evaporating dish (g):	________	________
3. Volume dispensed from volumetric pipette (mL):	________	________
4. Mass of evaporating dish and saltwater (g):	________	________
5. Mass of saltwater solution (g):	________	________
6. Mass of evaporating dish and dried salt (g):	________	________
7. Mass of remaining salt NaCl (g):	________	________
8. Mass percent of salt NaCl in solution (%):	________	________
9. Moles of salt NaCl (moles):	________	________
10. Molarity of saltwater solution (M):	________	________

There is no post-laboratory assignment this week

Lab 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 is known as a **neutralization reaction:**

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

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 can determine the concentration of a solution with 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 volume of a solution with 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.

Determined from known concentration of analyte

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

Determined from endpoint of titration

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.

In this lab, we will first use HCl as the titrant and NaOH as the analyte to determine the concentration of the NaOH solution to 4 significant figures. Then we will use the standardized NaOH solution as our titrant and a solution of acetic acid (vinegar) as our analyte and determine the concentration of the acetic acid solution.

Purpose

Students will learn skills and apply knowledge necessary to **perform an acid-base titration and accurately determine the molarity of two solutions of unknown concentration to 4 significant figures.**

Skills:

1. Use volumetric glassware for accurate volume measurements
2. Conditioning a buret with titrant
3. Accurate use of a buret for finding ΔV
4. Determination of a neutralization endpoint using a pH indicator

Knowledge:

1. Determine the balanced chemical equation for an acid and base neutralization
2. Determine moles of titrant using molarity and volume titrated
3. Determine moles of analyte using the balanced chemical equation
4. Molarity: mols of solute / L of solution

What you will do, briefly:

A. Set up a titration apparatus using a conditioned buret with a known concentration of HCl as the titrant and a solution of NaOH as the analyte and titrate to the endpoint, as indicated by a color change from the pH indicator. Use ΔV, the concentration of HCl and the volume of NaOH which was titrated to determine the molarity of the NaOH to 4 significant figures.

B. Utilize the titration apparatus using a conditioned buret with the standardized NaOH from part A as the titrant and a solution of acetic acid as the analyte and titrate to the endpoint, as indicated by a color change from the pH indicator. Use ΔV, the concentration of standardized NaOH and the volume of acetic acid which was titrated to determine the molarity of the acetic acid to 4 significant figures.

Criteria for Success:

1. Correct significant figures and units must be used when reporting each measurement, which reflects the uncertainty of the specific instrument or glassware used.

2. The calculations leading to determining the molarity of NaOH and acetic acid were carried out correctly and the final concentrations have 4 significant figures.

3. One or more specific errors are identified which feasibly explain significant variation from the expected concentrations of NaOH and acetic acid.

Name: ______________________ Experiment: ____

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.

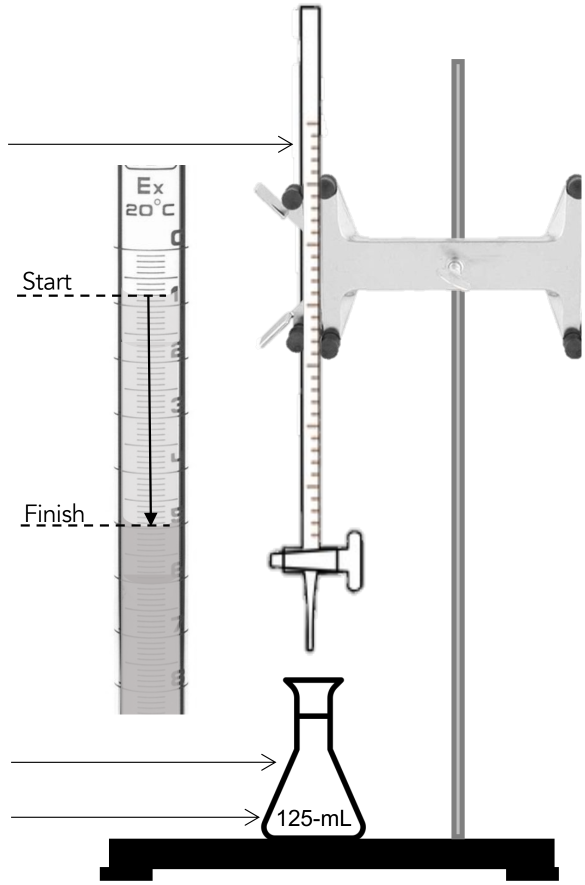
3. How many moles of NaCl are in 1.85 L of a 0.458 M NaCl solution?

4. How many liters are required to obtain 0.182 moles of NaNO3 from a 0.824 M NaNO3 solution?

Image Guide and Procedure

Name: ____________________ Experiment: ____

A Standardization of a NaOH solution



1. Clean and condition with HCl (aq) a 25.00 mL volumetric pipette as instructed.
2. Obtain________mL of HCl solution and record concentration:________________________
3. Pipette 25.00 mL of HCl into a 125 mL Erlenmeyer flask and remember to add:

   ____
4. Clean and condition the burette as instructed and fill it with:____________and record initial volume to an uncertainty of:____________
5. Dispense the titrant (______) to the flask waiting for the endpoint:______________

Note: As you begin to see pink, slow the additions until the lightest pink persists. Record the final volume.

6. The total titrant (________) added can be calculated: ________

   ________________________________ Then convert this value to L.
7. The moles of HCl (aq) added to the flask can be calculated:

   ____
8. The moles of NaOH consumed can be determined by the following reaction:______________________________________
9. The molarity can be calculated:______________________________

   Do the reaction twice and take an average.

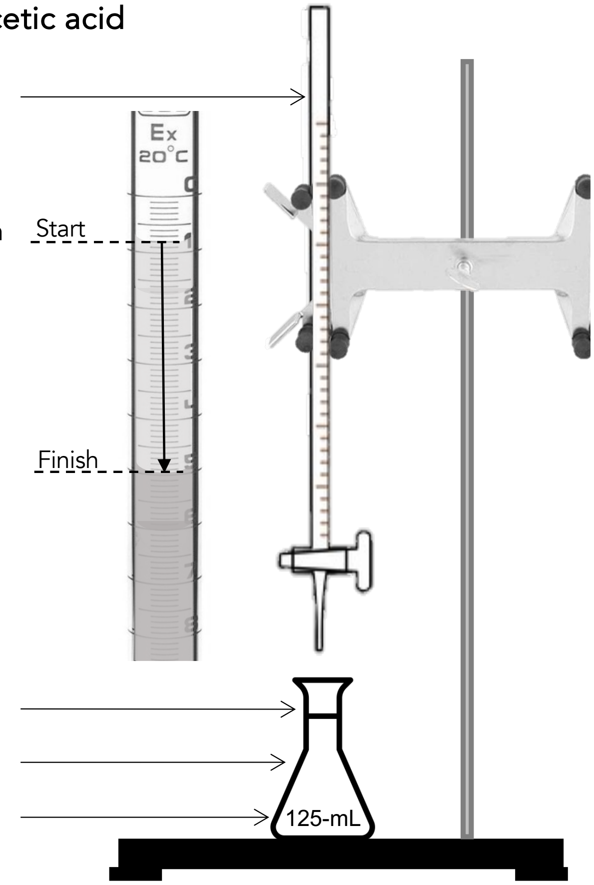
Additional Notes:

Safety Considerations:

Image Guide and Procedure

Name: ____________________ Experiment: ____

B Determination of concentration of acetic acid



1. Clean and condition with CH_3COOH (aq) a 5.00 mL volumetric pipette as instructed.
2. Obtain ________ mL of CH_3COOH (aq) solution
3. Pipette 5.00 mL of CH_3COOH (aq) into a 125 mL Erlenmeyer flask, add: ________________

   and remember to add: ________________
4. Prepare the the burette and fill it with: ________

   of a known concentration of: ________________

   Record the initial volume of titrant.
5. Dispense the titrant (________) to the flask waiting for the endpoint: ________________

Note: Swirl as you titrate, slow as pink starts to appear, stop when a faint pink persists.

Additional Notes:

6. The total titrant (________) added can be calculated: ________

   ________________________________ Then convert this value to L.
7. The moles of NaOH (aq) added to the flask can be calculated:

   ____
8. The moles of CH_3COOH consumed can be determined by the reaction: ________________________________
9. The molarity can be calculated: ________________________

   Do the reaction twice and take an average.

Safety Considerations:

Name: ____________________ Experiment: ______________________________________

Data / Observations

A Standardization of a NaOH solution

	Trial 1	*Trial 2*
1. Molarity of HCl (aq) solution (M):	________	
2. Initial burette reading (mL):	________	________
3. Final burette reading (mL):	________	________
4. Volume dispensed from burette (mL):	________	________
5. Volume dispensed converted to liters (L):	________	________
6. Moles of hydrochloric acid (mol):	________	________
7. Moles of sodium hydroxide (mol):	________	________
8. Molarity of NaOH (aq) solution (M):	________	________
9. Average molarity of NaOH solution (M):	________	

Name: ____________________ Experiment: ______________________________________

Data / Observations

B Determination of concentration of acetic acid

	Trial 1	*Trial 2*
1. Molarity of NaOH (aq) solution from part A (M):	____________	
2. Initial burette reading (mL):	________	________
3. Final burette reading (mL):	________	________
4. Volume dispensed from burette (mL):	________	________
5. Volume dispensed converted to liters (L):	________	________
6. Moles of sodium hydroxide (mol):	________	________
7. Moles of acetic acid (mol):	________	________
8. Molarity of CH_3COOH (aq) solution (M):	________	________
9. Average molarity of CH_3COOH solution (M):	____________	

There is no post-laboratory assignment this week

Appendices

Unit Conversion Guide

$$\text{Given unit} * \left(\frac{\text{Desired unit}}{\text{Given unit}}\right) = \text{Desired unit}$$

If a number is in the numerator, multiply; if a number is in the denominator, divide.

Physical Quantity	Unit Name	Unit Abbr.
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol

English Unit	SI Unit
1 mile	1.609 km
1.094 yds	1 meter
3.281 ft	1 meter
1 inch	2.54 cm
35.31 ft^3	1 m^3
1 gallon	3.785 dm^3
1.057 quarts	1 dm^3
1 fluid oz	29.57 cm^3
2.205 lbs	1 kg

Common "Pass Through" Conversion Factors			
Density	**Molar Mass**	**Molarity**	**Cubic Volume**
$d = \frac{\text{mass (g)}}{\text{volume (mL)}}$	$M.M. = \frac{\text{mass* (g)}}{\text{mole (mol)}}$ **Mass comes from sum of atomic masses*	$M = \frac{\text{moles (mol)}}{\text{volume (L)}}$	1 cm^3 = 1 mL 1 dm^3 = 1 L 1 m^3 = 1000 L
Mass ↔ volume	**Mass ↔ moles**	**Volume ↔ moles**	**Cubic length ↔ volume**

Prefix	Symbol	Conventional Notation	Exponential Notation	Example
tera	(T)	1,000,000,000,000	1×10^{12}	1 teragram (Tg) = 1×10^{12} g
giga	(G)	1,000,000,000	1×10^{9}	1 gigagram (Gg) = 1×10^{9} g
mega	(M)	1,000,000	1×10^{6}	1 megagram (Mg) = 1×10^{6} g
kilo	(k)	1,000	1×10^{3}	1 kilogram (kg) = 1×10^{3} g
hecto	(h)	100	1×10^{2}	1 hectogram (hg) = 1×10^{2} g
deka	(da)	10	1×10^{1}	1 dekagram (dag) = 1×10^{1} g
	____	1	1×10^{0}	
deci	(d)	0.1	1×10^{-1}	1 decimeter (dm) = 1×10^{-1} m
centi	(c)	0.01	1×10^{-2}	1 centimeter (cm) = 1×10^{-2} m
milli	(m)	0.001	1×10^{-3}	1 millimeter (mm) = 1×10^{-3} m
micro	(μ)	0.000001	1×10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano	(n)	0.000000001	1×10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico	(p)	0.000000000001	1×10^{-12}	1 picometer (pm) = 1×10^{-12} m
femto	(f)	0.000000000000001	1×10^{-15}	1 femtometer (fm) = 1×10^{-15} m

Nomenclature Guide

To begin nomenclature analysis, start with the question:

Is there a metal?

Yes – *Ionic species*

No – *Molecular species**

Note: Greek prefixes are **only** used when no metals or polyatomic ions are present

Do we know it's charge?

Yes: Name the metal then use the anion name with the suffix replaced with –ide i.e. sodium fluoride (NaF)

No: Name the metal, place parenthesis with the metal's charge in roman numerals, then use the anion name with the suffix replaced with –ide i.e. iron (II) chloride ($FeCl_2$)

No metal: Name using Greek prefixes to indicate the number of each element present. The only exception is when the first element has one, the mono- can be dropped. The last element uses the suffix –ide. i.e. dinitrogen pentoxide (N_2O_5) carbon monoxide (CO)

Cations we know:
Alkali metals (group 1) = +1
Alkaline earth metals (group 2) = +2
Aluminum (group 13) = +3
Ag is always +1; Cd, Zn are always +2

Anions we know:
Halogens (group 17) = -1
Oxygen and sulfur (group 16)= -2
Nitrogen and phosphorous (group 15)= -3

Roman Numerals	
1	I
2	II
3	III
4	IV
5	V
6	VI
7	VII

Greek Prefixes

Mono-	1	Hexa-	6
Di-	2	Hepta-	7
Tri-	3	Octa-	8
Tetra-	4	Nona-	9
Penta-	5	Deca-	10

Ionic species always pair up to balance the ionic charges, so they sum to zero.

For example:

Na^+ Cl^-	Ca^{2+} Cl^-	Al^{3+} O^{2-}
x1 x1	x1 x2	x2 x3
$\underbrace{NaCl}_{+1\ -1}$	$\underbrace{CaCl_2}_{+2\ -2}$	$\underbrace{Al_2O_3}_{+6\ -6}$

The following species form diatomic molecules in nature: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

The presence of H in the front of a species followed by (aq) indicates an ionic species with H^+: this is an **acid**. Name acids by considering the anion paired with H^+

If the anion has oxygen:
Name the anion and replace the suffix
If the anion suffix was –ate, change it to –ic
If the anion suffix was –ite, change it to –ous
Then add the word acid
i.e. HNO_3 (aq) NO_3 = nitrate, so **nitric acid**

If the anion does not have oxygen:
Start with the prefix hydro-
Then base of the anion with the suffix -ic
Then add the word acid
i.e. HCl (aq) hydrochloric acid
HCN (aq) hydrocyanic acid

*A species may be ionic without metal with the presence of a polyatomic cation. More on the next page

Nomenclature Guide

Some molecules gain or lose electrons to form polyatomic ions. While no metals may be present, these species behave like ions. When naming, just name the polyatomic species without changing any prefix or suffix

Here is a list of common polyatomic ions:

Polyatomic ion	Formula	Polyatomic ion	Formula
Ammonium ion	NH_4^{1+}	Acetate ion	$C_2H_3O_2^{1-}$
Hydronium ion	H_3O^{1+}	Permanganate ion	MnO_4^{1-}
Hypochlorite ion	ClO^{1-}	Hydroxide ion	OH^{1-}
Chlorite ion	ClO_2^{1-}	Nitrite ion	NO_2^{1-}
Chlorate ion	ClO_3^{1-}	Nitrate ion	NO_3^{1-}
Perchlorate ion	ClO_4^{1-}	Chromate ion	CrO_4^{2-}
Carbonate ion	CO_3^{2-}	Dichromate ion	$Cr_2O_7^{2-}$
Hydrogen carbonate or bicarbonate ion	HCO_3^{1-}	Hydrogen sulfite or bisulfite ion	HSO_3^{1-}
Hydrogen phosphate or biphosphate ion	HPO_4^{2-}	Hydrogen sulfate or bisulfate ion	HSO_4^{1-}
Phosphate ion	PO_4^{3-}	Sulfite ion	SO_3^{2-}
Cyanide ion	CN^{1-}	Sulfate ion	SO_4^{2-}

Increasing oxygen

A trend in oxo acids:

Prefix	Suffix	Example:
-hypo	-ite	hypochlorite (ClO^{1-})
N/A	-ite	chlorite (ClO_2^{1-})
N/A	-ate	chlorate (ClO_3^{1-})
Per-	-ite	perchlorate (ClO_4^{1-})

Note: the number of oxygen changes but the charge does not.

Anion \ Cation	NH_4^+	Li^+	Na^+	K^+	Mg^{2+}	Ca^{2+}	Ba^{2+}	Fe^{2+}	Fe^{3+}	Co^{2+}	Ni^{2+}	Cu^{2+}	Ag^+	Zn^{2+}	Pb^{2+}	Al^{3+}
F^-	Sol (aq)	Slight Sol	Sol (aq)	Sol (aq)	Insol (s)	Insol (s)	Slight Sol	Slight Sol	Slight Sol	Sol (aq)	Slight Sol	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Slight Sol
Cl^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Sol (aq)	Insol (s)	Sol (aq)
Br^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Sol (aq)	Slight Sol	Sol (aq)
I^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	--	Sol (aq)	Sol (aq)	--	Insol (s)	--	Insol (s)	Sol (aq)
ClO_3^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	--	Sol (aq)	Sol (aq)	Sol (aq)	--	Sol (aq)	--	Sol (aq)	Sol (aq)
OH^-	--	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Slight Sol	Sol (aq)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Slight Sol	Insol (s)	Insol (s)	Insol (s)
SO_3^{2-}	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Insol (s)	--	--	Insol (s)	Insol (s)	--	Insol (s)	Insol (s)	Insol (s)	--
SO_4^{2-}	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Slight Sol	Insol (s)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Slight Sol	Sol (aq)	Insol (s)	Sol (aq)
CO_3^{2-}	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	--	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	--
NO_2^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	--	--	--	Sol (aq)	--	Insol (s)	--	Sol (aq)	--
NO_3^-	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)	Sol (aq)
PO_4^{3-}	Sol (aq)	Insol (s)	Sol (aq)	Sol (aq)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)	Insol (s)

Legend for solubility table

Sol (aq)	Resultant ionic combination will be soluble, an aqueous solution	Insol (s)	Resultant ionic combination will be insoluble, a solid precipitate will form	Slight Sol	Resultant ionic combination will be slightly soluble, formation of precipitate depends on concentration and heat

Precipitate: In solution chemistry, a precipitate forms when aqueous ions combine to form an insoluble, solid substance. A precipitate may be observed as solid material collecting in the reaction vessel or simply as cloudiness forming, obscuring light and visibility through the reaction vessel.

Solubility	Volume of water as mL/g of substance
Very soluble	less than 1
Easily soluble	from 1 to 10
Soluble	from 10 to 30
Sparingly soluble	from 30 to 100
Slightly soluble	from 100 to 1,000
Very slightly soluble	from 1,000 to 10,000
Practically insoluble	more than 10,000

Source: Sigmaaldrich.com; USP.org

Standard levels of solubility used for analysis. These definitions were established by United States Pharmacopeia (USP)

	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									

H	SYMBOL
1	ATOMIC NUMBER
1.008	ATOMIC WEIGHT
Hydrogen	NAME