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GENERAL CHEMISTRY LABORATORY SAFETY RULES

Know these rules for any time that you are in a laboratory. Memorize them for the quiz to be given the first week of class.

Always wear the appropriate personal protective equipment (PPE). KEEP YOUR EYES PROTECTED BY WEARING SAFETY GOGGLES WHENEVER ANYONE IS DOING AN EXPERIMENT IN THE LABORATORY. If you wear contact lenses, consider leaving them at home and wearing your glasses when you have an experiment to perform. Keep your feet and skin protected by wearing impermeable close-toed shoes and sensible clothes. Consider wearing a lab apron or lab coat. Keep your hair protected by tying it back and keeping it from touching anything in the lab. All of these are PPE.

1. Never taste any chemical in the laboratory ever. This includes chemicals for the experiments, food, liquids, tobacco, anything. Consider this: student equipment is probably contaminated and unsafe to use for anything to be ingested! Your hands will also become contaminated by working in the laboratory, so you don't want to use your hands for food or beverages while in the laboratory. Never put anything in your mouth in lab, including your pen or pencil, your fingers, food, beverages, tobacco, etc.

2. Use caution when trying to identify the odor of any chemical. Waft the vapors toward your nose with a fanning motion of your hand and cautiously inhale. Try to avoid breathing fumes of any kind. If you (or another student) are doing any lab work that is creating an odor, move into the fume hood.

3. READ ALL LABELS CAREFULLY BEFORE USING CHEMICALS. Read the name and formula to make sure that both are correct for the experiment, and read the concentration of solutions to confirm the correct one also.

4. NEVER RETURN UNUSED CHEMICALS TO THE REAGENT BOTTLES, WHETHER SOLID OR LIQUID. Discard the chemical properly or give the excess to another student who needs it.

5. Dispose of all excess chemicals using directions given by the lab manual and/or the instructor. If no special directions are given: place all solids (including matches, litmus paper, etc.) in waste cans, and pour liquids down the sink while flushing with large amounts of water. You, the student, will neutralize acids and bases to a pH range of 3 - 11 prior to disposal down the drain. There are special containers in the fume hoods for disposal of certain chemicals. Use these only as directed by the instructor.

6. If you spill chemicals on your skin or clothing, wash immediately with large amounts of water. You should help one another when this happens, even to

the point of getting a student to the eye wash or the safety shower and helping with clean-up. NOTIFY THE INSTRUCTOR AS SOON AS POSSIBLE.

You should always wash your hands at the end of every lab period and whenever you leave the lab room for any reason.

7. Report any spills on the counters or floor to the instructor to receive instructions on clean-up. You are responsible for any and all clean-ups, including those in the balance room or in the fume hoods. Clean up your work station at the end of each lab period. Throw away trash and wipe the counter with wet paper towels.

8. Never heat liquids in a test tube over a burner. Heat water in a beaker and place the test tube in the hot water. This is called a hot water bath.

9. IN CASE OF FIRE: notify the instructor immediately, but do not wait to extinguish the fire. Cover the container if possible, or use one of the two fire extinguishers in the room. Learn how to find and use the extinguishers and the fire blanket.

10. In case of earthquake, turn off all burners and all supplies of natural gas. Take equipment off any ring stands and wait for instructions. Do not run outside the building. The college has assigned gathering areas for each building, which your instructor will tell you about during class.

11. If you need to make any glassware, be sure to fire-polish all glass tubing and stirring rods and any sharp edges. If you need to insert tubing or glass thermometers into rubber stoppers, lubricate the stopper with glycerol first.

12. If you have long hair, always wear it tied up or back for your time in the laboratory. Even if you are not using chemicals, other students have been in the laboratory and the bench tops and other areas are probably contaminated with chemicals. Don’t let the ends of your hair brush the countertops.

13. Everyone should be wearing close-toed shoes of leather or plastic, i.e., some impermeable substance. This keeps spills from contaminating your skin and also protects you from broken glassware, etc.

14. Do not wear shorts unless you have a very long apron to protect the skin of your legs. Have or bring a long-sleeved shirt or lab coat to protect your arms and clothing. Your clothing should be nonsynthetic, like cotton or wool.

15. Gloves are required for certain experiments. You must provide your own gloves, just like you provide your own safety goggles and lab coat. If you don’t have your personal protective equipment (PPE), you cannot perform the

experiment and you will be counted as "absent" during the lab period. This is done to emphasize to you the importance of PPE.

16. Do NOT perform unauthorized experiments. This includes deviations from the procedure as written. Never work alone in a laboratory.

17. Learn the location of the safety equipment. Find the eyewash, shower, fire extinguishers, fire blanket, first aid kit, and telephone. (At our college, you can call 911 directly by just hitting the 911 keys.)

18. Never pour water into acid, always pour acid into water.

How to handle an accident:

First, do not panic. Always notify the instructor as soon as possible, but still deal with the accident itself right away.

In case of serious injury: do take care of the student while contacting 911 immediately.

Chemical spills: if only a small portion of your skin has been in contact with chemicals, walk to the large sink and turn on the cold water. Rinse the area for at least five minutes. If a large area of your skin has been in contact with strong acid or strong base, walk to the safety shower and pull the lever. Send another student for the instructor. Don't worry about your clothing – there are lab coats that can be borrowed by you to replace your clothing temporarily.

Chemicals in the eye: this should never happen since you are wearing safety goggles, but if it does, have another student walk you to the eye wash station and help hold your head in the water stream for at least five minutes. Have another student get the instructor.

Minor burns: if you pick up a piece of hot equipment and have a minor burn on your hand, walk to the large sink and turn on the cold water. Run cold water over the burn for at least five minutes and send another student for the instructor. If necessary, a plastic beaker of ice and water can be set up to hold your hand in while waiting for transportation to a medical facility.

Minor cuts: flush the cut with water and send for the instructor. If small, you may be allowed to treat it yourself. If not, you may be sent to the college nurse or to a medical facility.

Cleanup of a chemical spill:

If the substance is volatile, flammable or toxic, warn other students and send for the instructor immediately. If you spill an acid in the fume hood or at your work station, get the neutralizer and spray excess quantities on the acid. Or if it's a large spill, get the bottle of $NaHCO_3$ and sprinkle it on the spill until gas no longer evolves. Then wipe up the liquid with large pieces of paper towel.

GENERAL ANALYTICAL TECHNIQUES

You will learn about the two general categories of chemical analysis during your chemistry laboratories: qualitative analysis and quantitative analysis. Qualitative analysis identifies what is present. Quantitative analysis deals with numbers, i.e., how much of something is present.

There are three general areas of quantitative analysis: gravimetric, volumetric and spectrophotometric. You will be learning to work in all three areas during your year of general college chemistry lab.

Gravimetric analysis involves the masses of substances. For example, determining the percent composition of elements in a compound is gravimetric analysis. Separating components of a mixture and identifying the quantity of each by mass is gravimetric analysis.

Volumetric analysis involves measuring the volumes of liquid solutions. You will usually work with a concentration measurement as well, typically Molarity, which is moles of substance per liter of solution. You will work with some other concentration measurements also, such as molality, moles of substance per kilogram of solvent.

Spectrophotometric analysis involves determining the transmittance and absorbance of specific wavelengths of light through a solution containing a certain concentration of a transition metal compound. The absorbance and concentration are related through an equation called Beer's Law.

Quantitative analysis has to be done as accurately and precisely as possible. You will be asked to do repetitive determinations and statistical analysis to determine their precision and accuracy.

LABORATORY TECHNIQUES

1. **Bunsen burner or other lab burner:** Look for the natural gas supply and connect the rubber hose to it. The burner should also have a gas valve, which you may have to open. Then turn the air flow adjustment for about half of the maximum airflow. Turn on the natural gas until you either hear (slightly) or feel with your hand that gas is flowing from the top of the burner. Light the burner immediately with a striker or a match. A hotter flame can be obtained by opening up the airflow. To turn off the burner, just shut off the gas supply.

2. **Electronic balances**: Solano College has simple but expensive three-decimal place top-loading balances and also the more expensive four-decimal place balances. Use caution! You should never place any chemical directly on the pan, but use containers, weigh boats or paper, and immediately clean up any spills using the brushes available. Most of the time, the balances will already be turned on, but if not, look for the on/off button and turn the balance on. Give it a few minutes to warm up. There will be a tare or zero button or command you should use prior to any mass measurements. When you place something on the balance to obtain the mass, give a few seconds or more for the balance to settle, then record your reading exactly as shown on the balance, including all zeroes. Also, return to exactly the same balance during any experimental procedure to minimize instrument error readings.

3. **Obtaining stock reagents:**
    1. **Solids:** Be sure to read the label first, so that you obtain the correct chemical! Never take the reagent bottle to your workstation; instead bring some container that you have labeled to put the reagent in with you to the cart or fume hood where the reagent has been set out for the experiment. Remove the lid or glass stopper and place it upside down on the counter, preferable on a paper towel. Use a scoopula to obtain the portion required. If a scoopula is not available, pour directly from the reagent bottle onto a weighing paper, using the "cement-mixer" style of turning the bottle and the reagent will pour out slowly. Do NOT tap hard on the end of the bottle to get the reagent to come out. PUT THE LID BACK ON TO KEEP THE CHEMICAL UNCONTAMINATED. NEVER RETURN UNUSED CHEMICAL TO A REAGENT BOTTLE – FIND ANOTHER STUDENT WHO NEEDS SOME OR ASK YOUR INSTRUCTOR FOR DISPOSAL GUIDELINES. If any chemical spills, go get the instructor, the whiskbroom and the pan for cleanup and ask for disposal guidelines.

    2. **Liquids:** Be sure to read the label first, so that you obtain the correct chemical in the correct concentration! Never take the

reagent bottle to your workstation; instead bring some container that you have labeled to put the reagent in with you to the cart or fume hood where the reagent has been set out for the experiment. Remove the lid or glass stopper and place it upside down on the counter, preferable on a paper towel. If it is a glass or plastic stopper with a holder on top, you are to hold it between your fingers – do NOT put this kind of stopper on the countertop. Pour from the reagent bottle into the beaker or flask or graduated cylinder that you brought up to the cart or hood. Pour slowly and gently and finish with a slight twist of the bottle to keep the last drop in the bottle! If a drop does go down the side of the reagent bottle, put the lid or stopper back on, then take the bottle to a sink. Rinse the outside of the bottle thoroughly without getting the top of the bottle or the stopper wet and dry the bottle thoroughly and then return it to the cart or hood. If larger spills occur, send someone for the instructor to get instructions on cleanup. If it is an acid, get the acid neutralizer bottle and spray thoroughly. If it is a large acid spill, get the dry sodium bicarbonate and sprinkle it on the acid until there is no more bubbling. PUT THE LID BACK ON TO KEEP THE CHEMICAL UNCONTAMINATED. NEVER RETURN UNUSED CHEMICAL TO A REAGENT BOTTLE – FIND ANOTHER STUDENT WHO NEEDS SOME OR ASK YOUR INSTRUCTOR FOR DISPOSAL GUIDELINES.

4. **Pipettes**: Obtain the correct size of pipette and pipette pump or bulb. If the pipette looks very clean and dry, you may start using it directly. If it looks wet, condition it prior to the first use by drawing up a small portion of reagent and rinsing around the inside of the pipette. Dispose of the reagent per instructions (your excess solutions beaker or the sink or a special collection container). Use the pipette pump or bulb to draw the reagent up into the pipette to the calibration mark. Hold it there and move the pipette quickly to the beaker or flask. Pull the pump or bulb off or press the release button on the pump to let the reagent drain out. DO NOT PUSH THE LAST PART OF THE REAGENT OUT OF THE TIP. Most pipettes are calibrated TD – to delivery.

5. **Burettes**: Always clean the burette barrel with a brush and rinse thoroughly. Follow with a deionized water rinse. Close the stopcock at the bottom of the burette. Then rinse it twice with 5-mL portions of the reagent that will be the titrant. Roll the barrel to contact it with the reagent, and then drain through the tip to clean the tip with the reagent. (Dispose of the rinse portions per the instructor's guidelines.) Obtain a burette clamp and put it on a ring stand. Close the stopcock again and fill the burette to above the 0.00 mL mark with the reagent. Open the stopcock and drain a few milliliters rapidly to push the air bubble out of the tip. Stop draining at 0.00 mL or beyond. Cover the top of the burette with a test tube to minimize exposure to air.

1. **Titration technique**: record the initial volume of the burette after giving the reagent at least 15 seconds to drain off the sides of the barrel. All reading should be taken with the meniscus at your eye level to avoid parallax. Read the volume at the bottom of the solution's meniscus. Most burettes are marked at 1-mL increments with lines around the barrel and at 0.1-mL increments with small lines. You should be able to estimate to the 0.01 mL when reading the burette. Read from the top down. You will have a flask or beaker with analyte in it. Put down a piece of white paper so you can see any color changes in the indicator more readily. Use one hand to work the stopcock and one hand to swirl the flask or beaker as you add titrant to the analyte. Stop titrating occasionally to rinse the tip of the burette and the walls of the flask or beaker with deionized water. As you approach the end point (color change of an indicator), slow down the addition of titrant. Try to go one drop or less at a time. If a color change occurs, swirl and see it persists for at least 30 seconds. If so, you have completed one titration, but if not, try to add one drop or less again. Record the final volume to two decimal places. Subtract the initial volume to obtain the volume required for the titration. At the end of the lab period, drain the titrant back into your container if you need to use it again. If not, then drain it to your excess solutions beaker and get instructions for its disposal. Rinse the burette with deionized water two or three times, including rinsing the tip. Replace the burette in storage.

6. **Making solutions**: You will probably be making some aqueous solutions during your general chemistry lab. You will usually need a volumetric flask. Solutions can be made by dilution or by dissolving the solid chemical in the solvent. Obtain the proper amount of solid or concentrated chemical. Start with the volumetric flask about ½ full of deionized water. Then add the solid or concentrated liquid slowly. Mix it well. Then add more deionized water until you are close to the mark on the neck of the volumetric flask. Switch to adding the deionized water with a clean eyedropper. DO NOT EXCEED THE MARK. If you do, dispose of the solution and start over!

7. **Decanting**: You may have formed a precipitate and want to separate the solid from the solution (called the supernatant). Let the precipitate settle to the bottom of the beaker or centrifuge tube (hopefully). Use a clean stirring rod and pour the solution slowly into another container, letting the solution trace down the stirring rod.

8. **Gravity filtration**. Get a piece of filter paper. Fold it in half, then fold it not in a perfect half again. Tear a small piece off the shorter corner. Put the folded filter paper in a glass funnel that is held by a burette clamp or set in a ring on a ring stand. Open it out in the larger part and press it against the top rim of the funnel. Wet it with the solvent, usually deionized water. Put a beaker or flask under the tip of the funnel, so that is touching the side near the

top. Decant most of the solution using your glass stirring rod, then swirl the remaining precipitate and solution and pour it quickly into the funnel without going over the top of the shortest part of the filter paper. Use the solvent to rinse the precipitate out of the container. This is gravity filtration, so you do have to give it time to drain. If you are interested in the solution, you can move on to the next step of your experiment. If you are interested in the precipitate, you may have to air-dry it in your locker or dry it in the oven.

9. **Vacuum filtration**. Obtain a Buchner funnel, funnel flask and rubber tubing. Get the filter paper that fits in the Buchner funnel. Set up the funnel flask first, using a utility clamp to hold it securely to a ring stand and place it near the sink. Attach the rubber hose to the water aspirator on the water tap at the sink. Put the Buchner funnel on the flask, check that the filter paper covers all the holes, and moisten the paper with the solvent, usually deionized water. Turn on the aspirator. Transfer the solution as described above for gravity filtration. Then transfer the precipitate and rinse the container. Let the water continue to draw suction on the filter paper for a few minutes. After that, disconnect the hose from the water aspirator prior to turning it off. Then turn off the water flow. (To prevent loss of the solution and/or contamination of the drains, you may be required to set up a trap. Consult with your instructor.)

10. **Centrifuges**: Use a centrifuge tube as designated by your instructor. Have a matching centrifuge tube filled to the same level as your tube. Place them in opposite sides of the centrifuge to balance the centrifuge. Turn it on to the speed required and leave it on for the time required. (If it makes a loud noise and/or vibrates a lot, turn it off and check your tubes locations and liquid levels. Do not let it continue to run.) Turn off the centrifuge and then WAIT for the centrifuge to come to a stop. Do NOT touch the spinning part to slow it down!!!

11. **Simple glass working**. You may be required to make a tube or bend a tube or make a stirring rod. This will require cutting off a specific length of solid glass rod or hollow glass tubing. You use a metal file to score at the measured length, by pushing the file away from you on the glass. Only make one mark. Pick up the glass rod or tube and turn the scored mark away from you. You can hold it with a paper towel if you want. Put your thumbs on either side of the mark and apply force quickly. The glass should snap at the mark. If you have some jagged edges on the glass you can rub them off with a wire gauze prior to fire-polishing the ends. You must fire-polish any piece of glass that has been cut or broken to prevent injuries in the use of the glass piece. Light the burner and place the end of the rod or tubing downward at about a 45° angle in the middle of the flame. Rotate it continuously and pull it out to check on your progress frequently. Hollow tubing will be done more quickly that solid gas rods. Put the hot piece of glass on a piece of wire gauze and let it cool for more than five minutes. NEVER set it on the counter or paper!

    You can bend glass tubing by obtaining a wing tip for the burner from the stockroom. You hold the glass tubing over the wing tip and rotate it constantly until you feel a softening. Take it out of the flame for 1-2 seconds, then bend it and hold it until it feels solid. Set it on a piece of wire gauze to cool. (Usually you will want a 90^{o} angle in your bend.)

If you need to insert glass tubing through a rubber stopper, first get a dropper bottle of glycerin (glycerol) and lubricate the tip of the glass tubing and the opening in the rubber stopper. Hold the stopper and the tubing with pieces of paper toweling. Hold the stopper in one hand and hold the glass tubing about one inch from the end to be inserted. Insert with a twisting motion with only moderate pushing. You keep moving your hand back about one inch until you have the length inserted that you need.

12. **Spectrophotometers**: There are various types of spectrophotometers that are used to determine the interaction of electromagnetic (EM) radiation with chemical species. The ultraviolet (below 400 nm), visible (400 to 750 nm) and infrared (above 750 nm) portions of the EM spectrum are the most commonly used wavelengths for this experimental determination. A spectrophotometer contains at least one light source, a slit, a diffraction grating, a sample holder (which holds small glass cuvettes), a photomultiplier, an amplifier, and a meter. A wavelength is selected by using the diffraction grating. The wavelength of light passes through the sample and may be absorbed by the chemical species as it is excited to a higher energy state. The amount of light that passes through is called transmitted light and is measured as percent transmittance compared to a blank sample in the cuvette. If a complete spectrum is run, the transmittance is measured at a variety of wavelengths. Percent transmittance is converted to absorbance: Abs = $\log (I_o/I_t) = \log (100/\%T)$. I_o is the intensity of the incident light and I_t is the intensity of the transmitted light. Another way to look at I_o is the intensity of transmitted light through the blank cuvette. Each chemical species will have its own unique absorbance spectrum. The area of maximum absorbance determines the color that the chemical species has.

    The actual absorbance is determined by the concentration of the species in solution, by the thickness and other properties of the cuvette, and by the molar absorptivity coefficient. These three are formulated in Beer's Law.

    $A = a \cdot b \cdot c$

    A is the absorbance, **a** is the molar absorptivity coefficient, **b** is the thickness of the cuvette and substance, and **c** is molar concentration. If Absorbance is plotted on the y-axis and concentration on the x-axis, then the slope is ab. The y-intercept should be zero. If absorbance is determined for a variety of known concentrations and graphed as just described, then we have a concentration curve that can be used to determine the concentration of an unknown from its absorbance.

    A spectrophotometer must be warmed up prior to use. It also needs to be calibrated each time it is used. The sample must not be contaminated and the cuvette must be kept very clean and dry on the outside. The same cuvette should be used for the entire experiment, since each has its own absorptivity. The cuvette should be rinsed several times with deionized water. Usually deionized water is used as the sample blank as well. Dry the outside of the cuvette only with KimWipes, which will not scratch the surface. Handle the cell only by its top, above where light is transmitted through the cuvette.

To calibrate: with no cuvette in the sample holder and the wavelength selected, set the reading at 0.0 percent transmittance. Then fill the cuvette about ¾ full with the blank solution and put it in the sample holder. Choose the wavelength of light required and set the reading at 100.0 percent transmittance.

When ready to use the solution under study: rinse the cuvette twice with the solution and discard. Then fill the cuvette about ¾ full with the solution. Make sure the cell is dry on the outside. Put it in the sample holder and take a reading.

13. **pH measurements and pH meters**:

    One way to determine the acidity of an aqueous solution is to measure its pH, which is the log of its [H_3O^+].

    The formula is: $pH = -\log[H_3O^+]$

    The pH can be measured by pH indicator paper, pH indicators, or a pH meter or probe.

    A generic pH meter contains a reference electrode and a glass or calomel electrode (which may be combined into one piece), wiring, and a scale. The tips of both electrodes have to be in the solution to complete the circuit and measure the electrical signal that is converted to pH units. Electrodes are very sensitive and expensive, so they must be handled with care. They must also be kept submerged in solution to prevent them from drying out. Don’t let the pH probe hit against the sides or bottom of a container.

    Most pH meters require calibration. They may also have a temperature setting or knob. Typically the pH 7 buffer is tested first and an adjustment is make, then either the pH 4 buffer or the pH 10 buffer is tested, depending on whether an acid or base is being measured.

14. **Sublimating**: some solids will go directly from the solid state to the gaseous state upon heating. This must be done slowly and with caution in the fume hoods.

Guidelines for Disposing of Laboratory Waste

Environmental legislation sets the rules for disposing of laboratory waste and the Solano Community College chemistry faculty and staff are committed to following the regulations. As directed by the laboratory experiments' procedures and your instructor, you will collect some chemicals in labeled containers that are then disposed of by the chemistry lab technician.

Unused reagents are sometimes returned to the stockroom to be used in other lab sections, but always consult with your instructor. NEVER put unused reagents back in reagent bottles.

You will be following your laboratory procedure to neutralize acids and bases before disposal down the drain. You should keep a large beaker labeled "Excess solutions" during your experiments for acids and bases. You can then adjust the pH to the desired range of 3 – 11 using HCl or H_2SO_4 to lower the pH and NaOH or $Mg(OH)_2$ to raise the pH. Make sure you don't pour solutions containing toxic metal ions into your excess solutions beaker. If your acid or base is greater than 10%-mass, pour into a large beaker of cold water to dilute it before neutralizing it.

Cations and Anions of Low Toxic Hazard: If soluble in water and less than 1-%-mass, these can go down the drain.

Cations: Al^{3+}, Bi^{3+}, Ca^{2+}, cerium, cesium, Cu, Au, Fe, lanthanides, Li^{+}, Mg^{2+}, Mo^{6+}, niobium(V), palladium, K^{+}, Rb^{+}, Sc^{3+}, Na^{+}, Sr^{2+}, tantalum, Sn^{2+}, Sn^{4+}, Ti^{2+}, Y^{3+}, Zn^{2+}, zirconium

Anions: HSO_3^{-}, BO_3^{3-}, $B_4O_7^{2-}$, Br^{-}, CO_3^{2-}, Cl^{-}, OCN^{-}, OH^{-}, I^{-}, O^{2-}, PO_4^{3-}, SO_4^{2-}, SO_3^{2-}, SCN^{-}

EXPERIMENT 01: PHYSICAL PROPERTIES OF SUBSTANCES

Objective 1

This first experiment introduces the student to a variety of laboratory techniques and to physical properties, such as solubility, freezing (melting) point, boiling point, and density.

Introduction

Physical properties, as opposed to chemical properties, can usually be observed without destroying the sample. There are many physical properties that each substance has that distinguish it from other substances. Appearance is just one area for observation, which would include the substance's phase, color, and general appearance (shiny or not, dull or sharp, crystalline or amorphous, powdery or chunky, etc.) Many other physical properties beyond those listed above are also used to identify a substance.

Solubility is designated by the maximum number of grams of substance per 100 grams of solvent, at a specific temperature. Dissolving a *solute* into a *solvent* makes a *solution*. Some solvents will dissolve more of the substance in question than others, depending on their physical structure as compared to the substance.

Melting point is the temperature at which the substance changes phases from solid to liquid or vice-versa (freezing point). If a substance stays at its melting point, it will reach an equilibrium where both liquid and solid are present and the amount of each no longer changes.

Boiling point is the temperature at which the substance changes phases from liquid to gas or vice-versa, it is condensing. If a substance stays at its boiling point, it will reach equilibrium where both liquid and gas are present and the amount of each no longer changes. Boiling point is very much affected by the atmospheric pressure, so it has to be recorded when determining the boiling point of a substance.

Density is defined as the mass of the substance per volume occupied. For liquids and solids, the density is usually reported in grams/milliliter or grams/cubic centimeter. (A milliliter is identical to a cubic centimeter.) Gases have density reported in grams/liter. Densities of solids and liquids change slightly with temperature, but density of a gas changes significantly with temperature. See the appendices for the density of water and some other substances.

Your instructor may want to review significant figures for reporting data and doing the calculations for this experiment. In the table below is a list of inherent instrument error.

Equipment Accuracy

This is a table of common chemistry equipment and their associated accuracy.

Equipment	Accuracy
Analytical balance	±0.0001 g
Top-loading balance	±0.001 g
Meter stick	±0.1 cm
Graduated cylinder	±0.1 mL
Pipette	±0.02 mL
Burette	±0.02 mL
Thermometer	±0.2°C

Examples

1. A sample metal rod with a mass of 45.872 g displaces 4.05 mL of water in a graduated cylinder. Determine the density of the metal and identify it from the table in the appendices.

D = m/V = 45.872g/4.05 mL = 11.33 g/mL. This is very close to Pb in the table.

2. A student is able to dissolve 5.355 g of NaCl in 15.000 g of water. What is the solubility of NaCl in water?

Solubility is expressed in grams of solute per 100 g of solvent. Set up a simple ratio:

$$\frac{5.355\text{g}}{15.000\text{g}} = \frac{x}{100.0\text{g}} \qquad x = 35.7\ \text{g}/100\text{g}$$

3. A student tests a red alcohol thermometer with freezing water and boiling water. The freezing water temperature is 2.5°C and the boiling point reads 105.0° Assume that barometric pressure is very close to 760 torr. Determine the equation of the line ($y = mx + b$) to correct these temperatures when using this thermometer. Then use it to correct the student's reading of room temperature of 24.0°C to the true room temperature.

You should know and remember that the true freezing point of water is 0.0°C and that the boiling point is 100.0°C.

Step one: determine the slope: $m = \Delta y/\Delta x = (100.0-0.0)/(105.0-2.5) = 0.9756$

Step two: determine the y-intercept: $100.0 = 0.9756(105.0) + b$; $b = -2.438$

Step three: $y = 0.9756x - 2.438$

Step four: $y = 0.9756(24.0) - 2.438$; $y = 21.0$°C

Prelaboratory Questions:

1. A student obtains an unknown metal in the shape of a small cylinder. She determines its mass to be 22.155 grams. Her graduated cylinder contains 31.22 mL of water and she adds the metal to it. The new volume is 33.33 mL. Determine the density and identity of the metal.

2. Calculate the boiling point of pure water at 653.7 torr.

Safety:

Always wear appropriate PPE (see page *i*).

Materials:

Chemicals: ice, deionized water, ethanol, vegetable oil, sodium chloride or sodium nitrate, 6 M acetic acid, glucose or sucrose solid, eggs, metal unknowns for density determination.

Equipment: small beakers, ten test tubes, red alcohol thermometers, hot plates, plus have a barometer available; (some instructors may choose to use Bunsen burners); 10-mL pipettes and pipette bulbs; three burettes set up to dispense 10 mL of each of the three solvents; dispensing bottle for each solvent set to dispense 10.0 g of each.

Methods:

You will probably be asked to work with a lab partner for this experiment. Each of you should fully participate in each step and each should record all *the* data. Part of your data is your lab partner's name!

Part A: Determine the boiling point and freezing point of pure water. Use the data and corrections to calibrate a red alcohol thermometer. Suspend the thermometer and read room temperature. (Note: your instructor may ask you to label and keep your thermometer once it is calibrated. If that is so, then each lab partner needs to calibrate a thermometer to keep.)

1. Obtain a thermometer and some deionized (DI) water. Using two relatively small beakers, put DI water in one with lots of ice cubes, and put the thermometer in the beaker for at least fifteen minutes. While you are waiting, put about 50 mL of DI water in the other beaker and put it on the hot plate until it boils. (You may even have time to set up other parts of this experiment.) Now go back to the ice-cold water and record the temperature to at least one decimal place. Let the thermometer sit at room temperature for a couple of minutes, then hold it in the boiling water for a couple of minutes. Record the temperature to at least one decimal place. NOTE: to properly use and record a temperature, set up the thermometer or hold it so the bulb or tip is not touching the bottom or sides of the container.

2. Either read or obtain the barometric pressure and correct the boiling point of water for the variation from standard pressure. True boiling point = $100.0 - (760 \text{ torr} - P_{bar}) * 0.037 ^\circ\text{C/torr}$).

3. Create a graph using your thermometer's freezing point vs. the true freezing point and your boiling point vs. the true boiling point of water from step 2 above. If possible, use the computer program to determine the equation of the line. This would be the equation to use the make corrections to temperature readings for the laboratory thermometer. Record the equation of the line somewhere in your laboratory manual so that you can always refer to it for future experiments, if you keep the same thermometer.

Part B: Determine the solubility of various substances in water, ethanol and vegetable oil. The substances might include glucose or sucrose, sodium acetate, sodium chloride, sodium nitrate, acetic acid, egg white, etc.

The three solvents are deionized water, ethanol and vegetable oil. They should be in labeled dispensing bottles set at the correct volumes to deliver 10.0 grams of each. The volume will NOT always be 10.0 mL for each solvent. See the instructor for the correct volume for each.

Weigh a 100- or 150-mL beaker and record the mass. Set up nine test tubes in your test tube rack. Label three sets of three test tubes for the three solvents (like W-1, W-2, W-3, E-1, E-2, E-3, O-1, O-2, O-3), then weigh each test tube. Put 10.0 g of each solvent in each test tube by dispensing the appropriate volume from the burette or dispensing bottle and weigh each test tube again by placing it in the preweighed beaker. Your instructor will tell you which three solutes to try to dissolve in each. Then add 1 g of solute to each at a time, until you reach a solubility limit. Record the mass just before reaching the solubility limit.

Dispose of all solutions and solvents down the drain, AFTER checking to see if another student can use your excess if you have any. Solids can be put in the wastebasket.

Part C: Density can be determined at a given temperature by measuring the mass and volume of a substance. Volume can be measured directly or by displacement of water. You will do both.

1. For each solvent in Part B above, pipette 10.00 mL into your <u>preweighed</u> small graduated cylinder. Record the mass of the graduated cylinder plus solvent. Use labeled or clean pipettes each time. Practice pipetting a few times prior to transferring the liquid to the small graduated cylinder. If you are working with a lab partner, show the instructor that both of you know how to use your pipette. (Alternatively, your instructor may ask you to use the previously set up burettes of the three solvents, and you will learn how to dispense exactly 10.00 mL from the burettes.) Use the mass and volume to determine density for each solvent.

2. Record room temperature.

3. Obtain a metal unknown from your instructor. Describe its appearance and record its number (if it has one). Weigh it and record the mass to three decimal places. Get out your large graduated cylinder and put about 30 mL of water in it. Record the volume to two decimal places. Remember to read the bottom of the meniscus and have it at eye level to avoid parallax. Carefully place the metal in the cylinder with the water and read the new volume to two decimal places. Determine the volume of water displaced and calculate the density. Look at the table of densities in the appendix and see if your density matches any of the metals listed there. If it does, then identify it by name. Return the metal unknown to your instructor.

Data Collection:

- Freezing Point of water _______
- Boiling Point of water_______
- Barometric Pressure _______
- Room Temperature reading _______

- Enter name of each solute in column one. Record the mass of each solvent and the maximum grams of solute under its heading in columns two, three, and four.

This is a table for students to list solutes and masses of solvents.

Solute/Solvent	**Water**	**Ethanol**	**Veg. Oil**

Observations:

- Record the mass of 10.00 mL of each solvent below.

Mass of small grad cyl _______

Mass of grad cyl + water _______ Volume reading _______

Mass of small grad cyl _______

Mass of grad cyl + ethanol _______ Volume reading _______

Mass of small grad cyl _______

Mass of grad cyl + veg. Oil _______ Volume reading _______

Unknown metal number _______

Observations of unknown metal:

Mass of metal unknown _______

Volume of water in cylinder _______

Volume of water + metal _______

Water displaced _______

Analyses/Calculations/Results:

Correct the boiling point of pure water for atmospheric pressure:

True boiling point = 100.0 – (760 – P_{bar})*0.037°C/torr

Corrected Boiling Point of water _______

Calculate solubility in grams solute per 100 grams of solvent and enter into the table.

Equation of the line:

Corrected Room Temperature: _______

This is a table to input calculated solubility

Solute/Solvent	Water	Ethanol	Veg. Oil

This is a table to input mass, density, and true density.

Solvent	Water	Ethanol	Veg. Oil
Mass in grams			
Density in g/mL			
True Density*, g/mL			

*At today's room temperature

Density setup and calculation:

Density ______ Possible identification ______

Post Lab Questions

1. Why is it necessary to calibrate instruments, such as balances, and measuring devices, such as thermometers and pipettes?
2. An object has of mass of exactly twelve grams on an analytical balance that has an accuracy of +0.1mg. How should the student record this mass, i.e., what number should be written down?
3. Use whatever resources necessary and find the density and melting point of osmium. Compare it to other metals and discuss these two physical properties.
4. Look up the density of water at your recorded room temperature. (See appendix.) Use the density and the mass of the pipetted (assumed) 10.00 mL of water as solvent in Part B to determine the actual volume of water delivered. Then determine your pipette’s (or your) error.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 02: SEPARATION OF COMPONENTS OF A MIXTURE

Objective 2

You will learn a number of laboratory techniques in this lab, such as decanting, extracting, etc.

Introduction:

Mixtures can be homogenous or heterogeneous in composition. Homogenous mixtures are called solutions. Mixtures of solids are usually heterogeneous. Either type can be separated by physical means back to its pure components. You will be assigned a mixture containing three chemicals: ammonium chloride, sodium chloride and sand. Ammonium chloride can be sublimed at 520°C, sodium chloride dissolves in water, and sand remains in the solid state.

You will need to review procedures for the following techniques: decanting, gravity filtering, and sublimating.

You will be determining the percent-by-mass of each component. Recall that percent is the part divided by the whole and multiplied by 100.

Example: A mixture of glucose and sand weighed 3.033 g. The glucose was dissolved in water and decanted from the sand. The water was evaporated and the glucose had a mass of 1.232 g. The sand was dried and had a mass of 1.811 g. The student needs to report the percent of each compound in percent mass.

Percent-mass = 100 * (1.232/3.033) = 40.62% glucose

Percent-mass = 100 * (1.811/3.033) = 59.71% sand

Although this adds up to slightly more than 100%, it is within tolerance.

Prelaboratory Questions:

1. Define the following: sublimation, decantation, dissolving, and evaporation.

2. In reading through the procedure, explain why you think the sand is rinsed more than once.

3. Identify each of the following as an homogenous mixture, heterogeneous mixture, pure compound or pure element:
    1. soil
    2. air
    3. a soft drink
    4. gold
    5. blood
    6. table salt

Safety:

Always wear appropriate PPE (see page *i*).

Materials:

Chemicals: sodium chloride, ammonium chloride, sand (silicon dioxide)

Equipment: evaporating dishes, watch glass, stirring rod, large graduated cylinder, ring stand, ring, clay or wire triangle, Bunsen burner, striker, tongs

Methods:

Work with one partner for this experiment. You'll be working in the fume hoods and have to share space.

1. Weigh a clean dry evaporating dish, record the mass, and then put about 2 grams of the unknown assigned to you in the dish and weigh it again. Find the unknown mass by difference and record it.

2. Find a spot in one of the fume hoods and set up your ring stand, triangle, Bunsen burner, etc. Put the evaporating dish containing the unknown on the triangle, and light the burner. Close the fume hood window most of the way. There will be white fumes. Use your stirring rod to stir once every five minutes for a total of 15 minutes. (You may have to move the burner away while you are stirring.) At the end of this sublimation step, turn off the burner and let the dish cool for at least ten minutes, letting it reach room temperature. Weigh the dish and contents. Calculate the difference in mass from step 1 above and record it as the mass of ammonium chloride.

3. Now dissolve the sodium chloride in the mixture in about 25-mL of deionized water. Mix for several minutes to make sure only sand remains as the solid. WEIGH ANOTHER clean dry evaporating dish, record the mass, and then transfer by decanting the sodium chloride solution from the first evaporating dish. If any of the sand ends up being transferred, pour it all back in the first dish and try again. Rinse the sand with another 10-mL of deionized water and decant, and repeat this rinse and decant step once more.

4. Place the evaporating dish containing the solution on the triangle above the Bunsen burner. HEAT SLOWLY at first – you don't want it to boil hard or spatter or you will lose NaCl. Cover it with a watch glass to prevent spattering when the liquid level has dropped more than 75%. DURING THIS TIME, you can dry the other dish containing the sand on another burner or a hot plate, again using a watch glass to prevent spattering. Start heating slowly and then turn it up. You may have to uncover it and break up lumps in the sand with your stirring rod. You will know when both are done and dry, because water vapor will stop appearing on the watch glass. Let the dishes cool to room temperature. Weigh the dish containing the NaCl. Subtract the mass of the evaporating dish you originally weighed, and record the difference as the mass of NaCl. Weigh the dish containing the sand and subtract the weight of the second evaporating dish to obtain the mass of the sand and record it. DO ALL THE CALCULATIONS BEFORE YOU CLEAN UP.

Disposal: (after showing your calculations to your instructor) wipe the sand out into a trash can. Rinse the ammonium chloride and sodium chloride out of the evaporating dishes down the sink.

Data Collection:

DATA TABLE AND OBSERVATIONS

Unknown Number __________

Observations:

Mass of evaporating dish __________

Mass of dish + unknown __________

Mass of unknown __________

Mass of dish + sand/NaCl __________

Mass of NH_4Cl __________

Mass of second evap dish __________

Mass of dish + NaCl __________

Mass of NaCl __________

Mass of dish + sand __________

Mass of sand __________

Mass of NH_4Cl, NaCl + sand __________

Difference from unknown mass __________

Analyses/Results/Calculations:

Determine the percent mass of each component of the mixture and create a results table showing the unknown number and the percent mass of each component. Also total the three masses and see if they are within 99-101% of the original mass. If not, you need to dry something a little more!

RESULTS TABLE

Unknown number _________

Table 1. Results

	Ammonium chloride	**Sodium Chloride**	**Sand**
Percent mass			

Post Lab Questions:

1. A student did this experiment and found her mixture was 15.05% NH_4Cl, 19.95% NaCl and 75.10% sand. What's wrong with these numbers and what might have caused the problem?

2. What major principles are demonstrated by this experiment?

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 03: CHEMICAL PROPERTIES OF A SUBSTANCE

Objective 3

This experiment introduces the student to a variety of laboratory techniques and to chemical properties when a variety of compounds react with each other.

Introduction:

Chemical properties are chemical reactions, and as such, they destroy the "sample." In contrast, measuring physical properties does not usually destroy the sample. By definition, a *chemical reaction* changes the reactants into something new – the products. There are many chemical reactions that can be used to distinguish one substance from another. Appearance of the reactants and products is the main area for observation.

How will you know if a chemical reaction occurred? There are five ways:

- a gas forms, detectable by observing bubbles or an odor
- a precipitate forms, either making the solution look cloudy or else actually appearing as a solid forming in the bottom of the reaction vessel
- a color change occurs, but this is difficult to use as a confirmation, because mixing a solution of blue food coloring with yellow food coloring would show a color change
- a temperature change, detected by thermometer or by touch, and either going up or down, indicating an exothermic or endothermic reaction
- a pH change, detectable using a pH meter, pH paper or pH indicator

Also, some chemical reactions take more time than others, so you must observe the reaction for more than a few minutes to be sure a reaction has occurred.

Examples:

When an acid and a base react, a ***neutralization reaction*** occurs. The reaction is ***exothermic*** and the reaction surroundings will increase in temperature.

When an aqueous solution of silver nitrate is mixed an aqueous solution of sodium chloride, both solutions being clear and colorless initially, white precipitate forms. If you look at the solubility rules or table, you'll see that silver chloride is insoluble in water. This is a ***precipitation reaction***.

When an acid reacts with any carbonate, carbon dioxide gas forms. This is a ***gas-forming reaction***.

Safety

Always wear appropriate PPE (see page *i*). You might want to wear gloves since you will be using an acid and a base.

Materials:

Chemicals: 2.0 M solutions of magnesium sulfate, hydrochloric acid, sodium hydroxide, sodium acetate, sodium chloride, sodium carbonate, ammonium chloride, magnesium nitrate, pH paper

Equipment: small test tubes or well plates, a large beaker for excess solutions

Methods:

You will first be observing the reactions of six solutions with three reagents, then testing two unknown solutions per student and identifying them as two of the six solutions.

Part A: You will be doing fifteen tests for reactions, either in small test tubes or in a well plate. If you use a well plate, scrub it with soap and water, rinse with tap water, and then rinse with deionized water. Put about five drops of each of the five solutions (sodium acetate, sodium chloride, sodium carbonate, ammonium chloride, and magnesium nitrate) in a grid of fifteen slots on the well plate. Then add five drops of one of the three reagents (magnesium sulfate, hydrochloric acid, and sodium hydroxide) to each of the five solutions, using up five of the slots. Observe carefully for formation of a gas, odor, precipitate, etc. Feel the bottom of each slot in the well plate to see if it is warmer or colder. Record your observations. Move on to adding five drops of the second of the three reagents to each of the five solutions, using up another five of the slots. Observe and record. Now do the last reagent.

If you have run out of slots on the well plate or if you have run out of small test tubes, then read ahead on disposal at the end of part B. Otherwise, continue to part B.

Part B: Ask your instructor which two of the five unknowns are assigned to you. Obtain them and record their unknown numbers or letters.

You will be doing six tests for reactions, either in small test tubes or in a well plate. If you still have room in your well plate, put about five drops of each of two unknown solutions in a grid of six slots on the well plate. Then add five drops of one of the three reagents to each of the two unknown solutions. Observe carefully for formation of a gas, odor, precipitate, etc. Feel the bottom of each slot in the well plate to see if it is warmer or colder. Record your observations.

Use your data from part A to identify your two unknowns. Getting the right answers here is worth points, so do this part as carefully as possible.

Disposal: pour all the slots of your well plate or small test tubes into your large excess solutions beaker. Add water to dilute the solutions. Get pH paper and determine the pH. If it is between 3 and 11, pour the excess solutions down the drain. If not, use 6 M HCl or 6 M NaOH to adjust the pH. Then pour the excess solutions down the drain.

Scrub your well plate or your test tubes with soap and water, rinse with tap water, then rinse with deionized water.

Prelaboratory Questions:

1. Fill in what you *think* will happen in a reaction grid below, using your textbook and a solubility table to assist you.

Solution/Reagent Table

Solution/Reagent	**$MgSO_4$**	**HCl**	**NaOH**
$NaC_2H_3O_2$			
NaCl			
Na_2CO_3			
NH_4Cl			
$Mg(NO_3)_2$			

2. You have been given the names in the list of chemicals and the actual chemical formulas in question 1 above. Now write the names and formulas together.

Data Collection:

Part A: Fill in your observations below. Use (g) to indicate a gas and then say bubbles or odor after it. Use (Ppt) to indicate a precipitate formed. DESCRIBE this precipitate!!! (Cloudy, milky, settled, clumpy, white, dirty white, yellow, etc.) Use T or T^- to indicate that the reaction caused a temperature increase or decrease. Use NR to indicate that no reaction occurred.

Reagents

Solution/Reagent	**$MgSO_4$**	**HCl**	**NaOH**
$NaC_2H_3O_2$			
NaCl			
Na_2CO_3			
NH_4Cl			
$Mg(NO_3)_2$			

Part B: Fill in your observations below just like you did in part A.

Observation Table

Solution/Reagent	**$MgSO_4$**	**HCl**	**NaOH**
Unknown _______			
Unknown _______			

Analysis/Calculations/Results:

Post Lab Questions:

Identification of unknowns:

Unknown _______ = ___________________________________

Unknown _______ = __________________________________

1. How do chemical properties differ from physical properties?

2. Write the chemical reactions for where they occurred in part A. Use subscripts to indicate phases.

3. How could you distinguish between the two gases that form in part A? One is carbon dioxide and the other is ammonia.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 04: FORMULA OF A HYDRATE

Objective 4

You will determine the elemental composition of at least one unknown hydrate and determine its formula in the hydrated form.

Introduction:

Hydrates are ionic compounds that have incorporated water molecules as part of their crystal structure. The ratio of water molecules to one unit of the ionic compound is constant, just as in other compounds. Some common hydrates are listed below with their common names and their IUPAC names.

Table 1. Common names, chemical names, and formulas.

Common Name	**Chemical Name**	**Formula**
Alum	Potassium aluminum sulfate dodecahydrate	$KAl(SO_4)_2 \bullet 12\ H_2O$
Borax	Sodium tetraborate decahydrate	$Na_2B_4O_7 \bullet 10H_2O$
Bluestone	Copper (II) sulfate pentahydrate	$CuSO_4 \bullet 5\ H_2O$
Gypsum	Calcium sulfate dihydrate	$CaSO_4 \bullet 2\ H_2O$
Washing soda	Sodium carbonate decahydrate	$Na_2CO_3 \bullet 10\ H_2O$
-	Iron (III) chloride hexahydrate	$FeCl_3 \bullet 6\ H_2O$

You will be determining the percent-by-mass of water in the unknown hydrate assigned to you.

Example

A hydrate showed that it was 14.7 %-mass water and the anhydrous molar mass of the compound was 208.2 g/mol. Assume 100.0 g of the hydrated salt. Then it would be 14.7 g of water and 85.3 g of anhydrous salt.

14.7 g/18.015 g/mol = 0.816 mol H_2O

85.3 g/208.2 g/mol = 0.410 mol salt

Mole/mole ratio = 0.816/0.410 = 1.99, which is within $\pm$0.05 of a whole number, so there are 2 moles of water for every mole of salt. The formula would be salt$\cdot 2H_2O$.

Prelaboratory Questions:

1. Why do you handle the crucible and lid only with crucible tongs?

2. Define a hydrate. What are the rules on naming a hydrate?

3. Why is the crucible covered by the lid while cooling, when before that you have to keep the lid slightly ajar?

Safety:

Wear appropriate PPE (see page *i*). Wear gloves if you need to use acid to clean your crucible. Do not wear the gloves when heating the crucible.

Materials:

Chemicals: one or more hydrates

Equipment: Bunsen burner, striker, ring stand, ring, triangles, crucibles and lids, crucible tongs, 250 mL beaker for excess solutions and rinsing.

Methods:

1. Look at the crucible and see if it's clean and that there are no cracks apparent in it. If it is cracked, trade it in for a better one. Clean out the crucible as best as you can. If you see solids dump it into the waste container in the hood. Use a spatula to scrape as much debris as you can from the crucible, adding some water can help loosen the debris. Dump everything into the waste container, then wipe the crucible dry. To dry further, use the Bunsen burner to heat the crucible slowly for a couple of minutes, then heat to red-hot for a full 5 minutes. Remove the Bunsen burner. Let the crucible and lid cool on the triangle for several minutes. Handle the crucible and lid with crucible tongs only from now on, to avoid contaminating the crucible with finger oils. Weigh the crucible, then the lid, then both together. Record all of these masses. (This is an insurance policy!)

2. Find out from your instructor which unknown hydrate you will be using and obtain 2.5 g of it. Put it in the crucible and weigh the crucible, lid and hydrate together. Determine the mass of the hydrate. Record all.

3. Set the crucible on the triangle. Set the lid on so it is just slightly ajar to let the water vapor escape. Light the burner and start heating slowly. You can turn up the heat some, but don't get the crucible red-hot or you might decompose the salt! Continue to heat at a higher intensity for ten minutes. Turn off the burner, adjust the lid so it completely covers the crucible and let the crucible cool with the lid on until at room temperature. Weigh the crucible, lid and salt. Record in your data table.

4. Repeat step 3. Record the mass. If the mass in this step is within 0.01 g, you are done. If not, repeat step 3 one more time.

5. Dump the solid salt from the crucible into your excess solutions beaker. Wipe out the crucible with Kim-wipes. Weigh the crucible, the lid, then both together. Repeat the entire experiment.

6. Dump the solid salt from the crucible into your excess solutions beaker. Wipe out the crucible with Kim-wipes. Rinse with water. Use 6 M NaOH to adjust the pH of the solution in the excess solutions beaker to bring pH to the 3 – 11 range. Pour down the drain.

Data Collection:

Unknown Number ____________

Observations:

Trial 1: **Trial 2:**

Mass of crucible ________ ________

Mass of lid ________ ________

Mass of crucible + lid ________ ________

Mass of crucible + lid + hydrate ________ ________

Mass of hydrate ________ ________

First drying:

Mass of crucible + lid + salt ________ ________

Second drying:

Mass of crucible + lid + salt ________ ________

Third drying: (if needed)

Mass of crucible + lid + salt ________ ________

Mass of anhydrous salt ________ ________

Mass of water ________ ________

%-mass water in hydrate ________ ________

Average %-mass water _______

Analyses/Calculations/Results:

Determine the mass of the hydrate, the anhydrous salt and water. Calculate the percent-by-mass of the water in the hydrate and show it to your instructor. Obtain the molar mass of the anhydrous salt from your instructor and calculate the mole/mole ratio of water to salt to find the formula of the hydrate, as shown in the example above. Determine the percent-error from the actual ratio of water to anhydrous salt.

Molar mass of anhydrous salt ________

Moles of anhydrous salt ________

Moles of water ________

Mole/mole ratio of water/salt ________

Formula of your hydrate ____________________

Post Lab Questions:

1. Calculate the percent-mass of water in iron (III) chloride hexahydrate.

2. Students can either heat the crucible too much, causing the anhydrous salt to decompose, or they don't heat enough and some water remains in the crucible. Discuss the effect of both of these on the students' calculation of percent water in the unknown hydrate.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 05: CHEMISTRY OF MAGNESIUM AND EMPIRICAL FORMULA

Objective 5

You will be reacting magnesium metal with oxygen to experience a combustion reaction that is also an oxidation-reduction reaction. You will also determine the empirical formula of the oxide of magnesium.

Introduction:

Chemical formulas are determined experimentally in a number of ways. We need chemical formulas in order to know the exact composition of compounds. Formulas generally indicate what elements are involved and their smallest whole number ratios (*empirical formula*) or their smallest ratio for a whole molecule (*molecular formula*), by using the symbols for the elements and subscripts to the right of the element symbol to indicate the number of atoms.

We also use the concept of mass for atoms, empirical formula units and molecules. Your periodic table shows the average atomic mass for the atoms in *amu*, or atomic mass units. However, if we are to use laboratory scale amounts of compounds, we need to know the mass of one mole, or an Avogadro's number of units. (Avogadro's number is 6.0221×10^{23} units/mole, where a unit can be an atom, formula unit, molecule, or whatever.) The mass of one mole, or *molar mass*, is expressed in grams/mole.

We also use the concept of *elemental composition* of compounds, which is the percent by mass of each element in a compound. If we know the formula and have the molar masses available, we can determine this easily. Pure compounds always have the same elemental composition.

If we don't know the chemical formula, we can determine the elemental composition experimentally, and use the composition to determine the empirical formula.

Example

A student finds that a chemical compound is 38.7 %-mass potassium, 13.9 %-mass nitrogen, and 47.5 %-mass oxygen.

Step 1: the student assumes she has 100 g of the compound and lets %-mass be grams of each element. She converts grams to moles by dividing by the molar mass of each element.

38.7 g/39.0983 g/mol K = 0.9898 mol K

13.9 g/14.0067 g/mol N = 0.9924 mol N

47.5 g/15.9994 g/mol O = 2.969 mol O

Step 2: the student divides each element's molar quantity by the smallest molar quantity, in this case, moles of K.

0.9898/0.9898 = 1.000 mol K/mol K

0.9924/0.9898 = 1.003 mol N/mol K

2.969/0.9898 = 3.000 mol O/mol K

Step 3: If any of the ratios are not within ±0.05 of a whole number, she will have to multiply all the ratios by a whole number that brings them all within ±0.05 of a whole number. In this case, all are within range, so the empirical formula is KNO_3.

In this experiment, you will be doing the combustion reaction of magnesium. There is a possible side reaction that can interfere with the results, so follow the instructions carefully.

Prelaboratory Questions:

1. How does empirical formula compare to molecular formula?

2. Write in your own words the Law of Definite Proportions.

3. Nicotine was extracted from tobacco leaves and analyzed. Its elemental composition turned out to be 74.03 %-mass carbon, 8.698 %-mass hydrogen and 17.27 %-mass nitrogen. Determine its empirical formula.

Safety:

Wear appropriate PPE (see page *i*). Use crucible tongs to handle the hot crucible and lid. Do not look at the combustion reaction. Wear gloves ONLY if you have to handle nitric acid for cleaning the crucible. It's best not to wear gloves around hot objects.

Materials:

Chemicals: magnesium ribbon, deionized water, 6 M HCl and 6 M NaOH in dropper bottles, pH paper.

Equipment: Bunsen burner, striker, ring stand, ring, triangles, crucibles and lids, crucible tongs, small pieces of steel wool to clean Mg ribbon, KimWipes, meter sticks, 250 mL beaker for excess solutions and rinsing.

Methods:

You will be working alone but sharing data with two other students, so you should make sure that they are all following the procedure carefully.

1. Look at the crucible and see if it's clean and that there are no cracks apparent in it. If it is cracked, trade it in for a better one and give the cracked crucible to your instructor. If there is residue, use DI water and rinse into your waste beaker and then wipe out with a Kim-wipe. Wipe the crucible inside and out until the crucible is dry. Handle the crucible and lid with crucible tongs only from now on, to avoid contaminating the crucible with finger oils. Weigh the crucible, then the lid, then both together. Record all of these masses. (This is an insurance policy!) Note: some of the metal crucibles will not get hot enough for the magnesium to rapidly react with oxygen.

2. Your instructor will tell you approximately what length of magnesium ribbon to cut in order to have 0.2 to 0.3 grams (usually 3-4 cm). Polish the ribbon with steel wool, wipe with a Kim-wipe. Do NOT handle again with your fingers to avoid contamination. Coil the ribbon to fit in the bottom of the crucible, but not too tightly. Weigh the ribbon and record it to 3 or 4 decimal places.

3. Put the ribbon in the bottom of the crucible and set the crucible on the triangle. Light the burner and start heating slowly. Check under the lid in a few minutes to let a little air in, but don't do it too much or too often. You don't want the magnesium to burn too fast, plus you shouldn't look at it. Continue to heat at a higher intensity until it all looks like ash inside the crucible and no more combustion is obvious. Take off the lid and continue to heat for a minute or two, red hot on the bottom of the crucible. If combustion starts again put the lid back on. After this, let the crucible cool with the lid off until at room temperature. Using an eyedropper, add a few drops of deionized water, spreading each drop around the bottom of the crucible. This will take care of the side reaction where magnesium nitride formed, and react it to make the oxide of magnesium. ($Mg_3N_{2(s)} + 3\ H_2O_{(l)} \rightarrow 3\ MgO_{(s)} + 2\ NH_{3(g)}$) Leave the lid off and heat the crucible with low heat for a few minutes, then let it cool without the lid on. Measure the mass of just the crucible and ash, no lid. Reheat for another minute, again with a low flame, and let it cool. Reweigh crucible and ash, no lid. If the mass is within 1% of the previous mass, you're done. If not, repeat from adding a few drops of water.

4. Put about 50 mL of water in your excess solutions beaker. Dump your "ash" into the beaker and rinse the crucible with water at least three times, pouring it each time into your excess solutions beaker. Wipe the crucible dry and put it away. Get some 6 M HCl and pH paper and a stirring rod. Check the pH of your excess solutions beaker and add the acid until the solution is within the 3 – 11 pH range. Pour down the drain and clean up.

5. After you have completed your calculations, get the data and calculations from two other students in lab. Try to determine an average mole to mole ratio for magnesium to oxygen and report this in your results table of the lab report.

Data Collection:

Observations:

Mass of crucible _______

Mass of lid _______

Mass of crucible + lid _______

Mass of Mg ribbon _______

Mass of crucible + ash _______

Mass of oxide of magnesium _______

Mass of oxygen _______

Analyses/Calculations/Results:

Determine the mole ratio of magnesium to oxygen and write the empirical formula. You have the mass of magnesium ribbon. You combusted the magnesium and thus added oxygen to it to make the oxide of magnesium. Its mass can be determined by taking the last mass measurement of the crucible and ash, then subtracting the mass of the crucible. The mass of oxygen in the compound is the difference between the final mass of the product and the initial mass of the magnesium. Use the molar mass of magnesium and oxygen from the Periodic Table and determine moles of each. Divide by the smaller number of moles, and apply a multiplier if needed to bring the ratio to within +0.05 of a whole number.

As stated above, collect data from two other students and calculate an average ratio of Mg to O.

Find the true formula for the oxide of magnesium and compare yours to it.

Convert mass of magnesium and mass of oxygen to moles. Determine the lowest whole number ratio. Write the empirical formula.

Moles of oxygen _______

Moles of magnesium _______

Ratio of Mg to O _______

Empirical formula _______

Other students' data:

Student's Name ___________________ Student's Name
____

Mass of oxide of magnesium _______ Mass of oxide of
magnesium _______

Mass of oxygen _______ Mass of oxygen _______

Mass of magnesium _______ Mass of magnesium _______

Moles of oxygen _______ Moles of oxygen _______

Moles of magnesium _______ Moles of
magnesium _______

Ratio of Mg to O _______ Ratio of Mg to O _______

Average empirical formula _______

True empirical formula _______

RESULTS TABLE (Create your own)

Post Lab Questions:

1. An analysis of a sulfide of copper showed it was 20.2 %-mass sulfur and 79.8 %-mass copper. Determine its empirical formula.

2. If you didn't "take care" of the side reaction by adding water to the crucible, and some Mg_3N_2 stayed in the crucible, how would that numerically affect your final result of the magnesium to oxygen ratio? Would it be high or low? Do some calculations.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 06: CHEMISTRY OF COPPER AND PERCENT YIELD

Objective 6

You will see the various reactions of copper, experience oxidation-reduction reactions, and determine percent yield after a number of reactions with copper. You will be graded on how much of the copper you recover from your initial mass of copper.

Introduction:

Doing the many chemical reactions and physical steps required in this lab will measure your developing skills as a chemistry student. You will be working with acids and bases and you are expected to know and use the safety precautions for working with them. You will be using the balance, decanting, drying, etc. You will be identifying the types of reactions that you perform in this experiment as single replacement or double replacement, and also determining if an oxidation-reduction reaction has occurred at each step.

You will also have to record all your observations of each reaction before and after the reaction has taken place. Be prepared to note the color of solutions and precipitates, the appearance of the precipitates, whether a reaction is exothermic or endothermic, etc.

The reactions are listed below.

1. $Cu_{(s)} + 4\ HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2\ NO_{2(g)} + 2\ H_2O_{(l)}$
2. $Cu(NO_3)_{2(aq)} + 2\ NaOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2\ NaNO_{3(aq)}$
3. $Cu(OH)_{2(s)} \rightarrow CuO_{(s)} + H_2O_{(g)}$
4. $CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$
5. $CuSO_{4(aq)} + Zn_{(s)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$ (To recover the copper (Zn or Mg))
6. $H_2SO_{4(aq)} + Zn_{(s)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$ (To react excess Zn or Mg)

The percent yield of copper after all these reactions will be the recovered mass of dry copper divided by the initial mass of copper.

$$\%\ \text{yield} = \frac{recovered\ mass\ of\ Cu}{initial\ mass\ of Cu} * 100\%$$

Prelaboratory Questions:

1. How will you determine, as an observation, whether the reactions in this experiment are exothermic or endothermic?

2. If you started with 0.50 g of copper, and ended with 0.35 g of copper, what would your percent yield be?

3. Define limiting reactant. Identify the limiting reactant in this experiment.

4. Suppose you had 0.0640 g of CuO in step 3. How many mL of 6 M H_2SO_4 would be required for the reaction? (Ignore significant figures this time.)

Safety:

Wear appropriate PPE (see page *i*). Wear gloves since you are using concentrated nitric acid, which is very corrosive. If you do get some on your gloves, go rinse them, and then dispose of them and put on a fresh pair of gloves. This is to prevent you from touching something else, like your skin, with the nitric acid on your gloves. Be cautious with the NaOH and sulfuric acid as well. Also, methanol and acetone are flammable. Keep away from open flames. Use them in the fume hoods as much as possible.

Materials:

Chemicals per student: about 0.04 g of copper wire (no. 16 or 18), 1 mL concentrated nitric acid (16 M), 1 mL 6 M NaOH, 2 mL 6 M H_2SO_4, Zn strips or Mg strips, 2 mL of methanol, 2 mL acetone, and boiling chips.

Equipment per student: hot plate, small beaker to use as a hot water bath, centrifuge tube, steel wool to clean the Mg strips, large beaker for excess solutions, pH paper for student to neutralize excess solution, KimWipes.

Methods:

You may be asked to work with one lab partner to minimize the amount of chemicals used, and to reduce the amount of fume hood space required for this experiment. Get a hot plate for your workstation and put a small beaker 2/3 full of deionized water on it on LOW for now. You may also want to put in boiling chips and 2-3 drops of 6 M HCl to prevent hard water deposits. Put another small beaker of deionized water on the hot plate to get warm for step 3.

Remember to record ALL your observations during the experiment.

1. Get about 0.04 g of no. 16 or no. 18 copper wire. Record the mass to four decimal places so that you have three significant figures. Coil up the wire so it fits in the bottom of a centrifuge tube. Put the wire in the clean dry centrifuge tube that you have also weighed. Measure the combined mass to four decimal places. Go to the fume hood and add about 20 drops of **concentrated** nitric acid. Keep the centrifuge tube in the hood to prevent the brown NO_2 gas from going into the lab room. Shake occasionally and wait for the brown gas to stop being evolved. If there is some copper wire remaining and no brown gas, add a few more drops of concentrated nitric acid. When the copper wire has completely reacted, add 20 drops of deionized water. NOTE and record all the reactions and color changes in these steps!

2. While stirring with a small stirring rod that fits in the centrifuge tube, add 30 drops of 6 M NaOH a few drops at a time. Centrifuge on high for 30 seconds to 2 minutes. Test for complete precipitation by adding 2 more drops of 6 M NaOH. If you don't see any more precipitate form, then it is complete. If precipitate does form, put in 10 more drops, stir, centrifuge and test again. Your solution should be clear and colorless and the precipitate a light blue.

3. Decant the supernatant liquid from the centrifuge tube into your excess solutions beaker. Hopefully your water bath is hot by now; if not, turn the

heat up. Put your centrifuge tube containing the precipitate in the water bath and watch for it to turn black. It should all turn black before going on. Pour about 8 mL of the hot deionized water into the centrifuge tube and stir. Centrifuge and decant into the excess solutions beaker. Put your hot plate in the fume hood for step 5 (drying).

4. Add 5 to 10 drops of 6 M H_2SO_4 while stirring. Add just enough to completely dissolve the black precipitate and make a pretty blue solution. Dilute this solution with water until the centrifuge tube is between $^1/_3$ and ½ full.

5. Obtain 10 cm Mg strip from your instructor, polish it with steel wool and cut into 5 roughly equal pieces (~ 2 cm each). You must stir and wait for the solution to become colorless. Stir to break the Cu forming on the Zn or Mg and let it settle to the bottom. When the solution is colorless, add about 5 drops of 6 M H_2SO_4 while stirring to react with any remaining Mg. You must have only Cu solid in the centrifuge tube. If you are still seeing bubbles of gas form, there is still unreacted Mg, so add more acid.

   Centrifuge and decant into the excess solutions beaker. Add 2 mL of deionized water to wash the Cu, centrifuge and decant. Rinse again, centrifuge and decant. Now go to a fume hood: add 2 mL methanol, stir, centrifuge and decant. Add 1 mL acetone, stir, centrifuge and decant. Pour the methanol and acetone into a beaker and take it to the red can in the fume hood.

   Put the centrifuge tube in the hot water bath in the fume hood for one minute. Take the centrifuge tube out of the bath and dry with paper towel and Kim-wipes. Air dry the tube for a few minutes more. Obtain the mass. If it is greater than your initial mass of centrifuge tube and copper, put the tube back in the hot water bath for one minute, dry and reweigh. Repeat until the mass stays the same to within 0.01 g.

6. If possible, get a 1000-mL beaker, fill it half-full of water and transfer your excess solution to it. Stir and check the pH with pH paper. It will probably be acidic. Add 6 M NaOH to bring the pH up to the 3 – 11 range and then pour the contents of the beaker down the drain. Dump your copper out on a paper towel and put the copper metal into a waste solids beaker provided by your instructor.

Data Collection:

Mass of copper wire ________

Mass of centrifuge tube ________

Mass of tube + copper initially ________

Mass of copper wire by difference ________

Mass of tube + copper finally ________

Mass of copper finally by difference ________

Percent yield (show calc below) ________

Observations:

Step 1:

Step 2:

Step 3:

Step 4:

Step 5:

Analyses/Calculations/Results:

Determine the percent yield of copper.

RESULTS TABLE

Table 1. Results

Initial mass copper	Final mass copper	Percent Yield

Post Lab Questions:

1. Look at all six reactions for this experiment and identify the type of each (combination, decomposition, single replacement, double replacement, or combustion). Then go back and identify which are oxidation-reduction reactions. Determine the oxidation number of copper on both sides in each of these reactions.

2. A student reports a percent yield of 107%? What's wrong with this and how could it happen? (It was NOT a calculation error.)

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 07: VOLUMETRIC ANALYSIS-TITRATION OF ACID AND BASES

Objective 7

This experiment introduces the student to preparation of solutions, the concept of pH, the use of indicators, and the use of titration as an analytical tool.

Discussion

You will be dealing with mostly aqueous solutions in general college chemistry. These solutions have various designations for their concentration of solute, the most common being *Molarity*. Molarity is defined as moles of solute per liter of solution. There are many experiments that require determining the molarity of a solution. Sometimes it is to provide a *Standard solution* for use in other experiments. A standard solution is one for which the molarity has been determined with accuracy and precision, usually to four or more significant figures. Sometimes the experiment is just to determine the molarity of a solution for determining the amount of a substance that was extracted from a sample.

A scientist or laboratory technician uses *volumetric analysis* to determine the concentration of solutions. This is just a way of saying that they are using precisely measured volumes rather than masses of substances. The technician will use a technique called *titration* to conduct this volumetric analysis. Titration involves using a *titrant*, usually a standard solution, to react with an analyte, the substance about which we want to know something. The titrant is dispensed from a *burette*, a volumetric device that allows you to determine what volume of titrant was used in the reaction with good precision.

If the reaction of titrant with analyte is an acid-base reaction, we have two methods for determining when the reaction is complete. One involves the use of an *indicator*, a weak organic acid molecule, which changes color when the reaction is stoichiometrically complete. The other method uses the tracking of the pH of the solution with a pH meter and plotting pH vs. the volume of titrant added. The pH will change sharply when the reaction approaches completion, which is what makes either method work. In this experiment you will be using the indicator method.

The acidity of a solution is determined by its pH, a logarithmic number representing the concentration of the hydronium ion in aqueous acid or base solutions. The formula for pH is:

$$pH = -\log [H_3O^+]$$

where the [ ] around an ion indicates its concentration in molarity.

In this experiment, where a weak acid is reacted with a strong base, the stoichiometric point occurs at a pH of around 8. When an indicator is used, the stoichiometric point is called the endpoint of the titration. When a pH meter is used, the stoichiometric point is called the equivalence point.

You will need to read the techniques for pipetting and for use of a burette. If this is your first time to do a titration, take your time and learn the technique well.

Example

A student boils 2.005 g of clover leaves in 5.00 mL of water to extract the natural organic acid known as oxalic acid, $H_2C_2O_4$. She filters the solution the solution after it cools and transfers it to an Erlenmeyer flask. She uses a burette filled with 0.1050 M KOH to titrate the solution of oxalic acid. She finds the volume of titrant used to be 21.15 mL. Her goal is to find the percent mass of oxalic acid in clover. This is solution stoichiometry, so the first step required is the balanced chemical reaction, which is:

$$H_2C_2O_{4(aq)} + 2\ KOH_{(aq)} \rightarrow 2\ H_2O_{(l)} + K_2C_2O_{4(aq)}$$

Step 2 is to convert the volume and concentration of KOH to moles:

$$0.02115\text{ L }(0.1050\text{ mol/L}) = 0.0022207\text{ mol KOH}$$

Step 3 is to convert moles of KOH to moles of $H_2C_2O_4$:

$$0.0022207\text{ mol KOH }(1\text{ mole }H_2C_2O_4/2\text{ mol KOH}) = 0.0011103\text{ mol }H_2C_2O_4$$

Step 4 is to convert moles of $H_2C_2O_4$ to grams:

$$0.0011103\text{ mol }H_2C_2O_4\ (90.0358\text{ g/mol}) = 0.09997\text{ g }H_2C_2O_4$$

Step 5 is to determine percent mass:

$$100 * (0.09997\text{ g }H_2C_2O_4/2.005\text{ g clover}) = 4.999\ \%\text{-mass}$$

Prelaboratory Questions:

1. Define molarity, primary standard, secondary standard, endpoint, and equivalence point.

2. A 5.773-g sample of oxalic acid dihydrate was dissolved in a 250-mL volumetric flask. A 25.00-mL portion of the acid took 32.18 mL of an unknown sodium hydroxide solution. Calculate the molarity of the NaOH solution.

Safety:

Always wear appropriate PPE (see page *i*). You might want to wear gloves since you will be working with acids and bases.

Chemicals per student: ~7 g solid oxalic acid dihydrate, ~200 mL of 0.1 M NaOH, ~ 50 mL store-bought vinegar, phenolphthalein indicator, 6 M HCl and 6 M NaOH.

Materials:

Equipment per student: ring stand, burette clamp, 50-mL burette, 25 mL and 5 mL volumetric pipette, two 250 mL volumetric flasks, graduated cylinders, several Erlenmeyer flasks, small beakers, large beaker for excess solutions, test tube.

Methods:

Part A: You will be making a solution of oxalic acid dihydrate to use as the primary standard, by which you will determine the concentration of a sodium hydroxide solution, which will become your secondary standard.

1. Obtain a 250 mL volumetric flask and clean it. The final rinse should be with deionized (DI) water. Determine the approximate mass of $H_2C_2O_4 \cdot 2H_2O$ to make a 0.050 M solution of oxalic acid. (Be sure to calculate the molar mass of the hydrate including the $2H_2O$) Record your calculated value. Weigh your smallest clean dry beaker, record the mass, and then add the approximate mass of oxalic acid dihydrate that you have determined above. Weigh the beaker with the acid in it and record. Determine the mass of the acid hydrate by difference and record it.
2. Transfer the solid acid to the volumetric flask. At this point, you will learn about quantitative transfer: keeping putting about 20 mL of the DI water in the beaker and swirling it to get all the acid then pour that water into the volumetric flask. Do this several times. Then add DI water directly to the volumetric flask. As you approach the line on the neck of the flask, switch to using an disposable pipet to add the last few drops of water. DO NOT go over this line or you will have to start all over again. Cover with parafilm and mix well. Make sure all the solid acid has dissolved.
3. Using the mass from step 1, calculate the exact molarity of your primary standard using appropriate significant figures.
4. Prepare a burette for the NaOH solution per the instructions in the techniques section of the lab manual using the 0.1 M NaOH already prepared in the hood. Fill your burette to past the zero mark with the NaOH solution and fully open the stopcock to let the air bubble out of the tip. Re-fill with NaOH solution to exactly the zero mark. Now you can get out three clean beakers or Erlenmeyer flasks to use as titration containers. Condition your volumetric pipet with the oxalic acid solution, then transfer 25.00 mL of the acid solution into each of the three containers (record the actual volume dispensed). Add 2-3 drops of indicator (phenolphthalein) to each. Record the initial volume reading of the NaOH burette and tart the titration of the first container. NOTE: all volume readings should be to two decimal places!!!

5. Repeat the titration as many times as it takes to get three good titrations that remain a pale pink for at least 1 minute and that agree within 0.1 mL of each other. Confirm with your instructor that the solution is the correct color pink. The solutions at the end of titration are essentially neutral and contain naturally occurring substances, so you can pour them down the drain and flush with tap water. Show your instructor each “good” titration you think you have performed. You will need to fill the burette with NaOH for additional titrations but you do not have to prepare and condition the burette as in the beginning. Make sure to record all of your initial and final concentrations to 2 decimal places.

6. Calculate the molarity for each good titration and determine the average molarity of NaOH. Show your calculations to your instructor and get them initialed. You should have four significant figures in your answer. Then calculate the standard deviation. Report these values in your results section.

Part B: You will use your secondary standard solution to determine the concentration of acetic acid in vinegar.

1. Obtain about 15-20 mL vinegar in a small beaker. Get at least three clean containers (beakers or flasks) ready to be the titration vessels. Obtain a 5-mL volumetric pipette. Practice pipetting per the instructions in the techniques section. Then pipette 5.00 mL of the vinegar into each of the three containers. Add DI water to each the 25 mL mark of each flask then add 2-3 drops of phenolphthalein. Prep and fill a burette with your standardized NaOH solution. Get rid of air bubble and re-fill to zero mark. Record the initial volume and perform the titrations. Remember to record volume to two decimal places. If you overshoot, you will have to do another titration.

2. Calculate the molarity for each good titration and determine the average molarity. Show your calculations to your instructor and get them initialed. You should have four significant figures in your answer.

3. Vinegar is labeled to show the percent-mass of acetic acid. You are going to convert your average molarity to percent-mass. Do this and show your calculations to your instructor. Also calculate the standard deviation. Report these values in your results section. Now read the label of the store-purchased vinegar for the actual percent-mass acetic acid (or ask your instructor) and determine your percent error.

Disposal: Combine all finished titrations, any excess oxalic acid, sodium hydroxide and vinegar should go in your excess solutions beaker. Use pH paper to determine the pH and add either 6 M HCl to lower the pH or 6 M NaOH to raise the pH to the range 3 – 11. It can then go down the drain.

Data Collection:

Part A:

Mass of beaker _______

Mass of beaker + oxalic acid _______

Mass of oxalic acid _______

Titration trials:

Trial No.	Vol. Acid, init	Vol. Acid, final	Vol. Acid used	Vol. Base, init	Vol. Base, final	Vol. Base used
1						
2						
3						
4						
5						

Part B: Titration Trials

Titration table

Trial No.	Vol of vinegar	Vol. Base, init.	Vol. Base, final	Vol. Base, used
1				
2				
3				
4				
5				

Analysis/Calculations/Results:

Part A: The primary standard is oxalic acid dihydrate, $H_2C_2O_4{\cdot}2H_2O$. Convert grams to moles and then determine the molarity, using the volume of the 250-mL volumetric flask.

The reaction for determining the concentration of the secondary standard is:

$$H_2C_2O_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ H_2O_{(l)} + Na_2C_2O_{4(aq)}$$

Use the known volume of the acid and its molarity to find moles of the acid. Do the mole/mole ratio to find moles of NaOH. Divide by the titration volume to find molarity for each titration. Calculate the average molarity. Then calculate the standard deviation. Report these values in your results section.

Part B: The reaction for determining the concentration of the acetic acid in vinegar is:

$$HC_2H_3O_{2(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaC_2H_3O_{2(aq)}$$

Use the known volume of the NaOH and its molarity to find moles of acetic acid. The volume of acetic acid for each titration is what you pipetted, or 3.00 mL. Calculate the molarity for each titration and then calculate the average molarity. Then calculate the standard deviation. Report these values in your results section.

Assume you have 100.00 mL of vinegar. Determine the number of moles of acetic acid using your average molarity, and then convert that to grams of acetic acid. The density of vinegar is close to 1.006 g/mL, so convert 100.00 mL to grams. Calculate the percent mass of acetic acid in the vinegar and record it. Calculate the percent error and record it.

Part A: Determine moles and molarity of oxalic acid.

Determine molarity of NaOH solution for each trial and find the average molarity.

Part B: Determine molarity of acetic acid in vinegar for each trial and find the average molarity

Determine the percent-mass of acetic acid in vinegar from your average molarity.

RESULTS TABLE

Net percent error of burette _______

Molarity of oxalic acid _______

Average Molarity of NaOH _______ SD ______

Average Molarity of acetic acid _______ SD ______

%-mass acetic acid _______

Percent Error _______

Post Lab Questions:

1. In Part A, a student didn't use the quantitative transfer technique to get all the oxalic acid from the beaker to the volumetric flask. What effect would this have on determining the molarity of the NaOH solution? Will the student's answer be high or low compared to the actual molarity?

2. Assuming that vinegar is exactly 5.00 %-mass acetic acid, determine percent error for your result from this experiment.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 08: VOLUMETRIC ANALYSIS: ANTACIDS ANALYSIS

Objective 8

This experiment introduces the student to the use of titration as an analytical tool for commercial preparations.

Introduction:

Review of Volumetric Analysis: You will be dealing with mostly aqueous solutions in general college chemistry. These solutions have various designations for their concentration of solute, the most common being *Molarity*. Molarity is defined as moles of solute per liter of solution. There are many experiments that require determining the molarity of a solution. Sometimes it is to provide a *Standard solution* for use in other experiments. A standard solution is one for which the molarity has been determined with accuracy and precision, usually to four or more significant figures. Sometimes the experiment is just to determine the molarity of a solution for determining the amount of a substance that was extracted from a sample.

A scientist or laboratory technician uses *volumetric analysis* to determine the concentration of solutions. This is just a way of saying that they are using precisely measured volumes rather than masses of substances. The technician will use a technique called *titration* to conduct this volumetric analysis. Titration involves using a *titrant*, usually a standard solution, to react with an analyte, the substance about which we want to know something. The titrant is dispensed from a *burette*, a volumetric device that allows you to determine what volume of titrant was used in the reaction with good precision.

If the reaction of titrant with analyte is an acid-base reaction, we have two methods for determining when the reaction is complete. One involves the use of an *indicator*, a weak organic acid molecule, which changes color when the reaction is stoichiometrically complete. The other method uses the tracking of the pH of the solution with a pH meter and plotting pH vs. the volume of titrant added. The pH will change sharply when the reaction approaches completion, which is what makes either method work. In this experiment you will be using the indicator method.The acidity of a solution is determined by its pH, a logarithmic number representing the concentration of the hydronium ion in aqueous acid or base solutions.

The formula for pH is:

$$pH = -\log [H_3O^+]$$

where the [ ] around an ion indicates its concentration in molarity. The sign of the log is changed to make it a positive number.

In this experiment, where an acidified solution of an antacid is reacted with a strong base, the stoichiometric point occurs at a pH of around 4. When an indicator is used, the stoichiometric point is called the endpoint of the titration. When a pH meter is used, the stoichiometric point is called the equivalence point.

You will need to read the techniques for pipetting and for use of a burette. If this is your first time to do a titration, take your time and learn the technique well.

Antacids: Human stomachs have a gastric juice of mostly HCl that has a pH range of 1 to 2, which starts the process of digestion of food. Some people suffer from acid indigestion or heartburn and take an over-the-counter remedy called an antacid; an antacid neutralizes the stomach's acid to relieve the symptoms. We know what neutralizes an acid -- a base. An antacid can contain any one of the bases in the table below. Some antacids are also made up to create something called a buffer system in the stomach or they are manufactured with a buffer system included. A buffer system is a weak acid and its conjugate weak base, which maintains a certain pH without much change in the pH.

Chemical Name	Chemical Formula
Aluminum hydroxide	$Al(OH)_3$
Calcium carbonate	$CaCO_3$
Magnesium carbonate	$MgCO_3$
Magnesium hydroxide	$Mg(OH)_2$
Sodium hydrogen carbonate	$NaHCO_3$
Potassium hydrogen carbonate	$KHCO_3$

Any of the carbonates or hydrogen carbonates above react with the dissociated strong acid with these net ionic reactions:

$$CO_3^{2-}{}_{(aq)} + 2\ H_3O^+{}_{(aq)} \rightarrow CO_{2(g)} + 3\ H_2O_{(l)}$$

$$HCO_3^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)}$$

Any of the hydroxides react with this net ionic reaction:

$OH^{-}_{(aq)} + H_3O^{+}_{(aq)} \rightarrow 2\ H_2O_{(l)}$

Antacids with carbonates create a buffer system in their reaction with the acid in the stomach, because the reaction is a two step process:

$CO_3^{2-}{}_{(aq)} + H_3O^{+}_{(aq)} \rightarrow HCO_3^{-}{}_{(aq)} + H_2O_{(l)}$

$HCO_3^{-}{}_{(aq)} + H_3O^{+}_{(aq)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)}$

The buffer system HCO_3^{-} and CO_3^{2-} is created in step 1.

Antacid	*Chemical ingredients*
Alka-Seltzer	$NaHCO_3$, $KHCO_3$, citric acid
Di-Gel	$Al(OH)_3$, $MgCO_3$, $Mg(OH)_2$
Gaviscon	$Al(OH)_3$, $MgCO_3$
Maalox, extra strength	$Al(OH)_3$, $Mg(OH)_2$
Phillips' Milk of Magnesia	$Mg(OH)_2$
Rolaids	$NaAl(OH)_2CO_3$
Tums	$CaCO_3$

We need to push the reaction past the buffer system that is created if a carbonate is the active ingredient of the antacid(s) assigned. An excess of HCl is used to completely react the buffer or antacid; the unused HCl is then titrated with NaOH. Since both the antacid and the NaOH are bases and HCl is an acid, the endpoint will come when:

moles of antacid base + moles of NaOH = moles of HCl

To avoid confusion antacid will refer to the antacid powder mixture while antacid base refers to the base that is in the antacid powder mixture. To determine the moles of antacid base:

moles of antacid base = moles of HCl – moles NaOH

Prelaboratory Questions:

1. Why are you using bromophenol blue indicator for this experiment?

2. Why do antacids make you burp?

3. A student pipettes 50.00 mL of 0.102 M $HCl_{(aq)}$ into a 0.422 g sample of antacid. The student finds that it takes 11.29 mL of 0.0825 M $NaOH_{(aq)}$ to reach the endpoint of the titration. How many moles of antacid base are in the sample?

Safety:

Always appropriate PPE (see page *i*). You might want to wear gloves since you will be working with acids and bases.

Chemicals per student: several bottles of store-bought antacid tablets and liquids, ~ 50 mL of ~0.1 M HCl, bromophenol blue indicator, ~100 mL of ~0.1 M NaOH solution, 6 M HCl and 6 M NaOH as needed.

Equipment per student: ring stand, burette clamp, a 50-mL burette, a 25-mL volumetric pipette, several Erlenmeyer flasks, small beakers, large beaker for excess solutions, test tube, mortar and pestle.

Methods:

You should work alone if there are enough work stations and chemicals. Pick an antacid and do three trials. First get your burette and prepare it for titration. Get about 125 mL of ~0.1 M NaOH, and record the exact molarity. Rinse the burette with 5-mL portions of this NaOH. Fill the burette up to the top with the NaOH solution and drain it through the tip to push out the air bubble. Then get a hot plate and set it on low to start with.

1. For your 3 trials obtain about 0.6 grams (0.2 g per trial) of the antacid, either liquid or tablet. If you chose a tablet, use a mortar and pestle to grind the tablet into powder. Weigh three 250-mL Erlenmeyer flasks, then add 0.2 g of antacid to each, then reweigh them. Record all the masses. Obtain a 25-mL pipette and use it to add 25.00 mL of ~0.1 M $HCl_{(aq)}$ to each flask. Read and record the actual molarity of the HCl solution. Mix the acid and antacid until the antacid is completely dissolved. You may need to heat the flasks on a hot plate to boiling and let boil for about one minute to help dissolve the solids.

2. Let the solutions in the flasks cool. Then add about 5 drops of bromophenol blue indicator. If the solutions turn blue, there is more antacid per tablet and you will need to add more HCl. Obtain a 5-mL pipette and 5-mL of the HCl to the flask(s) that are blue. Keep doing this until they are no longer blue. If you had to add HCl, keep track of the volume and record it.

3. Have the instructor check your burette set-up. Record the initial volume and start titrating the first antacid sample. Watch for the blue endpoint and go slowly. If the blue stays for 30 seconds, you are done with that sample. Record the volume and then titrate the second sample

Disposal: titrated solutions should have a pH of around 4 and can be poured directly down the drain. Any excess sodium hydroxide and hydrochloric acid should go in your excess solutions beaker. Use pH paper to determine the pH and add either 6 M HCl to lower the pH or 6 M NaOH to raise the pH to the range 3 – 11. It can then go down the drain. Leftover antacid can be washed down the drain.

Data Collection:

	Trial 1	**Trial 2**	**Trial 3**
Mass of beaker			
Mass of beaker+ antacid			
Mass of antacid			
Volume of NaOH used			

Volume of hydrochloric acid pipetted in each trial	
Molarity of HCl	
Molarity of NaOH	
Cost of antacid	
Mass of antacid preparation	

Instructor initials _______

Analyses/Results/Calculations:

The volume of NaOH used per sample along with its molarity (given) allow you to determine moles of NaOH. Use the total volume and molarity of HCl (given) to determine the moles of HCl in the sample. Subtract the moles NaOH from moles HCl to find the moles of antacid base. Divide the moles of antacid base by the mass of the antacid used in the trial. Repeat for the second and third trials and then find the average value for moles of antacid base/gram of antacid.

To compare costs of the antacids: determine the cost per gram of antacid by reading the cost and label on the bottle of antacid. Invert your answer to find the grams of antacid per penny. Multiply the average moles of antacid base/gram of antacid by the grams of antacid/penny to find the moles of antacid/penny. Compare results to other groups, the antacid with the higher number is the most cost-effective antacid.

ID of Antacid

	Trial 1	Trial 2	Trial 3
Moles of HCl			
Moles of NaOH			
Moles of Antacid Base			
Mole antacid base/g antacid			
Average moles antacid base/g antacid			
g antacid/penny			
Moles antacid base/penny			

Most cost-effective antacid _________________________________

Post-Lab Questions:

1. Why did you have to boil the antacid samples after you added the HCl?

2. Write the chemical reactions between the bases in Di-Gel and stomach acid, HCl.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 09: DETERMINATION OF WATER HARDNESS

Objective 9

You will determine the hardness of tap water from the lab sink as well as from other areas of Solano County or elsewhere.

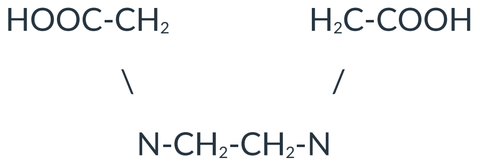
Discussion:

Water hardness has to do with the quantity of calcium and/or magnesium ions in the water. If you have already done the soap-making experiment, you learned that these ions precipitate the soap anions and create bathtub rings, etc.

The calcium and magnesium ions actually derive from the calcium carbonate or magnesium carbonate in the underground sources of our water. These ionic compounds have a very limited solubility in water, but it doesn't take much of either cation to create the hardness effect. There is usually more calcium ion than magnesium ion, so laboratories report water hardness as parts per million in mass of calcium carbonate in water. (There are other ions that contribute to water hardness, but these two are the major sources.)

You will be using a titration method to determine the amount of calcium and magnesium ions in water, but it is a reaction that involves the formation of a complex ion of the metal cations with EDTA. This ligand, ethylenediaminetetracetato ion, has a negative four charge. It attaches to a metal ion at six places on its structure where the ligand has lone pairs of electrons. The polyatomic ion actually folds around the central metal ion. EDTA is also called a chelating agent.

The reagent is the weak acid form of the ligand, ethylenediaminetetracetic acid, which looks like four acetic acid molecules attached to the two "end nitrogen atoms" of ethylenediamine.



$$\begin{array}{ccc} HOOC\text{-}CH_2 & & H_2C\text{-}COOH \\ \backslash & & / \\ & N\text{-}CH_2\text{-}CH_2\text{-}N & \end{array}$$

/ \

HOOC-CH_2 H_2C-COOH

The two nitrogen atoms each have one lone pair of electrons that can make a coordinate covalent bond with a central metal ion to form a complex ion. As the acidic hydrogens are removed from the –COOH groups, the singly-bonded oxygen is left with three lone pairs, one of which can make a coordinate covalent bond with a central metal ion. That makes the six sites at which EDTA can attach to a central calcium or magnesium ion.

The summary reaction will be:

$$Ca^{2+}_{(aq)} + EDTA^{4-}_{(aq)} \rightarrow [Ca(EDTA)]^{2-}_{(aq)}$$

You can standardize the EDTA solution using dry solid calcium carbonate or standard solution of calcium carbonate. The indicator will be a metallochromic one called Eriochrome Black T, which reacts with magnesium ions to form a dark red complex ion. As the EDTA reacts to form complex ions of the calcium and magnesium ions, it will pull the magnesium ions from the indicator. The indicator will then turn light blue. That is the end point color that you will be watching for.

If the water solution doesn't turn dark red with the addition of the indicator, you will be instructed to add a very small amount of magnesium ion.

The acid form of EDTA is referred to generically as H_4Y. It's not very soluble in water, and we want the anion form as a chelating agent, so the disodium salt of the acid is used to make the EDTA solution. The EDTA or Y^{4-} appears at about pH 10, so the titration will take place in a pH 10 buffered solution containing NH_3/NH_4^+.

Prelaboratory Questions:

1. What color is the indicator prior to the titration for water hardness? What color change will you be looking for?

2. If a 25.00-mL aliquot of 0.00998 M $CaCO_3$ solution takes 22.45 mL of EDTA to reach the endpoint, then what is the concentration of the EDTA solution?

3. If a 100.00-mL aliquot of tap water from the laboratory takes 32.35 mL of the above EDTA solution to reach the endpoint, then determine the water's hardness in units of parts-per-million by mass of $CaCO_3$.

Safety:

Wear appropriate PPE (see page *i*).

Materials:

Chemicals per student: ~100 mL of standardized 0.01 M calcium carbonate solution, tap water, ~5 mL Eriochrome Black T indicator, ~250 mL 0.01 M EDTA solution, ~100 mL pH 10 NH_3/NH_4Cl buffer solution (6 M NH_3 that is also 1.1 M NH_4Cl), ~150 mL of boiled deionized water, and possible some 0.30 M $MgCl_2$.

Equipment per student: burette, burette clamp, ring stand, 25-mL pipette, 50-mL pipette, pipette bulb, flasks, 10-mL graduated cylinder.

Methods:

1. Obtain about 250 mL of the ~0.01 M EDTA solution. Obtain about 100 mL of the standardized ~0.01 M calcium carbonate solution and record its actual molarity from the reagent bottle. Prepare a 50-mL burette for the EDTA solution. Prepare a 25-mL pipette for the standardized calcium carbonate solution. Have three clean 125- or 250-mL Erlenmeyer flasks ready. Set up for your titration and obtain your instructor's approval. Pipette 25-mL aliquots into each of your flasks. Add ~10 mL of the pH 10 buffer (NH_3/NH_4^+) to each flask using a small graduated cylinder. Add about 50 mL of boiled deionized water to each flask. Then add 4 to 5 drops of the Eriochrome Black T indicator

to each flask. If the solution does not turn dark red, then add 10 to 15 drops of 0.030 M $MgCl_2$, so there are Mg^{2+} ions to make the colored ion with the indicator.

2. Record the initial and final volumes for each of your three titrations, estimating the second decimal place when reading the burette. Remember you are looking for the solution to turn from dark red to pale blue.
3. Rinse the flasks out into the sink drain and clean with deionized water.
4. Calculate the molarity of the EDTA solution for each of your three titrations. The mole to mole ratio of Ca^{2+} to EDTA is one to one. Use the molarity of the standardized calcium carbonate solution to determine moles of calcium ion in each 25.00 mL aliquot. The moles of EDTA will be the same and can be divided by the volume of EDTA in each trial. Determine the molarity of EDTA for each trial. Use your three data points to find the average molarity, standard deviation and 90% confidence limit. Obtain your instructor's approval before continuing and using the EDTA as a secondary standard solution.
5. Set up three flasks again and refill your burette. Pipette 50.00 mL of tap water or your assigned water into each of the three flasks. Again add ~10 mL of the pH 10 buffer to each. Add 5 or more drops of the Eriochrome Black T indicator, until you see the dark red color. If the solution does not turn dark red, then add 10 to 15 drops of 0.030 M $MgCl_2$, so there are Mg^{2+} ions to make the colored ion with the indicator.
6. Titrate the three samples to the pale blue endpoint and record all the volumes.
7. Rinse the flasks out into the sink drain and clean with deionized water.

Data Collection:

DATA TABLE AND OBSERVATIONS

EDTA Analysis:

Volume of standardized $CaCO_3$ in each trial _______

Molarity of standardized $CaCO_3$ _______

Titration trials:

Trial No.	Vol. EDTA, init.	Vol. EDTA, final	Vol. EDTA, used
1			
2			
3			
4			
5			

Water Hardness:

Water sample name ____

Volume of water pipetted in each trial _______

Trial No.	Vol. EDTA, init.	Vol. EDTA, final	Vol. EDTA, used
1			
2			
3			
4			
5			

Analysis/Calculations/Results:

Water Hardness: the stoichiometry is a 1:1 mole ratio. Determine the moles of EDTA from its concentration (from above) and the volume titrated. Set that equal to the moles of calcium ion. Use the volume of water, 50 mL, to determine the molarity. Since it is water and the density is 1 g/mL, ppm-mass of $CaCO_3$ will be the same as milligrams of $CaCO_3$ per liter. Multiply the molarity by the molar mass of $CaCO_3$ and convert mass to milligrams.

RESULTS TABLE (Create your own)

Post Lab Questions:

1. Draw the structure of the ligand, ethylenediaminetetraacetato ion, showing all the lone pairs of electrons, around its central metal ion, Ca^{2+}.
2. Explain how ion-exchange resins work to give people “soft” water in their homes. (Look in your textbook or search the Internet.)

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 10: PREPARATION AND PROPERTIES OF ASPIRIN

Objective 10

This experiment introduces the student to simple organic synthesis, purification, yield, limiting reactant, etc. It will take about two laboratory periods to synthesize and dry the aspirin, etc.

Introduction:

The active ingredient of aspirin, salicylic acid, is actually a 10,000-year old remedy for fever, aches and pains. However, more than 100 years ago in 1899, the Bayer Company of Germany developed aspirin, an ester of the acid, because it would not be so hard on the stomach!

The synthesis of aspirin is classified as an esterification reaction in organic chemistry. This is a substitution reaction. An ester is the product of a reaction between an organic acid (acetic acid from acetic anhydride) and an alcohol (the –OH group on the salicylic acid). Before you jump to conclusions, in the commercial production of aspirin, the salicylic acid **is** the alcohol. Acetic acid is the acid part of the ester.

The method of preparation for your experiment uses the anhydride of acetic acid, which will react like acetic acid does. The reaction is shown below.

$$C_6H_4(OH)(COOH)_{(s)} + CH_3COOOCCH_{3(l)} \rightarrow C_6H_4(OOCCH_3)(COOH)_{(s)} + CH_3OH_{(l)}$$

Salicylic acid Acetic Anhydride Aspirin Methanol

Most students are familiar with esters without knowing what they are! They are present in our everyday life. Methyl salicylate, for example, is oil of wintergreen. Isopentenyl acetate is oil of banana or pear, while butyl butanoate smells like pineapple. The most common ester used as a medicine is aspirin (acetylsalicylic acid). There are other drugs that are also esters.

To experience other forms of esters, you will also be making oil of wintergreen, or methyl salicylate, by reacting salicylic acid with methanol. The distinctive odor of wintergreen lets you know your esterification reaction is successful, where the synthesis of aspirin doesn't give that immediate feedback.

You will also do one test to see if any salicylic acid remains after you have made your aspirin. A 1% $FeCl_3$ solution turns a dark purple color if the –OH group is still available on the salicylic acid. You'll compare your aspirin, pure salicylic acid and commercial aspirin to see what colors appear. If your initial salicylic acid reacted completely to produce aspirin, the phenolic group has been replaced with the acetate (acetyl) group, so the iron test should be negative. Of course, students will not get 100% yield, so some salicylic acid remains to make their sample purple in the iron test. You should expect that the pure salicylic acid should turn dark purple; the student-prepared sample should be somewhat purple (due to unreacted salicylic acid), and the commercial aspirin should be a pale orange color (no salicylic acid present).

Example:

A student used 2.022 g of salicylic acid and 3.517 g of acetic anhydride to synthesize her aspirin. At the end of the experiment she had 1.599 g of aspirin after correcting for water used for rinsing. Which reactant is limiting? What is the theoretical yield? What is the percent yield?

2.022 g (1 mole/138.1g)(1mole aspirin/1 mole salicylic acid) = 0.01464 mole aspirin
****THEREFORE SALICYLIC ACID IS THE LIMITING REACTANT.

3.517 g (1 mole/102.1 g)(1mole aspirin/1 mole acetic anhydride) = 0.03445 mole aspirin

Theoretical yield is based on the limiting reactant: 0.01464 mol aspirin (180.2 g/mol) = 2.638 g

Percent yield = 100 * (1.599g/2.638g) = 60.61% yield

Prelaboratory Questions:

1. What is an ester? How is an ester synthesized?
2. What is an organic acid? What structural component identifies them as an acid?
3. Why does your synthesized aspirin have to be recrystallized?
4. A student obtained 1.200 g of aspirin from 1.500 g of salicylic acid and excess acetic anhydride. What is her percent yield?

Safety:

Always wear appropriate PPE (see page *i*). You might want to wear gloves since you will be using an acid and a base.

Materials:

Chemicals: salicylic acid, acetic anhydride, concentrated sulfuric acid, ethanol, 1% iron (III) chloride solution, methanol, deionized water, ice, commercial aspirin tablets.

Equipment: large test tube or combustion tube, several test tubes, 125 mL Erlenmeyer flask, stirring rod, hot plate, small beaker for ice water, large beaker for hot water bath, large beaker for ice bath, test tube, vacuum flask and Buchner funnel, #5 filter paper to fit Buchner funnel, thermometer.

Methods:

Part A: Preparation of methyl salicylate

1. Put about 1 gram of salicylic acid in a large test tube and add about 5 mL of methanol. Go to the concentrated acids hood and add about 3-4 drops of concentrated sulfuric acid as a catalyst.

2. Heat the test tube in a hot water bath at 60°C for about 30 minutes (while you are setting up the aspirin synthesis).

3. Cool the test tube to room temperature with a "scrunchie" of paper towel in the opening. You should see a white solid form. Record its appearance and its odor, just as you remove the paper towel from the test tube.

4. Add one drop of the 1% $FeCl_3$ solution and note the color change.

Part B: Preparation of aspirin

1. Put about 3 grams of salicylic acid in a preweighed 125 mL Erlenmeyer flask. Obtain the exact mass and record. Determine the mass of salicylic acid. Add about 6 mL of acetic anhydride (use caution because the vapor can be very irritating to your eyes) and weigh the flask again. Record the mass. Determine the mass of acetic anhydride. Now go to the concentrated acid hood and add about 8 drops concentrated sulfuric acid SLOWLY. Swirl the flask for several minutes. Then put the flask in a beaker of boiling water for at least 20 minutes, and make sure to swirl or stir every few minutes. While this is going on, put ice and deionized water in your small beaker so you have ice water for rinsing later. Also put ice and tap water in a large beaker as an ice bath for your flask later.

2. After the 20 minutes has passed in step 1, take the flask out of the boiling water bath and add about 15 mL of the ice-cold water. (Record the volume of water used.) The water cools the aspirin and decomposes any excess acetic anhydride to acetic acid. Put the flask in the large beaker of ice and water and stir it every few minutes. (If you're not getting any crystals of aspirin and oil appears instead, put the flask back in the boiling water bath for another 20 minutes.) While your aspirin is cooling, set up a Buchner funnel for vacuum filtration of the aspirin. Put the filter paper in the funnel, turn on the water, and wet the paper with deionized water from a wash bottle.

3. If there is any liquid in your 125-mL beaker, pour it out slowly. Then pull the rest of the slurry out with your stirring rod and dribble it in the center of the filter paper. Put another 15 mL of ice-cold water in your flask and use it to

wash out the remaining aspirin crystals. Repeat this step, using the water to also wash the crystals on the filter paper. Remember to record the volume of water used. (If aspirin makes it through the filter paper, collect the filtrate in a beaker and put it in the ice bath. Vacuum filter again.)

4. Your aspirin is NOT pure at this point. It will need to be recrystallized by dissolving it in ethanol at its boiling point, and then letting it cool and recrystallize. Carefully take the filter out of the funnel and scrape the aspirin into a clean dry 125-mL Erlenmeyer flask. Also scrape out any crystals that are sticking to the funnel. Put the flask in a 60°C water bath on a hot plate for this next step. Add absolute ethanol a few mL at a time with stirring until all the crystals have dissolved. Do not go beyond 20 mL of ethanol and try to stay well below that. (If you make a mistake at this point, you will have to evaporate some of the ethanol off the solution in a fume hood.) Give it a couple of minutes, remove from the hot water bath and add 50 mL of 50°C water to the flask. If any crystals form now, put the flask back into the hot water bath until they dissolve. Remove the flask from the hot water bath. Now cover the flask with a small watch glass and let the solution cool to room temperature.

5. Weigh a second piece of filter paper and record the mass. Make sure you have a watch glass large enough to hold this filter paper. Use the Buchner funnel to filter your recrystallized aspirin. Pour 10 mL of ice-cold water in the flask and swirl. Pour it out around the solid on the filter to wash the crystals. Repeat this step. Let air run over the solid for a few minutes. Transfer the filter paper and all crystals of aspirin to your watch glass and set it aside for the next lab period.

NEXT LAB PERIOD: Weigh the filter paper and aspirin and record the mass. Determine the mass of aspirin formed.

6. Set up three small test tubes. Put a few grains of your aspirin in one, a few grains of your salicylic acid in the second and a few grains of ground commercial aspirin in the third. Add a couple mL of deionized water and dissolve the crystals in each test tube. Now add 1 drop of $FeCl_3$ solution to each. Let the color develop for a few minutes. Record the colors and comment on your aspirin compared to commercial aspirin.

Disposal: put your aspirin and methyl salicylate in water to dissolve and pour down the drain. Pour any extra ethanol down the drain and rinse down with water.

Data Collection:

Part A: Make observations about each reactant and the product.

Color change with $FeCl_3$:

Part B:

Mass of 125 mL flask ________

Mass of flask + salicylic acid ________

Mass of salicylic acid ________

Mass of flask, acid and anhydride ________

Mass of anhydride ________

Water used: (keep a running log)

Mass of filter paper ________

Mass of filter paper + aspirin ________

Mass of aspirin ________

Table 1. Tracking Color Change

Sample	**My aspirin**	**Salicylic acid**	**Commercial aspirin**
Color change			

Analyses/Calculations/Results:

Part B: Determine the theoretical yield and percent yield of your aspirin. When determining actual yield of your aspirin, correct it by adding 0.25 g per 100 mL of water used. Hopefully you kept track of the water as you used it for rinsing!

Calculate limiting reactant, actual yield, corrected actual yield, theoretical yield, and percent yield.

Table 2. Results

Mass Limiting Reactant	Mass Aspirin, corrected	Theoretical Yield	Percent Yield

Post Lab Questions:

1. A student reported a 103.5% yield. What's wrong with this? What should the student have done to correct for this?

2. Review the form and structure of an ester. If an ester has the condensed formula of $C_3H_6O_2$, what is its structure and what acid and alcohol were involved in its synthesis?

3. Find the organic chemistry structures for acetaminophen, ibuprofen and aspirin. Draw them. Do these compounds have similarities? What are the differences?

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 11: IDEAL GAS LAW

Objective 11

In this laboratory experiment, you will react magnesium with hydrochloric acid to produce hydrogen gas and magnesium chloride, as the balanced equation shows below. The gas is collected in a eudiometer (gas collecting tube) and its volume, temperature and pressure are used to learn about the gas laws.

$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow H_{2(g)} + MgCl_{2(aq)}$

Introduction:

There are several gas laws relating quantity of a gas, temperature, pressure, and/or volume. In this experiment, you will be doing several conversions, and applying Dalton's Law of Partial Pressures ($P_T = P_A + P_B + \ldots$), the Ideal Gas Law ($PV = nRT$) and the practice of stoichiometry. Students should also learn how to read the barometer and record atmospheric pressure.

When a gas is collected over water or through water by displacement, the gas becomes saturated with water vapor, and the pressure of the water vapor can be found in a table. The total pressure of the gases inside the eudiometer will be the sum of the collected gas pressure, such as hydrogen, and the pressure due to water vapor. To find the collected gas pressure, you must know the total pressure, which will usually be the same as atmospheric pressure, and you must look up the vapor pressure of water. A sample calculation is shown in the example below.

Examples:

A student reacts a strip of aluminum metal with nitric acid as the balanced equation shows below. The volume of gas collected is 94.5 mL at 23.0°C and 756.0 torr with a liquid height of 21.0 cm in the eudiometer. What is the mass of the aluminum metal?

$$2\ Al_{(s)} + 6\ HNO_{3(aq)} \rightarrow 3\ H_{2(g)} + 2\ Al(NO_3)_{3(aq)}$$

Step 1: Find the equivalent head pressure in mm Hg for the liquid in the column:

21.0 cm (10 mm/1cm)(1.05/13.6) = 16.2 mm Hg equivalent or torr

Look up the vapor pressure of water at 23.0°C in the table (it is 21.08 torr).

Use Dalton's Law of Partial Pressures:

756.0 torr = P_{H2} + 21.08 torr + 16.2 torr, therefore P_{H2} = 718.7 torr

Step 2: Use the Ideal Gas Law (PV=nRT) to find moles of Hydrogen, after converting volume to liters, pressure to atmospheres and temperature to Kelvin.

94.5 mL (1L/1000mL) = 0.0945 L

718.7 torr (1 atm/760 torr) = 0.94566 atm

23.0°C + 273.15 =296.15 K

Rearrange to n = PV/RT = (0.0945 L*0.94566 atm)/((0.082057 L-atm/mol-K)*296.15 K) = 3.68×10^{-3} moles of hydrogen

Step 3: Use stoichiometry to get moles and mass of aluminum:

3.68×10^{-3} moles H_2 (2 mol Al/3 mol H_2) (26.9815g/mol Al) = 0.0662 g Al

Prelaboratory Questions:

1. Write definitions for each of these: Dalton's Law of Partial Pressures, STP, Ideal Gas Law
2. How will you know when the magnesium ribbon has finished reacting?
3. A sample of H_2 gas is collected over water at 21.5°C and atmospheric pressure is 765 torr. The height of liquid in the eudiometer is 13.0 cm. What is the partial pressure of the H_2 gas?
4. List the safety cautions that you will observe in this experiment.

Safety:

Always wear appropriate PPE (see page *i*). Gloves should be worn during set-up and clean up of this experiment as you are dealing with an acid.

Materials:

Chemicals per student: ~10 mL 6 M HCl, 3-4 cm magnesium ribbon, ~10 inch thin copper wire to make the cage, 100 mL of deionized water, 1 L tap water, solid $NaHCO_3$ and/or 6 M NaOH to neutralize acid.

Equipment, general: ring stands, burette clamps, meter sticks or rulers, barometer, steel wool, scissors.

Equipment per student: eudiometer, large beaker (600 to 1000 mL), 1-hole rubber stopper to fit eudiometer, 110°C red alcohol thermometer, and graduated cylinder.

Methods:

1. Use a meter stick to obtain the exact length of a piece of magnesium ribbon assigned by the instructor, usually 3 to 4 cm. Record the length with at least 3 significant figures. Clean the Mg ribbon with steel wool and wipe off with Kim Wipes.
2. Fill your beaker about $^1/_2$ full with tap water. Coil the magnesium ribbon and wrap it with a 10-inch length of copper wire to make a “basket” around it, leaving a “tail” 2 to 3 inches long. This tail will go through a hole in a one-hole rubber stopper, be pinched around the wider end, and bent at an angle as demonstrated by the instructor.
3. Carefully pour about 10 mL of 6 M HCl (or 20 mL of 3 M HCl) into a clean eudiometer (expensive gas-collecting tube) at the sink at your station. At a 45° angle, slowly fill it the rest of the way with deionized water. Try to prevent any mixing of acid and water. SEE DEMONSTRATION BY INSTRUCTOR. Fill the tube until water is over the top, then insert the rubber stopper with the three inch tail of copper wire holding the Mg metal in the Cu basket. Let water come through the hole in the stopper. Cover hole with your finger, quickly invert the tube, and get the opening under the water level in your beaker. (The acid moves down quickly because it is a higher density, and the reaction with Mg starts immediately.) Note in your data and observations if there were any air bubbles left in the tube, as they will affect your volume reading and calculations.

4. Let the Mg react completely. If the small piece of Mg sticks to side, drops, etc., notify instructor. Cool for at least 20 minutes after reaction is complete. During this cooling time, record room temperature from a mercury thermometer and read/record barometric pressure from the barometer in the lab room. Toward the end of the cooling period, TAP LIGHTLY with a wooden pencil on the side of the tube to loosen all the gas bubbles. After the gas has cooled, read and record gas volume to **two** decimal places. Measure height of liquid remaining in the tube with a meter stick, from surface of beaker water to meniscus of tube liquid. Convert to centimeters to millimeters and record.

Disposal: Your large beaker now contains the dissolved salt $MgCl_2$ and some excess HCl. Slowly add sodium bicarbonate to the solution while stirring until the solution doesn't bubble very much, and then pour down the sink. Alternately, you can obtain pH paper from your instructor and add 6 M NaOH solution until the pH range is between 3 and 11, and then pour the contents of the beaker down the sink.

Data Collection:

Length of magnesium ribbon, cm ____________

Mass of magnesium ribbon, g ____________

Volume of Hydrogen gas, mL ____________

Temperature of gas, °C ____________

Atmospheric pressure, torr ____________

Height of solution in tube, mm ____________

Vapor pressure of water, torr ____________

Analyses/Calculations/Results:

1. Determine the partial pressure of hydrogen gas in the mixture of gases in the eudiometer: both $H_{2(g)}$ and $H_2O_{(g)}$ are present because the gas was collected through an aqueous solution.

$P_{total} = P_{barometric} = P_{H2} + P_{H2O} + P_{liquid\ column}$

P_{H2O} is vapor pressure of water at room temperature, found in a table in your lab manual.

$P_{\text{liquid column}}$ must be converted to mm Hg using the density ratio 1.05/13.6. Multiply this ratio by the liquid height in millimeters.

Now solve for the partial pressure of H_2. ($P_{H2} = P_{\text{barometric}} - (P_{H2O} + P_{\text{liquid column}})$)

3. Use the Ideal Gas Law to solve for moles of Mg and its mass, by solving first for moles of H2.
4. Compare this to the true mass of the Mg ribbon. Calculate your percentage error:

$$\frac{|\text{True value - calculated value}|}{\text{True value}} \times 100 = \text{Percent Error}$$

1. Partial pressure of Hydrogen gas:

2. Mass of Mg ribbon:

3. True mass and percent error:

RESULTS TABLE

Mass of Mg ribbon	True Mass	Percent Error

Post Lab Questions:

1. You are given an envelope containing a piece of Magnesium and a piece of Zinc with a total mass of 0.0833 grams. The volume of gas collected at 25°C with P_{atm} at 755 torr is 59.74 mL. Liquid column height is only 15 mm. Calculate the mass of each piece of metal in the envelope, showing all work. (Hint: two equations, two unknowns.)
2. Show the relationship between the Ideal Gas Constant, R, and the Standard Molar Volume of a gas.

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 12: MOLAR MASS OF A VOLATILE LIQUID – DUMAS METHOD

Objective 12

To learn about the gas laws and to determine the molar mass of a volatile liquid.

Introduction:

A volatile liquid is one with a low boiling point and high vapor pressure, which can be easily vaporized. Once in the vapor phase, we can apply the gas laws and determine much about the compound. You will be using the Ideal Gas Law, PV = nRT, where R is the Ideal Gas Constant with a value of 0.082057 L * atm/mol * K. Pressure must be in atmospheres, volume in liters, and temperature in Kelvin. The Dumas method is not the most accurate method for determining molar mass, but it is easily done. If you only need an approximation of molar mass, this method is sufficient. For example, if you already know the empirical formula of an unknown organic substance, finding its approximate molar mass will lead you to its molecular formula.

Example:

A student used this procedure to find the molar mass of an unknown volatile liquid. She found the volume of her Erlenmeyer flask to be 229.5 mL, the mass of her unknown was 0.578 grams, it completely vaporized at 41.5^{o}C and the barometric pressure was 764.5 torr.

Rearrange the Ideal Gas Law to solve for moles: n = PV/RT.

Thus, n = (1.006 atm*0.2295 L)/(0.082057*314.65 K) = 8.942x10^{-3} moles

Molar mass is mass/moles = 0.578 g/8.942x10^{-3} moles = 64.6 g.

Prelaboratory Questions:

1. Derive the four simple gas laws from the Ideal Gas Law.

2. What safety precautions would you tell the class about prior to conducting this experiment?

3. A student used combustion analysis to determine that a volatile organic compound was 52.14% carbon, 13.13% hydrogen and 34.73% oxygen. She used the Dumas method using 0.185 g of the compound and collected data at 90.5^{o}C, 764.5 torr, with a flask volume of 121.0 mL. Determine the molar mass and the molecular formula of this compound.

Safety:

Always wear appropriate PPE (see page *i*). Gloves are optional since the organic alcohol unknowns have been selected for their very low toxicity. However, they are volatile, so keep away from open flames and work in the fume hoods.

Materials:

Chemicals: possible unknowns for your experiment are methanol, ethanol and 2-propanol; dropper bottles of 6 M HCl.

Equipment per student: 125 mL Erlenmeyer flask, hot plate, thermometer (calibrated), ring stand and utility clamp, aluminum foil, thin copper wire suitable for wrapping. For class: small needles, barometer, a few 200 mL graduated cylinders.

Methods:

Your instructor may ask you to work in pairs if fume hood space is limited. You may also be required to calibrate your thermometer. (See earlier experiment.)

1. Cut a large enough aluminum foil square to extend about 1 inch around the top of your 125-mL Erlenmeyer flask. Trim it into a circle. Also cut about 20 cm of thin soft copper wire to use to hold it in place. Make sure you have a clean and dry flask. DO NOT WASH A FLASK FOR USE TODAY! Just choose your cleanest flask. Weigh this flask and record the mass to at least three decimal places. Set up a hot plate in a fume hood and put a 600-mL beaker about 2/3 full of tap water on **low** so it starts to heat up. Put some boiling chips and 3 or 4 drops of 6 M HCl in the beaker of water prior to heating. Have a ring stand set up by the hot plate to hold your flask and a thermometer. Also have a large cold water beaker standing nearby to cool your flask later.

2. Put the aluminum foil on the cardboard of your lab manual or notebook and use a small needle to make a tiny hole in the center of the foil. Just barely touch the foil with the tip of the needle! Now put the flask, foil and copper wire on the balance and record the mass of these to at least three decimal places.

3. Get your unknown volatile liquid from your instructor. It should be about 2 mL of liquid in a sealed test tube or small vial. Record the unknown number. DO NOT OPEN THE SAMPLE UNTIL YOU ARE READY TO USE IT.

4. Put the all of the ~2-mL sample into your flask and seal it with the foil. Wrap the copper wire tightly around the neck a few times and twist the ends tightly. (Do NOT weigh the unknown yet.) Close the utility clamp around the neck of the flask and attach it to the ring stand. Slowly lower the flask into the beaker of water so that it is immersed as much as possible. You may have to adjust the water level by adding or removing some water. Hang your thermometer in the beaker of water so it is measuring the water temperature and not touching anything else. Turn the hot plate to medium Do not rush the evaporation or you will miss the boiling point of the unknown.

5. Watch the liquid in the flask. When it starts to boil, record the temperature of the water. This will be the approximate boiling point of your unknown liquid. (The water will continue to get hotter itself, so you have to catch the liquid when it just starts to boil.) At some point during the experiment, read the barometer and record the atmospheric pressure.

6. Continue to watch the flask closely for about 20 minutes. As soon as all the liquid has vaporized, and no more is condensing in the neck of the flask where

it is cooler, remove the flask using the utility clamp, and set it in the cold water beaker. Just before doing this, record the temperature of the water, since the gas will be at that temperature. When your flask is cool (about five minutes), remove it from the cold water beaker, take off the utility clamp and wipe it dry with paper towels. Also carefully lift the foil and dry the foil and flask. Give it all a few minutes to air dry after this. Then weigh the flask, foil, wire, and unknown. Record the mass to at least three decimal places. Determine the mass of the liquid by difference.

7. Pour the unknown liquid down the drain, rinse the flask and then fill it to the very brim with water. Pour the water into a large graduated cylinder to determine the actual volume of the flask and record to at least one decimal place. If you are working with a partner, have your partner repeat this procedure so you can average the two readings for volume. (An alternate method is to fill the flask, weigh it, and convert mass to volume by using the density at the temperature of the water.)

Data Collection:

Unknown Number ________

Mass of 125 mL flask ________

Mass of flask, foil, wire ________

After evaporating: ________

Mass of flask, foil, wire, unknown ________

Mass of unknown ________

Temperature at which unknown started to boil ________

Temperature at which unknown completely vaporized ________

Atmospheric pressure ________

Volume of flask - first time ________

Volume of flask – second time ________

Average volume of flask ________

Analyses/Calculations/Results:

Calculate the molar mass of your unknown using the Ideal Gas Law. State the molar mass of unknown, unknown number, your guess as to its identity.

Post Lab Questions:

1. If 2 mL of your unknown were not enough to fill the flask with the vapor of the unknown, how would this affect your calculation of its molar mass? (Hint: what else would be present?)

2. If the flask isn't immersed in the water while vaporizing the volatile liquid, what effect will that have on the calculation of molar mass?

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 13: THERMOCHEMISTRY

Objective 13

You will determine the enthalpy of formation for MgO using calorimetry and Hess' Law.

1. $Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2(g)}$
2. $MgO_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)}$

You will learn to use statistics to evaluate your data as a class and as individuals.

Introduction:

Thermochemistry is the study of energy exchanged between a chemical reaction system and its surroundings. When energy transfers from the system to the surroundings, the reaction is exothermic. When energy transfers from the surroundings to the system, the reaction is endothermic. The amount of energy transferred at constant pressure is called the *enthalpy of reaction*, ΔH_{rxn}. There are many types of reactions, so the subscript can change to the type of reaction being studied, such as the formation of a compound from its elements or ΔH_f. For a combustion reaction we use ΔH_{comb}. Thermochemical equations include the enthalpy of reaction and the balanced chemical equation. See the examples below. Formation reactions are written for the formation of one mole of a compound from its elements.

$\frac{1}{2} N_{2\,(g)} + \frac{1}{2} O_{2\,(g)} \rightarrow NO_{(g)}$ $\Delta H_f = +181.8$ kJ

$CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(l)}$ $\Delta H_{comb} = -890.$ kJ

Calorimetry is the method used to measure the energy transferred by a chemical or physical process. It's called that because we used to use calories as the energy unit, although now we typically use the metric unit for energy, the Joule. The conversion factor is 1.000 cal = 4.184 Joules. In calorimetry, the calorimeter becomes the surroundings and the reaction system under study is the system. A thermometer is used to measure the change in temperature. A calorimeter is an insulated container that minimizes the heat transferred to the laboratory bench and room. A heat transfer calculation is used to determine the ΔH_{rxn}.

$q = m*c_p*\Delta T$, where q is heat transfer, m is mass, c_p is the specific heat capacity of a

substance, and ΔT is $T_{final} - T_{initial}$.

Specific heat capacity of a substance is a physical property that represents how much energy is required to raise the temperature of 1.000 g of a substance by exactly one degree Celsius from 14.5°C to 15.5°C. In this lab, we will be using the specific heat capacity of water, which is 4.184 J/g°C.

The calorimeter itself absorbs or releases some of the heat transferred, so you usually have to determine the heat capacity, C, of the calorimeter, which will be expressed in J/°C.

You will also be applying Hess' Law, which states that the sum of the ΔH_{rxn} for two or more reactions will be the ΔH_{rxn} for the summary reaction. First, you will determine the ΔH_{rxn} in kJ/mol for reactions (1) and (2) in the objectives above in parts B and C of the procedure. Then, using the following reaction (3) below, determine the enthalpy of reaction for the formation of MgO from its elements, $Mg_{(s)} + \frac{1}{2} O_{2(g)} \rightarrow MgO_{(s)}$.

3. $H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H_f = -286$ kJ/mol

There is an optional part of this experiment, part D, where you may be asked to determine $\Delta H_{solution}$ of a salt. Salts are held in the solid phase by ionic bonding. To dissolve in water, the ions that make up the salt must be strongly attracted to the water and give up their attraction to other ions. The energy holding the salt in the solid phase is called the lattice energy, ΔH_{LE}. The attraction of the ions to water is called hydration energy, $\Delta H_{hyration}$. The $\Delta H_{solution}$ will be the sum of these two energies. It can be slightly positive (endothermic) or slightly negative (exothermic). You will determine the value in kJ/gram for your assigned salt.

Examples:

1. Heat transfer: How much energy must be transferred to 100.0 g of water to raise the temperature from 21.2°C to 31.2°C?

$q = 100.0\text{ g} * 4.184\text{ J/g°C} * (31.2 - 21.2)\text{°C} = 4184\text{ J}$

2. Hess' Law: Given the three reactions below, find ΔH_{rxn} for $C_2H_{2(g)} + 2\ H_{2(g)} \rightarrow C_2H_{6(g)}$

Rxn 1. C_2H_6 (g) + $^{7}/_{2}$ O_2 (g) → 2 CO_2 (g) + 3 H_2O (l) ΔH_{rxn} = -1560 kJ

Rxn 2. 2 C_2H_2 (g) + 5 O_2 (g) → 4 CO_2 (g) + 2 H_2O (l) ΔH_{rxn} = -2599 kJ

Rxn 3. H_2 (g) + $^{1}/_{2}$ O_2 (g) → H_2O (l) ΔH_{rxn} = -286 kJ

Reverse Rxn 1. 2 $CO_{2(g)}$ + 3 $H_2O_{(l)}$ ® $C_2H_{6(g)}$ + $^{7}/_{2}$ $O_{2(g)}$ ΔH_{rxn} = +1560 kJ

Halve Rxn 2. $C_2H_{2(g)}$ + $^{5}/_{2}$ $O_{2(g)}$ → 2 $CO_{2(g)}$ + $H_2O_{(l)}$ ΔH_{rxn} = ½ (-2599 kJ)

Double Rxn 3. 2 $H_{2(g)}$ + $O_{2(g)}$ → 2 $H_2O_{(l)}$ ΔH_{rxn} = 2(-286) kJ

Notice that with canceling what appears on both sides, we have the reaction we wanted. Now just add up the three ΔH_{rxn}'s to get ΔH_{rxn} = -312 kJ.

3. A student followed the procedure for determining the heat capacity of her calorimeter. She had 50.0 g of warm water at 37.9°C and 50.0 g of cold water at 20.9°C. The final temperature, extrapolated, was 29.1°C. Find the heat lost by the warm water, heat gained by the cold water, and the heat capacity of her calorimeter. In the equations below, "C" is the heat capacity of the calorimeter.

q_{lost} = 50.0 g * 4.184 J/g°C * (29.1 – 37.9)°C = -1841 J

q_{gained} = 50.0 g * 4.184 J/g°C * (29.1 – 20.9)°C = +1715 J

$q_{lost} = -(q_{gained} + C^{*}\Delta T)$ $[\Delta T = T_{f,both} - T_{i,cold}]$

-1841 J = -(+1715 + C*(29.1 – 20.9)°C)

C = 15.4 J/°C

Prelaboratory Questions:

1. Define specific heat capacity of a substance, heat capacity of a calorimeter, and enthalpy of reaction.

2. If the specific heat capacity of methanol is 2.51 J/g°C, how much energy is required to raise the temperature of 80.0 g of methanol from 21.5°C to 33.7°C?

Safety:

Wear appropriate PPE (see page *i*).

Materials:

Chemicals per student: deionized water, ~ 500 mL 1 M HCl, ~0.4 g Mg ribbon, ~ 1 g MgO solid, various salts (NaCl, NH_4Cl, Na_2SO_4, $(NH_4)_2SO_4$).

Equipment per student: two Styrofoam cups and cardboard lid, thermometer or thermistor, ring stand, utility clamp, large beakers, hot plate.

Procedure:

You will probably need to work with one partner so that you can read the temperature and time and record them. You may be using red alcohol thermometers, which are not very accurate, but since you are interested in changes in temperature, they should be good enough. READ ALL TEMPERATURES TO ONE DECIMAL PLACE. If you use thermistors, they will have to be calibrated.

Part A: You will be determining the heat capacity of each calorimeter **if** your instructor tells you to do this part. Otherwise, the instructor should provide an average calorimeter heat capacity. For our experiment, the calorimeter is a simple set of two Styrofoam cups with a cardboard lid that is set in a beaker.

1. Set up your calorimeter. Get two of the same size Styrofoam cups and a cardboard lid to make a calorimeter. Record the mass to three decimal places. Find a beaker that your calorimeter can fit in for stability and set it in there. You'll need a glass stirring rod and a thermometer, or your instructor might choose to use a thermistor and the Vernier LabPro program. Neither the thermometer nor the thermistor should touch the bottom of the calorimeter, so you may have to use a one-hole rubber stopper, a utility clamp and a ring stand to hold the thermometer in place. Add about 50 mL of tap water using a graduated cylinder. Reweigh the calorimeter with the water in it. Stir the calorimeter water with the thermometer for two minutes and record the temperature as the T_i for the "cold" water.
2. Get another 50 mL of hot tap water if it's available, or else heat the available tap water to 35 ± 2°C in a beaker. Turn off the Bunsen burner, stir the beaker for two minutes, during which time you record the temperature of this warm water every 30 seconds. Immediately pour the warm water from the beaker into the cold water in the calorimeter. Cover and stir and record the temperature every 15 seconds for at least 2½ minutes.

3. Weigh the calorimeter with all the water in it and record. Plot the temperatures vs. the time as directed by the instructor, i.e., either on graph paper or on a computer spreadsheet. Do the following calculations:

$q_{lost} = -(q_{gained} + C^*\Delta T)$ [$\Delta T = T_{f,both} - T_{i,cold}$]

$q_{lost} = mass_{warm}{}^*(4.184 J/g°C)^*(T_{f,both} - T_{i,warm})$

$q_{gained} = mass_{cold}{}^*(4.184 J/g°C)^*(T_{f,both} - T_{i,cold})$

Solve for C, which is the heat capacity of the calorimeter. Put your result with your initials on the blackboard. When entire class has results on board, perform the Q-test on any suspected outliers. Find the average calorimeter heat capacity and its standard deviation.

Part B: Finding ΔH_{rxn} for reaction 1 in objective section above.

1. Find two of the same size Styrofoam cups and a cardboard lid to make a calorimeter. Record the mass to three decimal places. Add 100.0 mL of 1 M HCl to your calorimeter and record the mass. Determine the mass of 1 M HCl by difference. Find a beaker that your calorimeter can fit in for stability and set it in there. You'll need a glass stirring rod and a thermometer, or your instructor might choose to use a thermistor and the Vernier LabPro program. Neither the thermometer nor the thermistor should touch the bottom of the calorimeter, so you may have to use a one-hole split rubber stopper, a utility clamp and a ring stand to hold the thermometer in place. (See instructor's setup.)
2. Obtain about 15-cm length of Mg ribbon and clean it thoroughly with steel wool. Wipe it with Kem-wipes and don't touch again with your fingers. Record its mass to three decimal places. (Or your instructor may have you record the exact length of the ribbon and use a conversion the instructor provides to determine the mass.) Roll it up loosely so it will start out sitting in the bottom of the Styrofoam cup. DON'T PUT IT IN YET!
3. Once everything is set up, read the initial temperature of the HCl in the calorimeter. Remember to read your thermometer to at least one decimal place. Start stirring with the lid on, then lift the lid and drop the Mg ribbon in quickly. Read and record the temperature every 10 to 15 seconds – you decide and record the time as a running total. <u>Continue for at least five minutes</u>, so that you can see when the temperature reaches its minimum or maximum and then slowly starts back toward room temperature. Once done, pour the contents of the calorimeter into your large excess solutions beaker, rinse it out and <u>repeat steps 1 through 3</u>.

Part C: Finding ΔH_{rxn} for reaction 2 in objective section above.

1. Find two of the same size Styrofoam cups and a cardboard lid to make a calorimeter. Record the mass to three decimal places. Add 100.0 mL of 1 M HCl to your calorimeter and record the mass. Determine the mass of 1 M HCl by difference. Find a beaker that your calorimeter can fit in for stability and set it in there. You'll need a glass stirring rod and a thermometer, or your instructor might choose to use a thermistor and the Vernier LabPro program. Neither the thermometer nor the thermistor should touch the bottom of the calorimeter, so you may have to use a one-hole rubber stopper, a utility clamp and a ring stand to hold the thermometer in place.
2. Obtain about 0.4 g of MgO. Record its mass to three decimal places. DON'T PUT IT IN YET!
3. Once everything is set up, read the initial temperature of the HCl in the calorimeter. Remember to read your thermometer to at least one decimal place. Start stirring with the lid on, then lift the lid and drop the MgO solid in quickly. Read and record the temperature every 10 to 15 seconds – you decide and record the time as a running total. <u>Continue for at least five minutes</u>, so that you can see when the temperature reaches its minimum or maximum and then slowly starts back toward room temperature. Once done, pour the contents of the calorimeter into your large excess solutions beaker, rinse it out and <u>repeat steps 1 through 3 of part C</u>.

Part D: (optional) Determine the enthalpy of solution, ΔH_{soln}, for a salt, using the following procedure. Several salts will be available: NaCl, NH_4Cl, Na_2SO_4, $(NH_4)_2SO_4$.

1. Weigh the calorimeter, add 100 mL of tap water and reweigh the total; record.
2. <u>Carefully</u> weigh out about 7 grams of one of the salts and record the mass to three decimal places. Stir the tap water for two minutes with the thermometer and then record the temperature. Then quickly add the salt, cover, and start stirring to dissolve it. Record the temperature every 30 seconds for <u>at least five minutes</u>. You may be able to feel undissolved salt in the bottom of the calorimeter. If so, keep stirring and reading the temperature. (Ask the instructor if you are done collecting temperature vs. time data. Do NOT assume so.)
3. Repeat the procedure to this point so you have a second trial.
4. Graph the temperatures for both trials and extrapolate back to the time of mixing. Use the total mass of water plus salt for the calculation of D Assume the specific heat of the solution is about the same as that of water. Write your result on the blackboard with your initials, under the column for your choice of salt. Record other students' data for your salt, and determine the average. Find the percent difference between your result and the average.

Data Collection:

Part A: Record time and temperature on a separate piece of paper or directly into a computer spreadsheet program. Show calculations below.

Mass of cold water _______

Temperature of cold water in calorimeter _______

Mass of warm water _______

Temperature of warm water _______

Maximum temperature, extrapolated _______

ΔT for cold water, °C _______

ΔT for hot water, °C _______

q_{lost} _______

q_{gained} _______

Heat capacity of calorimeter _______

Part B: Record time and temperature on a piece of paper or directly into a computer spreadsheet program. Show calculations below.

	Trial 1	Trial 2
Mass of calorimeter	_______	_______
Mass of calor. + HCl	_______	_______
Mass of HCl	_______	_______
Mass or length of Mg ribbon	_______	_______
T_i of HCl in calorimeter	_______	_______
T_f, extrapolated	_______	_______
Heat transferred	_______	_______
Moles of Mg	_______	_______
ΔH_{rxn}, kJ/mol	_______	_______
Average ΔH_{rxn}, kJ/mol	_______	

Part C: Record time and temperature on a piece of paper or directly into a computer spreadsheet program. Show calculations below.

	Trial 1	Trial 2
Mass of calorimeter	_______	_______
Mass of calor. + HCl	_______	_______
Mass of HCl	_______	_______
Mass of MgO	_______	_______
T_i of HCl in calorimeter	_______	_______
T_f, extrapolated	_______	_______
Heat transferred	_______	_______
Moles of MgO	_______	_______
ΔH_{rxn}, kJ/mol	_______	_______
Average ΔH_{rxn}, kJ/mol	_______	

Optional Part D: Name of assigned salt ___________________________

Record time and temperature on a piece of paper or directly into a computer spreadsheet program. Show calculations below.

	Trial 1	Trial 2
Mass of calorimeter	_______	_______
Mass of calor. + H_2O	_______	_______
Mass of H_2O	_______	_______
Mass of salt	_______	_______
T_i of H_2O in calorimeter	_______	_______
T_f, extrapolated	_______	_______
Heat transferred, q	_______	_______
$\Delta H_{sol'n}$, kJ/gram	_______	_______
Average $\Delta H_{sol'n}$, kJ/gram	_______	

Calculations/Analysis/Results:

You have done two trials for each reaction. Your instructor will ask you to plot temperature vs. time for each trial either on graph paper or on a computer spreadsheet. You will have to determine the maximum temperature for each trial by extrapolating to time zero on your graph. This will take into account the heat lost to the outside of the calorimeter if necessary. If your temperature line slants down, draw a straight line through it back to time zero. Note: our little calorimeters are so good, you may not have to do this! But show all your graphs to your instructor and get them initialed. IF YOU DO THIS ON THE COMPUTER, PRINT ONLY ONE SET OF GRAPHS AND ATTACH TO ONE PERSON'S LAB REPORT.

Calculate ΔH_{rxn} for each trial for Reaction 1 as follows, which includes the heat absorbed by the calorimeter from part A:

$$q_{rxn} = [(\text{total of mass Mg} + HCl_{(aq)}) * 4.184 J/g\text{-}K * (T_f - T_i)] + [C^*(T_f - T_i)]$$

This is q for however many grams of Mg you had in each trial. Convert grams to moles of Mg and convert q to ΔH in kJ/mol for each trial. Determine the average of your two trials.

Calculate ΔH_{rxn} for each trial for Reaction 2 as follows, which includes the heat absorbed by the calorimeter from part A:

$$q_{rxn} = [(\text{total of mass MgO} + HCl_{(aq)}) * 4.184 J/g\text{-}K * (T_f - T_i)] + [C^*(T_f - T_i)]$$

This is q for however many grams of MgO you had in your trial. Convert grams to moles of MgO and convert q to ΔH in kJ/mol for each trial. Determine the average of your two trials.

Use your average ΔH_{rxn} for reaction 1 and reaction 2, and the given ΔH_{rxn} for reaction 3 to find ΔH_f for the formation of MgO: $Mg_{(s)} + ½ O_{2(g)} \rightarrow MgO_{(s)}$.

Report your average ΔH_{rxn}/mol for both reactions and your calculated ΔH_f for MgO on the blackboard by your initials. Copy all class data for statistical analysis. Perform the Q-test on any suspected outliers. Calculate average for each reaction from the class data. Also calculate the standard deviation for each, as well as the relative deviation.

Determine the percent error between your average and the actual ΔH_f for MgO.

Optional: calculate $\Delta H_{solution}$ of your salt. The salt itself and the water are either absorbing or releasing heat as a result of the physical process of dissolving the salt. Find the heat transferred to or from the water and salt for your experiment. Change the sign to determine the $\Delta H_{solution}$ for the process and then divide that by the initial mass of your salt to report it in kJ/gram.

$$q = mass_{water} * c_{p,water} * (T_{f,water} - T_{i,water}) + mass_{salt} * c_{p,salt} * (T_{f,water} - T_{i,water})$$

$$\Delta H_{solution} = -q/mass_{salt}(1\ kJ/1000\ J)$$

Specific heat capacities: NaCl = 0.854 , NH_4Cl = 1.57 , Na_2SO_4 = 0.903 , $(NH_4)_2SO_4$ = 1.41 unit J/g°C.

	Part A	Part B	Part C	ΔH_f, MgO	Part D (your salt)
Class averages	_______	_______	_______	_______	_______

Determination of ΔH_f, MgO: show all work using your results for part B and part C. Repeat the calculation using the class averages for parts B and C.

Percent difference and percent error:

Post Laboratory Questions:

1. Rewrite reactions (1) and (2) as net ionic equations. Then look up the heats of formation for the reactants and products and determine the true values for ΔH_{rxn}. Why were your values and the class' values not the same as the true values? What could we have done to improve our accuracy? Calculate your percent error for each.

2. Look up the specific heat of pure aluminum metal. If 20.00 grams of metal is heated to 100.0°C and then added to 100.0 grams of tap water at 24.5°C in **your** calorimeter, what will the final temperature be? (Did you include the heat lost to the calorimeter? You should!)

Discussion:

Conclusion:

Error Analysis:

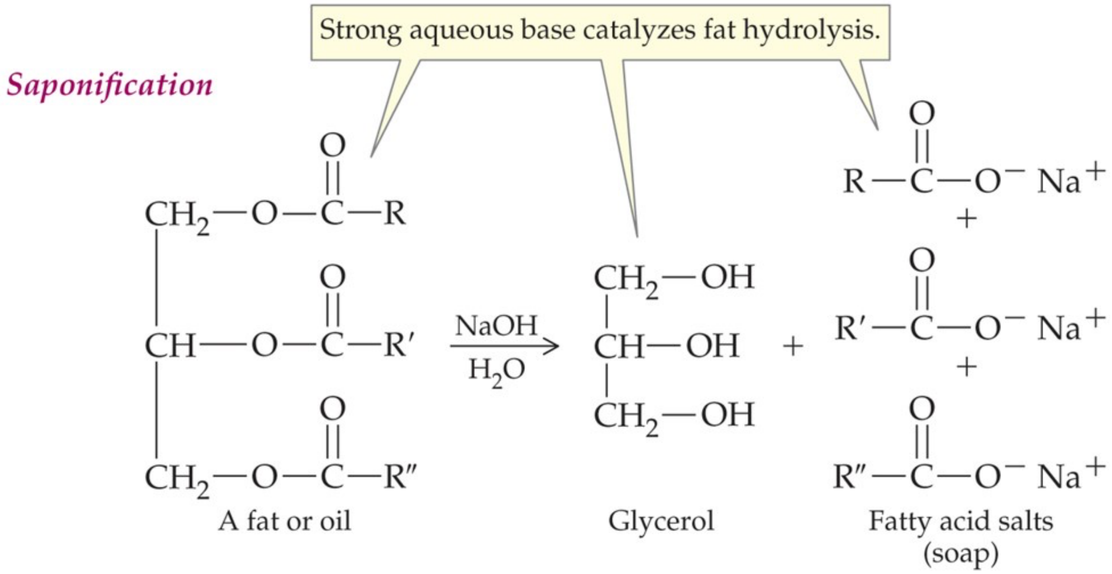
EXPERIMENT 14: PREPARATION AND PROPERTIES OF SOAP

Objective 14

This experiment introduces the student to synthesis and properties of soap, as well as the interparticle forces between molecules of soap, dirt or oil, and water.

Introduction:

Soaps are sodium or potassium salts of long chain fatty acids, created by the base hydrolysis of oils. Simple fats, such as vegetable oils, are generally triglycerides, which are really "triesters" of glycerol ($CH_2OHCHOHCH_2OH$) and various fatty acids (R-COOH). The long chain fatty acid found in many soaps is stearic acid, which has the condensed structural formula $CH_3(CH_2)_{16}COOH$. An ester can be hydrolyzed, i.e., broken into its acid and alcohol components with water. This is the first step in making soap. The second is to add the base and make the sodium salt of the fatty acid. The two steps are summarized as one below called saponification:



The triglyceride is broken into a triol (glycerol) and its three sodium salts of fatty acids. The process of making soap is called saponification. Glycerol is sometimes kept in the soap to soothe your skin.

How do soaps work? "Dirt" is composed of organic materials, which tend to be nonpolar. Oils and greases are also organic and nonpolar for the most part. Water is usually the solvent when we're trying to clean hands, clothes or dishes, and water is very polar. Water alone will not wash bacon grease out of the frying pan! But the sodium salt of a fatty acid has a long nonpolar end and a short polar end. When soap is put in water, the nonpolar ends find each other, leaving the polar ends and their sodium ion facing "out" toward the water. This roughly spherical glob of soap in water is called a micelle. Micelles are dispersed in the water, but NOT individual salts of fatty acids. This is called a suspension, not a solution, and makes a soap solution look cloudy. Oil and dirt molecules are taken in to the center of the micelle, because their interparticle forces (London dispersion forces) are very similar. The oil and dirt want to get away from the water, because their forces are so dissimilar – water molecules experience the interparticle force called Hydrogen bonding. So one end of soap is called hydrophobic – it doesn't like water, and the other end is called hydrophilic – it likes water.

You have probably all experienced a disadvantage that soap has – it makes bathtub rings. That's because the sodium and potassium salts of the fatty acids will form the suspension, but Ca^{2+} and Mg^{2+} salts of fatty acids precipitate and attach to the sides and bottom of the container.

This led to the development of "artificial" soaps called detergents, which are made from petroleum products rather than a natural reactant like vegetable oil or bacon grease. One of these detergents is an alkylbenzene sulfonate group. These do not form precipitates with Ca^{2+} and Mg^{2+}.

Prelaboratory Questions:

1. What is an ester? How is an ester synthesized?

2. What is an organic acid? What structural component identifies them as an acid?

3. What is hydrolysis?

4. What makes soap different from a detergent?

5. Identify and define all the interparticle forces at work during this experiment. Read it thoroughly! Use the new words you have learned.

Safety:

Always wear appropriate PPE (see page *i*). You might want to wear gloves since you will be using a base.

Chemicals per student: ~ 20 g vegetable oil, ~20 mL ethanol, ~20 mL 6 M NaOH, ~100 mL saturated NaCl solution, a commercial soap, a commercial detergent, mineral oil, test solution of hard water (7×10^{-4} M $CaCl_2$ and $MgCl_2$), 3 M HCl.

Equipment per student: 125-mL Erlenmeyer flask, 250-mL beaker, stirring rod, hot plate, small beaker for ice water, other beakers, Buchner funnel, pH paper.

Methods:

Part A: Preparation of soap

1. Turn a hot plate on LOW so that it is warmed up by the time you have combined your soap ingredients. Also get a small beaker filled with ice and add deionized water to it, so you have ice-cold water for another step. Now weigh a 125-mL Erlenmeyer flask and then add about 20 grams of vegetable oil to it. The exact mass is not important, although you should record it. Add about 15 to 25 mL of ethanol to the flask and shake it with the oil. Add 19 to 21 mL of 6 M NaOH while swirling the flask. Transfer the contents to a 250-mL beaker. Set the beaker on the hot plate and use a stirring rod to keep mixing the layers. At least every three minutes you should stir the mixture,

because it will be in two layers, with the vegetable oil on top. NOTE: ethanol is flammable, so your instructor may ask you to do this in a fume hood. Make sure there are no open flames nearby. Don't stir so much that your solution gets sudsy. It should turn into one cloudy layer that thickens in 20 to 30 minutes. Take the beaker off the hot plate and let it cool to room temperature. (Put a 400-mL beaker of deionized water on the hot plate to warm up for Part B.)

2. Add 100 mL of saturated NaCl solution to the soap mixture and stir for a couple of minutes. This step is called "salting out" and helps separate the water from the soap. Break up any remaining chunks so that they mix with the NaCl solution. Use a Buchner funnel to collect your soap. Wash the soap in the funnel with 25 mL of ice-cold water. Let the suction filtration continue to run so that it draws air across the soap to help dry it. Set up Part B while you are waiting for your soap to dry. Once you have taken out the 1 g of your soap and tested its pH, take your soap out of the funnel and mush it down into a plastic weigh boat. If your instructor has told you ahead of time, you can add a couple of drops of perfume by massaging it into the soap before you put it in the weigh boat. Press it down in the weigh boat and put it in your locker. Use a spatula to mark your initials in your soap and show your instructor. After two weeks, if you remember to do this, you can mold the soap into a bar or a ball, but don't plan on using this soap at home! This is old-fashioned "granny" soap made with lye (NaOH) and is pretty harsh on your skin. But it should be good for show and tell.

Part B: Properties of soap

1. Get about 1 g of a commercial soap and about 1 g of a commercial detergent and put them in their own labeled 250-mL beakers. Do the same with 1 g of your air-dried soap. Add about 100 mL of warm deionized water to each beaker.

2. Use a stirring rod (and rinse it between uses), test the pH of each solution you prepared in step 1 by touching a drop of the solution from your stirring rod to the pH paper and comparing the color to the scale for the paper. Record the pH of each. If your homemade soap has a pH greater than 8 or 9, break it up in the Buchner funnel and wash it again with 25 mL of ice-cold water. Let it air dry for 10 minutes and make up a new solution as directed in step 1.

3. Get out four test tubes and put about 10 mL of each of your soap or detergent solutions in its own test tube. Let the fourth one hold 10-mL of deionized water. Add 10 drops of mineral oil to each. Shake each individually

and observe the suspension or layering that results. Watch for about five minutes so that you can determine how much time elapses before separate layers appear, if at all.

4. Set up four test tubes again like you did in step 3, but don't add mineral oil. Instead, add about 2 mL of a solution labeled "Hard Water" to each. As you do, shake each and look immediately for cloudiness or precipitates. Watch for about five minutes so that you can determine if any cloudiness appears at all.

5. Set up four test tubes again like you did in step 3, but this time add 5 drops of 3 M HCl to each. Shake each one and see if it remains sudsy. Then add 10 drops of mineral oil to each, shake, and observe like you did in step 3. Compare your observations with step 3 and comment on the differences.

6. Repeat step 4 using regular tap water instead of the "hard water" and observe like you did in step 4. What can you conclude about our water supply?

7. Keep your gloves on and try "washing your hands" with your soap. How does it compare to the soap you are used to using at home?

Disposal: put your soaps and detergents in water to dissolve and pour down the drain. Pour any extra ethanol down the drain and rinse down with water. Pour any excess HCl or NaOH in your excess solutions beaker. Test pH with pH paper and add 6 M HCl or 6 M NaOH to bring the pH to the 3 to 11 range; pour down the drain with water.

Data Collection:

Part A:

Mass of vegetable oil _______

Volume of ethanol _______

Volume of saturated NaCl _______

Appearance of soap:

Part B:

Data Table

Beaker ID	**My soap**	**Commercial soap**	**Commercial detergent**
pH of solution			

Emulsification with mineral oil:

Data Table

Test tube ID	**My soap**	**Co mm. Soap**	**Comm. Det.**	**Water**
Initial Observations				
After 5 minutes				

Reaction with Hard Water:

Data Table

Test tube ID	**My soap**	**Comm. Soap**	**Comm. Det.**	**Water**
Initial Observations				
After 5 minutes				

Effect of HCl on sudsiness:

Data Table

Test tube ID	**My soap**	**Comm. Soap**	**Comm. Det.**	**Water**
Initial Observations				
After 5 minutes				

Effect of HCl on emulsification with mineral oil:

Data Table

Test tube ID	**My soap**	**Comm. Soap**	**Comm. Det.**	**Water**
Initial Observations				
After 5 minutes				

Reaction with regular tap water:

Data Table

Test tube ID	**My soap**	**Comm. Soap**	**Comm. Det.**	**Water**
Initial Observations				

After 5 minutes

Comment on differences between tap water and hard water:

Instructor initials _______

Post Lab Questions:

1. How can there be hydrophobic and hydrophilic parts on the same compound?

2. Look up glycerol and report on its properties. What is it good for?

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 15: CLAUSIUS-CLAPEYRON EQUATION

Objective 15

This experiment will enable you to determine the enthalpy of vaporization for water, using the Clausius-Clapeyron Equation and data collected regarding the change in the vapor pressure of water with temperature.

Introduction:

The vapor pressure of a volatile liquid will increase with temperature, but the relationship is not linear. It is in fact a curve and a graph of pressure vs. temperature is called a vapor pressure curve. However, graphing the natural log (*ln*) of the pressure vs. the inverse of the temperature in Kelvin does produce a linear relationship, shown by this equation:

Equation 1: $\ln P_{vap} = (^{-\Delta Hvap}/_{R})(^{1}/_{T}) + A$, which has the form y = mx + b, i.e., the equation of the line.

The $\ln P_{vap}$ is y, 1/T is x, and the slope, m, is $-\Delta H_{vap}/R$. Slope is also the change in y divided by the change in x or slope is $^{(y1-y2)}/_{(x1-x2)}$. (A is just the y-intercept.) Equation 2 shows the solution for finding the slope:

Equation 2: $(\underline{\ln P_{vap1} - \ln P_{vap2}}) = -\Delta H_{vap}/R * (^{1}/_{T1} - {^{1}/_{T2}})$

Note that *ln* a – *ln* b is the same as *ln*(a/b). If we use that as a substitution and then rearrange Equation 2, we have the Clausius-Clapeyron Equation:

Equation 3: $\ln (^{Pvap1}/Pvap_{2}) = (^{-\Delta Hvap}/_{R})\ (^{1}/_{T1} - {^{1}/_{T2}})$

(Note: ΔH_{vap} does change slight with temperature, but this works fairly well.)

Example

A student has recorded a volume of 3.30 mL at 5.0°C and a volume of 9.95 mL at 80.0°C. The barometric pressure is 745.0 torr. The vapor pressure of water at 5.0°C is 6.54 torr. Correct the volume readings, convert temperatures to Kelvin, adjust the dry air pressure for water vapor pressure, convert pressure to atmospheres, and then find the moles of air in the sample. Use that to find the vapor pressure of water at 80.0°C.

Using n = PV/RT, n = 0.9717 atm * 0.00310 L/(0.082057*278.15K) = 1.32×10^{-4} moles of dry air.

P_{air} = nRT/V = 1.32×10^{-4} mol * 0.082057 * 353.15K / 0.00975 L = 0.392 atm or 298.2 torr

P_{vap} = 745.0 torr – 298.2 torr = 446.8 torr

Prelaboratory Questions:

1. Write definitions for equilibrium vapor pressure, normal boiling point and enthalpy of vaporization.

2. Look at the vapor pressure data table for water in the appendix. Select two temperatures at least ten degrees apart, and use the pressure/temperature data to calculate DH_{vap}.

Safety:

Always wear appropriate PPE (see page *i*). Use caution handling hot plate, hot beaker and hot water.

Materials:

Chemicals: Deionized water and ice

Equipment, per pair of students: 10 mL graduated cylinder, 1000 mL beaker that covers the graduated cylinder, thermometer, a hot plate, utility and/or burette clamp, ring stand, large plastic tray

Methods:

In this experiment, you will be monitoring the volume of air and water vapor in an inverted 10-mL graduated cylinder at various temperatures. You will need to prepare a data table to collect the temperature and volume data or you may be collecting it in a computer program. You should work with one partner since you have to stir, read the temperature and volume, and record the data, all at the same time.

Your instructor may require you to calibrate your thermometer, or if you use Vernier LabPro, you will have to calibrate the thermistor. Use boiling water to set the high temperature and ice water to set the low temperature. You'll have to determine the equation of the line for the calibration of an alcohol thermometer, and then use it to correct all of your temperature readings.

1. Fill a tall beaker so that it covers your narrow-ended 10-mL graduated cylinder. Have a utility clamp or burette clamp ready to help hold the graduated cylinder down below the surface of the water in the beaker. Put about 9 mL of deionized water in the cylinder, then put your finger over the end to close it. Invert and submerge the cylinder in the water in the beaker. You should see that there are about 4 mL of air now at the closed end of the cylinder. Make sure the water in the beaker covers the cylinder and that the cylinder will stay in the beaker. Have your instructor inspect your setup at this point.

2. Using a hot plate and stirring often, heat the water in the beaker to about 85 °C, or whatever is required to increase the size of the air bubble so that it is in the unmarked portion of the volume scale of your cylinder. Turn off the hot plate.

3. Stir the water continuously, and when the air bubble is in the marked region of the cylinder, start recording volume and temperature about every 2 degrees. Try to read the volume to greater precision than 0.1 mL. You still read the bottom of the meniscus to obtain the volume. Have your instructor check on how you are reading the inverted graduated cylinder.

4. When the temperature drops below 50°C, put the beaker in a large plastic tray to catch water overflow, since you'll be adding ice cubes. Alternately, you can put the beaker right next to the sink and turn the rim so that excess water flows into the sink.

5. Add ice <u>a little at a time</u> and stir continuously. Continue to record volume and temperature every two degrees until the temperature is at or below 5.0 °C. PLEASE NOTE that the temperature change will occur much faster now, so try not to miss any readings.

6. Read the barometric pressure and record it.

7. Since the graduated cylinder is calibrated for reading right side up, you will have to correct your volume readings by subtracting 0.20 mL from all volumes recorded.

Data Collection:

You may attach the spreadsheet and graph or else note here which student has the spreadsheet and graph with their report.

Barometric pressure ________

Analysis/Calculations/Results:

Use a computer spreadsheet program to prepare a table with headings as shown below:

T, °C; V_{gas}, mL; T, Kelvin; Corr V_{gas}, mL; $P_{dry\ air}$, torr; P_{vap}, torr; *ln* P_{vap}; $1/T_{(Kelvin)}$

Enter your data in the spreadsheet (with the corrected temperature if required) and the volume. Put in an equation to convert the corrected temperature to Kelvin. (Note that all of these calculations can be entered as equations in the spreadsheet, letting the spreadsheet program do some of the work for you.)

Calculate the moles of dry air at 5.0°C (or your lowest temperature reading). The vapor pressure of water at 5.0°C is 6.54 torr. You may have to extrapolate to determine the vapor pressure at your temperature. Use the ideal gas law to solve for moles of dry air using your temperature reading, the barometric pressure, your corrected volume at the lowest temperature reading, and the value of R to at least four significant figures. REMEMBER! T must be in Kelvin and P in atmospheres.

$$n_{air} = P_{air} * V/RT$$

The total moles of gas inside the cylinder include the dry air and water vapor. Therefore the total pressure inside is the sum of these two partial pressures. Although there is some pressure from the height of the liquid water inside the cylinder, go ahead and set the two gas partial pressures equal to the barometric pressure.

Calculate the pressure due to the air at each temperature and volume, using the ideal gas law to solve for P. Convert pressure to torr.

$$P_{air} = n_{air}*RT/V$$

The partial pressure due to water vapor can be determined by subtracting the partial pressure of dry air from the barometric pressure, and again, use the units of torr.

$$P_{vap} = P_{atm} - P_{air}$$

Now determine the natural log of the water vapor pressure and the reciprocal of the temperature.

Use the spreadsheet to create a graph with *In* P_{vap} (in units of torr) as the y-axis and the reciprocal T (in units of Kelvin^{-1}) as the x-axis. Use linear regression to determine the slope and y-intercept. Write the equation of the line somewhere on your graph. The value of the slope is $-\Delta H_{vap}/R$, with R as 8.314 J/mol-K. Solve for ΔH_{vap}. Before you print the graph, change the background color to “None” and add more gridlines. Have the instructor OK the graph on the computer before printing it.

Print only ONE copy of the spreadsheet and one copy of the graph. One person of the team will put it in their lab report and the other will note in their report which team member has the copies.

Look up the true value of ΔH_{vap} for water at room temperature and calculate your percent error.

Collect the class data and determine the class average for ΔH_{vap} for water and the standard deviation.

Result Table

Calculate DH$_{vap}$	**True DH$_{vap}$**	**Percent Error**	**Class Average DH$_{vap}$**

Post Lab Questions:_

1. What experimental steps and measurements could lead to error in this experiment?

2. Acetone has a normal boiling point of 56.2 °C and a ΔH_{vap} of 31.97 kJ/mol. Calculate the equilibrium vapor pressure at 20.0 °C. (Hint: what does "normal" boiling point mean?)

Discussion:

Conclusion:

Error Analysis:

EXPERIMENT 16: FREEZING POINT DEPRESSION

Objective 16

In this experiment you will determine the molality and quantity of a solute in an unknown solution using freezing point depression data.

Introduction:

Every pure substance has a specific freezing point (melting point) at standard atmospheric pressure. This normal freezing point is the temperature at which the liquid and solid are both present and at equilibrium. The molecules have "lost" enough kinetic energy that they can experience intermolecular forces strong enough to form a crystal structure.

If a solute is present, whether ionic or covalent, the freezing point of the pure substance is depressed, i.e., it is lower than the pure substance. The phase change from liquid to solid will not occur at constant temperature nor will it start at the freezing point of the pure substance. The solute interferes with the formation of the crystal structure of the solvent. When the solution does freeze, the solid which forms is pure solvent, and the remaining liquid has a higher concentration of solute in it, so there is more interference and the temperature must continue to decrease in order to freeze more solvent.

This freezing point depression is directly related to the number of solute particles present, not on whether the solute was ionic or covalent. Each solvent has its own freezing point depression constant which relates the number of degrees the freezing point is lowered to the molality of the solute particles, not to the type of solute. To find the “depressed” freezing point of a solution, you need two calculation steps and you must look up the freezing point of the pure solvent and its freezing point depression constant. You calculate the change in temperature in step 1 and then subtract this change from the pure solvent’s freezing point in step 2.

Step 1: $\Delta T_f = K_f * \text{molality}$

Step 2: new $T_f = T_f^{o} - \Delta T_f$

Water as a solvent has a freezing point depression constant of 1.858°C/molal. Any aqueous solution which is 1.0 molal in dissolved particles will freeze 1.86 °C below the normal freezing point of water. A 1.0 molal solution of glucose will freeze at -1.86°C, while a 1.0 molal solution of sodium chloride will freeze at -3.53°C, which is almost twice the change that occurs for glucose, because one mole of sodium chloride provides almost two moles of particles. Ionic compounds have been tested for their effect on freezing point and boiling point, and a factor, called the Van't Hoff factor, is used to determine their colligative molality.

***Example*:**

1. What volume of ethylene glycol must be added to 15.0 mL of water to produce a solution with a freezing point of –13.0°C?

First solve for the molality of the solution: ΔT_f = 13.0°C = K_f*molality, therefore molality = 13.0°C/(1.858°C/molal) = 7.00 moles ethylene glycol/1 kg of water.

Find actual moles of ethylene glycol: 15.0 mL(1.00g/mL)(1kg/1000g)(7.00mole/kg) = 0.105 mole.

Convert moles to mass and then volume: 0.105 mol(62.07g/mol)(1mL/1.11g) = 5.87 mL of ethylene glycol.

2. What mass of NaCl must be added to 15.0 mL of water to produce a solution with a freezing point of –13.0°C?

First solve for the molality of the solution, remembering to use the Van't Hoff factor:

ΔT_f = 13.0°C = K_f*molality**i*, therefore molality = 13.0°C/(1. 858°C/molal)/(1.9) = 3.68 moles NaCl/1 kg of water.

Find actual moles of NaCl: 15.0 mL(1.00g/mL)(1kg/1000g)(3.68mole/kg) = 0.0552 mole.

Convert moles to mass: 0.0552 mol(58.4425g/mol) = 3.23 grams.

Prelaboratory Questions:

1. Write definitions for colligative properties, molality, colligative molality, and freezing point depression.

2. What volume of ethylene glycol must be added to 4.0 gallons of water to obtain a ΔT_f of 20.0^{o}C?

Safety:

Ethylene glycol is an organic chemical. Wear appropriate PPE (see page *i*) and possibly gloves if doing that procedure.

Disposal: Ethylene glycol **CANNOT** go down the drain. Pour your sample into the designated beaker for collection for this experiment.

Glucose and NaCl solutions may be poured down the drain.

Pour your salt/water/ice solution down the drain.

Materials:

Chemicals: Unknowns. Deionized water and ice. Lots of rock salt.

Equipment, per pair of students: 800-mL beaker, thermometer and/or thermistor. You may use a computer/lab interface to collect temperautre vs time data (Your instructor will demonstrate the equipment and set up).

Methods:

There three alternatives solutes for this lab experiment. One creates hazardous waste (ethylene glycol) and the others do not (glucose or NaCl) Your instructor will tell you which one you are using. By determining the freezing point, one can find the molality of the solution. (Remember, molality = moles of solute per kilogram of solvent.) If your solute is ethylene glycol, you will use the density of ethylene glycol (1.12 g/mL), the density of water, and the molar mass of ethylene glycol to calculate the volume of ethylene glycol in your 20.0 mL unknown sample. If your solute is glucose or sodium chloride you will use freezing point data to determine the molality, and in the case of sodium chloride, apply the Van't Hoff factor.

It is recommended that you work with a lab partner for this experiment. The very first thing you should do is put about 100 mL of deionized water in your large beaker, then add rock salt and stir, continuing to add rock salt until no more will dissolve (i.e. a saturated solution). The solubility of NaCl is 35.7g/100g, so put in plenty of rock salt. Then fill the beaker about halfway up with ice.

1. Obtain your unknown sample from the instructor. Record the unknown number. Also obtain a thermometer and a 800-mL or larger beaker. You may use a computer/lab interface to collect temperature vs time data (Your instructor will demonstrate the equipment and set up).

2. Determine the freezing point of water. Put ice in a small beaker with enough deionized water to cover the thermometer bulb. Stir carefully for at least five minutes, then read the thermometer (or thermistor) with the bulb held away from the bottom and sides of the beaker. Record this temperature to at least one decimal place.

3. Next pour your entire unknown sample (20.0mL) into a large clean dry test tube. Immerse the test tube with your sample in the beaker of salt/water/ice, making sure the liquid level inside the test tube is below the liquid level in the beaker.

4. Now hold the thermometer inside your test tube with the bulb in the solution. Stir the solution occasionally with the thermometer and stir the ice bath with a glass stirring rod. Use a separate thermometer to monitor and maintain the ice water bath temperature at -10^{o} Record the temperature and time every 10 seconds until the solution appears to be frozen or very slushy. Remove the test tube from the ice bath and warm it up with your hands to melt it again. Repeat the freezing process at least two more times. KEEP YOUR SAMPLE until you have shown the instructor your graph of temperature vs. time for all three trials.

Data Collection:

You may attach the spreadsheet(s) and graph(s) or else note here which student has them with their report.

Unknown Number ________ Freezing Point ________

Unknown Number ________ Freezing Point ________

Freezing point of pure water ________

Observations:

Analysis/Calculations/Results:

Plot temperature on the y-axis and time on the x-axis. You will probably have experienced supercooling before the solution was frozen. (The temperature will dip below the freezing point without the liquid freezing. As soon as crystallization starts, the temperature will go back up to the freezing point and then remain constant or drop slowly until all the liquid is frozen.) Draw your best straight line through the points on the declining part of your graph, and another straight line through the points on the level or nearly level part. The intersection of the two lines is the "true freezing point" of the solution. Do this for all three trials and average the three freezing points. To find the ΔT_f, take the difference between the freezing point of pure water on your thermometer and the freezing point of your unknown.

Use the freezing point depression constant of water (1.858°C/molal), the freezing point of your unknown, the density (1.11 g/mL) and molar mass of ethylene glycol (62.07 g/mol) to calculate the volume of ethylene glycol in your 20.0 mL of unknown. (Assuming the volumes of water and ethylene glycol are additive to the 20.0 mL you received.)

Use the freezing point depression constant of water (1.858°C/molal) and the freezing point of your unknown to calculate the molality of both unknowns. The Van't Hoff factor for NaCl is 1.9.

Post Lab Questions:

1. Given 1.0 molal solutions of sucrose, NaCl and $CaCl_2$, choose the one with the:
    A. highest freezing point
    B. lowest freezing point
    C. highest boiling point
    D. lowest boiling point

2. In many parts of the United States, families make homemade ice cream every summer using a container with a mixer for the ice cream ingredients. The container then rotates (by hand-cranking or a motor) in a bucket which contains rock salt and ice. Why do they need the rock salt?

Discussion:

Conclusion:

Error Analysis:

DRY LAB I: GRAPHING

Objective I

You will learn to make various types of graphs on the chemistry laboratory computers.

Introduction:

You will be using a spreadsheet program to enter data and learn the graphing functions.

There are various types of graphs: straight lines, curves, scatter plots, trends, etc. If data is collected over time, time is almost always on the x-axis. Most of your graphs in this lab should start out as x-y scatter plots with lines.

All graphs should have a title at the top in capital letters, a title for each axis, legends if needed, scale used, etc. You will have to scale your graphs manually so that the data points use up the whole scale. You should also make the plot area have no color – it uses up large amounts of ink from any printer.

Prelaboratory Questions:

none

Safety:

no safety concerns for this lab

Materials:

computer and printer

Procedure:

Work with a partner for this dry lab. Use Microsoft Excel if that is what is available on the lab computers (or Google Sheets). Prepare the tables of data listed below. Learn to use the graphing function. NOTE: Excel will assume that the first column is the x-axis and that the second and other columns are on the y-axis. Each graph (with the exception of Graphs 4 & 5) needs a title, axis titles, a fit of the plot (trendline), display of both equation and R^2. When you are finished with the graphs, call your instructor to your workstation to review your graph and initial your lab report sheet.

Graph 1: Create a graph with temperature in **Kelvin** on the x-axis and pressure in torr on the y-axis. To do this efficiently enter in all the temperatures into column A then have Excel do the conversion to Kelvin in column B (ask instructor to show you). Enter the pressure data into column C. Highlight both columns B and C and create a scatter plot (scatter only, no line).

VAPOR PRESSURE OF WATER (data for Graph 1)

Temperature, °C	**Pressure, torr**	**Temperature, °C**	**Pressure, torr**
5	6.54	30	31.824
10	9.209	31	33.7
15	12.79	32	35.7
16	13.6	33	37.7
17	14.5	34	39.9
18	15.5	35	42.20
19	16.5	40	55.324
20	17.535	50	92.59
21	18.66	60	149.4
22	19.84	70	233.7
23	21.08	80	355.3
24	22.39	90	525.8
25	23.756	98	707.3
26	25.21	98.5	720.2
27	26.76	99	733.2
28	28.37	99.5	746.5
29	30.06	100	760.0

At its freezing point, VP = 4.579 torr. At 110°C, VP = 1074.6 torr.

Graph 2: Plot the data below as scatter plot (scatter only, no line). Then determine the equation of the line for a Beer's Law plot. Figure out how to get Excel to determine the slope for you. IMPORTANT: In Excel you must enter 7.00×10^{-5} as "7.00E-5" (no spaces) or Excel will not recognize that number.

This is data to plot for Graph 2

$[MnO_4^-]$	7.00×10^{-5}	1.00×10^{-4}	2.00×10^{-4}	3.50×10^{-4}
Absorbance	0.175	0.250	0.500	0.875

Graph 3: Time should be the first column in your spreadsheet and also be the X-axis of your graph. Determine the instantaneous rate of reaction at 1300 seconds after creating the graph using the reaction rate data for the reaction $2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$

This is data to plot for Graph 3

$[N_2O_5]$	1.00	0.88	0.78	0.69	0.61
Time, seconds	0	200	400	600	800
$[N_2O_5]$	0.54	0.48	0.43	0.38	0.30
Time, seconds	1000	1200	1400	1600	2000

Graph 4: Determine the equivalence point from the titration curve of an unknown monoprotic acid. The generic reaction is $HA_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(l)} + A^-{}_{(aq)}$

This is data to plot for Graph 4

mL NaOH	**pH**	**mL NaOH**	**pH**
0.00	3.08	20.98	5.15
0.50	3.35	22.43	5.31
2.45	3.80	23.78	5.50
3.70	4.00	24.45	5.65
4.42	4.05	24.77	5.75
5.88	4.20	25.13	6.00
7.31	4.28	25.82	6.30
8.91	4.39	26.03	6.75
10.75	4.49	26.20	7.45
12.88	4.62	26.32	8.85
14.61	4.72	27.01	10.40
15.83	4.78	27.38	10.55
17.45	4.88	28.25	10.85
19.95	5.05	28.98	10.95

Graph 5: Create the “line spectra” for the element, mercury, with wavelength measured in nanometers. First make the scatter plot with wavelength as the x-axis, then switch to the column chart.

This is data to plot for Graph 5

Intensity	Mercury Wavelengths
10	404.6
50	407.7
70	435.8
100	546.0
50	576.9
50	579.0

Graph 6: Plot the data points (-40.00, -40.00), (0.00, 32.00), and (100.00, 212.00). Determine the slope and the y-intercept for the straight line using tools in Excel. Write the equation of the line and identify what it is for a chemistry class.

Data Collection:

no data to be collected

Analyses/Calculations/Results:

Graphs should go here:

Post Lab Questions:

For Graph 6, write the equation of the line and give an explanation of what it represents

DRY LAB II: PRACTICE WITH NOMENCLATURE

Objective II

Your will spend at least one lab period practicing with nomenclature. Fill out the tables below and check with the instructor periodically to make sure you're on the right track.

Introduction:

One postulate of Dalton's Theory is that atoms combine in ratios of whole numbers in units called molecules or formula units. The chemical formula represents the atoms in their ratios to each other. A compound has <u>new</u> physical and chemical properties, which are <u>not</u> like any of its constituent elements.

For example: $H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)}$

This is a chemical equation, written to show the chemical units that went into making the compound water. Both reactants are gases; it is a combustion reaction and the product is liquid water at standard thermodynamic conditions. The common name of H_2O is water but the proper IUPAC nomenclature is called dihydrogen monoxide. IUPAC or International Union of Pure and Applied Chemistry is an international federation that has worked to standardize the naming of chemical elements and compounds. By using IUPAC naming scientists all around the world can communicate chemical names and symbols even though their spoken and written language vary greatly.

You will learn how to name 4 classes of compounds: ionic, covalent, binary acids, and oxoacids. Each of the classes have their own sets of naming rules in which you will have to memorize. Just as when you learn a new language you need to learn the vocabulary and grammar in order to communicate effectively in that language. You can think of nomenclature as the language of chemistry where the polyatomic ions and element names are the vocabulary and the rules of naming is the grammar respectively. It may seem tedious to memorize these but just think of how inefficient it would be to always carry a dictionary with you and have to look things up to communicate. Knowing nomenclature by heart will help in streamlining the problem solving process in this chemistry course and beyond.

In order for compounds to be stable they have to be electrically neutral. For example, in ionic compounds such as CaO, the Ca^{2+} is balanced out by the O^{2-} (meaning 2+ + 2– = 0). Keeping this in mind will help when trying to determine the correct chemical formulas.

Ionic Compound Naming Rules

How to spot an ionic compound: cation + anion (metal + nonmetal)

1) The cation gets the element name or ion name
2) a) If the anion is monatomic, use "root" + "-ide" suffix (Table 1), b) if the anion is a polyatomic use the polyatomic ion name (Table 2)
3) If the metal is a variable-charged metal, use Roman numeral in parentheses (Table 3) to denote the charge of the metal ion, for example if charge is 3+ then (III)
4) When writing chemical formulas, the cation + anion charges must sum up to 0.

Table 1. Monatomic Anion Naming

Anion	Root	Name
O^{2-}	ox-	oxide
N^{3-}	nitr-	nitride
S^{2-}	sulf-	sulfide
P^{3-}	phosph-	phosphide
C^{4-}	carb-	carbide
I^{-}	iod-	iodide
Cl^{-}	chlor-	chloride
Te^{2-}	tellur-	telluride

Table 2. Polyatomic Anion Naming

Polyatomic Ion	Name
NO_3^{-}	nitrate
NO_2^{-}	nitrite
SO_4^{2-}	sulfate
PO_4^{3-}	phosphate
HPO_4^{3-}	hydrogen phosphate
ClO_4^{-}	perchlorate
$HBrO$	hypobromite
CO_3^{2-}	carbonate

Table 3. Roman Numerals

Number	Roman #
1	I
2	II
3	III
4	IV
5	V
6	VI
7	VII
8	VIII

Fixed Charged Metals

- All metals in Group 1A and 2A
- Al^{3+}, Zn^{2+}, Cd^{2+}, Ag^{+}

**Note: all other metals are variable-charged

Examples:

1) Name NaCl

   Strategize: First spot that NaCl is an ionic compound because Na is a metal and Cl is a nonmetal. Identify whether the metal Na is a fixed- or variable-charged metal. Since Na is in Group 1A it is a fixed-charge metal so no need for Roman numeral. Then look at the anion, Cl is a monatomic ion (single atom ion).

   Execute: The metal ion gets the element name so Na^{+} is called sodium (Rule 1), the monatomic anion Cl^- is called chloride (Rule 2a). **NaCl is sodium chloride**.

2) Name K_2SO_4
   Quick Strategize: ionic compound, K is fixed charged metal, SO_4 is actually SO_4^{2-} the polyatomic ion (many atoms ion) called sulfate (Rule 2b)

   Quick Execute: K_2SO_4 is potassium sulfate

3) Name $FeCl_3$
   Strategize: ionic compound, Fe is a variable-charged metal so we need to figure out its charge and use Roman numeral in parentheses. To figure out a metal's charge look at the nonmetal's charge (which is always known either by looking at periodic table for monatomic ions or the polyatomic ions' charge, memorize those!). Cl is monatomic anion in Group 7A so its charge is 1–, there are 3 Cl's total so total charge is 3–. Therefore, the Fe must be 3+ to keep the compound electrically neutral. Since Fe is variable-charged metal we must write iron(III) (Rule 3)

   Execute: $FeCl_3$ is iron(III) chloride

4) Name $FeCl_2$
   Quick Strategize: ionic compound, Fe is a variable-charged metal that is Fe^{2+} (since there are 2 Cl^-). Cl^- is a monatomic anion called chloride.

   Quick Execute: $FeCl_2$ is iron(II) chloride

   **NOTE: In examples 3 & 4, we can see why it is important that the Roman numeral is used. Since Fe can exist as Fe^{2+} or Fe^{3+} (variable charge), we must denote the charge when we name to distinguish between $FeCl_3$ and $FeCl_2$ which are completely different compounds. Just saying or writing "iron chloride" will not suffice.

5) Give the formula for lithium nitride
   Strategize: This is an ionic compound. First we need to figure out the charges of the cation and anion. Lithium is in Group 1A so Li^+, N is in Group 5A so N^{3-}. Remember that we need to write an electrically neutral chemical formula. Figure out how many Li and N needed to make the cation and anion "balanced". We will need 3 Li and 1 N.

   Execute: lithium nitride is Li_3N (always write the metal first)

6) Give the formula for cobalt(II) phosphate
   Quick Strategize: ionic compound, cobalt(II) is Co^{2+} and phosphate is PO_4^{3-}. Will need 3x Co^{2+} (6+ total) and 2x PO_4^{3-} (6– total).

   Quick Execute: cobalt(II) phosphate is $Co_3(PO_4)_2$

   **NOTE: Use parentheses when there are more than one polyatomic in the formula

Covalent Compound Naming Rules

How to spot a covalent compound: nonmetal + nonmetal

1) The 1st atom in formula gets element name
2) The 2nd atom gets "root" + "ide" (just like a monoatomic anion)
3) a) Use prefixes to denote how many there are of each atom (Table 4)
   b) EXCEPTION: "mono" prefix does not get used for the first element

Table 4. Prefixes

Number	Prefix
1	mon(o)
2	di
3	tri
4	tetr(a)
5	pent(a)
6	hex(a)
7	hept(a)
8	oct(a)
9	non(a)
10	dec(a)

**NOTE: the last vowel of the prefix can be dropped if the following letter in the name is a vowel.

Example: instead of pentaoxide → pentoxide

Examples:

1) Name N_2S_5
   Strategize: N is nonmetal, O is nonmetal so this is a covalent compound. There are 2 N's so use "di" (Rule 3a) and the element name (Rule 1), there are 5 S's so use "penta" (Rule 3a) and sulfide (Rule 2).

   Execute: N_2S_5 is dinitrogen pentasulfide

2) Name CO
   Quick Strategize: covalent compound, DON'T use "mono" for 1st atom C (Rule 3b), 1 O so use "mono"

   Quick Execute: CO is carbon monoxide

3) Give the formula for tetraphosphorus decoxide
   Quick Strategize: covalent compound, tetra is 4, dec is 10

   Quick Execute: tetraphosphorus decoxide is P_4O_{10}

4) Give the formula for iodine heptafluoride: IF_7

Acid Nomenclature

The definition of an acid is any compound that donates a proton, H^+, and forms H_3O^+ (hydronium ion) in water. When writing a chemical formula for an acid, the proton(s) is/are always written first. Since each proton is 1+, the number of protons will be equal to the charge of the anion of the acid. HCl (hydrochloric acid) is H^+ and Cl^- and sulfuric acid is H_2SO_4 (2 H^+ balanced by SO_4^{2-}) instead of HSO_4 for example. Hence, we can think of acids as having an ionic component (charged) even though they are considered more of a covalent compound. Thinking of acids as ionic can a be helpful when writing chemical formulas and acid-base reactions.

Note that an acid only has its acidic properties when dissolved in water. Therefore, when it is in the aqueous state it will get the acid name. If the compound is in the gaseous state, the covalent naming is used. For example HBr(*aq*) is called hydrobromic acid but HBr(*g*) is called hydrogen bromide.

Binary Acid Naming Rules

How to spot a binary acid: Formula leads with H. Binary means 2, so only 2 types of atoms in formula.

1) Use the prefix “hydro” followed by “root” + “-ic acid”

Examples:

1) Name HCl
   Strategize: The formula leads with H so it must be an acid. There are only 2 types of atoms H and Cl so it must be a binary acid. The prefix is always “hydro”, the root for chlorine is “chlor” then add “-ic acid”

   Execute: HCl is hydrochloric acid

2) Name HBr
   Quick Strategize: binary acid, root is “brom-“

   Quick Execute: HBr is hydrobromic acid

3) Give the formula for hydrosulfuric acid
   Quick Strategize: has the prefix “hydro” so must be binary acid, “sulfur-“ is root for sulfur which is in Group 6A so the charge is 2–. Therefore, 2 H^+ are needed.

4) *Quick Execute*: hydrosulfuric acid is H_2S

Oxoacid (Oxyacid) Naming Rules

How to spot an oxoacid: Formula leads with H. Contains O. All oxoacids are derived from the polyatomic anions

1) Identify the polyatomic ion that the acid is derived from.
   a) Keep the polyatomic name but "-ate" → "-ic acid"
   b) Keep the polyatomic name but "-ite" → "-ous acid"

Examples:

1) Name HNO_3
   Strategize: Formula leads with H so it's an acid. It is oxygen containing so it is an oxoacid. The polyatomic it is derived from is NO_3^-, nitrate, "-ate" → "-ic acid"

   Execute: HNO_3 is nitric acid

2) Name H_2SO_4
   Quick Strategize: oxoacid, sulfate, "-ate" → "-ic acid"

   Quick Execute: H_2SO_4 is sulfuric acid

3) Name H_2SO_3
   Quick Strategize: oxoacid, sulfite, "-ite" → "-ous acid"

   Quick Execute: H_2SO_3 is sulfurous acid

4) Give the formula for phosphoric acid
   Strategize: The name does not start with "hydro" it must be an oxoacid. The ending is "-ic acid" so the polyatomic must end in "-ate" so phosphate, PO_4^{3-}. We need 3 protons, H^+, to make the compound electrically neutral.

   Execute: phosphoric acid is H_3PO_4

5) Give the formula for hypoiodous acid
   Strategize: The name does not start with "hydro" (don't confuse with hypo) so oxoacid. The ending is "-ous" so polyatomic must end in "-ite" so hypoiodite, IO^-. Only 1 H^+ is needed to balance the charge.

   Execute: hypoiodous acid is HIO

IONIC

Fill in the following table. If the formula is given, write name; if name is given, write formula

a. $CdSO_4$	b. barium iodide
c. $FeCl_3$	d. aluminum carbonate
e. $CoCO_3$	f. lead(II) acetate
g. $Ba(NO_3)_2$	h. iron(III) hydroxide
i. $Zn(ClO_3)_2$	j. magnesium phosphate
k. $(NH_4)_2S$	l. bismuth(III) cyanide
m. SnO_2	n. potassium nitrite
o. $Al(C_2H_3O_2)_3$	p. aluminum phosphate
q. $Sb(OH)_3$	r. ammonium hydrogen phosphate
s. FeC_2O_4	t. tin(II) fluoride

COVALENT

Fill in the following. If the formula is given, write name; if name is given, write formula

a. NH_3 (+ common name)	b. hydrogen fluoride
c. N_2O_4	d. bromine pentafluoride
e. PCl_3	f. silicon tetrachloride
g. SF_6	h. carbon monoxide
i. H_2O (+ common name)	j. silicon disulfide
k. SO_3	l. hydrogen peroxide
m. P_2S_5	n. boron trifluoride

ACIDS Fill in the following. If the formula is given, write name; if name is given, write formula

Name of Acid	**Formula of Acid**	**Formula of Anion**	**Name of Anion**
Nitric			
			Sulfate
		Cl^-	
	H_2S		
			Carbonate
Hydrobromic			
	$HC_2H_3O_2$		
		SO_3^{2-}	
			Nitrite
	$HClO_3$		

ALL TOGETHER NOW!

Using IUPAC rules, write the chemical formula or name for each of the following:

1. barium sulfide	2. $Al_2(SO_4)_3$
3. manganese(III) iodide	4. P_2O_3
5. ammonium dihydrogen phosphate	6. $Zn(OH)_2$
7. carbon disulfide	8. Ag_2CrO_4
9. lead(II) sulfate	10. $Cu(C_2H_3O_2)_2$
11. magnesium carbonate	12. CoI_2
13. potassium permanganate	14. $Cu_2Cr_2O_7$
15. silver bicarbonate	16. Na_2O_2
17. bismuth(III) bromide	18. N_2O_3
19. copper(II) phosphite	20. $HClO_4$
21. chromium(III) chlorite	22. $Co(NO_2)_3$
23. ammonia	24. $BaCl_2$
25. tin(II) thiosulfate	26. $HClO$

27. copper(I) sulfite	28. H_2SO_3
29. sodium bicarbonate	30. HBr
31. nitric acid	32. CCl_4
33. sodium acetate	34. HIO_4
35. iron(III) dihydrogen phosphate	36. $HBrO_3$
37. iodic acid	38. $Cr_3(PO_4)_2$
39. mercury(II) perchlorate	40. $H_2S_{(aq)}$
41. nickel(II) borate	42. $HBrO_4$
43. cadmium thiocyanate	44. $HF_{(g)}$
45. ammonium sulfide	46. HBrO
47. hydrochloric acid	48. Hg_2Cl_2
49. hydrogen chloride	50. $Fe(CN)_3$

DRY LAB III: OXIDATION-REDUCTION REACTIONS

Objective III

Your will balance RedOx reactions during at least one lab period. Write your answers on a separate piece of paper and turn them in to the instructor by the end of the lab period.

Introduction:

Oxidation state or number is the designated number of electrons an atom has lost, gained or used to make a compound or polyatomic ion. However, oxidation number is not the same as charge!! Also, oxidation numbers must be whole numbers.

There are rules for assigning oxidation numbers: the highest rule in list applies in a conflict.

1. Free elements in their natural state are always assigned an oxidation number of 0, even diatomic elements.

2. The total of all oxidation numbers in a molecule or formula unit must be 0. Total of all oxidation numbers in a polyatomic ion must be equal to its charge.

3. Group IA and Group IIA metals are always assigned their charge number for oxidation number.

4. In covalent compounds, H is always assigned +1 and F is assigned -1 always.

5. Then assign O to be -2, if it works. There are exceptions, such as H2O2 and F2O.

6. In binary compounds with metals, Group VIIA is assigned -1, Group VIA is assigned -2, and Group VA is assigned -3. Hydrogen with a metal is assigned –1.

7. Many elements have more than one oxidation number possible. In that case, we use other atoms in the molecule, formula unit or polyatomic ion to determine what their oxidation numbers are.

For example, look at the oxidation number of Fe in these three compounds: FeO, Fe_2O_3 and Fe_3O_4. Recall that Single Replacement reactions look like A + BC → B + AC. An unbalanced single replacement reaction is given below:

$$\_\_\_Cu_{(s)} + \_\_\_AgNO_{3(aq)} \rightarrow \_\_\_Ag_{(s)} + \_\_\_Cu(NO_3)_{2(aq)}$$

Now we make it a net ionic equation:

___$Cu_{(s)}$ + ___$Ag^{+}_{(aq)}$ → ___$Ag_{(s)}$ + ___$Cu^{2+}_{(aq)}$

Look at the reactants and products and find the oxidation number for the two metals on both sides of the equation. (Cu metal is an element in its natural state, so its O.N. is 0. Ag^{+} has an O.N. of +1, while Ag metal is an element in its natural state again, so its O.N. is 0. Cu^{2+} has an O.N. of +2.)

What is happening to copper? In order to change oxidation number, it is "losing" two electrons or being oxidized. When an oxidation number increases, the process is called oxidation.

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ e^{-}$ (called the oxidation half-reaction)

What is happening to silver? Its oxidation number is decreasing and it has gained an electron. When an oxidation number decreases, the process is called reduction.

$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$ (called the reduction half-reaction)

Notice that copper is losing two electrons, but silver is only gaining one. The number of electrons lost has to equal the number of electrons gained; therefore two silver ions will gain two electrons. The balanced net ionic equation is:

$$Cu_{(s)} + 2\ Ag^{+}_{(aq)} \rightarrow 2\ Ag_{(s)} + Cu^{2+}_{(aq)}$$

This is called an oxidation-reduction reaction, because it involves both processes. The shorthand that is usually used is to say "redox" for oxidation-reduction. The main thing to remember is that REDOX REACTIONS INVOLVE THE TRANSFER OF ELECTRONS. You will learn to separate a reaction into its component half-reactions and determine which process is occurring for each half-reaction. If one "species" is oxidized, another must be reduced.

Use this memory aid: LEO says GER. (Loss of Electrons is Oxidation, Gain of Electrons is Reduction)

Other definitions: an oxidizing agent causes oxidation; therefore it is reduced. Vice versa, a reducing agent causes reduction therefore it is oxidized.

TO BALANCE REDOX REACTIONS:

Always remember that electrons lost must equal electrons gained per the Law of Conservation of Matter. There are two methods: the "oxidation number change method" and the Half-Reaction Method (also called the ion-electron method).

OXIDATION NUMBER CHANGE METHOD: You have to know if an acid or base is involved and which side of equation it's on. The steps are outlined below.

1. Identify all oxidation numbers and find which "species" changed oxidation numbers.
2. Use arrows to connect those which are changing and determine if the "species" was oxidized or reduced. Then determine the number of electrons lost or gained per atom in the "species."
3. Determine the multipliers needed for each to make electrons lost equal to electrons gained and insert the multipliers as coefficients in the reaction.
4. Check that all other atoms are balanced.

Example: $\_\_\_I_{2(s)} + \_\_\_Br_{2(aq)} + \_\_\_H_2O_{(l)} \rightarrow \_\_\_HIO_{3(aq)} + \_\_\_HBr_{(aq)}$

O.N. 0 0 +1, -2 +1,+5,-2 +1,-1

Steps 1 and 2: The iodine atom in its elemental state changed to the iodine atom in the acid with its oxidation number changing from 0 to +5. It had to lose five electrons per iodine atom to do that, therefore it was oxidized. The bromine atom in its elemental state changed to the bromide ion in the acid, HBr. Its oxidation number changed from 0 to –1, so it gained one electron per bromine atom. A gain of electrons is reduction.

Step 3 is a bit tricky for this equation. Each iodine atom needs to lose five electrons, but elemental iodine is diatomic. That means that 10 electrons have to be lost and that a coefficient of 2 has to be put in front of the HIO_3. If 10 electrons are lost, then 10 electrons have to be gained, so there has to be a total of 10 bromine atoms gaining one electron each. Since bromine is also diatomic, 10 bromine atoms will be provided by 5 Br_2. Then we need a stoichiometric coefficient of 10 in front of the HBr.

Step 4 requires a coefficient to be found to balance the H_2O. The coefficients have already been determined for the products, and the 2 HIO_3 means that there are six O atoms. Therefore, the coefficient for water will be six.

The balanced equation is:

$I_{2(s)} + 5\ Br_{2(aq)} + 6\ H_2O_{(l)} \rightarrow 2\ HIO_{3(aq)} + 10\ HBr_{(aq)}$

Review the half-reaction method that is in your lecture text. Also review an example step by step that involves an acid. Shown below is method for the half-reaction method with bases. It is all the same steps as for the half-reaction method in acid, but it is adjusted at the end to make it basic.

IN BASE: $\_\_\_MnO_4^{-}{}_{(aq)} + \_\_\_C_2O_4^{2-}{}_{(aq)} \rightarrow \_\_\_MnO_{2(s)} + \_\_\_CO_3^{2-}{}_{(aq)}$

$MnO_4^{-} \rightarrow MnO_2$
$MnO_4^{-} \rightarrow MnO_2 + 2\ H_2O$
$MnO_4^{-} + 4H^{+} \rightarrow MnO_2 + 2\ H_2O$
$MnO_4^{-} + 4H^{+} + 3\ e^{-} \rightarrow MnO_2 + 2\ H_2O$

Reduction half-reaction

$C_2O_4^{2-} \rightarrow \mathbf{2}\ CO_3^{2-}$
$C_2O_4^{2-} + 2\ H_2O \rightarrow \mathbf{2}\ CO_3^{2-}$
$C_2O_4^{2-} + 2\ H_2O \rightarrow \mathbf{2}\ CO_3^{2-} + 4H^{+}$
$C_2O_4^{2-} + 2\ H_2O \rightarrow \mathbf{2}\ CO_3^{2-} + 4H^{+} + 2\ e^{-}$

Oxidation half-reaction

Multiply reduction half-reaction by two and multiply oxidation half-reaction by three to get 6 electrons transferred. Then add up the half-reactions to make the balanced chemical equation:

$$2\ MnO_4^{-} + 3\ C_2O_4^{2} + 8\ H^{+} + 6\ H_2O + 6e^{-} \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 4\ H_2O + 12\ H^{+} + 6e^{-}$$

Cancel common species:

$$2\ MnO_4^{-} + 3\ C_2O_4^{2} + 2\ H_2O \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 4\ H^{+}$$

Now add OH^{-} equal to the number of H^{+}, but add to both sides. Then let the H^{+} and OH^{-} react to form water. Then cancel the water that appears on both sides.

$$2\ MnO_4^{-} + 3\ C_2O_4^{2-} + 2\ H_2O + 4\ OH^{-} \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 4\ H^{+} + 4\ OH^{-}$$

$2\ MnO_4^{-} + 3\ C_2O_4^{2-} + 2\ H_2O + 4\ OH^{-} \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 4\ H_2O$

$2\ MnO_4^{-} + 3\ C_2O_4^{2} + 4\ OH^{-} \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 2\ H_2O$

Disproportionation: sometimes one species is both oxidized and reduced. It may appear as a reactant or a product. Just use it twice and be sure to combine the results when done.

Methods:

PART A: Practice doing oxidation numbers: find the oxidation number of every atom in all the "species" in the following list. Show your answers on a separate piece of paper.

S_8, Al_2O_3, MnO_4^{-}, H_2O_2, CH_4, C_2H_6, C_6H_6, H_3PO_4, MgO, BaF_2, $Hg(NO_3)_2$, $Sr(ClO_4)_2$, N_2O, PCl_3, $NaHCO_3$, SO_4^{2-}, SO_3^{2-}, $SO_{3(g)}$ (not the ion)

PART B: OXIDATION-REDUCTION REACTIONS

Balance the following using the "oxidation number change method." Show your work and answers on a separate piece of paper.

1. $\_\_\_P + \_\_\_HNO_3 + \_\_\_H_2O \rightarrow \_\_\_H_3PO_4 + \_\_\_NO$
2. $\_\_\_H_2SO_4 + \_\_\_HI\ \ \_\_\_ \rightarrow H_2S + \_\_\_I_2 + \_\_\_H_2O$
3. $\_\_\_Sb + \_\_\_HNO_3\ \ \_\_\_ \rightarrow Sb_2O_5 + \_\_\_NO + \_\_\_H_2O$

4. ___NO_2 + ___H_2O → ___HNO_3 + ___NO

5. ___$K_2Cr_2O_7$ + ___H_2O + ___S → ___SO_2 + ___KOH + ___Cr_2O_3

6. ___$KMnO_4$ + ___HCl → ___Cl_2 + ___KCl + ___$MnCl_2$ + ___H_2O

Balance the following using the half-reaction method. Show your work and answers on a separate piece of paper. NOTE: all of these are initially unbalanced.

In Acid:

7. $PbO_2 + Cl^- \rightarrow Cl_2 + PbCl_2$

8. $Al + SO_4^{2-} \rightarrow Al^{3+} + SO_2$

In Base:

9. $MnO_4^- + IO_3^- \rightarrow MnO_2 + IO_4^-$

10. $N_2H_4 + Cu(OH)_2 \rightarrow N_2 + Cu$

11. $[Fe(CN)_6]^{4-} + CrO_4^{2-} \rightarrow [Fe(CN)_6]^{3-} + Cr_2O_3$

Disproportionation: in base

12. $P_4 \rightarrow H_2PO_2^- + PH_3$

DRY LAB IV: ATOMIC EMISSION SPECTRA

Objective IV

You will determine the relationship between atomic line spectra and the atomic structure.

Introduction:

Scientists have been studying the phenomenon of light for over 300 years. Light is really electromagnetic (EM) radiation, ranging from gamma rays (deadly) to radio waves (harmless - so far). EM radiation is measured by its wavelength (designated by lambda, l) and frequency (designated by nu, n). The range of wavelength goes from 1 picometer to meters long, while frequency goes from 10^{20}/sec to 10^{8}/sec. Visible light is a small portion of the EM spectrum with a wavelength range of 400 to 700 nm. Frequency and wavelength are related through the speed of light, c, which is 2.998×10^{8} m/s. We use the relationship:

$c = \lambda\nu$ in other words Speed of light = wavelength x frequency

Many of us are used to seeing continuous spectra in the visible region in the sky on a rainy day - it's called a rainbow. However, sometimes we find what are called discrete spectra or line spectra, which have only certain wavelengths of light rather than all of them.

Over 100 years ago, two scientists discovered that elements have specific and characteristic line spectra. When metal ions in solution are passed through a flame, we see a certain color. If we put that light through a prism or spectroscope, we see only certain wavelengths or colors appear. When gases in tubes are charged with electrical energy, we see a certain color, and again if we look at the light that is emitted through a prism or spectroscope, we see only certain wavelengths or colors. In both these cases, the atoms of the element are absorbing energy and then releasing energy in the form of visible light. We can actually use these line spectra to identify elements, even from distant stars.

Many scientists worked at explaining atomic line spectra. One equation we continue to use for explaining the line spectra of the Hydrogen atom is the Rydberg equation:

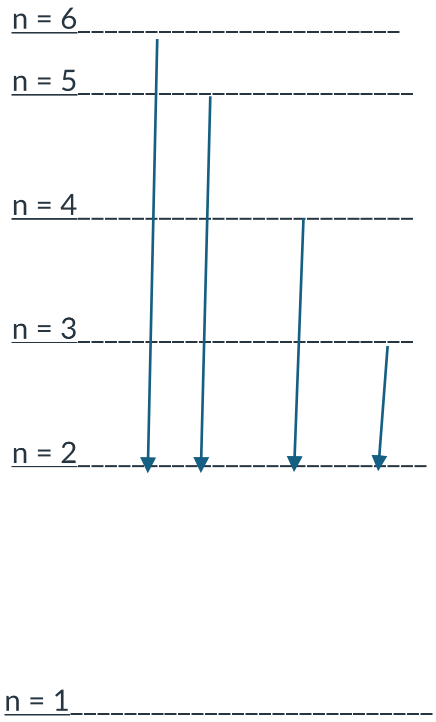
$$1/\lambda = R(1/n_f^2 - 1/n_i^2)$$

where n_f is the home energy state that the electron relaxes down to and n_i is the initial higher energy state. R is the Rydberg constant $1.097 \times 10^{7}\ m^{-1}$. Actually, n_f is always equal to 2 for the line spectra in the visible region, which leads to the Balmer series. Other n_f's are possible, and lead to other series in the ultraviolet and infrared regions.

Max Planck, Niels Bohr and other scientists theorized that the electrons in the atoms of the gases were absorbing energy from the flame or electricity and then releasing energy (DE) of only certain values. Planck had previously determined the relationship for EM radiation, which he said traveled in little packets of energy called photons:

$$\Delta E = h\nu \text{ or } \Delta E = hc/\lambda$$

where h is Planck's constant, or 6.626×10^{-34} J•s.



Balmer series: visible wavelengths as electron drops in Hydrogen atom

The Balmer series for the Hydrogen atom (3→2, 4→2, 5→2, 6→2) shows wavelengths in the visible region at 410 nm, 434 nm, 486 nm and 656 nm. It turns out that each of these relates to the release of a photon of light energy when an electron drops from a higher energy level to the 2nd energy level. Bohr used this data to create his model of the Hydrogen atom, where the electron "orbits" the nucleus at only certain allowed radii, which means that each "orbit" has its own discreet energy level or allowed energy state. Each state then has a defined ΔE from other states, and it takes this much energy to move an electron up, or the electron has to release this much energy to move down to a lower allowed energy state.

Bohr and other scientists developed the Quantum Theory of atoms, which uses a complex equation involving partial differential equations and assigned Quantum Numbers to define the probable distance and energy level for each electron in an atom.

Example

1. If ultraviolet light has a wavelength shorter than the minimum for visible light, what is its minimum frequency?

For example, if $\lambda = 350$ nm, then $\nu = c/\lambda = 2.998\text{x}10^{8}\ \text{m/s} / 3.50 \text{ x } 10^{-7}\ \text{m} = 8.57 \text{ x } 10^{14}\ \text{s}^{-1}$ or Hertz

What is the energy of the wavelength above?

$E = hv = 6.626\text{x}10^{-34}\ \text{J}\bullet\text{s}\ (8.57\text{x}10^{14}\ \text{s}^{-1}) = 5.68\text{x}10^{-19}\ \text{J}$

2. A Hydrogen atom has an electron excited up to the 7th level, and it drops back to the 2nd level, releasing a photon. Find the wavelength of the photon released and find the energy of it.

$1/\lambda = R(1/n_f^2 - 1/n_i^2) = 1.097\text{x}10^{7}\ \text{m}^{-1}\ (1/2^2 - 1/7^2) = 1.097 \text{ x } 10^{7}\ \text{m}^{-1}\ (0.2295918) = 2.5186 \text{ x } 10^{6}\ \text{m}^{-1}$

Then $= (2.5186 \text{ x } 10^{6}\ \text{m}^{-1})^{-1} = 3.97 \text{ x } 10^{-7}$ m or 397 nm.

$E = hv = hc/\lambda = 6.626\text{x}10^{-34}\ \text{J}\bullet\text{s}\ (2.998 \text{ x } 10^{8}\ \text{m/s})/3.97 \text{ x } 10^{-7}\ \text{m} = 5.00 \text{ x } 10^{-19}\ \text{J}$

PRELABORATORY QUESTIONS

1. List the colors of visible light from 400 nm to 700 nm.

2. As seen in the movie "Phenomenon" scientists think earthquakes are preceded by extra-low frequencies, around 100 s^{-1} or Hertz. What wavelength would this have? Calculate its energy. Is it high energy or low energy?

Safety

Wear appropriate PPE (see page *i*). Wear gloves while doing flame test portion.

Materials

Chemicals: about 50 mL each of 0.1 M aqueous solutions of chlorides of Cu^{2+}, Ca^{2+}, Ba^{2+}, K^+, Na^+; 50 mL of 6 M HNO_3, deionized water

Equipment: flame test loops (one per metal ion solution), Bunsen burners, beakers to hold metal ion solutions; flashlights or else borrow the white light source from Physics for the night labs or rainy days.

Disposal: The last group to do the flame tests should clean up. The beakers containing metal ion solutions are probably cross-contaminated and should be poured down the drain with a water flush. The 6 M HNO_3 used to clean the flame test loops needs to be poured in a **large** beaker of water and neutralized **slowly** with 6 M NaOH. It can then be poured down the drain.

Methods

There should be lamps available with tubes filled with various gases, and diffraction gratings which you will use to look at each gases' atomic line spectra. While doing so, answer the following questions on your lab report form.

For each gas or light listed below:

A) What overall color do you see when current is passed through the tube?

1. ordinary light bulb filament
2. hydrogen gas
3. mercury vapor
4. neon gas
5. oxygen gas
6. nitrogen gas
7. helium gas
8. daylight (sunlight)

B) What lines of color do you see through the diffraction grating or spectroscope for each. Identify each line as intense or dim.

1. ordinary light bulb filament
2. hydrogen gas
3. mercury vapor
4. neon gas
5. oxygen gas
6. nitrogen gas
7. helium gas
8. daylight (sunlight)

C) Explain the difference between a continuous spectrum and a discreet spectrum.
D) Now go back and divide the light sources above into continuous and discreet.

Where a discreet spectrum was observed above:

1) How were the observed wavelengths of light produced by the atoms of gas?

2) Why are the emitted light waves for each element of such very specific wavelengths? (Discuss in terms of electron energy levels.)

E) For the hydrogen atom, visible wavelengths are emitted by atoms whose electrons are excited into higher energy levels, which then drop back to the second energy level. If an electron were excited to energy level six, how many visible wavelengths of light would you expect to see? Show a picture to determine this.
F) Your instructor will assign you an energy level in a hydrogen atom.

1. Calculate the visible wavelength of light emitted as the electron drops back to the second energy level.
2. Then calculate the energy released by that electron.
3. How much energy would be released if a mole of hydrogen atoms did this?

G) Why does each element have its own unique atomic line spectrum, when all atoms have energy levels labeled 1s, 2s, 2p, 3s, etc.?

H) Now go to one of the Flame Test Stations set up for you in lab. Wearing your safety goggles, have one member of your group test each of the solutions by the following procedure. Record the solution formulas (should be: Cu^{2+}, Ca^{2+}, Ba^{2+}, K^{+}, Na^{+}) and the color(s) you see in the flame test. (NOTE: only one solution should be a yellow-orange color. If others show this, it is because of sodium contamination of the solutions.) Dip the platinum flame test loop in the 6-M nitric acid cleaning solution provided and then burn the liquid off in the burner flame. Now put a drop of the test solution on the flame test loop. Hold the loop just above the blue cone of the burner flame and observe very quickly and carefully what color(s) you see. As a challenge, you should try looking through a diffraction grating while someone in your group holds the flame test loop of each metal ion in the flame. Again, look quickly and observe the various lines. Try to report at least the dominant colors of lines you see.

Disposal: The last group to do the flame tests should clean up. Check with your instructor. The beakers containing metal ion solutions are probably cross-contaminated and should be poured down the drain with a water flush. The 6 M HNO_3 used to clean the flame test loops needs to be poured in a **large** beaker of water and neutralized **slowly** with 6 M NaOH. It can then be poured down the drain.

Post Lab Questions

1. If you know that one of the lines in the spectrum for boron ion is a wavelength of 518.0 nm, what energy does that represent? Calculate it.

2. If an electron in hydrogen is excited to energy level 7 and drops back to energy level 3, would you be able to observe it using the diffraction grating? Prove your answer with a calculation of its wavelength.

Discussion:

Conclusion:

Error Analysis:

DRY LAB V: MOLECULAR STRUCTURES AND THEORIES

Objective V

To draw Lewis electron dot structures and make models of molecules. To determine electron pair geometry, molecular geometry, bond angles, dipole moment, and hybridization of central atoms.

Introduction:

Covalent molecules follow certain "rules" for forming bonds, and they can be represented with the Lewis electron dot structures.

RULES for writing Lewis Electron Dot structures for covalent compounds:

1. Write the Lewis structure for each atom, then total the number of valence electrons. CO_2 would have a total of 16 valence electrons to use in its structure as bond (shared) pairs and lone pairs.
2. Put the atom that is furthest to the left and/or lowest in the periodic table as the central atom. In CO_2, C is central. Arrange the other atoms around it symmetrically. Then draw one single covalent bond to each, either by putting in two dots to represent the shared pair of electrons that make the bond or by drawing a single line, which also represents the bond. Subtract the total number of electrons in bonds from the total found in step 1. For CO_2, 16 - 4 = 12.
3. Put the remaining electrons around the terminal atoms as lone pairs of electrons, except for H. (Why?) If there are still some electrons left after the octet rule is satisfied for the terminal atoms, the leftover electrons go on the central atom as lone pairs.
4. See if each atom has satisfied the octet rule. In CO_2, the Oxygen atoms are happy but the Carbon atom is not. You must take one lone pair from an Oxygen atom and make it a bond pair between that Oxygen atom and the Carbon atom. You take from other terminal atoms symmetrically until the octet rule is satisfied. Don't go beyond the octet rule. (This is what makes <u>multiple bonds</u> - double or triple is possible, quadruple is not!)
5. Check again that you have a total of 16 valence electrons in the structure and that all atoms have satisfied the octet rule.
6. If you are drawing a structure for a polyatomic ion, you must account for the charge by adding in extra electrons in step 1 if it is an anion, or subtracting if it is a cation, like

NH_4^+. A large bracket must be placed around the structure and the charge indicated outside the bracket.

The octet rule is not always followed; there are exceptions. Some atoms in a molecule or polyatomic ion will be electron deficient, such as N in NO_2. Some atoms will have more than eight electrons around them; these are the larger atoms and can have expanded valence shells.

Sometimes it is difficult to tell which is the best Lewis structure when multiple bonds are involved, or when you are unsure which is the central atom. Rule #2 above should decide that most of the time. For example with HCN, Carbon is left of Nitrogen, so it must be central. With multiple bonds, you should consider how many single bonds each atom would usually make. See the table below.

Preferred number of bonds for each group number.

Group Number	**IVA**	**VA**	**VIA**	**VIIA**	**VIIIA**
No. of bonds preferred	4	3	2	1	0

This is just a starting point. Once you are in Period 3 or beyond, other arrangements are possible. Also remember that H and F always make only one covalent bond.

You must also use *Valence Shell Electron Pair Repulsion Theory* to determine the geometry of the electron pairs around a central atom and the molecular geometry around a central atom. In this theory and in valence bond theory, primary covalent bonds are called sigma bonds, represented by the Greek s, and second and third covalent bonds are called pi bonds, or p. The VSEPR geometry is determined by the total of the number of sigma bonds and the number of lone pairs of electrons that are around a central atom.

Summary of VSEPR

The elctronic and molecular geometry.

TYPE	**EP Geometry**	**Molecular Geometry**	**Bond Angles, °**	**Examples**
AB_2	LINEAR	LINEAR	180	CO_2, HCN
AB_2E	TRIG PLANAR	BENT	120	SO_2, $SnCl_2$
AB_2E_2	TETRAHEDRAL	BENT	109	H_2O, H_2S
AB_2E_3	TRIG BIPYR	LINEAR	180	XeF_2, I_3^-
AB_3	TRIG PLANAR	TRIG PLANAR	120	H_2CO, CO_3^{2-}
AB_3E	TETRAHEDRAL	TRIG PYR	109	NH_3, $SOCl_2$
AB_3E_2	TRIG BIPYR	T-SHAPED	90, 180	ClF_3
AB_3E_3	OCTAHEDRAL	T-SHAPED	90, 180	
AB_4	TETRAHEDRAL	TETRAHEDRAL	109	NOF_3
AB_4E	TRIG BIPYR	SEE-SAW	90,120,180	SeF_4
AB_4E_2	OCTAHEDRAL	SQUARE PLANAR	90, 180	ICl_4^-
AB_5	TRIG BIPYR	TRIG BIPYR	90,120,180	PCl_5
AB_5E	OCTAHEDRAL	SQUARE PYR	90, 180	$XeOF_4$
AB_6	OCTAHEDRAL	OCTAHEDRAL	90, 180	SF_6

A represents the central atom, B is a bonded atom and E is a lone pair of electrons.

Valence bond theory predicts that central atoms actually undergo a hybridization of their atomic orbitals into equal shape and energy hybrid orbitals when making bonds with other atoms. Use the table below to predict the hybridization of central atoms be drawing the Lewis structure and then adding up the number of s bonds and lone pairs on the central atom.

Hybridization of Orbitals

Hybridization depending on the electronic geometry

Hybrid type	No. of s bonds and LP's	e- pair geometry	Bond angle	Examples
sp	2	Linear	180	C_2H_2
sp^2	3	Trig. planar	120	H_2CO,BF_3
sp^3	4	Tetrahedral	109.5	CH_4
dsp^3	5	Trig bipyr	90,120	PCl_5
d^{2}sp^3	6	Octahedral	90	SF_6

Electronegativity is a relative measurement of the attraction of an atom in a covalent bond for the electrons in that bond. We use electronegativity differences to determine if a covalent bond is a polar covalent bond, i.e., is one end of the bond slightly negative and the other end slightly positive. In general, if the difference is between 0.5 and 2.0, we say the bond is a polar covalent bond. If it is below 0.5, we say the bond is nonpolar covalent. If it is greater than 2.0, the bond is probably an ionic bond. See your textbook or large periodic table for electronegativity values to use in this lab.

Dipole moment: Molecules with polar covalent bonds can, by their shape, have a dipole moment, unless the polarity vectors cancel each other. If the atoms or groups of atoms attached are the same and are in a symmetric arrangement, there is no net dipole moment, and the molecule is nonpolar (CCl_4, PCl_5, SF_6). If they're not the same atoms or group of atoms and they are not symmetric, they will probably have a dipole moment (BrF_3, OF_2, NH_3, SF_4)

Examples

Draw the Lewis electron dot structure for $CFCl_3$ using the rules above. There are 32 valence electrons. Carbon is to the left and it likes to make more bonds than the halogens, so it will be the central atom. The four halogens are arranged symmetrically around the carbon, with one bond to each for a total of eight electrons. That leaves 24 electrons to go on the terminal atoms as lone pairs. These are used up.

If you made the model, you would notice that its molecular geometry was tetrahedral. There are no lone pairs on the central atom, so the electron pair geometry is also tetrahedral. Carbon is making four single bonds, which are s bonds, and Carbon has no lone pairs. That means the hybridization is sp^3. The C-Cl bonds have an electronegativity difference of 0.5, with Cl having more attraction for the electrons in the bond; the C-F bond is very polar with an electronegativity difference of 1.5. Since the molecule is not symmetrical, it will have a dipole moment, so we say it is a polar molecule.

Materials:

model kits or online simulators

Methods:

Part A:

First you will draw the Lewis structures for the molecules listed. Then use the model kits to make the covalent molecules. (See the rules on the model kits for determining which color is which atom.) Compare the completed model with your drawing. Does the model show the double or triple bonds if they were in the Lewis structure? You should use the long flexible bonds to make multiple bonds. Use the shorter stiff bonds for single bonds.

Write all your answers on blank paper and turn in at end of lab period. Your paper should have for each molecule or polyatomic ion:

the Lewis structure

the 3-D structure

the number of s bonds on the central atom(s) the number of lone pairs on the central atom(s)

the hybridization of the central atom(s)

the electron pair geometry

the molecular geometry

identification of the molecule as polar or nonpolar

answers to questions for each

Your instructor will demonstrate how to draw 3-D structures. Put as much of the molecular model in the same plane as possible. Continue to use atomic symbols to represent the atoms. Use straight lines to indicate bonds that are in the same plane. Use elongated triangles to represent bonds to atoms above the plane and use disappearing wedged lines to indicate bonds to atoms below the plane

1. Diatomic fluorine.
2. a. CF_4, b. NF_3 and c. OF_2. Why do carbon, nitrogen and oxygen make different numbers of bonds with fluorine?
3. a. CO_2, b. SO_2. How do you know when to make double or triple bonds?
4. a. N_2H_4 and b. N_2O_4. What do you notice about the bonding in these two molecules?
5. a. $AsCl_3$ and b. $AsCl_5$. What is different between them? What is different about the second one compared to all the other molecules above?
6. a. H_2O_2 and b. C_2H_2. What do you notice about the bonding in these two molecules? Notice there is more than one central atom.
7. a. $TeCl_6$ and b. XeF_4. What is different about these?
8. a. NO_2^- and b. NH_4^+.

Questions for Part A:

1. Draw the Lewis structures for SiOS and CIN. How do you know which is the central atom and how many bonds to make to the terminal atoms?

2. What is the difference between empirical formula and molecular formula? Why do we need both?

Part B:

You will draw Lewis structures and make models of organic compounds, to understand the groupings of organic molecules and the functional groups that can be attached. First then, for each step below, you draw the Lewis structures for the molecules named. Then use the model kits to make the covalent molecules. (See the rules on the model kits for determining which color is which atom.) Compare the completed model with your drawing. Does the model show the double or triple bonds if they were in the drawing?

Write all your answers on blank paper and turn in at end of lab period. Your paper should have the Lewis structures, the 3-D structures, the hybridization, the molecular geometry, and answers to questions for each number.

1. Make CH_4, C_2H_6, C_3H_8, and C_4H_{10}. Is there another way you could have put the C_4H_{10} molecule together? (HINT: look up the word *isomer*.)

2. Make C_2H_4, C_2H_2, C_3H_6, C_3H_4, C_4H_8, and C_4H_6. Is there another way you could have put the last two molecules together?

3. Make C_2H_5OH and CH_3OCH_3. Would their chemical and physical properties be different? What is making the difference? Name their functional groups.

4. Make C_2H_5OH and CH_3COOH. Would their chemical and physical properties be different? What is making the difference? Name their functional groups.

5. Make $CH_3COOC_2H_5$. "React" this with a water molecule, without adding any more atoms or taking any away. It should make an organic acid and an alcohol. Name them.

6. Draw Lewis structures first; then make four isomers of $C_3H_6Cl_2$.

7. Draw and make benzene, C_6H_6, which has a ring structure, but still follows all the Lewis rules. Keep your benzene ring together. Make cyclohexane, C_6H_{12}. How does it compare to benzene?

Questions for Part B:

1. Define molecular formula, structural formula and condensed structural formula. Choose an organic molecule to demonstrate each type of formula.
2. What is an isomer?
3. Look up the difference between structural and geometric isomers and explain here.
4. How many compounds can you make that have just one carbon and one oxygen and as many hydrogens as needed for each? Draw them all. Can you name them?

Discussion:

Conclusion:

APPENDIX A: Statistics and Data Analysis

Accuracy: how close a measurement or result is to its true value.

Precision: reproducibility of replicate measurements.

Determinate (Systematic) Error: is a result of some flaw in a measurement, including instrument error and operator error.

Indeterminate (Random) Error: is associated with every measurement because of the estimate made for the last digit in any measurement.

Average value: $x = x_i/n$, the sum of the measurements taken divided by the number of measurements taken. The average is sometimes called the mean. It is not the median.

Median value: in a list of measurements, the median is the midpoint on the list. If the number of data points is even, then the median is between the two midpoints on the list.

Using *the Q-test* to find the outliers in data. A data point may look like it is quite far from the rest of the data points in its value and you may want to omit it when analyzing the data. The Q-test must be done if there are a small number of measurements, as there usually are in general chemistry labs. The data point can be rejected if the quantity, Q, defined by the equation below is greater than the critical Q_c for the specific number of measurements. See table of critical Q_c values below.

$$Q = \frac{|\text{suspected value – value closest to it}|}{\text{Highest value – lowest value}}$$

Table of N and Q Values

N	3	4	5	6	7	8	9	10
Q_c	.94	.76	.64	.56	.51	.47	.44	.41

Standard Deviation, $= [\ S(x - x_i)^2/(n-1)]^{1/2}$

Average Deviation, $x_{avg} = (1/N)\ _{t=1}{}^{N}\ |x_i - x|$

Relative Deviation has to do with how the data are arranged around the average. It takes several steps to find relative deviation, starting with finding the average value for the data. The average value has to be based on good data, i.e., outliers have already been eliminated. The factor is 1000 so that relative deviation is reported in parts per thousand.

Relative deviation, $d = 1000 \text{ x average of } |x - x_i|/x$

Confidence Limits: these indicate the precision of the data collected. The value calculated here is called the 90% confidence limit or interval.

90% Confidence Limit $= t/N^{1/2}$

The t value varies with the number of data points, as shown in the table below.

Table of N and t values

The t value	6.314	2.920	2.353	2.132	2.015
N	2	3	4	5	6

Results are generally reported as the average, x, plus or minus the 90% confidence limit.

Percent Error: $\%\text{-error} = \frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100$

Experimental Uncertainty is determined by taking into account the indeterminate error associated with each measurement in a calculation. For example, when determining the density of an object, both the mass and the volume are measured and have some uncertainty. The density range must be calculated based on the greatest mass and lowest volume compared to the least mass and greatest volume. The final result is given with the actual result plus/minus the difference to the greatest and least densities calculated.

Example: a student collected the following data for density calculations of water at 25°C. She used a 10.00-mL with a pipette each time with a relative uncertainty of 0.01 mL. She used a three-place balance with a relative uncertainty of 0.001 g. Her density results were: 0.9970 g/mL, 0.9955 g/mL, 0.9939 g/mL, 0.9991 g/mL, and 0.9988 g/mL.

The average value, x = 0.99686 g/mL, which can be reported to four significant figures at 0.9969 g/mL.

The Q-test result indicates that there are no outliers, with a Q of 0.3077 vs. a Q_c of 0.64.

The standard deviation, = 0.002203.

The average deviation, x_{avg} = 0.001728.

The relative deviation, d = 1.733 ppt.

The 90% confidence limit turns out to be 0.002100.

The percent error from the true value of 0.997044 g/mL is 0.01845 %.

If the student applied the relative uncertainties in the measurements of volume and mass, the range of density would be 0.9929 g/mL to 1.000 g/mL, so she would report the average plus or minus this range: 0.9969 g/mL ± 0.00395 g/mL.

Many times you must report the results as the average plus or minus the 90% confidence limit. If that were the case here, she would report 0.9969 g/mL + 0.002100 g/mL

APPENDIX B: Conversion Factors and Scientific Constants

MASS

453.6 g = 1 lb **

1 g = 0.03527 oz.

1 kg = 2.205 lb

1 metric ton = 1 x 10^6 g

LENGTH

2.540 cm = 1 in **

1 m = 39.37 in

1 angstrom = 0.1 nm = 1 x 10^{-8} m

~20 drops = 1 mL

VOLUME

1 cm^3 = 1 mL **

1 L = 1.0567 qt **

1 gal = 4 qt

PRESSURE

1 atm = 760 torr = 760 mm Hg = 14.696 lb/in^2 = 101.325 kPa **

1 Pa = 1 N/m^2

TEMPERATURE

$^oF = 1.8\ ^oC + 32$ **

$K = {^oC} + 273.15$ **

ENERGY

4.184 J = 1 cal **

1 J = $9.48\text{x}10^{-4}$ BTU

1 L • atm = 101.3 J = 24.2 cal

1 eV = 1.602×10^{-19} J

1 kw • hr = 3413 BTU = 3.601×10^{6} J

CONSTANTS

Velocity of light in a vacuum, c = 2.9979×10^{8} m/s **

Ideal Gas Constant, R = 0.082057 L • atm/mol • K = 8.3145 J/mol • K **

Avogadros' Number, $N_a = 6.0221 \times 10^{23}$ units/mol **

Planck's constant, $h = 6.6261 \times 10^{-34}$ J • s/photon

Faraday constant, F = 96485 C/mol e-

Charge on an electron, 1.602×10^{-19} Coulombs

Mass of an electron, 9.109×10^{-28} grams

Mass of a proton, 1.6726×10^{-24} grams

Mass of a neutron, 1.6749×10^{-24} grams

Atomic mass unit, amu, 1.660×10^{-24} grams

MEMORIZE THESE

APPENDIX C: COMMON NAMES AND IUPAC NOMENCLATURE

Table of common names, IUPAC nomenclatures, and chemical formulas.

Common Name	Chemical Name	Formula
Alcohol	ethanol (ethyl alcohol)	CH_3CH_2OH
Aqua regia	mixture of conc. Nitric + Hydrochloric acids	HNO_3 + 3 HCl by volume
Aspirin	acetylsalicylic acid	$CH_3COOC_6H_4COOH$
Baking soda	sodium hydrogen carbonate	$NaHCO_3$
Banana oil	Pentyl acetate (n-amyl acetate)	$CH_3COOC_5H_{11}$
Bauxite	hydrated aluminum oxide	$Al_2O_3 \cdot xH_2O$
Bleaching powder	calcium chloride hypochlorite	$Ca(ClO)_2$, $Ca(ClO)Cl$
Blue vitriol	copper (II) sulfate pentahydrate	$CuSO_4 \cdot 5H_2O$
Borax	sodium tetraborate decahydrate	$Na_2B_4O_7 \cdot 10H_2O$
Brimstone	sulfur	S_8
Calamine	zinc oxide	ZnO
Calcite	calcium carbonate	$CaCO_3$
CalgonÔ	polymer of sodium metaphosphate	$(NaPO_3)_n$
Calomel	mercury (I) chloride	Hg_2Cl_2
Carborundum	silicon carbide	SiC
Caustic soda	sodium hydroxide	$NaOH$
Chalk	calcium carbonate	$CaCO_3$
Chile saltpeter	sodium nitrate	$NaNO_3$

Copperas	iron (II) sulfate heptahydrate	$FeSO_4 \cdot 7H_2O$
Cream of tartar	potassium hydrogen tartrate	$KHC_4H_4O_6$
DDT	dichlorodiphenyltrichloroethane	$(C_6H_4Cl)_2CHCCl_3$
Dextrose	glucose	$C_6H_{12}O_6$
Epsom salts	magnesium sulfate heptahydrate	$MgSO_4 \cdot 7H_2O$
Fool's gold	iron pyrites	FeS_2
Freon	dichlorodifluoromethane	CCl_2F_2
Glauber's salt	sodium sulfate decahydrate	$Na_2SO_4 \cdot 10H_2O$
Glycerin	propan-1,2,3-triol (glycerol)	$CH_2OHCHOHCH_2OH$
Green vitriol	iron (II) sulfate heptahydrate	$FeSO_4 \cdot 7H_2O$
Gypsum	calcium sulfate dihydrate	$CaSO_4 \cdot 2H_2O$
Hypo	sodium thiosulfate pentahydrate	$Na_2S_2O_3 \cdot 5H_2O$
Invert sugar	mixture of glucose and fructose	$C_6H_{12}O_6$ (diff forms)
Laughing gas	dinitrogen oxide (nitrous oxide)	N_2O
Levulose	fructose	$C_6H_{12}O_6$
Limewater	calcium hydroxide	$Ca(OH)_2$
Lye	sodium hydroxide	$NaOH$
Magnesia	magnesium hydroxide	$Mg(OH)_2$
Marble	calcium carbonate	$CaCO_3$
Marsh gas	methane	CH_4
Milk of lime	calcium hydroxide	$Ca(OH)_2$
Milk of magnesia	magnesium hydroxide	$Mg(OH)_2$
Milk sugar	lactose	$C_{12}H_{22}O_{11}$

Mohr’s salt	iron (III) ammonium sulfate hexahydrate	$Fe(NH_4)(SO_4)_2 \cdot 6H_2O$
Moth balls	naphthalene	$C_{10}H_8$
Muriatic acid	hydrochloric acid	$HCl_{(aq)}$
Oil of vitriol	sulfuric acid	H_2SO_4
Oil of wintergreen	methyl salicylate	$C_6H_4(OH)COOCH_3$
Oleum	sulfur trioxide (fuming sulfuric acid)	SO_3 ($H_2S_2O_7$)
Paris green	copper (II) acetate + copper (II) arsenite	$Cu(C_2H_3O_2)_2 + 3\ Cu(AsO_2)_2$
Plaster of Paris	calcium sulfate hemihydrate	$CaSO_4 \cdot \frac{1}{2}\ H_2O$
Potash	potassium carbonate	K_2CO_3
Quartz	silicon dioxide	SiO_2
Quicklime	calcium oxide	CaO
Rochelle salt	potassium sodium tartrate	$KNaC_4H_4O_6$
Rouge	iron (III) oxide	Fe_2O_3
Sal ammoniac	ammonium chloride	NH_4Cl
Salt (table salt)	sodium chloride	$NaCl$
Saltpeter	potassium nitrate	KNO_3
Silica	silicon dioxide	SiO_2
Sugar (table sugar)	sucrose	$C_{12}H_{22}O_{11}$
Teflon	polymer of tetrafluoroethylene	$(C_2F_4)_n$
Washing soda	sodium carbonate decahydrate	$Na_2CO_3 \cdot 10H_2O$
White lead	basic lead (II) carbonate	$PbCO_3 \cdot Pb(OH)_2$
Wood alcohol	methanol (methyl alcohol)	CH_3OH

APPENDIX D Density of Water Vs Temperature:

Table of density of water vs temperature.

T, $^{\circ}$C	D, g/mL	T, $^{\circ}$C	D, g/mL	T, $^{\circ}$C	D, g/mL	T, $^{\circ}$C	D, g/mL
0.00	0.99987	15	0.999099	22	0.997770	30.00	0.99567
4.00	1.000000	16	0.998943	23	0.997538	40.00	0.99224
6.00	0.99997	17	0.998774	24	0.997296	50.00	0.98807
8.00	0.99988	18	0.998595	25	0.997044	60.00	0.98324
10.00	0.99973	19	0.998405	26	0.996783	70.00	0.97781
12.00	0.99952	20	0.998203	27	0.996512	80.00	0.97183
14.00	0.99927	21	0.997992	28	0.996232	90.00	0.96534

Density of Some Organic Compounds at 20.0°C in Grams/Milliliter:

Density of Some Organic Compounds at 20.0 C in Grams/Milliliter

Chloroform, $CHCl_3$	1.492
Diethyl Ether, $CH_3CH_2OCH_2CH_3$	0.714

Ethanol, CH_3CH_2OH	0.7893
2-propanol, $CH_3CHOHCH_3$	0.785
Toluene, $C_6H_5CH_3$	0.867
Acetone, CH_3COCH_3	0.79
Cyclohexane, C_6H_{12}	0.78
Hexane, C_6H_{14}	0.66
Methanol, CH_3OH	0.79
Naphthalene,	1.15
Ethyl acetate, $CH_3COCH_2CH_3$	0.901
Cyclohexene, C_6H_{10}	0.81
1-Hexene, C_6H_{12}	0.67

Density of Some Metals at 25^{o}C in Grams/Milliliter:

Density of Some Metals at 25 C in Grams/Milliliter

Al	Cd	Cu	Fe	Pb	Ni	Zn	Hg*	Au	Ag	Sn	Mg
2.70	8.65	8.96	7.87	11.34	8.90	7.13	13.6	19.32	10.5	7.31	1.73

*Hg is a liquid at this temperature.

Density of Some Compounds at 20oC

GASES, D in g/L

O_2	H_2	NaCl
.084	.021	2.16

SPECIFIC HEAT CAPACITY OF VARIOUS ELEMENTS AND COMPOUNDS, J/g^{o}C

Ag	.235
Al	.903
As	.329
Au	.129
$C_{(graphite)}$	.711
Cr	.447
Cu	.385
Fe	.4498
Hg	.138
Pb	.128
Rh	.243
Sn	.227
U	.329
Zn	.388
H_2O	4.1796
CH_3CH_2OH	2.45
olive oil	2.0
$CaCO_3$	.821

APPENDIX E: Solubility Tables and Rules

Solubility table

IONS	Na^+, K^+	NH_4^+	Li^+	Ag^+	Hg_2^{2+}	Hg^{2+}	Pb^{2+}	Cu^+	Cu^{2+}	Ba^{2+}	Ca^{2+}	Bi^{3+}	Sn^{4+}	Sr^{2+}
CH_3COO^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Br^-	S	S	S	I	I	S	I	I	S	S	S	S	S	S
CO_3^{2-}	S	S	I	I	I	I	I	I	I	I	I	I	I	I
ClO_3^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Cl^-	S	S	S	I	I	S	I	I	S	S	S	S	S	S
CrO_4^{2-}	S	S	S	I	I	I	I	I	S	I	S	I	I	I
CN^-	S	S	I	I	I	I	I	I	I	I	I	I	I	I
F^-	S	S	S	S	I	I	I	I	I	I	I	I	I	I
OH^-	S	S	S	I	I	I	I	I	I	S	I	I	I	S
I^-	S	S	S	I	I	S	I	I	S	S	S	I	I	S
NO_3^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
NO_2^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
$C_2O_4^{2-}$	S	S	I	I	I	I	I	I	I	I	I	I	I	I

O^{2-}	S	S	S	I	I	I	I	I	I	S	I	I	I	S
ClO_4^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
MnO_4^-	S	S	S	S	S	S	S	S	S	S	S	S	S	S
PO_4^{3-}	S	S	I	I	I	I	I	I	I	I	I	I	I	I
SO_4^{2-}	S	S	S	s	S	I	I	S	S	I	S	S	S	I
S^{2-}	S	S	S	I	I	I	I	I	I	S	S	I	I	S
SO_3^{2-}	S	S	I	I	I	I	I	I	I	I	I	I	I	I
SCN^-	S	S	S	I	I	S	I	S	S	S	S	S	S	S
$S_2O_3^{2-}$	S	S	S	I	S	S	I	S	S	S	S	S	S	S

S = soluble, s = slightly soluble, I = poorly soluble

Rules for Water-Soluble Salts:

All Na^+, K^+ and NH_4^+ salts are soluble.

All CH_3COO^-, NO_3^-, ClO_3^-, ClO_4^- and MnO_4^- salts are soluble.

All Cl^-, Br^- and I^- salts are soluble, EXCEPT those with Ag^+, Hg_2^{2+}, Pb^{2+}, Cu^+, and Tl^+. Iodides of Bi^{3+} and Sn^{4+} are insoluble.

All SO_4^{2-} salts are soluble EXCEPT those with Ba^{2+}, Sr^{2+}, Pb^{2+}, and Hg^{2+}.

All Group IA and Group IIA HCO_3^- salts are soluble.

Many SCN^- and $S_2O_3^{2-}$ salts are soluble, EXCEPT Ag^+ and Pb^{2+} for both anions; Hg_2^{2+} is not soluble with SCN^-.

Rules for Poorly Soluble Salts in Water:

All F^- salts are insoluble EXCEPT Ag^+, NH_4^+ and Group IA cations.

Most salts of O^{2-} and OH^- are poorly soluble, EXCEPT those of the Group IA cations, Ba^{2+}, Sr^{2+} and NH_4^+. $Ca(OH)_2$ is slightly soluble. If an oxide salt is soluble, it produces OH^- in water.

Most salts of S^{2-} are poorly soluble, EXCEPT those with NH_4^+, Al^{3+}, Group IA cations and Group IIA cations.

Most salts of CrO_4^{2-} are poorly soluble, EXCEPT those with NH_4^+, Ca^{2+}, Cu^{2+}, Mg^{2+}, and Group IA cations.

Most salts of CO_3^{2-}, PO_4^{3-}, BO_3^{3-}, AsO_4^{3-}, AsO_2^-, CN^-, $C_2O_4^{2-}$, and SO_3^{2-} are poorly soluble, EXCEPT those with NH_4^+ and Group IA cations.

Most salts of SiO_3^{2-} are poorly soluble EXCEPT those with Group IA cations.

Organic Compounds Solubility In Various Solvents:

Organic Compounds Solubility In Various Solvents

	WATER	ETHANOL	ACETONE	BENZENE
Acetone	s	s	-	s
Acetamide	s	s		
Acetanilide		s	s	s
Anthracene			s	s
Benzamide	s	s		
Benzoic acid	(ss)	s	s	s
2-Butanone	s	s	s	s
Cyclohexane		s	s	s
Cyclohexene		s	s	s
Ethanol	s	-	s	s

Ethyl acetate	s	s	s	s
Heptane		s	s	s
n-Hexane		s		
1-Hexene		s		s
Methanol	s	s	s	s
Naphthalene		s	s	s
1-Propanol	s	s	s	s
2-Propanol	s	s	s	s

APPENDIX F: Vapor Pressure Of Water

Table of the vapor pressure of water at different temperatures.

Temperature, °C	Pressure, torr	Temperature, °C	Pressure, torr
5	6.54	30	31.824
10	9.209	31	33.7
15	12.79	32	35.7
16	13.6	33	37.7
17	14.5	34	39.9
18	15.5	35	42.20
19	16.5	40	55.324
20	17.535	50	92.59
21	18.66	60	149.4
22	19.84	70	233.7
23	21.08	80	355.3
24	22.39	90	525.8
25	23.756	98	707.3
26	25.21	98.5	720.2
27	26.76	99	733.2
28	28.37	99.5	746.5
29	30.06	100	760.0

At its freezing point, VP = 4.579 torr. At 110°C, VP = 1074.6 torr.