Three Rivers COMMUNITY COLLEGE

CHE 111: Concepts of Chemistry

Laboratory Exercises

Compiled by: James E. Copeland, Professor of Biology & Chemistry
Edited: June 2013

Chemistry 111 - Exercise 1: Laboratory Safety Procedures and Policies Lab

SAFETY LAB:

Conduct in the laboratory is less formal than in the lecture portion of the course, however, loud talking, playing, the drinking of beverages, chewing gum, the consumption of food, smoking cigarettes, and the application of makeup is **prohibited.** Remember ….. lab work will be as safe as **you** make it. The following is a list of precautions that **must be taken** to minimize the occurrence of serious situations in the lab.

1. Locate all available safety devices in the lab and become acquainted with how they operate. (Shower, eyewash, fire extinguishers, fire blanket, gas fume hoods, spill centers, and first aids kits.) Also locate the nearest telephone for dialing 911 and other emergency numbers. *Know what to do in an emergency*.

2. Always wear approved eye protection in the lab. *Contact lenses are only Allow in lab when worn in concert with approved eye protection when performing experiments.* Never rub your eyes with your hands they may be contaminated. (If a chemical should get in your eyes, wash your eyes for at least twenty minutes in flowing water or the eye wash fountain.) Please use a buddy system and help your classmate to the eyewash fountain if he or she has difficulty seeing. Then report the incident to your instructor!

3. Wear protective clothing such as a lab coat or lab apron. Long sleeves or hair should be secured to prevent contact with chemicals or fire. Cotton clothing or other natural fabrics is preferable to synthetic fabrics, because many synthetic fabrics may dissolve if organic solvents are spilled on them and some synthetic fabrics are flammable! Arms and legs should be covered. Gloves should be used as appropriate when working with chemicals. ***(Special note: watches, bracelets, and rings should not be worn while performing experiments.)***

4. Leather shoes must be worn at all times. Shoes protect the feet against accidental chemical spills, dropped objects, and broken glassware.
   Open toed shoes, sandals, and canvas shoes are not allowed in lab!!

5. Take care-handling glassware. Cuts resulting from insetting glass tubing into rubber stoppers are the most frequent type of injury in the laboratory. Never force glass tubing into the stopper. Lubricate the tubing and the stopper, protect your hands, and insert tubing in a gentle twisting motion. If you sustain a cut, wash the wound with antiseptic soap and water, seek first-aid and report the incident to your instructor. Immediately clean up broken glassware.
   Place broken glass in the appropriate container.

6. All chemicals are potentially dangerous. Chemicals may enter the human body in three ways: (1) Ingestion – never perform mouth pipetting, please use suction bulbs or calibrated pipette tools; never taste any chemical in lab. All eating, drinking, and gum chewing is prohibited in lab. (2) Inhalation – avoid inhaling fumes and vapors form chemicals and their reactions by working in a well-ventilated space. Use the gas fume hood to perform experiments that produce dangerous toxic vapors or unpleasant vapors. And never directly smell any chemical. (3) Absorption – some chemical substances can penetrate the skin without any noticeable feeling; therefore if any chemical should spill on your hands immediately wash your hands with water, then report the incident to your instructor. Chemical burns may occur if a chemical is allowed to remain in contact with the body for any length of time. Please try to avoid all spills especially acids and bases! ***If you need to mix an acid with water, to prevent splatter pour the acid into the water.*** If a small area of your body comes in contact with an acid, wash the area with a dilute solution of (5%) sodium hydrogen carbonate (sodium bicarbonate), then large amounts of water. If the chemical substance is basic (hydroxides), wash the area with a 3% boric acid solution, then large amounts of water. If any chemical spill is over a large portion of your body, use the safety shower and remove contaminated clothing immediately.
   To avoid possible contamination, never place the top, lid, or cap of a chemical container on the work bench. Hold it in your hands properly.

7. All chemical spills should be cleaned up immediately even water. **Remember** only **water** down sinks, caustic or corrosive liquids must be placed in a safety disposable container (ask your instructor for this container). For mercury spills use the mercury spill kit. For acid spills, the application of a safety solution of sodium hydrogen carbonate (sodium bicarbonate) is recommended or use a chemical spill pillow. For bases (hydroxides), the application of vinegar or boric acid is recommended or use a chemical spill pillow. All solid must be placed in an appropriate safety container.
   Never place solids in the sink.
   Never carry large containers of chemicals in your hands. Use a cart!

8. FIRE! FIRE! FIRE! If your clothing catches on fire **walk do not run** to the safety shower, stand underneath, pull chain to extinguish the flames. Or smother the fire with the fire safety blanket. To reduce the chance of serious burns PLEASE help your fellow student if he or she is on fire. If a small fire starts in your working area shout fire! Then use the appropriate powder from the small fire station. Or use a carbon dioxide extinguisher to put the fire out. If a fire occurs in the lab, all students must turn off all heating device and quickly and calmly exit the lab and building.

Treat burns immediately by putting the burn area under cold water for at least Fifteen minutes. Cold water reduces the pain and blisters.
Report the incident to your instructor.

9. **NEVER** work alone in the laboratory for safety. Please do not perform any unauthorized experiments.

10. Be properly prepared to do the experiment. Read the written procedures before coming to class and understand what you are to do. Lack of familiarity wastes time and is a major cause of injury. Know the hazards before you do the experiment.

11. Perform the experiment as directed. Do not do anything, which is not a part of an approved experimental procedure. Follow all instructions given by your professor or lab instructor.

12. Treat all chemicals with the respect they deserve. Know the hazards before you handle the material. Read the chemical labels very carefully. Read the label when you pick up the chemical; read it again before you use the chemical; read the label after you finish with the chemical. Review MSDS's.
Many mistakes – some dangerous – results from mixing the wrong chemicals.

13. Never return unused reagent to the reagent bottle*. Be careful only to take the quantity that you actually need. Do not contaminate the reagents.
(Unless told to do so by you instructor.)

14. Never leave an ongoing experiment unattended. Turn off your Bunsen burner or other heat source whenever you are not using it. Never let heating devices operate unattended.

15. Never take chemicals, supplies, or equipment out of the laboratory.

16. Clean any and all glassware that contained any chemical at the end of the lab period. Wash down your workbench with a mild cleaning solution at the end of the lab period. If you had a small chemical spill while weighing your chemical for an experiment, clean the chemical balance (centigram or electronic) and/or work area immediately.

17. Put away all equipment and reagents, and wash your hands at the end of each work session.

*** For your protection, please inform your instructor of any medical condition that may create an emergency in lab (diabetic, epileptic, fainting spells, asthmatic, allergies etc.) and whether you regularly receive medication for control of any disorder.

Personal Hygiene and Safety Precautions

Everyone working in a science lab should be aware of the dangers of ingesting chemicals, being in direct contact with chemical and other hazardous materials. These common sense precautions will minimize the possibility of such exposure:

a) **Do not** bring any unnecessary items to the laboratory **do not place** any personal items:(pocketbooks, purses, book bags, coats, hats or caps, umbrellas, etc.) on the laboratory tables or near your feet.
b) **Do not** prepare, store (even temporarily), or consume food or beverages, or chew gum in the laboratory.
c) **Do not** smoke in the laboratory (Additionally, be aware that tobacco products in opened containers can absorb chemical vapors.
d) **Do not** apply cosmetics while in the laboratory.
e) **Do not** carry chemical containers around the room use a cart when necessary.
f) **Never** pipette by mouth, always use a suction bulb or pipetting device.
g) **Never** wear, or bring lab coats or aprons into areas were food is consumed.
h) **If** small amounts of chemicals should spill on you hand, or small areas of your body, immediately wash the area with water for at least fifteen minutes, the report the incident to your instructor. For spills over large area of your body use the safety shower for fifteen to twenty minutes and remove contaminated clothing.
i) **Wash hands** thoroughly before leaving the laboratory, even if gloves were worn.
j) **Wash** lab coats or jackets separately from or personal laundry.

Data Sheet - QUESTIONS:

1. In the space below sketch an outline of the laboratory, indicating on the drawing the location of the exits, fume hoods, safety shower, eyewash station, spill center, fire extinguishers, fire blanket, first aid kits, and your work station.

2. Why is eating and drinking not allowed in lab?

3. List four conditions of dress and/or types of clothing that are not permitted in the laboratory and state the reason that is would not be acceptable.

4. What is the proper way to mix acids and water?

5. How would you extinguish a fire caused by a combustible metal?

Laboratory Policies and Procedures

The primary purpose for laboratory work is to experimentally emphasize the nature of science (Chemistry) by providing you the student with exercises designed to -enhance your understanding of the theories, facts, and laws of science. Laboratory gives you an opportunity to search and discover concepts and principles of chemistry through experimentations. Understand, however, that chemistry like life cannot always be controlled or characterized by rules that work 100% of the time.
Good rules work about 90% of the time; the other 10% are call exceptions.
Remember, the purpose of lab is to experiment, observe, think critically, analyze, and decide. The following is a list of general procedures that should be followed each and every laboratory period.

1. Read the assignment in the lab manual provided to you by your instructor before coming to class. Preparation outside of class (laboratory) cannot be overemphasized. You must read the protocol or procedure for performing the experiment and know what equipment you need and when you need to use it, also what chemicals are needed and how to handle them safely. As well as, what steps that are to be taken, and when to carry out these steps for successful completion of the experiment. Lack of preparation may lead to careless handling of chemicals and equipment, which in most cases is dangerous. More importantly, if you wait until the laboratory period to organize and assimilate the procedure or protocol for the experiment, you will lose the precious time needed to think about and analyze what you are doing.
   At the beginning of each lab the instructor will briefly engage you in a discussion about the experiment and its purpose.

2. Before you begin an experiment, become familiar with all of the equipment that is to be used and the function or purpose of the equipment for the procedure. Discuss the process of the experiment with your lab partner before you begin the procedures.

3. Work in your own area. (**SAFELY**).

4. If you need extra equipment or reagents, ask your instructor.

5. Record all necessary data as you perform your experiment.

6. Return all reagents to the proper storage area, clean all glassware and equipment then return them to the cart or storage area.

7. Wash down your work area with a mild cleaning solution, and then dry the surface.

8. Make sure the sink is free of trash. Turn off the water and gas.

9. Turn in the Laboratory Report Sheet as required.

NOTES:

Familiarize yourself with the location of fire extinguishers and other safety equipment. When you need them, there is no time to ask.

Experiments involving toxic or noxious gases are best carried out in the fume hood.

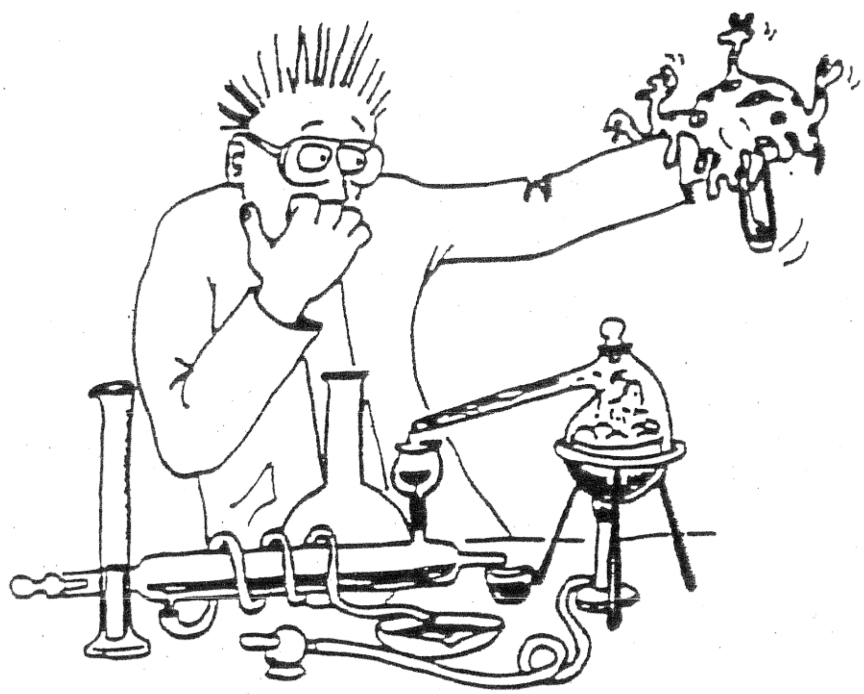
Wear protective goggles at all times! You, of course, will be careful and never have an accident, but you never know when your neighbor is going to squirt you in the eye.

When pouring liquids from reagent bottles, hold the stopper between two fingers.

Never point the mouth of a heated test tube at your neighbor, He may retaliate!

Do not take reagent bottles to your desk. You may receive harsh words from the fellow who just spent fifteen minutes looking for the reagent.

Keep your bench top clean and dry. Spillage is avoided by keeping containers and apparatus at the back of the bench.



Please don't experiment on your own yet!

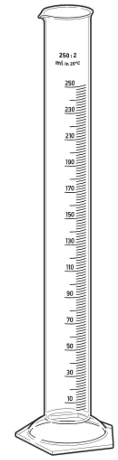
CHE 111: CONCEPTS OF CHEMISTRY

COMMON LABORATORY EQUIPMENT

THREE RIVERS COMMUNITY COLLEGE

LABORATORY EQUIPMENT

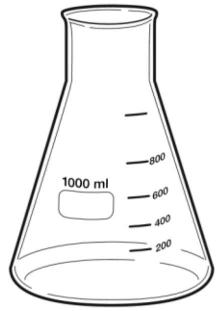
Beaker Cover



Graduated Cylinder

Beaker

Bottle, Wide Mouth

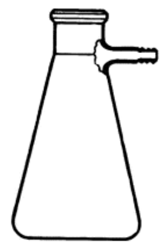


Erlenmeyer Flask

Florence Flask

Watch Glass

Evaporating Dish

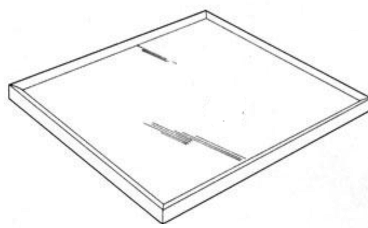


Filter Flask

Funnel

Test Tubes

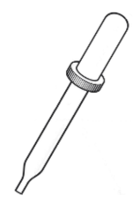
Test Tube Brush



Glass plate

Weighing Dish, Aluminum

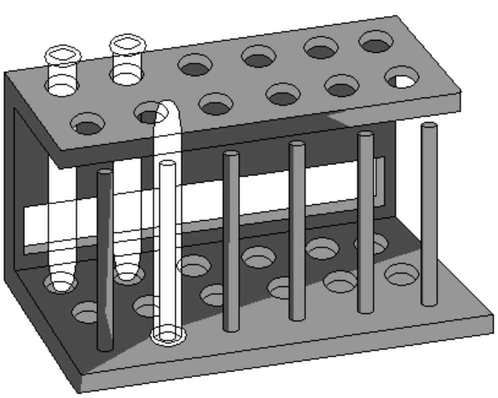
Stirring Rod



Medicine Dropper

Stirring Rod with
Rubber Policeman

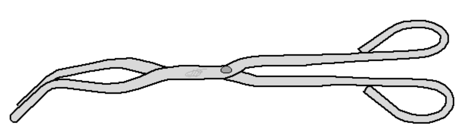
Scoopula



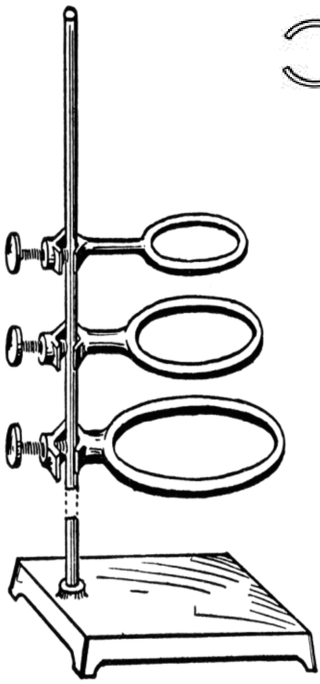
Test Tube Support

Pipet

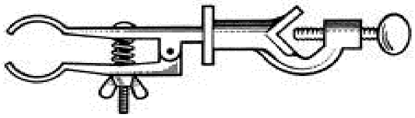
Test Tube Clamp



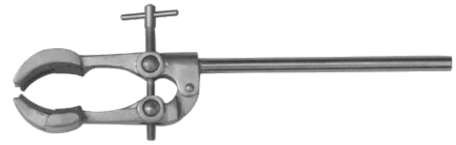
Tongs, Crucible



Ring Stand

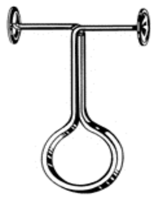


Burette or
Utility Clamp



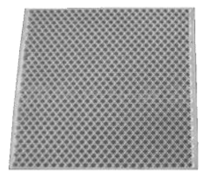
Extension Clamp

Screw Clamp

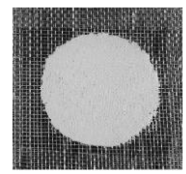


Pinch Clamp

Clamp Holder

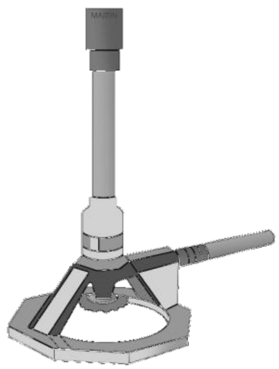


Ceramic Mat



Wire Gauze Square
Ceramic Center

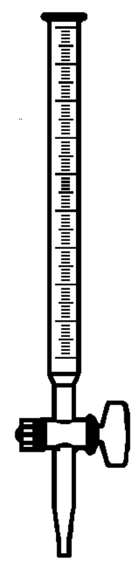
Ring Stand with
Double Buret Clamp



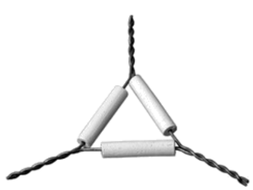
Burner

Flame Spreader
or Wing Top

Separatory Funnel



Buret with
Glass Stopcock



Clay Triangle

Crucible and Cover

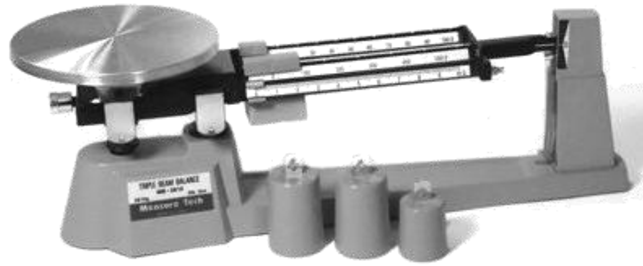
Mortar and Pestle

Rubber Stoppers

Electrical Centrifuge

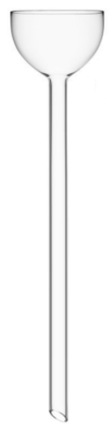
Triangular File

Triple-Beam Balance

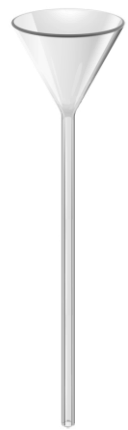


Laboratory Balance

Porcelain Spatula



Thistle Tube



Funnel Tube

Buchner Funnel

Forceps

Spatula

Nichrome Wire

Rubber Stoppers

Graduated Pipet

Volumetric Pipet

Combustion Spoon

Chemistry 111 – Exercise 2:   Measurements in Chemistry

JEC

The two basic measures of matter are mass and volume. In addition, linear measurements may also be made of a sample of matter. In chemistry mass is measured with a chemical balance (scale); volume can be measured with a graduated cylinder, buret, or pipette; and length is commonly measured with a ruler. Many experiments carried out in chemistry are quantitative. They involve assigning numbers to such units as length, volume, mass, and temperature.

Length – The simple measuring device found in every general chemistry laboratory is the meter stick. The meter stick is divided into ten parts called decimeters. Each decimeter in divided into ten parts called centimeter, because there are one hundred of these units in a meter. And each centimeter is divided into ten parts called millimeters, because there are one thousand of these units in a meter.

Volume – Units of volume in chemistry uses the S.I. system based on metric units. The metric volume units are closely related to linear units. The cubic centimeter (cc) or (cm^3) represents the volume of a cube one-centimeter on edge. A larger unit is the liter, which are exactly $1000cm^3$. A milliliter is equal to $1/1000^{th}$ of a liter and has the same volume as a cubic centimeter.

Mass – The mass of a sample is a measure of the quantity (amount) of matter it contains. In the metric system, mass is measured in terms of grams. In the S.I. system mass is measured in terms of kilograms (1000 grams).

Temperature – Thermometers used in chemistry are marked in degrees Celsius. The freezing point of pure water is taken as 0^o Celsius and the boiling point of pure water in taken as 100^o Celsius at one atmosphere of pressure. In the S.I. system temperature scale is measured in degrees Kelvin (K).

MATERIALS NEEDED:

Centigram scale, Electronic scale, 15 cm ruler, Meter stick, Medium size test tube, Crucible, Forceps, 50ml or 100ml Beaker, 1000ml Beaker, Buret w/stopcock, 10ml Graduated cylinder, 100ml Graduated cylinder, Dropping bottles w/distilled water, Weight boats, Copper pellets, Regularly shaped metal object, Size 11 or 12 Rubber stopper, Rocks, Ring stand, Buret clamp, Heavy irregularly shaped glass (Pestle), and 50ml Erlenmeyer flask w/rubber stopper.

CHEMICAL NEEDED:

Distilled water, Liquid A, Liquid B.

PROCEDURE 1: Length

Using a 15cm scale or a meter, determine the length of this sheet of paper.

Paper _________.

Work with a lab partner, to determine your height in:

centimeters _________.

meters _________.

feet _________.

inches _________.

PROCEDURE 2: Volume

Step 1: Fill a medium size test tube to the very top with tap water, then pour the contents of the test tube into a 100ml graduate cylinder and record the volume on the data table. "***NOTE***" the point on the graduate cylinder that coincides with the bottom curved surface of the liquid (meniscus) is read as the volume. *Empty & dry the graduated cylinder.*

Step 2: Fill a crucible to the very top with tap water, then pour the contents of the crucible into a 100ml graduate cylinder and record the volume on the data table. *Empty & dry graduated cylinder.*

Step 3: Fill a small beaker (50ml) to the very top with tap water, then pour the contents of the beaker into a 100ml graduated cylinder and record the volume on the data table. *Empty & dry graduated cylinder.*

Step 1: Using a 10ml graduated cylinder and a micropipette from the dropping bottle of distilled water, determine the number of drops of water that equals one milliliter of water. Record your results on the data table.
Empty & dry graduated cylinder.

Step 2: Repeat step 1 twice more.

Step 1: Attach a buret to a ring stand with a buret clamp and fill to the measured (calibrated) capacity with tap water. (Make sure the stopcock is closed.)
Record the total volume of water in the buret.

Step 2: Place an Erlenmeyer flask beneath the buret.

Step 3: Slowing open the stopcock and displace 20ml of water from the buret into the flask. Record the final volume of water in the buret.

DATA TABLE

Volume of test tube ______(ml).

Volume of crucible ______(ml).

Volume of beaker _______(ml).

Number of drops of water.

1st trial _____________

2nd trial _____________

3rd trial _____________

Average number of drops. _____________

Initial volume of the buret _______; Final volume of the buret ____________.

PROCEDURE 3: Mass

Step 1: Place a *weigh* boat on a centigram balance and weigh it to the nearest 0.01 of a gram and record the mass on the data table below.

Step 2: Using forceps place one copper pellet into the *weigh* boat and weigh the combination *weigh* boat and pellet on the centigram balance to the nearest 0.01 of a gram and record the mass on the data table below.

Step 3: Subtract the mass of the *weigh* boat from the mass of the *weigh* boat and pellet to obtain the mass of the pellet. Record the mass of the pellet on the data table below.

DATA TABLE

Mass of the *weigh* boat and pellet _________________(g)

Mass of the *weigh* boat _________________(g)

Mass of the copper pellet _________________(g)

Step 1: Place a 50ml beaker on a centigram balance and weigh it to the nearest 0.01 of a gram and record the mass on the data table below.

Step 2: Using forceps place a copper pellet into the beaker and weigh the combination – beaker and pellet on the centigram balance to the nearest 0.01 of a gram and record the mass on the data table below.

Step 3: Subtract the mass of the beaker from the mass of the beaker and pellet to obtain the mass of the pellet. Record the mass of the pellet on the data table below.

DATA TABLE

Mass of the beaker and pellet ________________(g)

Mass of the beaker ________________(g)

Mass of the copper pellet ________________(g)

Measuring mass using the Acculab V – 200 Electronic Balance

Step 1: Obtain an Electronic Balance from the cabinet and press the On/Memory key to turn the unit on and wait until it reads 00.00.

Step 2: Place a weigh boat on the tray top, wait for a stable reading and record the mass of the weigh boat on the data table below.

Step 3: Using forceps place a copper pellet into the weigh boat, wait for a stable reading and record the mass of the weigh boat and pellet on the data table below. *Press the Off key to turn the unit off.*

Step 4: Subtract the mass of weigh boat from the mass of the weigh boat and pellet to obtain the mass of the pellet. Record the mass of the pellet on the data table below.

DATA TABLE

Mass of the weigh boat and pellet _____________(g)

Mass of the weight boat _____________(g)

Mass of the pellet _____________(g)

NOTE: *The "tare" feature on the electronic balance is designed to allow the user to reset the balance to zero at any time. Tare can also be used to eliminate the value of the container (weight boat, watch glass, weight paper, etc.), from the weighting procedure and display only the net weight of the material. Place the container on the weighting tray wait for a stable reading and then press tare. The unit returns to zero and the weight value of the container is permanently removed from the remainder of the procedure.*

Step 1: Press the ON/Memory key to turn the unit on.

Step 2: Place a weight boat on the tray top and wait for a stable reading, then press "tare" to return the balance to zero and permanently remove the weight of the container from the weighting procedure.

Step 3: Using forceps place one copper pellet into the weigh boat and wait for a stable reading, then record your results in the data table below.
(SPECIAL NOTE: When weighting chemical for experiments slowly add the chemical into the container until the desired mass is achieved.)

Step 4: Press the Off key to turn the unit off. Return the electronic balance to the cabinet.

DATA TABLE

Mass of the Pellet ______________(g)

DENSITY

Density – Equal volumes of different substances may have very different masses.
The mass of a substance in relation to its volume is density. The term density refers to the mass of a substance per unit volume of that substance or D=M/V.

The volume of a regularly shaped object (box, cylinder, sphere) can be obtained by certain linear measurements. Example: V = length x width x height in cubic units.
The volume of irregularly shaped objects cannot be measured in this manner.
A convenient way of determining the volume of a solid, regardless of its shape, is to submerge it in a measured quantity of water in a large graduated cylinder if possible. The increase in the volume, as measured by the rise in the water level in the container, gives the volume of the solid.

PROCEDURE 4: DENSITY (Density equals the Mass divided by the Volume)

Step 1: Place a number 11 or 12 rubber stopper on the centigram balance and weigh it to the nearest 0.01g and record the mass on the data table.
Step 2: Put 600ml of tap water in a 1000ml beaker.
Step 3: Gently place the rubber stopper into the beaker of water. (The water level will rise.)
Step 4: With the stopper remaining in the beaker pour water from the beaker into a 100ml graduated cylinder back to its original water level. (600ml)
Step 5: Record the volume of water in the graduated cylinder on the data table. *(This volume is equivalent to the volume of the rubber stopper.)*
Step 6: Calculate the density of the rubber stopper.

DATA TABLE

Mass of the Rubber Stopper ____________(g)

Volume of the Rubber Stopper ____________(ml)

Density of the Rubber Stopper _____________(g/ml)

Step 1: Place an irregular shaped glass object (pestle) on the centigram balance and weigh it to the nearest 0.01g and record the mass on the data table.
Step 2: Put 1000ml of tap water in a 1000ml beaker.
Step 3: Carefully and slowly place the glass object into the beaker of water. (The water level will rise.)
Step 4: With the glass object remaining in the beaker pour water from the beaker into a 100ml graduated cylinder back to the original water level. (1000ml)
Step 5: Record the volume of the water in the graduated cylinder on the data table. *(This volume is equivalent to the volume of the glass object.)*
Step 6: Calculate the density of the irregular shaped glass object.

DATA TABLE

Mass of the Irregularly Shaped Glass ________________(g)

Volume of the Irregularly Shaped Glass ________________(ml)

Density of the Irregularly Shaped Glass ________________(g/ml)

Step 1: Place a rock on the centigram balance and weigh it to the nearest 0.01g and record the mass on the data table.
Step 2: Put 600ml of tap water in a 1000ml beaker.
Step 3: Carefully and slowly place the rock into the beaker of water. (The water level will rise.)
Step 4: With the rock remaining in the beaker pour water from the beaker into a 100ml graduated cylinder back to the original water level. (600ml)
Step 5: Record the volume of the water in the graduated cylinder on the data table. *(This volume is equivalent to the volume of the rock.)*
Step 6: Calculate the density of the rock.

DATA TABLE

Mass of the Rock _______________(g)

Volume of the Rock _______________(ml)

Density of the Rock _______________(g/ml)

Step 1: Weigh a metal cubed object on the centigram balance to the nearest 0.01g and record the mass on the data table.
Step 2: Using a 15cm ruler measure the length, width, and height of the object and record these measurements on the data table.
Step 3: Calculate the volume of the metal object.
Step 4: Calculate the density of the metal object.

DATA TABLE

Mass of the Metal Object _________________(g)

Length of the Metal Object _________________(cm)

Width of the Metal Object _________________(cm)

Height of the Metal Object _________________(cm)

Volume of the Metal Object _________________(cm^3)

Density of the Metal Object _________________(g/cm^3)

Repeat this part using the water display procedure and compare your answers.

PROCEDURE 5: Density of Liquids

Step 1: Place a 50ml Erlenmeyer flask w/rubber stopper in it on a centigram balance and weigh it to the 0.01g and record the mass on the data table.
Step 2: Pour 50ml of distilled water into a 100ml graduated cylinder. Then pour the water from the graduated cylinder into the 50ml Erlenmeyer flask and replace the rubber stopper.
Step 3: Weigh the flask, stopper and water on the centigram balance to the nearest 0.01g and record the mass on the data table.
Step 4: Obtain the mass of the distilled water by subtracting the mass of the flask w/stopper from the mass of the flask, stopper and water.
Step 5: Calculate the density of the distilled water. (*Clean & dry the flask.)*

DATA TABLE

Mass of the Flask, Stopper, and Water ________________(g)

Mass of the Flask & Stopper ________________(g)

Mass of the Water ________________(g)

Volume of the Water ________________(ml)

Density of the Water ________________(g/ml)

Step 1: Weigh a 50ml Erlenmeyer flask w/rubber stopper in it on a centigram balance to the nearest 0.01g and record the mass on the data table.
Step 2: Pour 50ml of either liquid A or B *(ask your instructor which liquid to use)* into a 100ml graduated cylinder. Then pour the liquid from the graduated cylinder into the 50ml Erlenmeyer flask and replace the stopper.
Step 3: Weigh the flask, stopper and liquid on the centigram balance to the nearest 0.01g and record the mass on the data table.
Step 4: To obtain the mass of the liquid subtract the mass of the flask and stopper from the mass of the flask, stopper and liquid.
Step 5: Calculate the density of the liquid.
Step 6: Pour the liquid into an appropriated container. (Ask your instructor which container to use.) Wash and dry the flask.

DATA TABLE

Mass of the Flask, Stopper, and Liquid ____________________(g)

Mass of the Flask & Stopper ____________________(g)

Mass of Liquid A or B ____________________(g)
(Circle the liquid you used)

Volume of the liquid ___________________(ml)

Density of Liquid A or B ___________________(g/ml)
(Circle the liquid you used)

Specific gravity is the ratio of the mass of a substance to the mass of an equal volume of water at the same temperature. The calibrated hydrometer is used to determine the specific gravity of some solids and liquids. To express the specific gravity of a substance, the density of water at 4^{o}C is used as the standard, which is 1g/ml. However, water at room temperature has a slightly lower density of 0.998g/ml. In calculating the specific gravity of a substance, use both densities in the same units. *REMEMBER: Specific gravity does not have units.*

$$\text{Specific Gravity} = \frac{\text{Density of a substance}}{\text{Density of water}}$$

Step 1: Calculate the specific gravity of liquid A or B by dividing its density by the density of water. Use both values from the experiment.
Step 2: Record the specific gravity of liquid A or B on the data table.

DATA TABLE

Specific Gravity of Liquid A or B ____________________
(Circle the liquid used)

CONVERSIONS FROM THE ENGLISH SYSTEM TO THE METRIC SYSTEM AND VICE VERSA.

Calculate each of the following conversion. (*Can be done as homework.)*

1. 12 inches to meters

2. 50 centimeters to inches

3. 90 grams to pounds

4. 125 pounds to kilograms

5. 3.6 liters to pints

6. 8 quarts to milliliters

1.06qt = 1 liter
2.54cm = 1 inch
2.2 lb = 1 Kg
454g = 1 lb

NOTES:

NOTES:

Chemistry 111 - Exercise 3: Percentage of Water in a Hydrate
JEC

Equipment needed: Ring and Ring Stand, Evaporating Dish, Ceramic Wire Gauze,
Glass Stirring Rod (Small), Electronic Balance,
Bunsen Burner or Hot Plate, Crucible Tongs,
Asbestos Heat Pad, 10ml Graduate Cylinder,
Gas Lighter, Solid Waste Container, Overflow Containers,
Plastic Weigh Boats, Spatulas or Spoons.

Chemical needed: Copper(II) Sulfate Pentahydrate – $CuSO_4 * 5H_2O$
Distilled Water

Procedure:

Step 1: Wash a glass stirring rod and evaporating dish in soapy water, rinse with large amounts of tap water then rinse both with distilled water. Dry the stirring rod and evaporating dish thoroughly with clean dry paper towels.

Step 2: Obtain an electronic scale from the cabinet; be sure to put the scale on a clean flat surface in your working area away from all flames or heat and chemicals. Press the on/memory key to turn the unit on and wait for a stable reading of 0.00; place your clean dry evaporating dish on the top tray of the scale and wait for a stable reading. Record the mass of the evaporating dish in the spaces provided on the data sheet to the nearest .01 of a gram. Remove the evaporating dish from the electronic scale and turn the unit off.

Step 3: Obtain a weigh boat, press the on/memory key to turn the electronic scale on and wait for a stable reading of 0.00. Place the weigh boat on the top tray of the scale and wait for a stable reading, now press the **tare** button and wait for the scale to return to 0.00. **Immediately** use a spatula or spoon to place small amount of Copper(II) Sulfate Pentahydrate into the weight boat until it reads 5.00grams. Remove the weigh boat and chemical from the scale and turn the unit off. Transfer the chemical from the weigh boat to the evaporating dish.

Step 4: Press the on/memory key on the electronic scale to turn the unit on and wait for a stable reading of 0.00. Place the evaporating dish containing the Copper(II) Sulfate Pentahydrate on the top tray of the scale; and weigh to the nearest .01 gram. Record this mass in the space provided on the data sheet. Remove the evaporating dish and contents from the scale and turn the unit off. Subtract the mass of the evaporating dish from the mass of the evaporating dish and contents and record the mass of the Copper(II) Sulfate Pentahydrate to the nearest .01 gram in the space provided on the data sheet. After you subtract; *If the mass of the Copper(II) Sulfate Pentahydrate is 5.00g you may begin the experiment. If it is not 5.00g but is only off by 0.01g you may begin the experiment (4.99g or 5.01g). If not repeat your weighing process until it is more accurate.*

Step 5: Place the evaporating dish and its contents on ceramic wire gauze and ring on the ring stand over a blue flame or a hot plate turned to its maximum. (Note any change in the $CuSO_4 * 5H_2O$ crystals). Continue gentle heating of the crystals with an occasional stirring with the glass rod until an almost white powder in formed. If any powder sticks to the stirring rod scrape it back into the evaporating dish. ***This change takes about ten to fifteen minutes***. Do not over heat or heat too fast!

Step 6: After an almost white powder is formed, turn the Bunsen burner off, if you use a hot plate leave it on. And carefully remove the evaporating dish and contents from the ring stand or hot plate with the crucible tongs, place it on the asbestos heat pad and allow to cool to room temperature. (About ten minutes.)

Step 7: After it has cooled weigh the evaporating dish and contents on the electronic scale to the nearest .01g and record the mass in the space provided on the data sheet. Subtract the mass of the dish from the initial weighing from the mass of the dish and powder and record the mass of the powder in the space provided on the data sheet.

Step 8: Place the dish and contents back on the ring stand or hot plate and reheat for five minutes. Remove the dish and contents from the ring stand or hot plate and place it on the heat pad to cool. (The re-heating is done to see if any more water is lost from the powder.) After the dish has cooled (about ten minutes), weigh the dish and contents to the nearest .01g on the electronic scale and record the mass in the space provided on the data table. Subtract the mass of the dish from the initial weighing from the mass of the dish and powder, and record the mass of the powder in the space provided on the data sheet. Compare the mass of the powder between the first and second heating, if they are the same the experiment is over. If the mass of the powder from the first and second heat differs by more than 0.10g, then heat a third time and weigh after it cools and record on the data sheet. (This allows you to obtain a constant mass.)

Step 9: Calculate your experimental percentage of water in Copper(II) Sulfate Pentahydrate. Calculate your percent yield.

Step 10: After you finish with the first trial, hold the cool dish and contains in the palm of your hand and pour 2.00ml of distilled water onto the powder in the dish. Be ready to remove the dish from your hands. Record all observations and answer all questions.

Step 11: Perform a second trial of this experiment by repeating steps 1 through 9.

Step 12: Place chemical waste in the appropriately marked waste container.

Step 13: Clean all glassware and equipment and then return them to the appropriate storage area.

Step 14: Wash down your work area and turn in lab report.

DATA SHEET

TRIAL 1	TRIAL 2
_______ Mass of dish and contents	_______ Mass of dish and contents
_______ Mass of dish	_______ Mass of dish
_______ Mass of Copper(II) Sulfate	_______ Mass of Copper(II) Sulfate

---

First Heating	First Heating
_______ Mass of dish and contents	________ Mass of dish and contents
_______ Mass of dish	________ Mass of dish
_______ Mass of powder	________ Mass of powder

Second Heating	Second Heating
_______ Mass of dish and contents	________ Mass of dish and contents
_______ Mass of dish	________ Mass of dish
_______ Mass of powder	________ Mass of powder

Final Heating	Final Heating
_______ Mass of dish and contents	________ Mass of dish and contents
_______ Mass of dish	________ Mass of dish
_______ Mass of powder	________ Mass of powder

The mass of the water is determined by subtracting the mass of the powder from the mass of the crystal.

_______ Mass of the crystals	_________ Mass of the crystal
_______ Mass of the powder	_________ Mass of the powder
_______ Mass of the water	_________ Mass of the water

Observations:

Describe the physical properties of the Copper(II) Sulfate crystals before heating.

Describe the physical properties of the anhydrous powder formed after heating.

Describe the physical properties of the powder after the distilled water was added.

Calculations:

The actual % of water in Copper(II) Sulfate Pentahydrate 36.00%.

Calculate the % of water of Copper(II) Sulfate Pentahydrate as determined by your experimental results.

$$\% \text{ Mass} = \frac{\text{Mass of Part}}{\text{Mass of Whole}} \times 100$$

Calculate the percent yield determined by this experiment.

$$\% \text{ yield} = \frac{\text{experimental yield (mass)}}{\text{theoretical yield (mass)}} \times 100$$

Questions:

1. What is a hydrate?

2. What process did you perform during this experiment that resulted in the formation of an anhydride?

   What is an anhydride?

3. When the distilled water was added to the anhydrous Copper(II) Sulfate powder in step 10 was the reaction endothermic or exothermic? ___________________________________.

4. Define:

   a) endothermic reactions:

   b) exothermic reactions:

NOTES:

Chemistry 111 – Exercise 4: The Nature of Chemical Substances Chemical Changes or Physical Changes

JEC

The three physical forms of matter are solids, liquids, and gases. The physical properties of matter include: odor, color, taste, solubility, density, specific heat, boiling point, and melting point. *Physical changes may occur in matter without a change in the chemical identity of that particular matter.*

The chemical properties of matter are concerned with the composition of the matter, what elements makeup the matter and in what proportions. *Chemical changes occur when a change in the composition of matter can be observed indicated by the formation of new substance(s). Usually, chemical reactions can be identified by the presence of one or more of the following: color change, odor change, the production of an odor, production of heat, or light, the production of gas, the formation of a precipitate, and the cooling of a substance.*

Purpose: To carry out and observe various procedures and determine if the observed reaction was a physical change in matter or a chemical change in matter.

Equipment needed: Evaporating dish, Mortar and Pestle, Glass funnel, 50ml Beaker, 250ml Erlenmeyer flask, Glass stirring rod, Test tube rack, Medium size test tube, Large Pyrex test tube, Test tube holder, Metal Cutter, Glass tubing, Glass cutter, # 1 rubber stoppers, # 4 One hole rubber stopper, # 4 Solid rubber stopper, Weigh boats, Large water troughs, Gas lighter, Crucible tongs, Universal clamp, Ring and ring stand, Ceramic wire gauze, Bunsen burner, Plastic spoon, and Spatula.
Other material needed: 15cm Filter paper, Book of matches, Straws, Wood splinters, Sterile cotton swaps.

Chemicals needed: Magnesium ribbon, Manganese dioxide, Copper wire, Charcoal powder, Sodium Chloride, 30% Hydrogen peroxide, Potassium chlorate, Lime water, and Distilled water.

Procedure 1

The element magnesium will unite with the element oxygen to form the compound magnesium oxide.

Step 1: Clean and dry an evaporating dish.
Step 2: Cut a short piece of magnesium ribbon from the roll.
Step 3: Hold the magnesium ribbon over the evaporating dish with a pair of crucible tongs.
Step 4: Heat the magnesium ribbon by striking a match and allowing the flames to touch the magnesium ribbon until a visible reaction occurs. Immediately open the tongs and allow the ribbon to fall into the evaporating dish. Observe and record the results.
Step 5: Discard the waste in the appropriate container – clean and dry evaporating dish and tongs.

Observation and Questions:

What happened when the match heated the magnesium ribbon? ___________________________

____.

Did a color change occur? _________; Did an odor change occur or was an odor produced? ________.

Was a gas produced? _________. Does the new substance look like the magnesium ribbon? ________.

What is the source of the oxygen that has united with the magnesium ribbon?

____.

What was the function of the lighted match? ____

____.

Was this a physical or chemical change? ____.

Procedure 2

Heating the element copper will cause it to combine with the element oxygen to form the compound copper oxide.

Step 1: Using the metal cutter – cut a small piece of bright copper wire from the roll.
Step 2: Heat one end of the copper wire in the flame of a Bunsen burner. While heating the wire note any color change.
Step 3: Remove the wire from the flame of the Bunsen burner and allow it to cool.
Step 4: Record your observations.
Step 5: Discard the waste wire in the appropriate container.

Observation and Questions:

What happened, as the copper wire was heated in the flames of the Bunsen burner?

____.

Was this a physical change or a chemical change? ________________________________.

After cooling does the end of the copper wire look as it did before cooling? ______________.

What is the color of the wire after cooling? ___________________________________.

Procedure 3

Step 1: Place two plastic chemical spoons full **or** one plastic tea spoon full of charcoal in a mortar, then an equal amount of (salt) sodium chloride likewise in the mortar and grind them together with a pestle. Then use a glass stirring rod to mix them together until a evenly disbursed mixture is formed.
Step 2: Add 125ml of distilled water to the powder in mortar and stir vigorously with the glass stirring rod for a few minutes.
Step 3: Place a glass funnel w/ filter paper into a 250ml Erlenmeyer flask.
Step 4: Slowly and carefully pour small amount of the solution from the mortar into the glass funnel with the filter paper and filter into the Erlenmeyer flask. If the filtrate is clear, decant into a clean dry 50ml beaker. (No more than 10 to 15 ml) *(If the filtrate is not clear, then clean all glassware and re-filter the filtrate with a new sheet of filter paper. Decant into a clean dry beaker50ml.)*
Step 5: Wash your hands and obtain a cotton swab from your instructor – dip the sterile cotton swab into the filtrate and place it on your tongue to taste it. Immediately put cotton swab into the regular trash. Record your observation.
Step 6: Setup ring & ring stand w/ceramic wire gauze over a Bunsen burner or use a hot plate.
Step 7: Wash and dry an evaporating dish. Then pour about 10ml of the filtrate into the evaporating dish.
Step 8. Place the evaporating dish and its contents onto the wire gauze and ring stand or hot plate and gently heat until all of the liquid has evaporated.
Step 9: If anything is left in the evaporating dish, after it cools saturate a sterile cotton swab with distilled water and place it onto the material in the dish and touch it to your tongue to taste it. (Record your observation.)
Step 10: Discard all materials into the proper container and clean and dry all glassware.
Step 11: Record your observations and answer all questions.

Observation and Questions:

When the evenly disbursed mixture of salt and charcoal was formed did you observe any characteristics of a chemical change? ___________.

Explain your answer:
____

____.

Was the color of the filtrate the same as the liquid poured from the mortar? _________;

What was the color of the filtrate? ________________.

What was the taste of the filtrate? ________________.

Was anything left in the dish after heating? _________. If so did it have a taste? ___________.

If anything was left in the evaporating dish after heating, what is that substance?

____.

Did a physical or chemical change occur during this procedure? ________________________.

Procedure 4

Hydrogen peroxide in solution decomposes slowly, forming water and oxygen gas.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

Step 1: Place 2ml of 30% hydrogen peroxide into a clean dry test tube and using a test tube holder (clamp) hold the test tube up to the light and observe for a few minutes. Record your observation.
Step 2: Holding the mouth of the test tube away from you, place a small sample (a pinch) of manganese dioxide (fill only the very tip of the spatula), into the hydrogen peroxide and observe the results.
Step 3: Touch the walls of the test tube with your hands and record the results.
Step 4: Discard all waste into the appropriate container and clean and dry all glassware.
Step 5: Record your observations and answer all questions.

Observation and Questions:

Were there any bubbles of gas observed during step 1? __________. If so what were those bubbles of gas? _________________________.

Was a gas produced when the manganese dioxide was added to the hydrogen peroxide?

____________.

Did the manganese dioxide have any effect on the hydrogen peroxide's rate of decomposition?

_____________.

Explain your answer. ____

____

____.

Was heat produced during the reaction after the manganese dioxide was added?

_________________.

Was this reaction a physical change or a chemical change?

____________________________________.

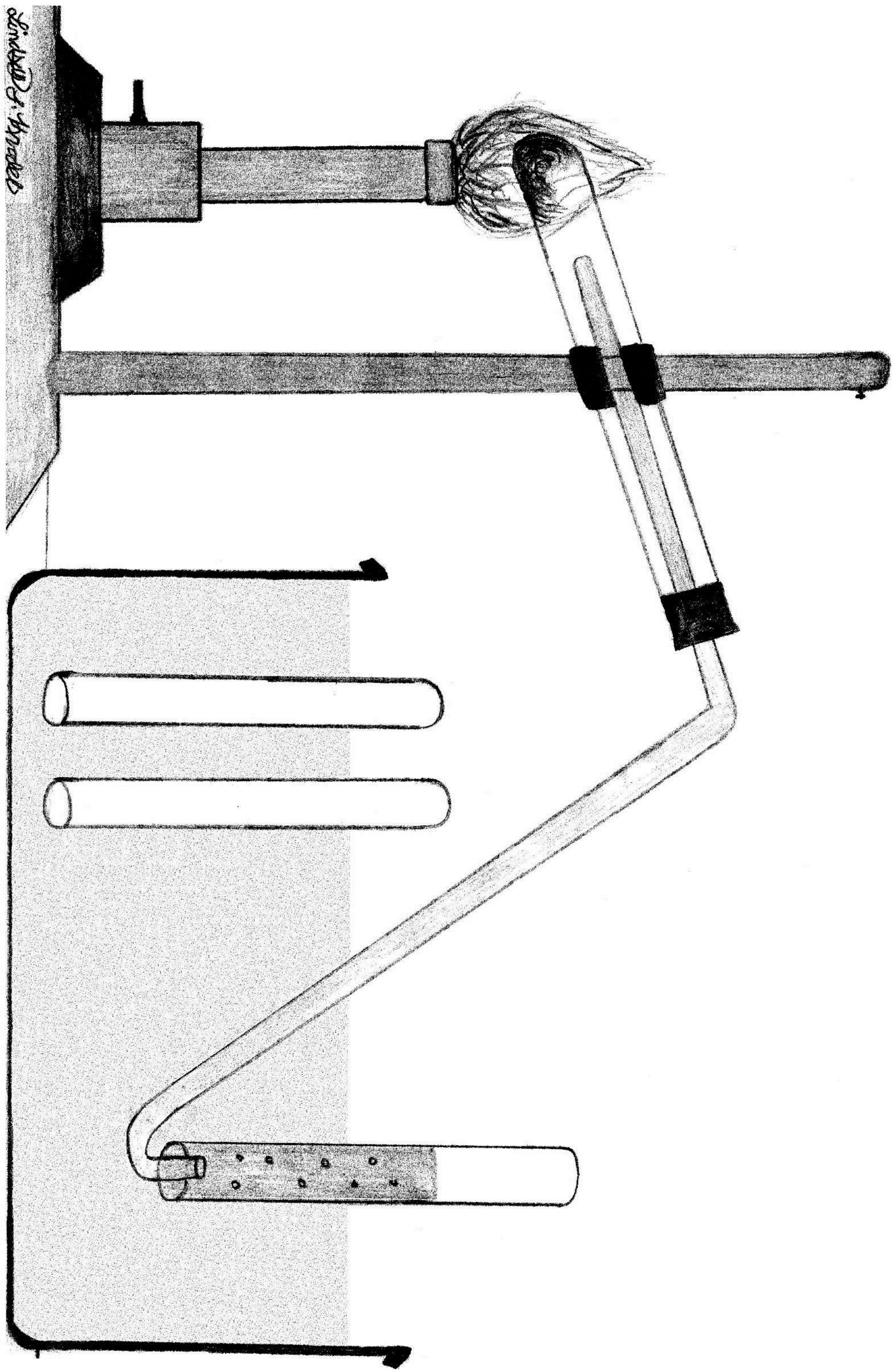
Has any apparent changes taken place in the manganese dioxide?

_____________________________.

Explain your answer: ____

____.

Define the term "Catalyst".



Procedure 5
(Preparation of Oxygen)
(Work in large groups – only two setup per work bench)

Step 1: Clean and completely dry a large Pyrex test tube.

Step 2: Cut and bend a piece of glass tubing as **illustrated by the instructor or according to the diagram** and insert it through a #4 one-hole rubber stopper. (*CAUTION: Do not force the tubing through the hole, moisten both the hole and the tubing, using a paper towel for safety, pass the tubing through the hole in a twisting motion.*)

Step 3: Clamp the Pyrex test tube w/attached glass tubing onto the ring stand (***Place the glass delivery tube into an large empty container for the water as illustrated by your instructor, to ensure proper setup of the apparatus, and to make sure that the end of the delivery will be below the water line when filled.) Afterwards remove the delivery tube from the large container for water.***

Step 4: Weigh 5 grams of potassium chlorate in a **weigh boat** on the centigram balance and place it in the large Pyrex test tube.

Step 5: Weigh 2 grams of manganese dioxide in a **weigh boat** on the centigram balance and place it in the large Pyrex test tube along with the potassium chlorate.

Step 6: Place a # 4 solid rubber stopper into the Pyrex test tube and shake the test tube until the two powders are well mixed.

Step 7: Remove the # 4 solid rubber stopper from the Pyrex test tube and place the # 4 one-hole rubber stopper, with the glass tubing attached, tightly into the pyrex test tube.

Step 8: Clamp the Pyrex test tube w/attached glass tubing onto the ring stand.

Step 9: With the end of the glass tubing on the outside of the container for water, fill the container to the top with tap water. Also, fill three medium size test tubes with water, cover their opening with your thumb to avoid spilling, invert them and place them in the container of water. Replace the end of the glass delivery tube into the large container of water (**make sure that the end of the delivery tube is below the water level)**.

Step 10: Immediately heat the Pyrex tube and contents gently over the Bunsen burner until bubbles of gas appears in the large container of water.

Step 11: Quickly slip one of the inverted test tubes of water over the delivery tube and collect the gas by the displacement of the water. **Hold on to the test tube as you do this to prevent it from rising and falling over.** When all of the water has been displaced by the gas, cork the test tube tightly with a # 1 rubber stopper beneath the surface of the water and remove it, then place the test tube of gas in an inverted position into the test tube rack.

Step 12: Quickly fill the other two test tubes with gas as described in **step 11** and place them in the test tube rack. (Steps 10, 11 and 12 will take about 30 to 40 seconds total.)

Step 13: ***Special Caution: Remove the end of the delivery tube from the large container of water and stop heating the powder. NOTE: If you stop heating before you remove the glass delivery tube, water will back into the Pyrex test tube and the pressure and chemical change could cause an explosion.***

Step 14: One member of the group should pick up one of the containers of gas – while another member of the group lights a wood splinter, then blow out the flame.

Step 15: Remove the rubber stopper from the test tube of gas and quickly thrust the glowing splinter well into the test tube without touching the wet slides. Record you observation.

Step 16: Quickly obtain 3ml of limewater in a 25ml or 50ml beaker.

Step 17: **Repeat steps 14 and 15, however when the reaction is over drop the splinter into the test tube and add the limewater, stopper immediately and shake vigorously. Record your observations.**
Step 18: Now pour 3ml of fresh lime water into a clean dry medium size test tube, using a straw gently blow your breath through the limewater for one to three minutes. Record your observation.
Step 19: Discard the chemical waste into the appropriate container, clean and dry all glassware, breakdown your equipment assembly and return to the proper storage. Wash down your workbench.
Step 20: Record your observations and answer all of the questions.
Step 21: Turn in your Lab Report.

What happened when the glowing splinter entered the test tube of gas? ____________________

____.

What is the gas that was liberated by this experiment? _____________________.

When carbon dioxide is passed through limewater an insoluble calcium carbonate is formed, which is indicated by the solution turning cloudy or milky.

Did the lime water turn cloudy or milky when it was added to test tube with the wood splinter in

Step 17? __________.

Did the limewater turn cloudy or milky when you blew your breathe through

it in Step 18? ________. How would you compare the results of Step 17 with Step 18?

____

____.

Did a physical or chemical change occur during this experiment? ________________________.

What was the function of the manganese dioxide? ____________________________________

____.

Why did you use a Pyrex test tube? ____

____.

NOTES:

NOTES:

Chemistry 111 - Exercise 5: Molecules, Formula Units and Compounds

Matter is anything that has mass and occupies space. Matter that is composed of entirely the same kind of atoms is an element. All elements can be placed in one of two basic groups: metal and nonmetals. The differentiating factor between metals and nonmetals are the physical and chemical properties of each element.

Physical Properties

METALS	NONMETALS
High metallic luster (shine)	Little luster - dull
Good electrical conduction	Poor electrical conduction
Good thermal (heat) conduction	Poor thermal (heat) conduction
Malleability – can be beat into shapes	Not malleable
Ductility – can be drawn into thin wire	No ductility
High density	Lower density
High melting points	Lower melting points
High boiling points	Lower boiling points
Hardness – most metals are hard	Soft

Chemical Properties

Metals do not combine chemically with each other to form molecules and compounds.
Metals combine chemically with nonmetals to form molecules and compounds.
Nonmetals combine chemically with metals and other nonmetals to form molecules & compounds.

The atom is the smallest part of the element that can undergo a chemical change. When two or more atoms join together they form a molecule or a formula unit. The part of the atom that is involved in forming molecules or formula units is the electrons. Electrons are arranged in shells (energy levels), the outer shell (energy level) of almost all metallic atoms have one, two, or three electrons. Whereas, almost all nonmetals have either five, six, or seven electrons in their outer shell (energy level). This difference in the outer shell electrons is the reason for the great difference in the chemical properties between metals and nonmetals.

The combination of atoms of metals and nonmetals tends to form ionic compounds for the most part, because metals tend to lose electrons and become positively charged atoms, and nonmetals tend to gain electrons and become negative charged atoms. The electrons that are lost or gained are called valence electrons. Valence is the tendency of atoms to lose or gain electrons. Metals lose electrons and have a positive charge (valence), nonmetals gain electrons and have a negative charge (valence).

The octet rule (the rule of eight) suggest that when elements combine to form compounds, each atom tries to get at least eight electrons in its outermost energy level. Since metals have fewer valence electrons they tend to be oxidized (lose electrons) and become cations. Nonmetals have a greater number of valence electrons so they tend to be reduced (gain electrons) and become anions. Cations are positively charged and anions are negatively charged. Two metals cannot form a compound chemically because they both form cations that would repel each other. This is why metals can only form compounds with nonmetals.

Because of covalent bonding (that is the sharing of valence electrons) between two or more adjacent atoms, nonmetals can combine chemically with other nonmetals to form molecules and compounds. However, in the study of chemistry, we find groups of atoms that are bound together yet they do not satisfy the octet rule and are not complete molecules. These tight groups of atoms, however, can form compounds (as if they were a single atom) and are called radicals or polyatomic ions. We write the formulas for radical or polyatomic ions in brackets to help us remember that act as a single unit: example $(PO_4)^{-3}$.

If radicals or polyatomic ions form compounds like single atoms, then we can expect they have valences. Most radicals or polyatomic ions are nonmetals that are grouped with other nonmetals, they are not molecules or compounds, but a combination of elements that still have a potential charge.

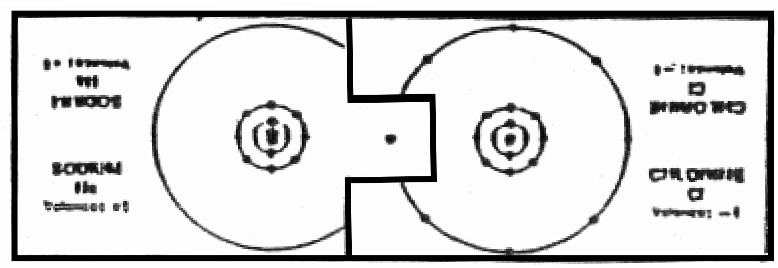
Example: $(OH)^{-1}$ a hydroxide ion, is composed of an atom of oxygen which needs two more electrons to complete its outer energy level, the oxygen can combine with one hydrogen atom and borrow its single electron. However oxygen's need for two electrons is only half satisfied and requires an additional electron. When the hydroxide ion (polyatomic) forms a compound it must gain one electron, and hence it has a valence of $^{-}1$. Another example: sulfate $(SO_4)^{-2}$, is composed of four oxygen atoms that need a total of eight electrons (two per oxygen atom) and a single sulfur atom acting as a metal lending its six valence electrons to the oxygen atoms. Instead of its common valence of $^{-}2$, sulfur assumes a valence of $^{+}6$. However, two more electrons are still needed to form a compound; the sulfate ion (polyatomic) must gain two electrons. Therefore, the sulfate ion $(SO_4)^{-2}$ has a valence of a negative two. In your molecule maker kit there are some cards that represent radicals or polyatomic ions. Pick out the cards for each radical or polyatomic ion, study the card, and fill out the chart below:

Chart of radicals or polyatomic ions

Radical or Polyatomic ion	**The Name of the Radical or Polyatomic ion**	**Valence**	**The number of atoms of each element on one card**
(OH)			______ atom(s) of **H** ______ atom(s) of **O**
(NO_3)			______ atom(s) of **N** ______ atom(s) of **O**
(SO_4)			______ atom(s) of **S** ______ atom(s) of **O**
(CO_3)			______ atom(s) of **C** ______ atom(s) of **O**
(PO_4)			______ atom(s) of **P** ______ atom(s) of **O**
(NH_4)			______ atom(s) of **N** ______ atom(s) of **H**

Procedure:

We can now build molecules or formula units for compounds from any metal and nonmetal we choose and keep a record of the compounds we make. Using the cards provided in the molecule maker kit, form a rectangle with the cards. EXAMPLE: Sodium, a metal, can only form compounds with nonmetals. Let us see how this could occur when sodium combines with chlorine to form the compound sodium chloride (salt). According to the octet rule, sodium tends to lose its single valence electron and chlorine already has seven valence electrons, and need just one more to make eight. Chlorine will borrow the one electron that sodium has to offer, and the two atoms will join to form a compound. Check this out with your molecule maker. Take a card marked sodium and notice it has a single projection which represents its one valence electron, now pick up a card marked chlorine and notice it has a single indentation which represent the number of electrons that it needs. Fit the two cards together as you would a jigsaw puzzle. If they form a rectangle then you have a compound. They do form a rectangle.



The electron lost by the sodium is gained by chlorine and placed into its outer energy level, thus making it complete with eight electrons. *There is only one combination of cards that will form a rectangle.* The reason why there is only one combination of cards to form a rectangle or form a compound is that the number of electrons lost must be the same as the total number of electrons gained. This is easily checked by adding the oxidation number (valences) of all of the cations and anions together, and if the sums of the oxidation numbers equal zero a complete compound has been formed. (Oxidation numbers are positive or negative whole numbers that gives the combining of the elements.) Example: Na^{+1} and Cl^{-1}; $+1 - 1 = 0$.

One way to keep a record of the compounds that are formed is to write the chemical formula for the compound. To write a formula, we write the chemical symbol for metal ion first, then the symbol of nonmetal ion second. To the right and slightly below the symbol we write a number as a subscript, which tells us how many atoms of each element that is needed to represent one molecule or one formula unit of the compound formed. Example: Ca_3P_2, this tells us that it will take 3 atoms of calcium to combine with 2 atoms of phosphorus to form the compound Calcium Phosphide. Whenever we see a chemical symbol that is not followed by a subscript number the number is understood to be 1 and we do not write the 1. Example: NaCl, indicates one atom of sodium and one atom of chlorine in one formula unit of Sodium Chloride.

Materials needed:

1. (1) Molecular maker kit that contains the following 48 cards:

Metal (blue cards)	Nonmetals (red {pink} cards)
3 Hydrogen	3 Fluorine
3 Sodium	3 Chlorine
3 Potassium	3 Oxygen
3 Magnesium	3 Sulfur
3 Calcium	2 Phosphorus
2 Aluminum	3 Hydroxide
3 Ammonium	3 Nitrate
	3 Carbonate
	3 Sulfate
	2 Phosphate

2. Pens or pencils

Step 1: Obtain a Molecule maker kit from your instructor.

Step 2: Follow the directions given on the three data table and complete each project.

Step 3: After you complete each Data Table, check the kit to make sure all of the pieces are accounted for and return the kit to the proper storage area.

Step 4: Obtain a Laboratory report sheet from your instructor and fill it out, without notes or help from your classmates, then bring the report to your instructor.

Step 5: Wait for instructions for part two of this lab.

Data Table 1

Directions: Use your Molecule Maker Kit to build the formula units and to supply you with the data for writing the correct chemical formula for the compounds. Remember these rules: (**1**) a compound is only correct when the cards representing the atoms of the elements form a rectangle; the sum of the oxidation numbers (valences) will equal zero; (**2**) cations symbols are written first in the formula and anions symbols are written second in the formula; (**3**) the subscript number for the atoms of each element is written in space to the right and slightly below the symbol for the element; (**4**) when the number of atoms is one only write the symbols the subscript is understood to be 1.

Compound made from:	Number of atoms of each element	Chemical Formula
Example: Magnesium and Phosphorus	__3__ atom(s) of Mg __2__ atom(s) of P	Mg_3P_2
Hydrogen and Sulfur	_____ atom(s) of H _____ atom(s) of S	
Calcium and Chlorine	_____ atom(s) of Ca _____ atom(s) of Cl	
Sodium and Phosphorus	_____ atom(s) of Na _____ atom(s) of P	
Aluminum and Sulfur	_____ atom(s) of Al _____ atom(s) of S	
Calcium and Fluorine	_____ atom(s) of Ca _____ atom(s) of F	
Aluminum and Chlorine	_____ atom(s) of Al _____ atom(s) of Cl	
Magnesium and Oxygen	_____ atom(s) of Mg _____ atom(s) of O	
Potassium and Fluorine	_____ atom(s) of K _____ atom(s) of F	
Sodium and Oxygen	_____ atom(s) of Na _____ atom(s) of O	
Calcium and Sulfur	_____ atom(s) of Ca _____ atom(s) of S	

Data Table 2

Directions: Using the cards from your Molecule Maker Kit, construct the following compounds using only the elements and radical or polyatomic ions asked for on this data sheet. Complete the data sheet and write the correct formula for each compound. Remember a radical or polyatomic ion is placed in brackets indicating that it is treated as a single atom; example: $Fe_2(SO_4)_3$. Also, if the subscript number in a formula is one, do not write the one; example: $Na^{+1} + (OH)^{-1} =$ NaOH Brackets are not necessary if you only need one of the polyatomic ions.

Compounds made from:	Number of cards for each element or group of elements	Chemical Formula
Potassium and Sulfate	______ card(s) of K ______ card(s) of SO_4	
Calcium and Nitrate	______ card(s) of Ca ______ card(s) of NO_3	
Magnesium and Phosphate	______ card(s) of Mg ______ card(s) of PO_4	
Ammonium and Phosphate	______ card(s) of NH_4 ______ card(s) of PO_4	
Ammonium and Chlorine	______ card(s) of NH_4 ______ card(s) of Cl	
Sodium and Sulfate	______ card(s) of Na ______ card(s) of SO_4	
Aluminum and Sulfate	______ card(s) of Al ______ card(s) of SO_4	
Calcium and Carbonate	______ card(s) of Ca ______ card(s) of CO_3	
Hydrogen and Nitrate	______ card(s) of H ______ card(s) of NO_3	
Magnesium and Hydroxide	______ card(s) of Mg ______ card(s) of OH	

Data Table 3

Directions: Use your Molecule maker Kit to construct the compounds from the elements and radicals or polyatomic ions listed. Use the cards to determine the oxidation states for each element or groups of elements, and write the correct chemical formula for the compound formed, as well as, to count the number of atoms of each element in one formula unit of the compound. Complete the data table.

Compound made from:	Number of cards for each element or polyatomic ion	Oxidation numbers + or -	Chemical Formula	Number of atoms of each element
Example: Aluminum & Sulfate	_2_ card(s) Al _3_ card(s) SO_4	2 (+3) = +6 3 (–2) = - 6 0	$Al_2(SO_4)_3$	_2_ atom(s) Al _3_ atom(s) S _12_ atom(s) O
Magnesium & Nitrate	____ card(s) Mg ____ card(s) NO_3			___ atom(s) Mg ___ atom(s) N ___ atom(s) O
Ammonium & Sulfate	____ card(s) NH_4 ____ card(s) SO_4			___ atom(s) N ___ atom(s) H ___ atom(s) S ___ atom(s) O
Calcium & Phosphate	____ card(s) Ca ____ card(s) PO_4			___ atom(s) Ca ___ atom(s) P ___ atom(s) O
Sodium & Hydroxide	____ card(s) Na ____ card(s) OH			___ atom(s) Na ___ atom(s) O ___ atom(s) H
Hydrogen & Carbonate	____ card(s) H ____ card(s) CO_3			___ atom(s) H ___ atom(s) C ___ atom(s) O
Aluminum & Hydroxide	____ card(s) Al ____ card(s) OH			___ atom(s) Al ___ atom(s) O ___ atom(s) H
Sodium & Sulfate	____ card(s) Na ____ card(s) SO_4			___ atom(s) Na ___ atom(s) S ___ atom(s) O
Hydrogen & Phosphorus	____ card(s) H ____ card(s) P			___ atom(s) H ___ atom(s) P
Magnesium & Fluorine	____ card(s) Mg ____ card(s) F			___ atom(s) Mg ___ atom(s) F
Aluminum & Oxygen	____ card(s) Al ____ card(s) O			___ atom(s) Al ___ atom(s) O

NOTES:

Chemistry 111 - Exercise 6: Molecular Representation – Covalent Bonding

Covalent bonding – a sharing of outer energy level electrons is the primary type of chemical bonding in organic compounds, as well as, inorganic compounds when you have two or more nonmetals combined together. The atoms in this kit are represented by colored enamel balls drilled to receive connecting bonds. The number of holes drilled in each ball indicates the bonding capacity for that element (the number of valence electrons the atom needs to share with another atom or atoms to form a molecule and/or a compound).

1. The yellow balls represent the element Hydrogen (H).
2. The green balls represent the element Chlorine (Cl).
3. The orange balls represent the element Sodium (Na).
4. The purple balls represent the element Potassium (K).
5. The red balls represent the element Oxygen (O).
6. The blue balls represent the element Nitrogen (N).
7. The black balls represent the element Carbon (C).

Linkage (bonding) of these balls is by wooden pegs or cadmium plated helical springs. To satisfy bonding you will insert a wooden pegs or cadmium plated helical spring into the drilled holes and connect the atoms (wooden balls) together. The wooden pegs are used for univalent bonding (when two adjacent atoms share a single bond between them); the *long pegs are used for univalent bonding between adjacent carbon atoms or carbon to nitrogen bonding (C – C – C etc.) or (C – N); the short pegs for all other univalent bonding between adjacent atoms.* The (2 inch) cadmium plated helical springs are used for multiple bonds between adjacent atoms, ring structural linkage, flexible bonds and unsaturated linkages such as: (C = C or C = O etc.).

Directions: Using the Molecule Maker Kit (balls) and the chemical formula for the compounds on the data table construct and study the molecule and compounds. Be sure to examine the molecular structure and record the shape of the molecule and/or compounds.

Objective: The student will be able to name and/or write the correct chemical formulas for the compounds contained in this data table.

Procedure:

Step 1: Obtain a Molecule maker kit from your instructor.
Step 2: Construct, study and learn the structure, name and formula for the compounds in the data table.
Step 3: Check to be certain that all parts to the model are put back in the box and return kit to the proper storage area.

DATA TABLE

MOLECULAR MODELS – MOLECULES, COMPOUNDS, AND FORMULAS

Name of the Molecular or Compound	Chemical Formula
1. Hydrogen gas	H_2
2. Oxygen gas	O_2
3. Water	H_2O
4. Hydrogen peroxide	H_2O_2
5. Carbon dioxide	CO_2
6. Hydrogen chloride	HCl
7. Carbon tetrachloride	CCl_4
8. Methane	CH_4
9. Methanol	CH_3OH
10. Ethane	C_2H_6
11. Ethylene	C_2H_4
12. Ethanol	C_2H_5OH
13. Potassium cyanide	KCN
14. Glycerol	$C_3H_5(OH)_3$
15. Glycine	CH_2NH_2COOH

JEC

NOTES:

Diagrams of Molecule Models:

Diagrams of Molecule Models:

Diagrams of Molecule Models:

Diagrams of Molecule Models:

Diagrams of Molecule Models:

Diagrams of Molecule Models:

Chemistry 111 – Exercise 7: Introduction to Qualitative Analysis
JEC

Qualitative analysis is the identification of unknown chemical substances by the use of specific reactions. When a chemical reaction occurs between two compounds in aqueous solution, sometimes a precipitate forms. Many metals ions can be identified in this way by the color of their precipitates when they react with Hydrogen sulfide. For example, when copper ions react with Hydrogen sulfide, Copper(II) sulfide (a precipitate) and nitric acid are formed. Since the Copper(II) sulfide is a precipitate (a solid) and has a specific color it is possible to identify it.

The equation for this reaction is: $Cu(NO_3)_2 + H_2S \rightarrow 2HNO_3 + CuS$

In this procedure Hydrogen sulfide is added directly to various metal ion solutions, and a visible reaction should occur. If a precipitate does not occur when the Hydrogen sulfide solution is added to one of the metal ion solutions, then the pH of the metal ion solution must be changed by adding ammonia water in order for precipitation to occur.

Equipment needed: Spot Plates w/depression.

Chemicals needed: Copper(II) nitrate, Cadmium nitrate, Antimony(III) chloride, Tin(II) chloride, Lead(II) nitrate, Bismuth(III) nitrate, Cobalt nitrate, Nickel(II) nitrate, Manganese(II) chloride, Zinc chloride, Iron(III) nitrate, Chromium (III) sulfate, Unknowns I, II, III, IV, V, Ammonia water, Hydrogen sulfide solution.

Important Note*:** Hydrogen sulfide gas is ***poisonous and it should be handled carefully. Hydrogen sulfide gas is soluble in water and to make the Hydrogen sulfide solution – the H_2S generator is heated for a few minutes allowing the generated gas to bubble through distilled water. The hydrogen sulfide is easily identified by a foul odor of rotten eggs. Because of the danger of this process, your instructor will conduct this operation.

Procedure 1:

Step 1: Place 5 drops of one of the metal ion solutions into one of the depressions in the spot plate.

Step 2: Immediately add two drops of the freshly prepared Hydrogen sulfide solution to the metal ion solution and note the results (the formation and the color of the precipitate).* Record the results on the data table provided.

* *If a precipitate is not formed, then it will be necessary to add one drop of ammonia water to the metallic ion solution before adding the Hydrogen sulfide solution. For those metal ion solutions you must use Procedure 1A below.*

Procedure 1A:

Step 1a: Place 5 drops of the metal ion solution into one of the depressions in the spot plate. Then immediately add one drop of ammonia water to the metal ion solution.

Step 2a: Now add two drops of the freshly prepared Hydrogen sulfide solution and note the results (the formation and color of the precipitate). Record the results on the data table provided.

Step 3 or 3a: Repeat the appropriate procedure until all of the metal ions solutions are tested, recording the results on the data table.

Procedure 2:

Unknowns: Using procedure 1 or 1A from above, whichever is appropriate, test all of the unknowns and record the results on the data table. Then compare the results of the color of the precipitates of the unknowns to that of the known metal ion solution. Using your observed comparisons, determine the name of the unknowns and write their names on the data table.

DATA TABLE 1

Metallic Ion Solution	Results: The Color of the Precipitate
Antimony(III) chloride $SbCl_3$	
Bismuth(III) nitrate $Bi(NO_3)_3$	
Cadmium nitrate $Cd(NO_3)_2$	
Chromium(III) sulfate $Cr_2(SO_4)_3$	
Cobalt nitrate $Co(NO_3)_2$	
Copper(II) nitrate $Cu(NO_3)_2$	
Iron(III) nitrate $Fe(NO_3)_3$	
Lead(II) nitrate $Pb(NO_3)_2$	
Manganese(II) chloride $MnCl_2$	
Nickel(II) nitrate $Ni(NO_3)_2$	
Tin(II) chloride $SnCl_2$	
Zinc chloride $ZnCl_2$	

DATA TABLE 2

Unknown Metallic Ion Solution	Results: The Color of the Precipitate
Unknown I	
Unknown II	
Unknown III	
Unknown IV	
Unknown V	

Based upon the information tabulated above and compared with the results in data table 1, what metallic ion is present in:

Unknown I ____

Unknown II ____

Unknown III ____

Unknown IV ____

Unknown V ____

Questions:

1. Hydrogen sulfide gas is (soluble , insoluble) in water.
   (Circle the correct answer.)

2. When hydrogen sulfide reacts with a metal ion a ______________________

   is formed which has a specific _____________________ and is a method for

   identifying the metal ion.

3. Which of the metal ions tested were basic (alkaline solutions)?

   ________________________ ______________________________

   ________________________ ______________________________

   ________________________ ______________________________

4. Complete and balance the following chemical equations:

   a) $Cu(NO_3)_2 + H_2S \rightarrow$

   b) $Cr_2(SO_4)_3 + H_2S \rightarrow$

   c) $ZnCl_2 + H_2S \rightarrow$

5. Write and balance a complete chemical equation from the following:

   a) the reaction between Tin(II) chloride and Hydrogen sulfide.

   b) the reaction between Lead(II) nitrate and Hydrogen sulfide.

   c) the reaction between Antimony(III) chloride and Hydrogen sulfide.

JEC

NOTES:

Chemistry 111 – Exercise 8: Determining Chemical Formulas – Quantitative Analysis

Introduction: Sometimes when two homogenous solutions are mixed together a precipitate forms. And the amount of the precipitate formed is dependent upon the correct ratio of ions needed to form the new compound. When the correct numerical ratio of these ions combine, the greatest amount of precipitate will be observed. This lab experiment is designed to determine the correct ratio in which certain ions will combine to form a stable compound.

Purpose: To experimentally determine the correct chemical formula for:
Procedure 1 - Calcium carbonate.
Procedure 2 - Copper(II) hydroxide.

Equipment Needed: Test tube rack w/ seven medium size test tubes, and a glass marker.

Solutions Needed: Procedure 1: Calcium chloride and Sodium carbonate.
Procedure 2: Copper(II) sulfate and Sodium hydroxide.

Procedure 1:

Step 1: Observe the physical properties of Calcium chloride such as the color of the solution and the state of matter. Record your observations on line 1 on Data Table 1. Observe the physical properties of Sodium carbonate such as the color of the solution and the state of matter on line 2 on Data Table 1.

Step 2: Label your test tubes 1 through 7 and place them in order in the test tube rack.

Step 3: Combine the Calcium chloride and the Sodium carbonate solutions into the seven test tubes using the information given in the following table. (*It is extremely important that you do this one test tube at a time, {place all the materials required into each test tube before moving on to the next}.) Try to do this with as little solution as possible coming into contact with the insides walls of the test tubes.*

Combination Table

	Test tube #1	Test tube #2	Test tube #3	Test tube #4	Test tube #5	Test tube #6	Test tube #7
Drops of $CaCl_2$	0	4	8	12	16	20	24
Drops of Na_2CO_3	24	20	16	12	8	4	0
Ratio of Ca : CO_3 ions	0 Ca : 24 CO_3	1 Ca : 5 CO_3	1 Ca : 2 CO_3	1 Ca : 1 CO_3	2 Ca : 1 CO_3	5 Ca : 1 CO_3	24 Ca : 0 CO_3

Step 4: Let the test tubes stand undisturbed for 20 minutes or until all of the precipitate has settled to the bottom of each test tube. *(Use this time to answer as many questions as possible in the observation and questions section.)*

Step 5: After all of the precipitate has settled to the bottom of the test tubes, carefully examine each test tube to determine which test tube has the greatest amount of precipitate, and record the number of that test tube on line 5 on Data Table 1.

Step 6: Observe the physical properties of the new substances formed as to the color and state of matter, and record this information on line 3 and line 4 on Data Table 1.

Step 7: Discard the chemicals in an appropriate container (ask your lab instructor). Then wash and dry all of your test tubes.

DATA TABLE 1

Line Number	Substance	Physical Properties
Line 1	Calcium chloride solution	
Line 2	Sodium carbonate solution	
Line 3	New liquid substance formed	
Line 4	New solid substance formed	
Line 5	Record the number of the test tube with the greatest amount of precipitate	
Line 6	The ratio of Ca ions to CO_3 in the above test tube (line 5) is:	
Line 7	Based on the information on line 6 write the formula for: Calcium carbonate.	

Procedure 2:

Step 1: Observe the physical properties of Copper(II) sulfate solution such as color and state of matter, record this information on line 1a on Data Table 2. Observe the physical properties of Sodium hydroxide such as color and state of matter, record this information on line 2a on Data table 2.

Step 2: Label your test tubes 1 through 7 and place them in order in the test tube rack.

Step 3: Combine the Copper(II) sulfate solution and the Sodium hydroxide solution into the seven test tubes using the information in the following table. (*Please do this one test tube at time as described in step 3 of procedure 1.)*

Combination Table

	Test tube #1	Test tube #2	Test tube #3	Test tube #4	Test tube #5	Test tube #6	Test tube #7
Drops of $CuSO_4$	0	4	8	12	16	20	24
Drops of NaOH	24	20	16	12	8	4	0
Ratio of Cu : OH ions	0 Cu : 24 OH	1 Cu : 5 OH	1 Cu : 2 OH	1 Cu : 1 OH	2 Cu : 1 OH	5 Cu : 1 OH	24 Cu : 0 OH

Step 4: Let the test tubes stand undisturbed for 20 minutes or until all of the precipitate has settled to the bottom of the test tubes. (*Use this time to answer as many questions as possible in the observation and questions section.)*

Step 5: After all of the precipitate has settled to the bottom of the test tubes, carefully examine each test tube to determine which one has the greatest amount of precipitate and record the number of that test tube on line 5a on Data Table 2.

Step 6: Observe the physical properties of the new substances formed as to color and state of matter, and record this information on line 3a and line 4a on Data Table 2.

Step 7: Discard the chemical in an appropriate container (ask your lab instructor). Then wash and dry all of your test tubes.

DATA TABLE 2

Line Number	Substances	Physical Properties
Line 1a	Copper(II) sulfate solution	
Line 2a	Sodium hydroxide solution	
Line 3a	New liquid substance formed	
Line 4a	New solid substance formed	
Line 5a	Record the number of the test tube with the greatest amount of precipitate	
Line 6a	The ratio of Cu ions to OH ions in the above test tube (line5a) is:	
Line 7a	Based on the information on line 6a write the formula for: Copper(II) hydroxide.	

Observations and Questions:

1. How many different elements are there in one formula unit of Calcium chloride? ________

2. List the elements and the number of atoms of each element in one formula unit of

   Calcium chloride: ___________; _____________ & ____________; ______________

3. What is the symbol and valence (ionic charge) for Calcium? ______________________

4. What is the symbol and valence (ionic charge) for Chlorine? ______________________

5. How many different elements are there in one formula unit of Sodium carbonate? ______

6. What is the symbol and valence (ionic charge) for Sodium? ________________________

7. What is the symbol and valence (ionic charge) for Carbonate? ______________________

8. Based on the true valences, write the correct chemical formula for Calcium carbonate.

   ______________________________.

9. Check your information on line 7 on **Data Table 1;** does the experimental results for the chemical formula for Calcium carbonate agree with your answer for question 8 above? Yes / No (**Circle your choice**)

10. If you answered **NO** to question 9 above, then list all of the possible sources of error in this experiment.

11. Write the correct chemical symbols and valences for the following:

    Copper(II) _________________; ________________.
    Sulfate _________________; ________________.
    Sodium ________________ ; ________________.
    Hydroxide ________________ ; ________________.

12. Based on the correct valence for Copper(II) ions and hydroxide ions; write the correct

    chemical formula for Copper(II) hydroxide. ___________________________________

13. Check your information on line 7a on **Data Table 2**; does the experimental results for the chemical formula for Copper(II) hydroxide agree with your answer for question 12 above? Yes / No (**Circle your choice**)

14. If your answer to question 13 is **No**; then list all possible sources of error for this experiment.

15. If the correct formula for Calcium carbonate **had been** Ca_2CO_3, which test tube would have

    contained the greatest amount of precipitate? ___________

16. If the correct formula for Copper(II) hydroxide **had been** $Cu(OH)_5$, which test tube would

    contained the greatest amount of precipitate? ___________

17. Experimentally what results would you expect to observe in order to determine the correct chemical formula for the new liquids formed?

18. Complete and balance the following equations:

a) $CaCl_2 + Na_2CO_3 \rightarrow$

b) $CuSO_4 + NaOH \rightarrow$

NOTES:

Chemistry 111 – Exercise 9: Stoichiometry Mass-Mass relationship in an Acid-base reaction

JEC

Objectives: To gain an understanding of the mass-mass relationship in a chemical reaction. Also, to determine and prepare an experimental yield of sodium chloride.

Discussion: Generally speaking a reaction that goes to completion occurs when one or more of the reactants is completely consumed. There are several ways in which a reactant may be consumed, resulting in an irreversible reaction: (1) if an insoluble precipitate is formed; (2) a weakly dissociated substance is formed such as water or a weak acid; (3) a gas is formed. In this experiment you will allow sodium hydrogen carbonate (sodium bicarbonate) (baking soda) to react with hydrochloric acid for the purpose of obtaining a high yield of sodium chloride according to the following reaction:

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

Using an accurately measured mass of sodium bicarbonate and enough six molar (6M) concentration of hydrochloric acid to react completely with it, you should be able to isolate the sodium chloride from the other products and determine its mass. The theoretical yield can be calculated using the mole/mass ratios obtained from the above chemical equation. The percentage yield can be calculated by dividing the mass of the experimental yield by the mass of the theoretical yield and multiplying by 100.

Equipment Needed: Evaporating Dish, Watch Glass, Ring and Ring Stand, Asbestos Heat Pad, Spatula or Spoon, Electronic Scale, Ceramic Wire Gauze, Bunsen Burner, Gas Lighter.

Chemical Needed: 6M Hydrochloric acid in glass dropping bottles
Distilled water in plastic dropping bottles
Sodium hydrogen carbonate (Sodium bicarbonate)

Procedure:

Step 1: Wash an evaporating dish and watch glass in soapy water, then rinse them with large amounts of tap water. Dry both the evaporating dish & watch glass with paper towels.

Step 2: Obtain an electronic scale from the cabinet, place it on a clean flat surface in your work area away from heating devices and water. Press the on/memory key to turn on the unit, and wait for a stable reading of 00.00. Place the combination of the evaporating dish and watch glass on the top tray of the electronic scale and wait for a stable reading. Record the mass of the evaporating dish and watch glass to the nearest .01 gram in the space provided on the data sheet.

Step 3: Remove the evaporating dish and watch glass from the electronic scale and turn the unit **off**. Remove the watch glass from the evaporating dish. Press the on/memory key on the electronic scale to turn the unit on and wait for a stable reading of 00.00, then place the evaporating dish on the top tray of the electronic scale and wait for a stable reading. Press the **tare button** and wait for the unit to read 00.00 **using a spatula or spoon immediately but slowly** add (sodium bicarbonate) Sodium Hydrogen Carbonate into the evaporating dish until it reads 2.00g. Remove the evaporating dish from the electronic scale, turn the scale **off** and place the watch glass (convex surface down) on top of the evaporating dish.

Step 4: Press the on/memory key on the electronic scale and wait for a stable reading of 00.00. Place the combination evaporating dish, sodium bicarbonate (Sodium Hydrogen Carbonate), and watch glass on the tray of the electronic scale and weigh the dish, sodium bicarbonate (Sodium Hydrogen Carbonate), and watch glass cover to the nearest .01 of a gram and record the mass in the space provided on the data sheet. Subtract the mass of the evaporating dish and watch glass from the mass of the evaporating dish, sodium bicarbonate (Sodium Hydrogen Carbonate), and watch glass and record the mass of the sodium bicarbonate (Sodium Hydrogen Carbonate) to the nearest .01 a gram in the space provided on the data sheet.

If the mass of the sodium bicarbonate (Sodium Hydrogen Carbonate) is 2.00g, you may begin the experiment.

If it is not 2.00g but is only off by no more than .01 gram (1.99g or 2.01g) you may proceed with the experiment. If it is greater than .01 repeat the weighing procedure until it is more accurate.

Step 5: Place the convex side of the watch glass down and slightly off center so that the lip of the evaporating dish is uncovered.

Step 6: Add the 6M hydrochloric acid solution drop wise (one drop at a time) down the lip of the evaporating dish onto the sodium bicarbonate inside the dish. Continue this process until all of the sodium bicarbonate is covered with acid and there is no longer a sign of a visible reaction occurring when one more drop of acid is added. At this point securely lift the dish, contents, and cover and gently swirl to ensure that all of the acid comes in contact with the sodium bicarbonate (Sodium Hydrogen Carbonate).

If a visible reaction occurs add one more drop of acid, lift and swirl.

Repeat this process until no visible reaction occurs when another drop of acid is added and you have swirled the contents.

Do not add excess hydrochloric acid.

Step 7: Lift the glass cover into a vertical position over the evaporating dish and use the plastic dropping bottle of distilled water to rinse the convex side of the watch glass, with a few drops of water at a time, collecting the washing in the evaporating dish.

Step 8: Place the evaporating dish, contents, and watch glass cover on the ceramic gauze on the ring stand and use a blue flame or use a hot plate to gently heat the dish and content until all of the liquid has evaporated and it is completely dry.

(This would include any condensation that develops on the watch glass.)

Turn the Bunsen burner **off** and allow the evaporating dish to cool for ten minutes on the ring stand. After ten minute use your hand to remove the dish, contents, and cover from the ring stand and place it on the asbestos heat pad and cool to room temperature. Or use heat resistant gloves to immediately remove the dish, contents, and lid from the hot plate and place it on the asbestos heat pad and cool to room temperature.

Step 9: After it has cooled, weigh the dish, contents, and glass cover on the electronic scale to the nearest .01 of a gram and write this mass down somewhere on your lab manual.

Step 10: Repeat **Step 8** by heating for five minutes to obtain a constant mass.

Step 11: After it has cooled, weigh the dish, contents, and glass cover on the electronic scale to the nearest 0.01 of a gram and write this mass down on a piece of scrape paper.

Step 12: Compare the two masses on the scrape paper, if they are the same or if they agree within 0.02 grams the experiment is over. If they differ by more than 0.02 grams repeat **Step 10 until you get two consecutive masses to agree within 0.02g.**

Step13: Record the mass of the evaporating dish, contents, and glass cover in the space provided on the data sheet. (From the last heating.)

Step 14: Subtract the mass of the evaporating dish and glass cover (from the initial weighing) from the mass of the evaporating dish, contents, and glass cover to determine the mass of the salt produced experimentally. Record the mass of the salt in the space provided on the data sheet.
Step 15: Calculate yield and record in the space provided on the data sheet.
Step 16: For second trial, repeat Steps 1 through Step 14.
Step 17: Clean and dry all glassware and laboratory equipment before return it to the proper storage.

Calculations:

1. Calculate the mass in grams of NaCl that should theoretically be produced in this experiment.

2. Calculate the percentage yield in this experiment.

$$\% \text{ yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100$$

Data Sheet

Weights and Measurements	Trial 1	Trial 2
Mass of the evaporating dish, $NaHCO_3$, and watch glass		
The mass of the evaporating dish and watch glass ___ (from the initial weighing in step 2)		
Mass of the Sodium bicarbonate ($NaHCO_3$)		
The mass of the evaporating dish, contents, & watch glass (*after heating and obtaining a constant mass)*		
The mass of the evaporating dish and watch glass ___ (from the initial weighing in step 2)		
The Mass of the salt (NaCl) produced experimentally		

****
****

Percent yield of salt as calculated from the experimental results	Trail 1	Trail 2

Questions:

1. Give two scientific explanations as to why the reaction in this experiment goes to completion and is irreversible.

2. List several possible sources of error that you think may have influenced the accuracy of your experimental results.

NOTES:

Chemistry 111 – Exercise 10: Solutions
JEC

Although there are many different types of solutions, only aqueous (water) solutions will be examined in this exercise. This is because water is the most important biological solvent. Many medications are given as water solutions. Solutes traveling to and from the cytoplasm of cells must exist in a watery medium. Minerals, as well as, many pollutants are dissolved in various bodies of water in the environment.

A solution is a weakly bound mixture of a solute and a solvent. In a solution the solute is the chemical substance that is to be dissolved and occurs in the lesser quantity, and the solvent is the dissolving medium and usually occurs in the greater quantity.

There are a variety of ways to express solution concentration - that is the ratio of solute to solvent.

Concentrated solution: *means a large amount of solute dissolved in a solvent.*

Saturated solution: *contains all of a particular solute that can be held at a given temperature and pressure. It is in dynamic equilibrium with the un-dissolved solute.*

Dilute solution: *means a small amount of solute is dissolved in a solvent.*

Unsaturated solution: *contains a smaller amount of a particular solute in solution that can be held in a saturated solution under the same conditions.*

Supersaturated solution: *contains an excess amount of a particular solute in solution that could be held in a saturated solution under the same conditions. Supersaturated solutions are unstable and if a surface for crystal formation is provided, the excess solute will separate from solution.*

Objectives:
1) To gain an understanding of solubility and factors that affect solubility.
2) To gain an understanding of factors that affects the rate of solubility.
3) To gain an understanding of solubility and saturation end points.
4) To learn about the characteristics of different types of solution.

EQUIPMENT NEEDED: Mortar & Pestle, 600ml Beaker, 250ml Beakers, Weigh boats, Filter paper, Glass Stirring rod, Test tubes, Test tube rack, Test tube holder, Rubber stoppers, Thermometers, Hot Plates, Plastic spoons, Metal spatulas, 10ml Graduated cylinder, 100ml Graduated cylinder, Centigram balance.

CHEMICAL NEEDED: Large Copper(II) Sulfate crystals, Salt(Sodium Chloride), Sugar (Sucrose or Dextrose), Sodium Carbonate, Calcium Carbonate, Sodium Hydroxide crystals, Ethanol, Sodium Thiosulfate, Ammonium Chloride, Cooking oil, Potassium Permanganate crystals, Distilled water, Saturated Salt (NaCl) Solution.

The Dissolving Process

Procedure: 1

Step 1: Put 100ml of distilled water into a clean dry 250ml beaker, and then place the beaker on a white sheet of paper in a secure area of your work bench.

Step 2: Using a metal spatula, carefully drop two or three small crystals of Potassium permanganate into the water. ***DO NOT STIR.*** Observe the beaker periodically over the next 60 minutes. Record your results.

Observation:

____

____

____

Possible explantation:

____

____

____

Effects of temperature, surface area and stirring on the rate of dissolution.
Procedure: 2

Step 1: Place a hot plate in a secure region of your work bench plug it into the electrical outlet on the desk, fill a 600ml beaker with 400ml of tap water, place the beaker on the hot plate and turn the unit on high enough to bring the water to a boil.

Step 2: Place several large crystals of Copper(II) sulfate in a mortar and grind into a pulverized powder with the pestle.

Step 3: Using a metal spatula evenly divide the pulverized powder into two test tubes. (***Small amounts 1/12th the volume of the test tubes.***) Then mark the test tubes #1 & #2. Place test tube #1 and #2 in a test tube rack.

Step 4: Find an ungrounded Copper(II) sulfate crystal that approximates the volume of the pulverized powder and put it into a third test tube and mark it #3, then place it in the test tube rack.

Step 5: Quickly, add equal amounts of distilled water to all three test tubes about ¾ full. Place all three test tubes into the boiling water simultaneously.

Step 6: Immediately begin stirring the contents of test tube #1, using a glass stirring rod and a test tube holder (clamp). As soon as the contents of test tube #1 dissolves, remove it from the boiling water and record the length of time it took for it to dissolve.

Step 7: Continue heating the contents of test tubes #2 and #3 until their contents dissolves. ***(DO NOT STIR THE CONENTS OF TEST TUBE #2 & #3)***

Step 8: When the contents of the remaining test tubes dissolve, remove the test tube from the boiling water and record the length of time that it took for the dissolution in each tube.

Observations and Questions:

Length of time for the dissolution of the pulverized Copper(II) sulfate with stirring. __________

Length of time for the dissolution of the pulverized Copper(II) sulfate without stirring. ________

Length of time for the dissolution of the Copper(II) sulfate crystal without stirring. ___________

Which test tube's contents dissolved first? _________ Possible explanation:

Which test tube's contents took the longest to dissolve? ________ Possible explanation:

Saturated Solutions
Procedure: 3

Step 1: Pour 3ml of a saturated Sodium chloride (NaCl) solution into each of two test tubes.

Step 2: Using a metal spatula add a small amount of Sodium chloride (a pinch) to one of the test tubes, place a rubber stopper securely in the mouth of the test tube and shake vigorously. Observe and record the results.

Step 3: Using a metal spatula add a small amount of sugar (Sucrose or Dextrose) (a pinch) to the other test tube, place a rubber stopper securely in the mouth of the test tube and shake vigorously. Observe and record the results.

Questions and Observations:

Did the added pinch of salt dissolve in the saturated salt solution? ____________

Did the added pinch of sugar dissolve in the saturated salt solution? __________

Give a possible explanation for the above results.

Solubility and Saturation End Points
Procedure: 4

Step 1: Weigh 1 gram of Sodium carbonate in weigh boat on a centigram balance, then place it into a clean dry 250ml beaker. Add 10ml of distilled water. Stir vigorously with a glass stirring rod.

Step 2: If the compound does not dissolve, add 10ml of distilled water at time, stir vigorously until it dissolves or until the beaker is full. (***Dissolution is complete if the solution is clear.***) Record your results.

Step 3: **<u>Repeat steps 1 & 2 of this experiment using 1 gram of Calcium carbonate.</u>**

Observations and Questions:

Did the Sodium carbonate dissolve completely? _____. The amount of water used. ________

Did the Calcium carbonate dissolve completely? _____. The amount of water used. ________

Based on the above results, what can you say about the solubility of these two substances?

Heat of Solutions
Procedure: 5

Step 1: Using a 10ml graduated cylinder measure 5ml of distilled water and pour the water into a test tube. Place the test tube into the test tube rack.
Step 2: Place a thermometer into the water, wait a few minutes until the thermometer reading stabilizes and record the temperature of the water.
Step 3: Using a clean, dry weigh boat, weigh 2 grams of Ammonium chloride on the centigram balance and carefully put the Ammonium chloride into the test tube of distilled water and allow it to dissolve – measure the change in temperature of the water as the chemical is dissolving.
Step 4: Touch the sides of the test tube and record your results.

Step 5: **Repeat Steps 1 through 4 Using a different test tube & 2 grams of Sodium hydroxide.**
(*CAUTION: Do not touch the Sodium Hydroxide with your fingers.*)

DATA

Temperature of the water in test tube 1. ________________________

Temperature of the dissolving Ammonium chloride. ________________________

Temperature change. ________________________

Temperature of the water in test tube 2. ________________________

Temperature of the dissolving Sodium hydroxide. ________________________

Temperature change. ________________________

Observations and Questions:

How did test tube 1 (Ammonium chloride feel to the touch? ________________________________

How did test tube 2 (Sodium hydroxide feel to the touch? ________________________________

Which of these reactions was exothermic? ____

Which of these reactions was endothermic? ____

Supersaturated Solutions
Procedure: 6

Step 1: Put 400ml of tap water into a 600ml beaker, place it on the hot plate and bring to a boil.

Step 2: Weigh 5 grams of Sodium thiosulfate in a clean – dry weigh boat – then place it into a clean – dry test tube. Add 15 drops of distilled water drop by drop. Place the test tube into the beaker of boiling water until the crystals are completely dissolved.

Step 3: Using a test tube holder (clamp) remove the test tube from the beaker of water, place a rubber stopper loosely in the mouth of the test tube and let the tube cool to room temperature in the test tube rack. (15-30 minutes) *Please turn off the hot plate.*

Step 4: After the test tube has cooled to room temperature, add one crystal of Sodium Thiosulfate to the test tube and record your observation.

Observation:

Liquids dissolved in Liquids

Determine whether or not the following pairs of liquids are soluble in each other by mixing 2ml portions of each liquids, and note whether or not layers of liquids forms or a complete mixture occurs. (**NOTE: Do not put oil or glycerin in the graduated cylinder – put 2 ml of water into the test tube and use a glass maker to measure the 2ml level on the test tube – pour the water out and pour the oil directly into the test tube.)**

Procedure: 7

Step 1: Place equal amounts of the two liquids in the data table into a test tube, stopper, shake vigorously. Let stand for about five minutes and record your results.

Data Table

1. Water and Ethanol ____________________________________
2. Water and Oil ____________________________________
3. Ethanol and Oil ____________________________________

NOTES:

Chemistry 111 – Exercise 11: The Properties of Acids and Bases
JEC

PURPOSE AND THEORY:

The purpose of this experiment is to become familiar with the properties of acids and bases. An acid can be defined as any chemical substance that is capable of donating hydrogen ion (H^+) to another substance. Acids can occur naturally in foods such as lemons, vinegar, and grapefruit. Acids have a sour taste, react with indicator dyes to form color changes. Acids react with metallic oxides and bases forming salt and water. Acids also react with certain metals to form hydrogen gas (H_2). A base can be defined as any substance that is capable of accepting a hydrogen ion (H^+). Bases have a bitter taste and feel slippery or soapy to the touch. Bases react with indicator dyes to form color changes. Bases will react with acids to form salt and water.

EQUIPMENT NEEDED: Lab lids, Medium test tubes, Plugs and delivery tubes, Plastic spoons, Metal spatulas, Wood splints, Matches, Chemplate depression dishes or Lab lids, Plastic stirring sticks, pH indicator paper, Test tube holder (clamp), Test tube rack.

CHEMICALS NEEDED: 6M Hydrochloric acid (HCl), 6M Acetic acid (CH_3COOH), 6M Potassium hydroxide (KOH), Unknown concentration of Acetic acid (CH_3COOH), 1M HCl - Hydrochloric acid, 1M HNO_3 – Nitric acid, 1M H_2SO_4 – Sulfuric acid, 1M CH_3COOH – Acetic acid, 1M KOH – Potassium hydroxide, 1M NaOH – Sodium hydroxide, 1M NH_4OH – Ammonium hydroxide, Limewater, Phenolphthalein, Copper metal filings, Magnesium metal shavings, Zinc metal (powder), Iron metal filings, Calcium carbonate chips.

PROCEDURE 1: ACIDS, BASES AND INDICATORS

Remove a piece of indicator paper from the vial, tear in half and place on a clean paper towel. Add a drop of one of the solutions listed below to the piece of indicator paper and record your results by comparing the color change with the chart on the vial of indicator paper. Repeat the procedure until all solutions are tested.

SOLUTION	COLOR OF PAPER	pH RESULTS
1M HCl – Hydrochloric acid		
1M HNO_3 – Nitric acid		
1M H_2SO_4 – Sulfuric acid		
1M CH_3COOH – Acetic acid		
1M KOH – Potassium hydroxide		
1M NaOH – Sodium hydroxide		
1M NH_4OH – Ammonium hydroxide		

PROCEDURE 2: EFFECTS OF ACIDS AND BASES ON METALS

A. Place a small piece or amount of each of the following metals in separate cavities of the Chemplate: (Using a Spatula) Iron, Zinc, Magnesium and Copper. Add enough drops of 6M HCl (10-15) to cover each metal. The reaction is the same in each case where a reaction occurs, differing only in the speed of the reaction. The relative reaction rates may be used to compare the activities of the metals. Rate the activity of metals reacting with HCl from (**1**) being the **most** active to (*4*) being the *least* active.

Zinc _____ ; Iron _____ ; Magnesium ______ ; Copper ______.

B. Repeat the test adding 6M CH_3COOH (10 to 15 drops) to cover the metals.

Zinc _____ ; Iron _____ ; Magnesium _____ ; Copper ______.

Compare the effect of the Acetic acid on the metals with that of Hydrochloric acid.

C. Repeat the test again adding 6M KOH (10 to 15 drops) to cover the metals. Compare the reaction of the metals with KOH and HCl in the table below.

DATA TABLE

METAL	REACTION WITH HCl	REACTION WITH KOH
Magnesium		
Copper		
Zinc		
Iron		

PROCEDURE 3: PREPARATION OF HYDROGEN GAS

1. Place 15 drops of the 6M HCl into a test tube. (Have a burning wood splint ready)

2. Add several pieces of Magnesium shaving to the test tube of acid.

3. As soon as a visible reaction occurs, hold the burning wood splint over the mouth of the test tube.

4. Record your results.

RESULTS:____

PROCEDURE 4: EFFECT OF ACID ON CARBONATES:

1. Place 15 drops of 6M HCl into section C of the lab-lid or a cavity in the Chemplate. (Be prepared to cover with the plug and delivery tube as soon as you add the calcium carbonate chips.)

2. Place 10 drops of fresh limewater to section D of the lab-lid or a cavity in the Chemplate.

3. Add two or three calcium carbonate chips to the test tube section or cavity containing the 6M HCl and plug immediately, placing the other end of the gas delivery tube directly into the limewater.

4. Note the bubbling of the limewater and record your results.

RESULTS: ____

NOTE: If limewater turns cloudy or milky, it indicates carbon dioxide was liberated.

PROCEDURE 5: NEUTRALIZATION OF ACIDS AND BASES

MEASURING THE CONCENTRATION OF AN ACID BY “TITRATION”:

1. Place exactly 10 drops of the unknown acid in the large cavity of the Chemplate or a test tube and add one drop of Phenolphthalein solution.

2. Add the 0.5M of NaOH one drop at a time. Carefully count the drops as they are added and stir with a new clean plastic stirring stick after each drop, until the acid solution turns pink and stays pink after stirring. (The acid at this point is neutralized.)

3. Calculate the concentration of the unknown acid.
   Volume (# of drops) of the base x concentration of base = Volume (#of drops of the acid) x concentration of the acid

$$\frac{\text{\# of drops of base}}{\text{\# of drops of acid}} = \frac{\text{concentration of acid}}{\text{concentration of base}} \quad \text{or} \quad \frac{\text{\# of drops of base}}{\text{\# of drops of acid}} = \frac{X}{\text{concentration of base}}$$

CONCENTRATION OF THE UNKNOWN ACID ________________.

Complete and balance the following equation.

$KOH + HCl \rightarrow \qquad\qquad + H_2O$

NOTES:

Chemistry 111 – Exercise 12: Titration of Commercial Vinegar (5%)

JEC

Discussion: Titration, a volumetric method of quantitative analysis, uses a standard solution (a solution of known concentration) to determine the concentration of another solution (a solution of unknown concentration). Titration is a method of choice for quantitative analysis of acids and bases. In this exercise, titration is used to calculate the weight percentage of acetic acid in vinegar. Acetic acid has a low pH. It reacts with sodium hydroxide to form water and sodium acetate. When all of the acetic acid has reacted to form sodium acetate, the pH of the solution is neutral (7.0). Any further addition of sodium hydroxide causes the pH of the solution to rise dramatically becoming basic. The Phenolphthalein changes from colorless to pink as the pH rises above 8. One mole of sodium hydroxide reacts with one mole acetic acid to form Sodium acetate, which is neutral: $NaOH + CH_3COOH \rightarrow NaCH_3COO + H_2O$ The mass of acetic acid can be calculated from the molarity of the acid measured by the titration.

EQUIPMENT NEEDED: Buret w/stopcock, Buret clamp, 10ml Volumetric pipette, 50ml Beaker, 100ml Beaker, 125ml Erlenmeyer flask and rubber stoppers, Small glass funnel, Watch glass, Ring stand, and Electronic Balance.

CHEMICAL NEEDED: 0.40M NaOH Sodium Hydroxide Solution, Vinegar Sample, Phenolphthalein Solution.

PROCEDURES:

Step 1: Wash, rinse and dry buret w/stopcock.
Step 2: Using a 50ml beaker and glass funnel pour 10ml of .40M NaOH in buret with the stopcock closed, then stopper and invert several times. Remove stopper and discard waste NaOH in the proper container.
Step 3: Using the buret clamp, clamp the buret to the ring stand. (Be sure the stopcock is closed).
Step 4: Using a 50ml beaker and glass funnel, fill the buret to the 50ml level with 0.40M sodium hydroxide solution.
Step 5: Record the buret reading to the nearest 0.01 ml on the data sheet.
Step 6: Obtain 80 to 100 ml of vinegar sample in a clean, dry beaker. Cover the beaker with a watch glass.
Step 7: Wash a 125ml Erlenmeyer flask & stopper in soapy water and rinse with lots of tap water then rinse with a small sample of vinegar and completely dry a 125ml Erlenmeyer flask and stopper.
Step 8: Weigh the Erlenmeyer flask and stopper to the nearest 0.01 gram on the electronic balance and record the reading on the data table.
Step 9: After rinsing out the 10 ml volumetric pipette with a small portion of vinegar, pipette exactly 20 ml of the vinegar sample (2 * 10ml = 20ml) into the clean dry 125ml Erlenmeyer flask and stopper the flask.
Step 10: Weigh the flask, stopper and contents on the electronic balance to the nearest gram and record the reading on the data table.

Step 11: Add 2 drops of phenolphthalein indicator to the flask and stopper the flask to prevent evaporation.

Step 12: Remove the stopper and immediately, place flask beneath the buret and slowly open the stopcock. Using a slow drop rate, titrate to the end point. (The point at which the solution changes from colorless to a faint pink that persists for at least 30 seconds). (A white sheet of paper under the flask aids in observing the color change.)

Step 13: Record the amount of 0.40M NaOH used.

Step 14: Repeat the titration two more times.

Step 15: Discard all chemical in the proper container / Clean and Dry All Glassware.

DATA TABLE

Mass of the flask, stopper and vinegar sample: ____________________

Mass of the flask, stopper: ____________________

Mass of the vinegar sample: ____________________

Molarity of sodium hydroxide: ____________________

Volume of vinegar weighed: ____________________

Volume of the vinegar titrated: ____________________

Molecular weight of acetic acid (Molar Mass): ____________________

Titration Results

Titration trial number	Volume added (ml)
(1)	__________________________
(2)	__________________________
(3)	__________________________

Average Volume (ml) added: ______________________________

Calculations:

a) Average volume of sodium hydroxide (base) added. ________(ml)
(From titration results)

b) Molarity of acetic acid = $\frac{\text{(Volume of the base) (Molar concentration of the base)}}{\text{Volume of the acid}}$ _____(M)

c) Density of vinegar: D = M/V ___________ (g/ml)

d) Grams/liter of acetic acid in vinegar = (Molarity of acetic acid) (Molar Mass of acetic acid)

________ (g/liter)

e) Grams/ml of acetic acid in vinegar = $\frac{\text{Grams/liter of acetic acid in vinegar}}{1000}$ _______ (g/ml)

f) Percentage of acetic acid in vinegar: % = $\frac{\text{Grams/ml of acetic acid in vinegar}}{\text{Density of vinegar}}$ (100) =

Percentage of acetic acid in vinegar = ________________%

QUESTIONS:

1. What is the purpose of rinsing the pipette with vinegar before using it?

2. What is the purpose of the phenolphthalein?

3. Why would you not want the vinegar to evaporate?

4. Describe the correct method for reading the volume of solution in a buret.

NOTES:

Chemistry 111 – Exercise 13: Demonstration only - Electrolytes
JEC

An electric current is the movement of electrically charge carriers. In metallic conductors, the charge carriers are electrons. In ionic substances in the molten state or in solution, current is conducted by ions. Positively charged ions (cations) migrate toward the negatively charged electrode (the cathode). Negatively charged ions (anions) move in the opposite direction toward the positively charged electrode (the anode).

Electrons can move through a crystal lattice (in a band of orbits called the conduction band), but ions cannot. Ions must be free to move in the liquid or gaseous state before they can conduct an electric current. Ionic conductors in general are called electrolytes.

In this experiment we will examine a number of substances to see whether they are electronic conductors, electrolytes, or nonconductors, non-electrolytes. From the previous remarks it is evident that a solid conductor must be electronic. If a solution is conductive, it must be ionic. The conductivity of solutions depends to large extent on the concentration of charge carriers or ions. If a substance is totally ionized in solution, the concentration of charge carriers is high, and the solution will be a good conductor. Such solutions are strong electrolytes. If the substance is only slightly ionized in solution, the concentration of ions is low, this solution will have low conductance and will therefore be a weak electrolyte.

PROCEDURE

The instructor will test a number of substances with a qualitative conductance device; a 12-volt circuit with a light bulb is constructed so that the test substances can be placed between the electrodes. If the substance is a good conductor, the light bulb will glow brightly. With a weak electrolyte, the bulb will glow dimly. If the substance is a non-electrolyte the bulb will not glow.

Record your observations in the data table provided. Note the physical state of the substance and whether it appears to conduct electricity well, weakly, or not at all.

DATA TABLE

CONDUCTION OF ELECTROLYTES

Conductance

TEST SUBSTANCE	STRONG	WEAK	NONE

QUESTIONS

1. What is an electrolyte?

2. How does reduction occur at the cathode?

3. How does oxidation occur at the anode?

NOTES:

Chemistry 111 - Exercise 14: Organic / Biochemistry

JEC

Purpose: To illustrate some of the chemical methods that are used to identify or test for the presence of various organic compounds, such as carbohydrates, lipids (fats), and proteins.
To demonstrate and establish the presence, nature and action of enzymes.

Equipment needed: Test tubes, Test tube rack, Test tube clamp (holder), 50ml Beakers, 600ml Beaker, Mortar and Pestle, Glass stirring rod, Glass marker, Knives, Cutting or chopping board, Filter paper, Hot plate.

Chemicals needed: Dropping bottles of the following: Benedict's solution; Biuret reagent; Potassium iodide solution; Sudan IV.
Distilled water, 4% Glucose solution (sugar solution),
2% Starch solution, .5% Protein solution, .5% Lipid solution,
and Food samples.

TESTING METHODS:

Sugar Test: *Benedict's* solution is a reagent that will test for the presence of sugar. Place test solution into a test tube and add benedict's solution. Then place the test tube containing the substances into boiling water. Leave the test tube in the boiling water anywhere from 30 seconds to 2 minutes. If a green color appears then sugar is present.
(Reducing sugars will change from green → yellow → orange → rusty red precipitate.)
If the solution remains blue or turns any color other than the one(s) describes above, then the test is negative for sugar.

Starch Test: *Potassium iodide* solution is a reagent that will test for the presence of starch. Place test solution into a test tube, then add the Potassium iodide solution, the appearance of a deep purple or bluish-black color indicate the presence of starch. (If its glycogen a starch-like polysaccharide found in animal tissue, then a dark mahogany red color will appear.) If there is not a nocticable change in the color of the iodine solution, then the test is negative for starch.

Lipid Test: *Sudan IV* is a reagent that will test for the presence of a fat (lipid). Sudan IV is fat soluble; place test solution into a test tube and add Sudan IV, the appearance of a bright red percipitate or color indicates the presence of a fat (lipid). If the solution reamin pale red, then the test is negative for lipids.

Protein Test: *Biuret's* solution is a reagent that will test for the presence of proteins. Place test solution into a test tube and add biuret's solution, the apperance of a pinkish-lavender (violet) color (a lavender or violet color with a pinkish tint), indicates the presence of a protein. *The reagent reacts with the peptides, the building units of proteins, because of the nature of this reaction it will reverve itself and the color will fade.* If the solution does not change from it orginal color, then the test is negative for proteins.

Procedure 1:

Step 1: Put 400ml of tap water into a 600ml beaker and place it on a hotplate and bring the water to a boil.

Step 2: Set up twenty test tubes into the test tube rack, then label them as follows:
a set of four test tubes labeled # 1; a set of four test tubes labeled # 2;
a set of four test tubes labeled # 3; a set of four test tubes labeled # 4;
a set of four test tubes labeled # 5.

Step 3: Put exactly 2ml of water into a 10ml graduate cylinder, then pour the water into one of the test tubes and use a glass marker to mark the 2ml level on the test tube. Pour the water out and use that test tube to mark the 2ml level for the other 19 test tubes.

Step 4: Using the appropriately marked 50ml beaker pour 2ml of each of the following solution into the first set of five test tubes:
Test tube # 1 – distilled water
Test tube # 2 – glucose (sugar) solution
Test tube # 3 – starch solution
Test tube # 4 – lipid solution
Test tube # 5 – protein solution

Step 5: Add 2ml (about two droppers full) of benedict's solution into each of the first five sets of test tubes and simultaneously place all of them into the boiling water bath for .5 to 1 minute. Watch for a color change. When a color change occurs in a single test tube immediately remove all of the other test tubes and place them in the rack and observe the remaining test tube until the reaction is complete. Record your results on the data table.

Step 6: Using the setup in step 4 add five to ten drops of potassium iodide solution into each of the set of five test tubes. Watch for a color change. Record your results on the data table.

Step 7: Using the setup in step 4 add five to six drops of sudan IV to each of the third set of five test tubes place a rubber stopper into each tubes and shake vigorously for a minute or two. Place each tube back into the rack and wait five minutes. Record your results on the data table.

Step 8: Using the setup in step 4 add ten drops of biuret's reagent into each of the last set of five test tubes. Watch for a color change. Record your results on the data table.

Step 9: Discard all solutions into the appropriate container and wash and dry all of the test tubes.

Procedure 2: (Food testing)

Step 1: For solid food cut and dice sample on a cutting board, then place the diced food into the mortar and crush and grind with the pestle. All other foods **except** for the raw egg can be added directly to mortar and crush and grounded with the pestle.
Step 2: Add distilled water until the mortar is one-third full and stir vigorously with a glass stirring rod for several minutes. Then decant the liquid into a 50ml beaker. Remove the solid material from the mortar and discard into the appropriate container, and wash and dry the mortar and pestle.
Step 3: Place 2ml of the liquid from the food sample into a set of four test tubes marked 1 through 4.
Step 4: Add 2ml of benedict's solution into the test tube marked # 1, then place the tube into the boiling water bath for 1 to 2 minutes. Record results on the data table.
Step 5: Add five to ten drops of potassium iodide solution into test tube marked # 2 and record the results on the data table.
Step 6: Add ten drops of Sudan IV into the test tube marked # 3 stopper, shake vigorously let stand for five minutes and record the results.
Step 7: Add ten drops of biuret's solution into test tube marked # 4 and record the results on the data table.
Step 8: Repeat steps 1 – 7 until all of the sample foods have been tested.

Data Table

Item #	Solution Or Food Sample	Results of the Sugar Test	Results of the Starch Test	Results of the Lipid Test	Results of the Protein Test
1.	Distilled Water				
2.	Glucose Solution (Sugar)				
3.	Starch Solution				
4.	Lipid Solution				
5.	Protein Solution				
6.					
7.					
8.					
9.					
10.					
11.					
12.					
13.					
14.					
15.					
16.					
17.					
18.					
19.					
20.					

Procedure 3: ENZYMATIC ACTION

Step 1: Salivate into a test tube and add 2ml of benedict's reagent to the saliva, then place it into the the boiling water bath for 1 to 2 minutes. If it does not turn green, then move to Step 2:

Step 2: Obtain a sugarless cracker from the instructor and chew the cracker for several minutes without swallowing any of the cracker. Then collect the mass in a 50ml beaker and add enough distilled water to make a solution. Stir vigorously for three or four minutes.

Step 3: Decant 2ml of the liquid into two separate test tubes. Discard the remaining materials into the appropriate container and wash and dry the beaker.

Step 4: Add 2ml of benedict's reagent to one of the test tubes, then place it into the boiling waterbath for two or three minutes and record your results.

Step 5: Add five drops of potassium iodide solution to the other test tube and record the results.

Step 6: Clean and dry all glassware and return to proper storage or to the cart. Wash and dry your work area, and discard all unused foods into the appropriate container.

Note: Salivary amylase (an enzyme) coverts starch into sugar.

Observation and Results:

NOTES:

Student Agreement

Many of the reagents (chemicals) and some equipment in a chemistry laboratory are potentially dangerous. Following rules of laboratory safety and using common sense throughout the course will enhance the learning experience and hopefully increase the student's confidence in his or her ability to safety use chemical and equipment.

I have read the laboratory safety rules and procedure, and/or have listened to the instructor's lecture on laboratory safety. I understand these rules and I agree to the follow safety rules and procedures. Also, I agree to abide by any additional instruction, written or verbal provided by my instructor, the laboratory manager or "Chemical Hygiene Officer".

___________________________________ ______________ _______________

(***Laboratory Course***) ***(Lab Day)*** ***(Time)***

____

(Please print your name legibly)

____ ______________________

(*Student's Signature Required)* (Date)

***(Please list below any medical condition or allergies that may create an emergency in the laboratory.)**

<u>Please turn this sheet in to your lab instructor.</u>

James E. Copeland, Chemical Hygiene Officer

Chemistry 111 - Exercise 1: Laboratory Safety Procedures and Policies
LABORATORY REPORT

Name: ______________________________ Date: _______________

Laboratory safety, procedures, and policies.

1. What do you think are the most common type of laboratory accidents and how would you prevent them?

2. Where would you carryout experiments that evolve toxic gases? Why?

3. If a large-scale fire erupts at your workstation, what must you do?

4. How would you extinguish a fire caused by a combustible metal?

5. In lab your neighbor's clothing has caught fire, what should you do?

6. How would you clean up: a) an acid spill? b) a base spill? c) broken glass?

7. How would you mix acids with water?

8. What is the number one cause of accidents in lab?

Chemistry 111 – Exercise 2: Measurements in Chemistry
LABORATORY REPORT

Name: ____ Date: ___________________

Measurements

List the skills, concepts and/or conclusions in a brief outline form that you have learned as a result of this exercise.

List the correct weighting procedure using the triple or quad beam balance.

What is the correct way to read the volume of a liquid in a graduated cylinder?

Chemistry 111 – Exercise 3: Percentage of water in a Hydrate

LABORATORY REPORT

Name: ____ Date: ___________________

Percentage of water in a hydrate.

1. What is a hydrate?

2. What process resulted in the formation of the anhydride in this experiment?
   (Be specific)

3. Define exothermic reactions.

4. Define endothermic reactions.

Chemistry 111 – Exercise 4: The Nature of Chemical Substances
Chemical Changes or Physical Changes
LABORATORY REPORT

Name: ____ Date: ____________________

1. What was the source of oxygen that united with the magnesium ribbon?

2. Was the burning of the magnesium ribbon a physical or chemical change?

3. Was the mixing of the charcoal with the sodium chloride to form a evenly disbursed powder a physical or chemical change?

4. Was the burning of the wood splinter a physical or chemical change?

5. What is meant by a chemical change?

6. What is meant by a physical change?

7. How would you compare the burning of the wood splinter (a fire) with the process of respiration (breathing)?

8. Why was a Pyrex test tube used in procedure 5?

Chemistry 111 – Exercise 5: Molecules, Formula Units and Compounds
LABORATORY REPORT

Name: ____ Date: ___________________

Write the correct chemical formula for the compounds formed by the combination of the following ions:

a) Potassium (K^{+1}) and Bromide (Br^{-1}) ________________________________

b) Mercury (Hg^{+2}) and Iodide (I^{-1}) ________________________________

c) Magnesium (Mg^{+2}) and Nitride (N^{-3}) ________________________________

d) Ferric (Fe^{+3}) and Chloride (Cl^{-1}) ________________________________

e) Cadmium (Cd^{+2}) and Oxide (O^{-2}) ________________________________

f) Calcium (Ca^{+2}) and Phosphide (P^{-3}) ________________________________

g) Lithium (Li^{+1}) and Oxide (O^{-2}) ________________________________

h) Silver (Ag^{+1}) and Nitrate (NO_3^{-1}) ________________________________

i) Sodium (Na^{+1}) and Phosphate (PO_4^{-3}) ________________________________

j) Aluminum (Al^{+3}) and Sulfate (SO_4^{-2}) ________________________________

k) Zinc (Zn^{+2}) and Chloride (Cl^{-1}) ________________________________

1) Bismuth (Bi^{+5}) and Oxide (O^{-2}) ________________________________

m) Cupric (Cu^{+2}) and Sulfide (S^{-2}) ________________________________

n) Magnesium (Mg^{+2}) and Cyanide (CN^{-1}) ________________________________

o) Ammonium (NH_4^{+1}) and Sulfite (SO_3^{-2}) ________________________________

q) Sodium (Na^{+1}) and Bicarbonate (HCO_3^{-1}) ________________________________

r) Ferric (Fe^{+3}) and Sulfide (S^{-2}) ________________________________

s) Calcium (Ca^{+2}) and Fluoride (F^{-1}) ________________________________

t) Ferrous (Fe^{+2}) and Phosphate (PO_4^{-3}) ________________________________

Chemistry 111 – Exercise 8: Determining Chemical Formulas – Quantitative Analysis
LABORATORY REPORT

Name: ____ Date: ___________________

1. Hydrogen sulfide gas is (soluble - insoluble) in water. (**Circle your choice**)

2. Metals ions react with Hydrogen sulfide solution forming _____________________ that has a specific __________________.

3. Ammonia water had to be added to _______________________ metal ionic solutions before they would react with the hydrogen sulfide solution.

4. In the reaction between Calcium chloride and Sodium carbonate which test tube had the greatest amount of precipitate?

   Test tube # _________. What was the ratio of Ca ions to CO_3 ions? ______ : ______

5. How could you experimentally determine the correct chemical formula for the new liquids formed?

Write balanced equations for the following:

6. the reaction between Nickel(II) nitrate and Hydrogen sulfide.

7. the reaction between Tin(II) chloride and Hydrogen sulfide.

8. Complete and balance the following equation:

   $CaCl_2$ + Na_2CO_3

Chemistry 111 – Exercise 9: Stoichiometry
Mass-Mass relationship in an Acid-base Reaction
LABORATORY REPORT

Name: ____ Date: ___________________

1. Give two reasons that the reaction in this experiment is irreversible.

2. Calculate your percent yield for this experiment.

Chemistry 111 – Exercise 10: Solutions

LABORATORY REPORT

Name: ____ Date: ___________________

1. Define the term solution.

2. Did the pinch of salt crystals dissolve in the saturated salt solution? ___________

   Did the pinch of sugar crystals dissolve in the saturated salt solution? __________

   Explain your answer.

3. How do temperature, surface area, and stirring affect the rate of dissolution?

4. Which liquids were miscible in each other?

Chemistry 111 – Exercise 13: Demonstration only - Electrolytes
LABORATORY REPORT

Name: ____ Date: ____________________

1. Did the metals react with the 6M HCl? ____________________

   Did the metals react with the 6M KOH? ___________________

2. Rate the activity of the metals with 6M HCl from (**1**) being the **most** active to (*4*) being the *least* active.

   Zinc _____ ; Iron _____ ; Copper ______ ; Magnesium ______.

3. When 6M HCl was added to Magnesium, the release of hydrogen gas was determined by

   ______________________________________, when tested with the burning wood splint.

4. Complete and balance the following equations:

   a) $Mg + \quad HCl \rightarrow \quad MgCl_2 + H_2$

   b) $KOH + \quad HCl \rightarrow \quad\quad\quad + H_2O$

5. Define the term electrolyte.

6. Name at **least** six strong electrolytes tested. ____________________________________;

   __________________________________; ______________________________________;

   __________________________________; ______________________________________;

   ____.

Chemistry 111 - Exercise 14: Organic / Biochemistry
LABORATORY REPORT

Name: ____ Date: __________________

1. Name the reagent that will test for the present of sugar. ____________________________

2. The occurrence of a deep blue-black color in a solution when tested with potassium – iodide solution indicates the presence of ____.

3. The appearance of a pinkish – lavender color in the biuret's reagent indicates the presents of _______________________________________.

4. If a lipid is present in a substance Sudan IV forms a ______________________________.

5. ________________________ is the enzyme in salvia that is capable of breaking down starch.