LAWRENCE HIGH SCHOOL
CHEMISTRY
LAB MANUAL
VOLUME 1

"Bunsen, I must tell you how excellent your study of chemical spectroscopy is, as is your pioneer work in photochemistry — but what really impresses me is that cute little burner you've come up with."

Table of Contents

Lawrence High School Science Laboratory Policy .....

Laboratory Equipment and Skills.....

Safety in the Chemistry Laboratory..... 1

Lab 1 Qualitative Observation of a Chemical Reaction..... 15

Lab 2 Quantitative Observations of a Chemical Reaction..... 19

Lab 3 Measuring Mass..... 23

Lab 4 Density Determination..... 31

Lab 5 Physical and Chemical Change..... 39

Lab 6 Mixture Separation..... 47

Lab 7 Heat of Fusion of Ice..... 51

Lab 8 Heat of Crystallization of Wax..... 57

Lab 9 Heating and Cooling Curves..... 61

Lab 10 Sub Atomic Particles..... 67

Lab 11 Periodic Trends in Atomic Radii..... 71

Lab 12 Nuclear Marbles..... 75

Lab 13 Electrons in an Atom..... 81

Lab 14 Electron Structure 85

Lab 15 Candium..... 91

Lab 16 Black Box..... 95

Lab 17 Flame Test..... 99

Lab 18 Emission Spectra..... 103

Lab 19 Half Life..... 109

Lab 20 Periodic Puzzle..... 115

Lab 21 It's in the Cards..... 117

Lab 22 Group 2A..... 125

Lab 23 Hydrogen..... 131

Lab 24 Lewis Dot Diagrams..... 135

Lab 25 Ionic Formula Writing..... 143

Lab 26 Bonds, Polarity, Solubility..... 149

Lab 27 Names and Formulas Ionic Compounds..... 155

Lab 28 Chemical Reactions..... 161

Lab 29 Conservation of Matter..... 167

Lab 30 Relating Moles to Coefficients..... 171

Lab 31 Mole-Mass Relationships..... 177

Lab 32 Mass-mass Relationships..... 181

Lab 33 % Water in Pop Corn..... 185

Lab 34 Ornament

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Lawrence High School Science Laboratory Policy

As stated in the new Standards from the New York State CORE Curricula Guides for all science regents courses: "as a prerequisite for admission to the Regents exams (in science) students must have successfully completed 1200 minutes (usually 30 labs each 40 minutes in duration) of laboratory experiences with satisfactory reports on file."

To help each student achieve a successful laboratory experience, the following is the mandated laboratory policy at Lawrence High School:

1. **All laboratory reports must be submitted during the marking period in which the laboratory experience is completed.** Laboratory reports will not be accepted after the end of that marking period.

2. **All students must complete a minimum of 70% of the laboratory reports required for each marking period to receive a passing grade in the course.** Students who fail to submit the 70% will receive a maximum marking period grade of 60, regardless of the academic work already achieved in that course.

3. Laboratory work will have a weight of at least 20% of the grade for each marking period.

4. A completed lab report does not mean it is "satisfactory" for a grade that can be used towards the state mandated 1200 minutes. The science teacher must accept the lab report as satisfactory with a passing grade. (Each teacher may establish his/her own policy regarding the time requirement for the submission of the report during a marking period. Further, there is no departmental policy regarding the formatting of reports.)

5. Laboratory reports that are submitted late during a marking period may be considered for Regent credit (towards the 1200 minutes), but may be considered too late for a passing grade.

6. Some teachers may require students to maintain their own laboratory folder throughout the school year. It is the student's responsibility to present the completed folder at the end of the year in order to receive the laboratory credit.

Date _____ Teacher_____

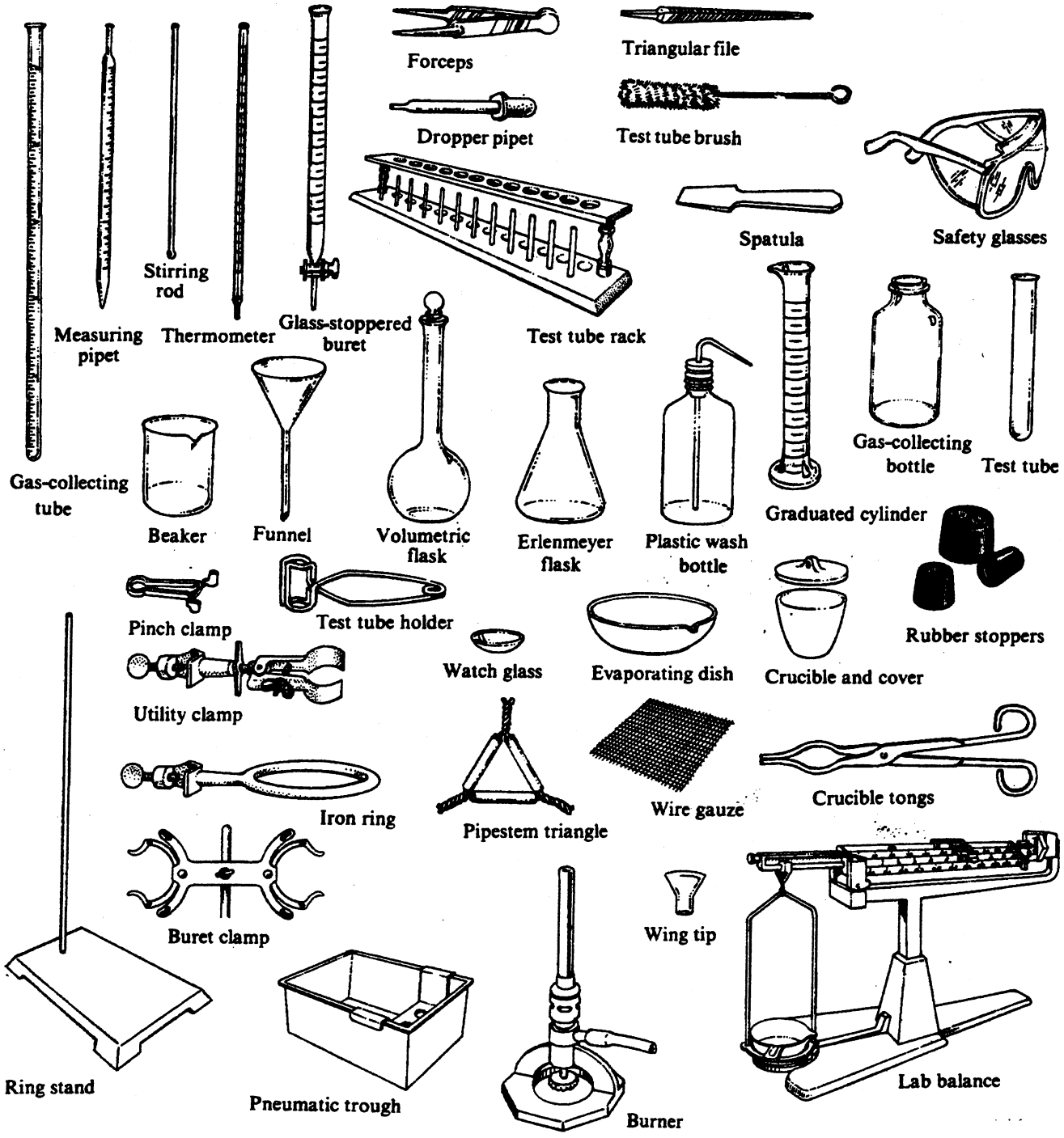
Student's signature_____

Parent's signature_____

Laboratory Equipment and Skills

RECOGNIZING LAB EQUIPMENT

The equipment you will be using most frequently in the laboratory is illustrated in Figure 1. Study this figure carefully and familiarize yourself with each item.



Name_____Date_____

Safety in the Chemistry Laboratory

Introduction

Some of your most exciting and rewarding high school experiences will be found in the chemistry laboratory. You will be actively involved in each lab experience. From your observations and analysis of the data you collect, you will draw conclusions from your experiment. You are going to be working with equipment and chemicals that can cause serious injury to yourself or others if you are not careful in handling them in a safe manner. Accidents do not happen by themselves. They are caused by carelessness, rushing to get done, and not adhering to safety rules. You must read the rules that follow and adhere to the rules during every lab experience you perform.

General safety guidelines

1. Be prepared to work as soon as you arrive to class.
2. Read the lab procedure before you perform a lab.
3. Never do any procedure not assigned or instructed by your teacher.
4. Work station must remain clean and uncluttered.
5. Loose clothing and exposed bodies are inappropriate and can result in your injury.
6. Long hair should be tied back and loose jewelry should be removed.
7. Follow all instructions carefully.
8. Safety goggles and aprons should be worn during each lab.
9. Keep flammables and combustibles away from open flames.
10. Never touch or taste anything that you have not been instructed to do so.
11. Never put your face near the mouth of a container holding chemicals or being heated.
12. When identifying an odor, use a wafting motion.
13. Any use of poisonous vapors will be conducted under a fume hood.
14. Dispose of waste materials as instructed.
15. Clean spills immediately.
16. Clean your area at the end of each lab.
17. Report all accidents to your instructor immediately.

Handling chemicals

18. Read each label twice before using any reagent.
19. When mixing an acid and water, ***always add the acid to the water.***
20. Do not touch chemicals with your hands.
21. If you touch any chemical with your hands, wash your hands right away.

Handling glassware

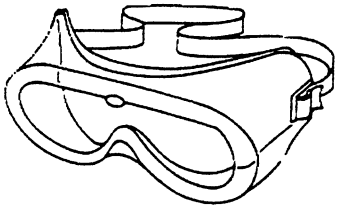
22. Never handle broken glass with your bare hands.
23. Dispose of broken glass as instructed by your teacher.
24. Allow hot glass to cool before touching it.

Heating substances

25. Be cautious when working with flames.
26. Turn the burner of if it is not being used.
27. Do not heat anything that you have not been told to heat.
28. Never point the mouth of the glassware you are heating towards anybody.
29. Never leave an open flame unattended.

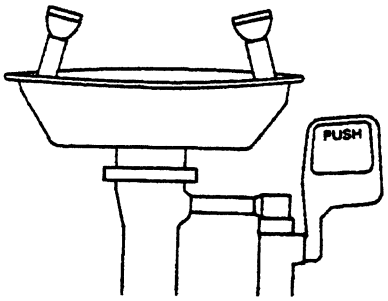
Protection for the eyes

Government regulations, as well as common sense, demand the wearing of protective eyewear while you are in the laboratory. Such eyewear must be worn even if you, personally, are not working on an experiment.



Although you may not use this particular type of goggle, your eyewear must include shatterproof lenses and side shields that will protect you from splashes. ***Safety glasses must be worn at all times while you are in the laboratory.*** Failure to wear safety glasses may result in your being failed or withdrawn from your chemistry course, or in some other disciplinary action being taken.

In addition to protective goggles, an eyewash fountain provides eye protection in the laboratory. Should a chemical splash near your eyes, you should use the eyewash fountain before the material has a chance to run behind your safety glasses.



The eyewash has a panic bar that enables the eyewash to be activated easily in an emergency. If you need the eyewash, don't be modest- use it immediately. It is critical that you protect your eyes properly.

Protection from fire

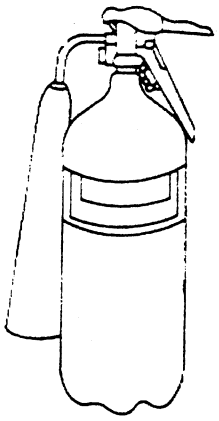
The danger of uncontrolled fire in the chemistry laboratory is very real, since the lab typically has a fairly large number of flammable liquids in it, and open-flame gas burners are generally used for heating. With careful attention, however, the danger of fire can be reduced considerably.

Always check around the lab before lighting a gas burner to ensure that no one is pouring or using any flammable liquids. Be especially aware that the vapors of most flammable

liquids are heavier than air and tend to concentrate in sinks (where they may have been poured) and at floor level. Since your laboratory may be used by other classes, always check with your instructor before beginning to use gas burners.

The method used to fight fires that occur in spite of precautions being taken depends on the size of the fire and on the substance that is burning. If only a few drops of flammable liquid is nearby, the fire can usually be put out by covering it with a beaker. This deprives the fire of oxygen and will usually extinguish the fire in a few minutes. Leave the beaker in place for several minutes to ensure that the flammable material has cooled and will not flare up again.

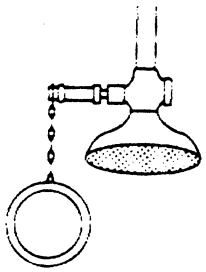
In the unlikely event that a larger chemical fire occurs, carbon dioxide fire extinguishers are available in the lab.



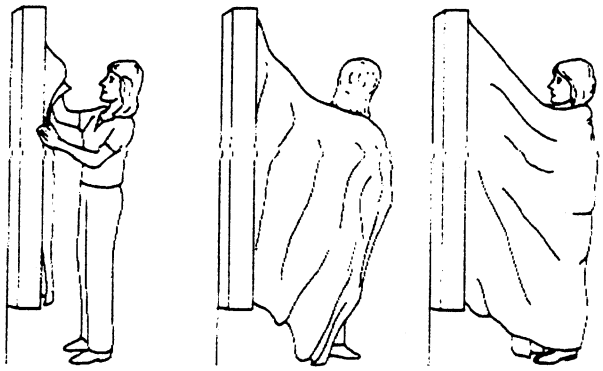
Before activating the extinguisher, pull the metal safety ring from the handle. Direct the output from the extinguisher at the base of the flames. The carbon dioxide not only smothers the flames, it also cools the flammable material quickly. If it becomes necessary to use the fire extinguisher, be sure afterward to turn the extinguisher in to be refilled. If the carbon dioxide extinguisher does not immediately put out the fire, evacuate the laboratory and call the fire department. Carbon dioxide fire extinguishers must ***not*** be used on fires involving magnesium or certain other reactive metals, since carbon dioxide may react vigorously with the burning metal and make the fire worse.

One of the most frightening and potentially tragic accidents is the igniting of a person's clothing. For this reason, certain types of clothing are not appropriate for the laboratory and must not be worn. Since sleeves are most likely to come in closest proximity to flames, any garment that has bulky or loose sleeves should not be worn in the laboratory. Certain fabrics should also be avoided; such substances as silk and certain synthetic materials may be highly flammable. Long hair also presents a clear danger, if it is allowed to hang loosely in the vicinity of the flame. Long hair must be pinned back or held with a rubber band.

In the unlikely event a student's clothing or hair is ignited, you should take prompt action to prevent severe burns. Most laboratories have two options for extinguishing such fires: the water shower and the fire blanket.



In the event that someone's clothing or hair is on fire, immediately push or drag the person to the shower and pull the metal ring of the shower. Safety showers generally dump 40-50 gallons of water, which should extinguish the flames. Be aware that the showers cannot be shut off once the metal ring has been pulled. For this reason, the shower cannot be demonstrated.



The fire blanket is intended to smother flames.

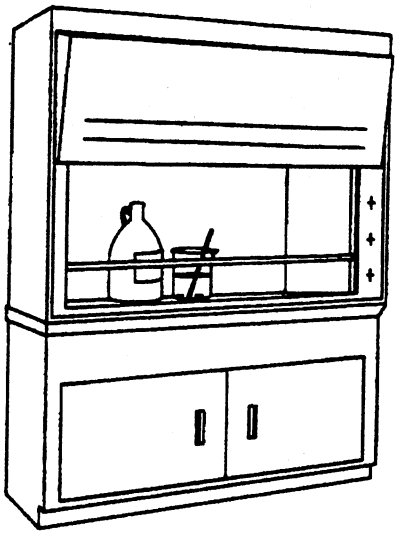
Protection from chemical burns

Most acids, alkali, oxidizing, and reducing agents must be assumed to be corrosive to skin. It is impossible to avoid these substances completely, since many of them form the backbone of the study of chemistry. Usually, a material's corrosiveness is proportional to its concentration. Most of the experiments in this manual have been set up to use as dilute a mixture of reagents as possible, but this does not entirely remove the problem. Make it a rule to wash your hands regularly after using any chemical substance and to wash immediately, with plenty of water, if any chemical substance is spilled on the skin. After working with a substance known to be particularly corrosive, you should wash your hands immediately even if you did not spill the substance. Someone else using the bottle of reagent may have spilled the substance on the side of the bottle. Do not make the mistake of thinking that because an experiment calls for dilute acid, this acid cannot burn your skin. Some of the acids used in the laboratory are not volatile, and as water evaporates from a spill, the acid becomes concentrated enough to damage skin. Whenever a corrosive substance is spilled on the skin, you should inform the instructor immediately.

In the event of a major chemical spill, in which substantial portions of the body or clothing are affected, you must use the emergency water shower.

Protection from toxic fumes

Many volatile substances have toxic vapors. A rule of thumb for the chemistry lab is "If you can smell it, it can probably hurt you." Some toxic fumes (such as those of ammonia) can overpower you immediately, whereas other toxic fumes are more insidious. The substance may not have had that bad an odor, but its fumes can do severe damage to the respiratory system. There is absolutely no need to expose you to toxic fumes. Most chemistry laboratories are equipped with fume exhaust hoods.



The exhaust hood has fans that draw vapors out of the hood and away from the user. The hood is also used when flammable solvents are required for a given procedure, since the hood will remove the vapors of such solvents from the laboratory and reduce the hazard of fire. The hood is equipped with a safety-glass window that can be used as a shield from reactions that could become too vigorous. Naturally, the number of exhaust hoods available in a particular laboratory is limited, but never neglect to use the hood if it is called for, merely to save a few minutes of waiting time. Finally, reagents are sometimes stored in a hood, especially if the reagents evolve toxic fumes. Be sure to return such reagents to the designated hood after use.

Protection from cuts and simple burns

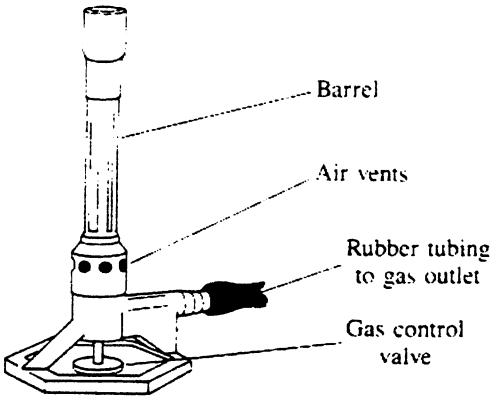
Perhaps the most common injuries to students in the beginning chemistry laboratory are simple cuts and burns. Glass tubing and glass thermometers are used in nearly every experiment and are often not used properly.

Most simple burns in the laboratory occur when a student forgets that an apparatus may be hot and touches it. Never touch an apparatus that has been heated until it has cooled for at least five minutes, or unless specific tongs for the apparatus are available.

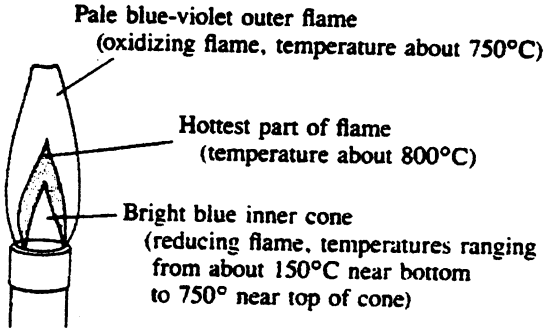
Report any cuts or burns, no matter how apparently minor, to the instructor immediately. What may seem like a scratch may be adversely affected by chemical reagents or may become infected; therefore, it must be attended to by trained personnel. Broken glass should be placed in a box marked for broken glass.

Proper use of the laboratory burner

The laboratory burner is the most commonly used apparatus in the general chemistry laboratory, and may pose a major hazard if not used correctly and efficiently.
The typical burner is correctly called a Tirrel burner (Bunsen burner). Compare the burner you will be using with the burner shown in the diagram.



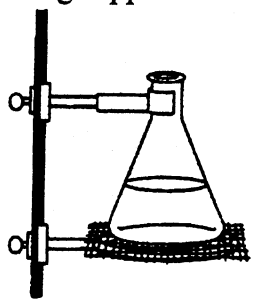
Most laboratories are supplied with natural gas, which consists primarily of the hydrocarbon methane (CH^4). A length of thin-walled rubber tubing should attach your burner to the gas main jet. If your burner has a screw-valve on the bottom for controlling the flow of gas, the valve should be closed (by turning in a right hand direction) before you light the burner. The barrel of the burner should be rotated to close the air vents. To light the burner, turn the gas main jet to the open position. If your burner has a screw-valve on the bottom, open this valve until you hear the hiss of gas. Without delay, user your striker to light the burner. After lighting the burner, the flame is likely to be yellow and very unstable. The flame at this point is relatively cool and is inefficient for heating. To make the flame hotter and more stable, open the air vents on the barrel of the burner slowly to allow oxygen to mix with the gas before combustion. This should cause the flame's size to decrease and its color to change. A proper mixture of air and gas gives a pale blue flame, with a bright blue cone-shaped region in the center. The hottest part of the flame is directly above the tip of the bright blue cone. Whenever as item is to be heated strongly, it should be positioned directly above this bright blue cone.



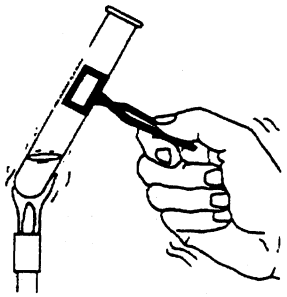
Protection from apparatus accidents

An improperly constructed apparatus can create a major hazard in the laboratory, not only to the person using the apparatus, but to his or her neighbors as well. Any apparatus you set up should be constructed exactly as indicated in this manual. If you have any questions as to whether you have set up the apparatus correctly, ask your instructor to check the apparatus before you begin the experiment.

Perhaps the most common apparatus accident in the lab is the tipping over of a flask or beaker while it is being heated or otherwise manipulated. All flasks should be clamped securely with an adjustable clamp to a ring support.



An all too common lab accident occurs when a liquid must be heated to boiling in a test tube. Test tubes hold only a few milliliters of liquid and require only a few seconds of heating to reach the boiling point. A test tube cannot be heated strongly in the direct full heat of the burner flame. The contents of the test tube will super-heat and blow out of the test tube like a bullet from a gun. Ideally, when a test tube requires heating, a boiling water bath in a beaker should be used. If this is not possible, then hold the test tube at a 45°angle a few inches above the flame and heat only briefly, keeping the test tube moving constantly (from top to bottom, and side to side) through the flame during the heating. ***Aim the mouth of the test tube away from yourself and anyone else.***



Caution Alert symbols

The symbols shown below are used throughout this lab manual at points where extra caution should be exercised. Whenever you see one of these symbols, stop, read the material carefully, and proceed with extra care. If you have any questions, ask your teacher before going on.

Safety Clothing

Safety Glasses

These two symbols appear at the beginning of each experiment. They are to remind you that safety glasses and a lab apron (or coat) are to be worn *at all times* when working in the lab. For some activities, your teacher may also instruct you to wear protective gloves.

Fire

This symbol indicates the presence of an open flame. Loose hair should be tied back or covered, and bulky or loose clothing should be secured in some manner.

Caustic Substance

This symbol indicates a caustic or corrosive substance — most frequently an acid. Avoid contact with skin, eyes, and clothing. Do not inhale vapors.

Breakage

This symbol indicates an activity in which the likelihood of breakage is greater than usual, such as working with glass tubing, funnels, etc.

Dangerous Vapors

This symbol indicates the presence of or production of poisonous or noxious vapors. *Use the fume hood* when directed to do so. Care should be taken not to inhale vapors directly. When testing an odor, use a wafting motion to direct the vapor toward your nose.

Explosion

This symbol indicates that the potential for an explosive situation is present. When you see this symbol, read the instructions carefully and *follow them exactly.*

Poison

This symbol indicates the presence of a poisonous substance. Do not let such a substance come in contact with your skin and do not breathe its vapors.

Electrical Shock

This symbol indicates that the potential for an electrical shock exists. Read all instructions carefully. Disconnect all apparatus when not in use.

Radiation

This symbol indicates a radioactive substance. Follow your teacher's instructions as to the proper handling of such substances.

Safety Quiz

1. Sketch your laboratory room, showing the locations of the exits, fire extinguishers, safety shower, eyewash fountain, fire blankets, and fume hoods.

2. Why must you wear safety goggles at all times while you are in the laboratory?

3. Suppose there is a spill of a flammable substance near an open flame, which is then ignited. How should this fire be extinguished?

4. What should you do if a larger fire were to break out in the room?

5. What should you do if a student's clothing caught fire?

6. List 5 types of clothing or footwear that are not acceptable.

7. What should you do if you accidentally pour acid on your clothing?

8. What should you do if you spill acid on your skin?

9. What are the most common injuries in the lab? How can they be prevented?

10. Where should reactions that produce toxic gases be performed?

11. Why should you clean up any spill right away?

12. Why should you never change the procedure for an experiment?

13. Why should apparatus always be clamped securely to a metal ring stand before starting a chemical reaction?

14 Why should you never eat or drink while you are in the laboratory?

15. Suppose the experiment you are to perform in a lab involves substances that are toxic or corrosive. What should you do before the laboratory to prepare yourself for using these substances safely?

16. Complete the paragraphs below, using words from the word list.

Word list

accident	fire
acid	fume
apron	instructions
burner	loose
combustible	reagent
dangling	return
disposal	spill
taste	water

When working in the laboratory you should perform only those activities assigned by your teacher. Carefully follow all _____ both written and oral. Wear safety glasses and a lab _____ at all times. Avoid apparel that presents safety hazards, such as _____ garments and _____ jewelry. Keep your hair, clothing, and all other _____ materials away from open flames.

Never touch or _____ any substance unless you are specifically instructed to do so by your teacher. Any activity that involves poisonous vapors must be carried out in a(n) _____ hood. Always read and double check the label on a (n) _____ bottle before removing anything from it, and never _____ unused material to the bottle. If you are directed to mix an acid and water, always be sure to add the _____ to the _____, never the reverse. Never use a gas _____ to heat or burn anything unless you are instructed to do so, and be sure to turn it off when it is not in use.

Clean up any _____ immediately. Follow your teacher's instruction for _____ of any waste material. Know the location of all emergency equipment, such as first aid kits and _____ extinguishers. Report any _____ to your teacher immediately.

17. Use the diagram on the next page to answer these questions.
    a. List 3 unsafe activities shown in the illustration and explain why each is unsafe.

    b. List 3 correct lab procedures.

    c. What should Bob do after the accident?

    d. What should Sue have done to avoid an accident?

    e. Compare Luke and Duke's lab techniques. Who is following the rules?

    f. What are 3 things shown in the lab that should not be there?

    g. Compare Joe and Carl's lab techniques. Who is doing it the correct way?

    h. What will happen to Ray and Tim when the teacher catches them?

i. List 3 items in the illustration that are there for the safety of the students in the lab.

Name_____ Date_____

1 Qualitative Observations of a Chemical Reaction

Introduction

The single thing that makes science different from all other disciplines is that scientists do experiments. One important part of any experiment is making careful observations. These observations must be objective. Being objective means putting aside any pre-conceived ideas. Scientists must study what *really* occurs; not what they *wish* would occur. After observations are made and data is collected, scientists must interpret the data and draw a conclusion.

In this investigation, you will try to solve a problem by doing a series of investigations in a logical way and by making *qualitative observations* of a chemical reaction. After you carefully describe and analyze what is happening, you will be able to come up with your own conclusion to the problem.

Problem

When a candle burns, what is actually burning?

Introductory Questions

1. What is a qualitative observation?

2. Why do scientific observations have to be objective

3. List all the parts of a lighted candle that could actually burn

4. What do *you* think is actually burning in a candle?

Equipment

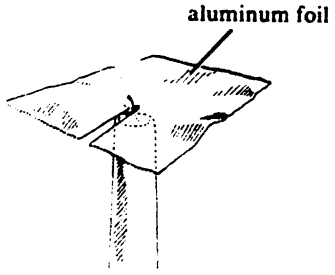
beaker, 250 mL
Erlenmeyer flask, 125mL
glass square
microspatula
metric ruler
microscope slide
rubber stopper
safety glasses

Materials

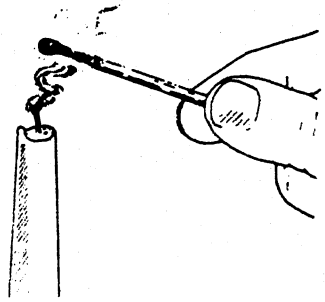
candle (2cm diameter)
matches, 6-8
toothpicks, 2
limewater solution
string, 4 cm long
aluminum foil
cobalt chloride paper

Procedure

1. Follow this procedure that will guide you to draw a conclusion about what is actually burning in a lighted candle. Record your observations at each step in the "Observations" section below.
2. Heat the bottom of a candle with a lighted match and secure it to a glass square on your lab table. Light the candle and allow it to burn for several minutes. Briefly describe the burning candle.
3. Place a piece of string on the glass square. Light it and observe what happens. Write your observations. Does it continue to burn like the wick in the wax does?
4. Light a toothpick with a match. Make and record observations about what happens. Does it burn like the wick does? Push a new toothpick into the top of the candle. Light it with another match and write your observations.
5. Use a microspatula to transfer a small amount of liquid from the bowl of the candle onto the glass square. Try to light it and note the result.
6. Make a slit in a small piece of aluminum foil and place it securely below the wick but above the wax, separating the wax from the wick. Light the wick and again make observations.



7. Remove the aluminum foil and light the candle again. Blow out the candle and *immediately* place a lighted match in the "smoke" about 2 cm above the wick. Note the result.



8. Invert a 250mL beaker over the lit candle. Note any liquid substance that collects on the inside of the beaker. Test this liquid with fresh (blue) cobalt chloride paper.
9. Invert a 125mL Erlenmeyer flask over the lit candle for several minutes. Remove the flask, turn it right-side-up, and add about 10mL of the clear limewater solution. Stopper and shake the flask. Note any color changes in the solution, beside the collection of soot from the candle.
10. Clean of the glass square and return all materials including the candle.

Observations

Place your observations for each step here.

2.

3.

4.

5.

6.

7.

8.

Conclusions and Questions

1. What does limewater test for?

2. What does cobalt chloride test for?

3. Give one example illustrating the difference between an observation and a conclusion.

4. What is the source of energy that is needed to light the candle?

5. List the steps in the scientific method.

6. What important part of a scientific experiment was missing from this lab exercise?

7. What is actually burning in a lighted candle?

8. Prove your conclusion

Name_____ Date_____

2 Quantitative Observations during the Combustion of a Candle

Introduction

Most of the experiments that you will be conducting require measurements. Measurements are quantitative observations. The measurements that you will be taking and recording are the data for this experiment.
It is important to make and record all initial data before you begin. The quantitative values that you have recorded often change during and after a reaction. We will be analyzing these changes in data to determine what changes have occurred and why. The data you record provides a record for your experiment.
You can only be as accurate as the equipment you are using to record your data. Your ability and skills in using the equipment will help to determine the accuracy of your results.

Introductory Questions

1. What is a quantitative observation?

2. Why should you make accurate readings?

3. What kind of measurements could change during a chemical reaction?

Problem

What conclusions can be determined from analyzing the quantitative measurements of a burning candle?

Equipment

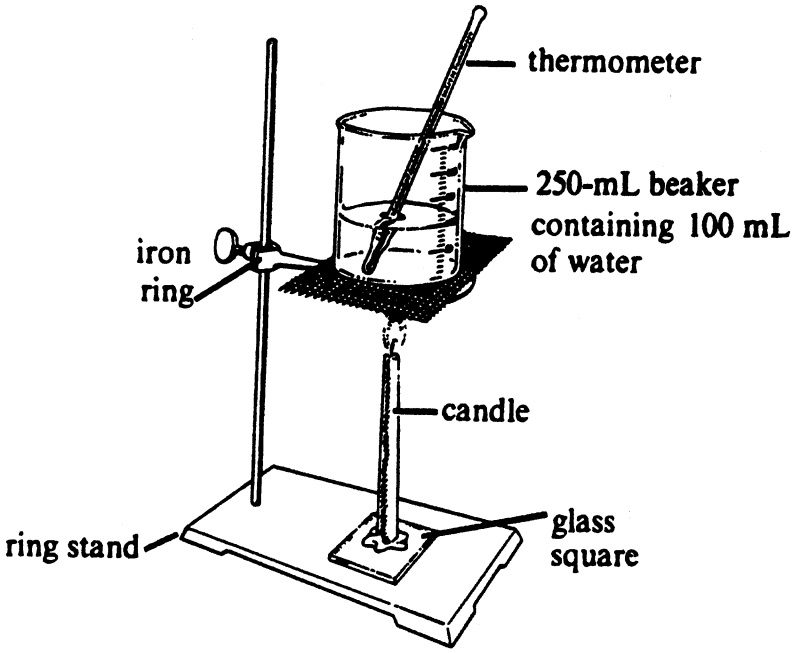
balance
ring stand
iron ring
wire gauze
thermometer
250-mL beaker
100-mL graduated cylinder
second hand clock
safety glasses
glass square

Materials

Candle (2-cm diameter)
Matches

Procedure

1. Place the candle on the glass square by lighting it with a match, dripping some wax onto the glass and placing the candle on the wax drippings. Blow out the candle.
2. Find the mass of the candle with the glass square and record it in the data table.
3. Place the candle under the iron ring and adjust the height of the ring 4 cm above the candle.



4. Accurately measure 100 mL of water in a graduated cylinder.
5. Pour the water into a 250-mL beaker.
6. Measure and record the temperature of the water. The thermometer bulb should be completely immersed in the water, but not touch the glass beaker.
7. Place the beaker on the wire gauze as shown in the diagram below.

8. Light the candle. Heat the water for exactly 10 minutes. Blow out the candle.
9. Measure the temperature of the water and record the temperature in the data table.
10. Find the mass of the candle with the glass square and record it in the data table.
11. Light the candle and repeat steps 8 – 10. This is your second trial.

Observations and data

	Trial 1	Trial 2
Mass of candle and glass square		
Mass of candle and glass square after heating		
Time candle burned and water heated		
Temperature of water at start		
Temperature of water after heating		

Calculations

Show work: Must show equations when needed, substitution with units, answer with units and correct number of significant figures.

1. Change in mass of candle
   Trial 1

   Trial 2

2. Rate of candle's change in mass per minute
   Trial 1

   Trial 2

3. Change in temperature of water
   Trial 1

   Trial 2

4. Rate of temperature change of water per minute
   Trial 1

   Trial 2

	Trial 1	Trial 2
Change in mass of candle		
Rate of candle's change in mass per minute		
Change in temperature of water		
Rate of temperature change of water per minute		

Conclusion Questions

1. What 3 quantitative measurements were made during this investigation?

2. Compare your calculations with 2 other teams. Were your results identical? Account for any differences.

3. What is the difference between rate and time?

4. What is the difference between heat and temperature?

5. How does the heat from the combustion of the candle relate to the temperature change of the water?

Name_____ Date_____

3 Measuring Mass

Introduction

You will need to learn how to measure mass using three different methods.

1. *Measuring mass directly.* This is the simplest type of measurement. The object whose mass is to be determined, is placed directly on the balance pan. If you are using a balance beam, move the appropriate riders into position until the pointer is balanced at the zero point. The mass of the object is read from the position of the riders.
2. *"Measuring out" a substance.* When measuring out solid crystals such as sodium chloride (salt), a piece of weighing paper is placed on the scale first. Move the riders until the pointer is balanced at zero. Set the riders to the desired mass (this must be equal to the mass of the paper and the mass to be measured out). Add the substance to be measured to the weighing paper until the pointer is balanced at the zero point.
3. *Determining mass by difference.* This technique involves subtraction. If you are finding the mass of a substance that is in a container, find the combined mass, then subtract the mass of the container. This technique is also used to determine changes in mass that occur during a reaction.

Introductory questions

1. When you begin an experiment, how could you determine which method you should be using?

2. Why is it wrong to measure some substances directly onto the balance pan?

Problem

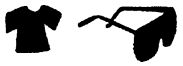
What are the three methods for measuring mass? What should be considered when choosing a method?

Equipment

lab balance
watch glass
beaker, 150-mL
beaker, 50-mL
microspatula
graduated cylinder, 100-mL
timer
safety glasses

Materials

sodium chloride (NaCl)
calcium chloride ($CaCl_2$)
weighing paper
balloon
coins, 1 penny & 1 nickel

Procedure 

Part I Measuring mass directly

1. Make sure your balance is properly "zeroed out". Ask your teacher for assistance if you are having difficulty adjusting the balance to zero.
2. Place a penny on the balance pan. Move the riders until the pointer is balanced at zero. Record the mass of the penny in the data table.
3. Repeat step 2 for all of the objects listed below. Record the mass of each object in the data table.
    a. A nickel
    b. A watch glass
    c. A 150-mL beaker
    d. A 100-mL graduated cylinder

Part II Measuring out a substance

4. Place a piece of weighing paper on the balance pan. Move the riders until the pointer is zeroed. Record this reading.
5. Your goal is to measure out exactly 6.50 g of sodium chloride (NaCl) onto the piece of weighing paper. Add 6.50 to the mass of the weighing paper and move the riders to this position on the balance. Record this setting.
6. Using a microspatula, add sodium chloride to the weighing paper until the pointer is zeroed. Discard the sodium chloride and weighing paper in the trash can.
7. Place a dry 50-mL beaker on the balance pan. Move the riders until it is zeroed. Record this reading.
8. Move the riders until they read exactly 23 g more than the reading from step 7. Record this reading.
9. In a 100-mL graduated cylinder, measure exactly 35 mL of cold tap water. Slowly and carefully pour the water from the graduated cylinder into the beaker on the balance until the pointer is zeroed. Note and record the volume of water remaining in the graduated cylinder. Discard the water and dry the glassware.

Part III Determining mass by difference

10. Measure and record the mass of a watch glass.
11. Using a microspatula, place 20-25 crystals of calcium chloride ($CaCl_2$) on a watch glass.
12. Measure and record the combined mass of the watch glass and calcium chloride.

13. Record the time this measurement was taken.
14. Using the microspatula spread the crystals out on the watch glass. Look at the crystals carefully and record your observations. Set the watch glass aside to be examined again later. At least 10 minutes.
15. Measure and record the mass of a piece of weighing paper.
16. Using the microspatula, place a heaping scoop of sodium chloride (NaCl) on the paper. Measure and record the combined mass of the weighing paper and NaCl.
17. Add a second scoop of NaCl to the paper. Measure and record the combined mass of the paper and 2 scoops of NaCl.
18. Add a third scoop of NaCl to the paper. Measure and record the combined mass of the paper and salt. Discard the paper and salt.
19. Measure and record the mass of a 150-mL beaker.
20. In a 100-mL graduated cylinder, obtain exactly 30 mL of cold tap water. Pour the water into the beaker. Measure and record the combined mass of the beaker and water.
21. Place a balloon on the balance. Find and record the mass of the balloon.
22. Inflate a balloon and tie off the open end so that no air can escape. Measure and record the mass of the inflated balloon. Puncture and discard the balloon.
23. Observe the calcium chloride crystals that you put aside on the watch glass. Record your observations.
24. Measure and record the combined mass of the watch glass and contents.
25. Record the time this measurement was taken. Discard the calcium chloride. Wash and dry the watch glass.

Observations and data

Part I

Step 2	Mass of a penny	g
Step 3	Mass of a nickel	g
	Mass of a watch glass	g
	Mass of a 150-mL beaker	g
	Mass of a 100-mL graduated cylinder	g

Part II

Step 4	Mass of weighing paper	g
Step 5	New setting (weighing paper & 6.50g)	g
Step 7	Mass of 50-mL beaker	g
Step 8	New setting (50-mL beaker & 23.0g)	g
Step 9	Volume of water remaining in graduated cylinder	mL

Part III

Step 10	Mass of watch glass	g
Step 12	Mass of watch glass and $CaCl_2$	g
Step 13	Time measurement was taken	
Step 15	Mass of weighing paper	g
Step 16	Mass of weighing paper & 1 scoop NaCl	g
Step 17	Mass of weighing paper & 2 scoops NaCl	g
Step 18	Mass of weighing paper & 3 scoops NaCl	g
Step 19	Mass of 150-mL beaker	g
Step 20	Mass of 150-mL beaker & 30 mL of water	g
	Mass of 30 mL of water	g
Step 21	Mass of deflated balloon	g
Step 22	Mass of inflated balloon	g
Step 24	Mass of watch glass & $CaCl_2$ (after sitting)	g
Step 25	Time measurement was taken	

Observations

Calcium chloride crystals

Initial observations from step 14

Final observations from step 23

Calculations

1. Calculate the volume of water added to the beaker in step 9.
   (Initial volume – volume remaining in graduated cylinder)

_____mL

2. Calculate the mass of 1mL of water.
   (Mass of water (step 8) ÷ volume of water)

_____g

3. Using the mass-by-difference technique, calculate the mass of:

a. $CaCl_2$ crystals added to the watch glass
   (Step 12 – step 10)

_____g

b. NaCl (1 scoop)
   (Step 16 – step 15)

_____g

c. NaCl (2 scoops)
   (Step 17 – step 15)

_____g

d. NaCl (3 scoops)
   (Step 18 – step 15)

_____g

e. 30 mL of water
   (Step 20 – step 19)

_____g

f. Moisture absorbed by $CaCl_2$ crystals
   (Step 24 – step 12)

_____g

4. Calculate the difference in mass between the inflated balloon and the deflated balloon.
(Step 22 – step 21)

_____g

5. Calculate how much time, in minutes, elapsed between the two measurements of the $CaCl_2$.
(Step 25 – step 13)

_____min

Conclusion Questions

1. Calculate the masses of the individual scoops of NaCl.

Scoop 1 _____g

Scoop 2 _____g

Scoop 3 _____g

2. Compare the masses of the individual scoops. Are the three values identical? Why not?

3. What measurements would you make to determine the average mass of a heaping microspatula of NaCl? How would you arrive at the average mass?

4. The difference between the mass of the balloon when inflated and its mass deflated is not an accurate measurement. Why? Should the measurement have been more or less?

5. Suppose you were asked to measure out 5 grams of $CaCl_2$. Describe how you would make this measurement.

6. Suppose you wanted to know the mass of a quantity of orange juice that was poured into a drinking glass. Describe how you would determine this mass.

7. A beaker contains a quantity of a liquid. You want to know the combined mass of the beaker and the liquid. Describe how you would make this measurement.

8. Suppose you were asked to compare the mass of a nickel and a sample of NaCl crystals.

a. Which method would you use to determine the mass of the nickel?

b. Which method would you use to determine the mass of the NaCl crystals?

c. Which of the two measurements is likely to be the more accurate measurement? Explain.

Name_____ Date_____

4 Density Determination

Introduction

Chemistry is the study of matter, which is defined as anything that has mass and takes up volume. During this experiment you will measure volumes directly and indirectly. The relationship between mass and volume will be used to determine density.

Volumes of liquids can be measure directly in a graduated cylinder. The volume of a regularly shaped solid can be determined by multiplying height, length and width. Many solids are irregularly shaped and can not be measured in this way. The volume of these objects can be measured indirectly by water displacement. The amount of water displaced is equal to the volume of the solid.

Density is an important property of matter. It can be used to help identify an unknown substance. Density is defined as the quantity of matter in a given unit of volume.

Density = mass ÷ volume or D = m ÷ v

Introductory questions

1. Describe the method you used in Earth Science when you found the volume of a rock or mineral.

2. What factors can change the density of a substance?

Problem

How can you determine the density of unknown solids and liquids?

Equipment

Laboratory balance	beaker, 50-mL
Centimeter ruler	scissors
Graduated cylinder. Various sizes	
thread	safety glasses

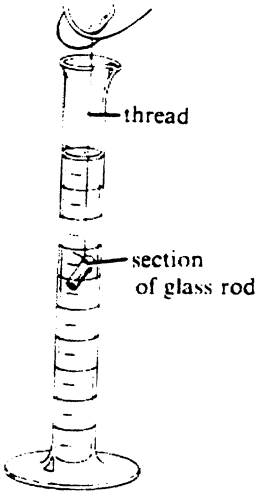
Materials

Distilled water
Regularly shaped metal object
Metal pellets
20 pre-1982 pennies
unknown liquids
1 irregularly shaped solid (rocks or metals)
section of glass rods (about 5cm long)
20 post-1982 pennies

Procedure

Part I Solids [Use the smallest graduated cylinder that your sample will fit into, without forcing it]

1. Measure and record the diameter and length of the section of glass rod. (±0.2mL)
2. Measure and record the mass of the section of glass rod.
3. Fill a graduated cylinder with tap water to the 7.0-mL mark. (±0.2mL)
4. Tie a piece of thread around the section of glass rod as shown.



Lower the glass into the water in the cylinder until it is completely under water. Do not allow it to touch the sides of the cylinder. Read and record the new water level.

5. Repeat steps 2-4 with the irregularly shape solid.
6. Obtain a regularly shaped solid. Measure and record its dimensions and mass.
7. Add exactly 30-mL of water to the graduated cylinder.
8. Immerse the metal object in the water and record the new water level.
9. Remove the object and refill the cylinder to the 30-mL mark. Set it aside until step 12.
10. Record the mass of a 50-mL beaker.
11. Add 50 metal pellets to the beaker. Measure and record the combined mass of the beaker and pellets.
12. Carefully pour the pellets into the water in the graduated cylinder (avoid splashing since this volume is recorded) Record the new water level. Remove the pellets. Spread them out on a paper towel to dry.

Part II Liquids at room temperature

13. Measure and record the mass of a clean, dry graduated cylinder.
14. Add exactly 10.00 mL of distilled water to the cylinder. Measure and record the combined mass of the cylinder and water. Dispose of the water. Clean and dry the cylinder.
15. Repeat step 14 for the two unknown liquids. DO NOT DISPOSE OF THE LIQUIDS. RETURN THE LIQUID TO ITS PROPER BOTTLE.

Part III Density of pennies

16. Determine the mass of 20 pennies minted prior to 1982 Record you measurements.
17. Pour 40 mL of water into the graduated cylinder.
Carefully, pour the pennies into the water. Record the new water level.
18. Repeat steps 16 and 17 with 20 pennies minted after 1982. Record your measurements.

Observations and data

Part I

1. Dimensions
    a. Glass rod: diameter_____cm length_____cm

    b. Regularly shaped object: shape_____

    Length _____cm, Width _____cm, Height _____cm

2. Mass and water displacement

Sample	Mass	Original water level	Final water level
Glass rod	g	mL	mL
irregularly shaped object	g	mL	mL
regular metal object	g	mL	mL
beaker	g	XXXXXXX	XXXXXXX
beaker & pellets	g	XXXXXXX	XXXXXXX
metal pellets	g	mL	mL

Part II

1. Mass of 10-mL graduated cylinder _____g
2.

Sample	Mass (cylinder & 10-mL sample)
distilled water	g
unknown liquid A	g
unknown liquid B	g

Part III

	Mass of 20 pennies (g)	Volume of 20 pennies (mL)
Pre-1982 pennies		
Post-1982 pennies		

Calculations

1. a. From the measured diameter and length of the glass rod, calculate its volume.

_____cm^3

b. From the measured dimensions of the regularly shaped object, calculate its volume.

_____cm^3

2. Determine the volumes of the following solid objects by calculating how much water they displace (final water level – original water level).

	final water level (mL)	original water level (mL)	volume of object (mL)
glass rod			
irregularly shaped object			
regularly shaped metal object			
metal pellets			

3. Calculate the mass of the metal pellets.

_____g

4. Calculate the densities of the section of glass rod and the regularly shaped metal object, using their measured masses and the volumes calculated from their dimensions.

Sample	Mass $_{(g)}$	Volume (cm^3)	Density (g/cm^3)
glass rod			
metal object			

5. Using their measured masses and the volumes calculated from their water displacements; calculate the densities of the solid samples.

Sample	Mass$_{(g)}$	Volume (mL)	Density (g/mL)
glass rod			
irregularly shaped object			
regularly shaped object			
metal pellets			

6. Using mass-by-difference, calculate the masses of the liquids.

Samples	Mass of cylinder & sample	Mass of cylinder	Mass of sample
distilled water			
unknown liquid A			
unknown liquid B			

7. Using the measured masses and the measured volumes, calculate the densities of the liquids.

Sample	Mass $_{(g)}$	Volume (mL)	Density (g/mL)
distilled water			
unknown liquid A			
unknown liquid B			

8. Using the measured masses and measured volumes of the pennies, calculate their densities.

Sample	Mass $_{(g)}$	Volume (mL)	Density (g/mL)
pre-1982 pennies			
post-1982 pennies			

Conclusions and Questions

1. Compare the value of the volume of the glass rod calculated from its dimensions to that attained by measuring its water displacement.

2. Make the same comparison for the values attained by these two methods for the volume of the regularly shaped object.

3. Of these two methods which is more accurate? Explain.

4. From your observations and data, are the metal object and the metal pellets composed of the same metal? Give evidence to support your answer.

5. The density of glass ranges from 2.4 to 2.8 g/cm^3. That of distilled water at 20° C is approximately 1.0 g/mL. Compare your results with these values. What errors might account for differences between your results and these accepted values?

6. Why is it necessary to indicate temperature when giving density values of liquids?

7. The table below lists the densities for several different substances. (Liquids are at 20°C). Using this table and your data, try to identify the following:

Regularly shaped metal _____
Unknown liquid A_____
Unknown liquid B_____

Sample	Density (g/cm^3)
lead	11.34
copper	8.93
aluminum	2.70
Sample	Density (g/mL)
ethanol	0.79
vegetable oil	0.92
ethylene glycol	1.11
glycerine	1.25

8. How did the densities of the two groups of pennies compare? How could you account for the difference?

9. Describe two methods for determining density.

10. Why is density an important property?

Name_____ Date_____

5 Physical and Chemical Change

Introduction

Chemistry is the study of matter and the changes that occur. A chemist often needs to describe matter very thoroughly. Physical properties that can be observed with our senses include size, shape, color, and phase. Quantitative physical properties include density, specific heat capacity, and boiling point. Changes that may occur that involve physical properties are tearing, dissolving melting, boiling (all phase change processes) expanding, etc. No new chemical bonds are formed, so the chemical composition remains the same.

Chemical properties of a substance include all of the chemical changes possible for that substance. The change will result in a new substance being formed. Chemical composition and properties will be different than the original substance. The atoms have been rearranged. Examples of chemical changes are burning, rusting, decomposing, etc.

Introductory questions

1. What observations can you make to determine if a change is physical or chemical?

2. Why are phase changes considered physical changes?

Problem

What distinguishes a physical change from a chemical change?

Equipment

Laboratory balance
Lab burner
5 test tubes (13 x 100)
Test tube rack
Test tube holder
Watch glass
microspatula
dropper pipette
mortar and pestle
glass square
magnet (in a baggie or saran-wrapped)
safety glasses

Materials

Copper sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)
Sodium chloride (NaCl)
Hydrochloric acid (6 M HCl)
Silver nitrate (0.1 M $AgNO_3$)
Sulfur $_{(s)}$, powdered
Iron filings (Fe)
Magnesium ribbon (Mg) 5-cm
paper (5 cm x 10 cm) long)
2 birthday candles
matches

Procedure

Fire **Caustic Substance**

Note and record all observations in your data table.

1. Place the birthday candle in the small clean test tube.
2. Gently heat the wax over the burner flame until the wax has melted.
3. Allow the wax to cool. Record your observations.
4. Light the candle and allow some of the wax to drip onto the glass square.
5. Blow out the candle and secure the candle to the melted wax on the glass square.
6. Light the candle and allow it to burn completely. Proceed with the experiment while the candle burns.
7. Tear the piece of paper into small pieces and place them on the watch glass. Record your observations.
8. Light the paper with a match and allow the paper to burn. Record your observations.
9. Add a microspatula of NaCl (sodium chloride) to a small quantity of water (about 5 mL) in a clean test tube.
10. Place your thumb over the mouth of the test tube and shake well. Record your observations.
11. Using the dropper pipette, add 10 drops of silver nitrate ($AgNO_3$) to the NaCl-water mixture. Record your observations.

12. Tear the piece of magnesium ribbon into 5 pieces. Place the pieces in a clean test tube and carefully add a few drops of 6 M HCL (hydrochloric acid). ***CAUTION: Use extreme care in handling this acid. It will cause severe burns if allowed to come in contact with the skin.*** Feel the bottom of the test tube with your finger tip. Record your observations.
13. Using a mortar and pestle, grind several crystals of $CuSO_4 \cdot 5H_2O$ into a powder. Record your observations.
14. Place the powder in a clean test tube with the microspatula.
15. Heat gently over a burner flame until you observe a color change. Allow the test tube to cool. Record your observations.
16. Add a few drops of water to the $CuSO_4 \cdot 5H_2O$. Record your observations.
17. Using a lab balance, measure out the following samples: 0.50g of iron filings and 0.50g of powdered sulfur. Test each sample with a magnet. Leave the magnet in the plastic wrap. Record your observations.
18. Place the iron filings in a test tube. Heat the test tube gently until the iron filings glow red. Test the sample with a magnet
19. Mix the two samples, iron and sulfur) thoroughly in a clean test tube. Run the magnet along the side of the test tube. Record your observations.
20. Heat the iron-sulfur mixture in a burner flame until the mixture "glows". Allow the sample to cool.
21. Examine the mixture by probing it with the microspatula and testing it with the magnet. Record your observations.

Observations and data

Step	Sample	Observations
3	candle after heating and candle after cooling	
6	Candle after burning	
7	tearing paper	
8	burnt paper	
10	NaCl in water	
11	NaCl in water and $AgNO_3$	
12	Mg and HCl	
13	grinding $CuSO_4 \cdot 5H_2O$	
15	$CuSO_4 \cdot 5H_2O$ after heating	
16	$CuSO_4 \cdot 5H_2O$ and water	
17	sulfur and iron	
18	Sulfur-iron mixture	
20	sulfur-iron after heating	

Conclusions and Questions

1. Classify the following properties of sodium metal as physical or chemical:

   a. silver metallic color _____

   b. turns gray in air _____

   c. melts at 98°C _____

   d. reacts explosively with chlorine _____

2. Classify the following changes as physical or chemical:

   a. steam condenses to liquid water on a cool surface _____

   b. baking soda dissolves in vinegar, producing bubbles _____

   c. mothballs gradually disappear at room temperature _____

   d. baking soda loses mass as it is heated _____

3. What are two indications that a chemical change has occurred?

4. The following changes are not always chemical changes. Explain how they might be physical changes.
   a. change of color

   b. apparent loss of mass

   c. apparent disappearance of a substance

5. How can substances in a mixture be separated? Give an example from the experiment.

6. How can substances in a compound be separated? Give an example from the experiment.

7. Indicate whether the following changes are physical or chemical. Support your conclusions

	Change occurred physical or chemical	Support your conclusion
melting candle wax		
burning a candle		
tearing paper		
burning paper		
dissolving NaCl		
mixing NaCl and $AgNO_3$		
tearing Mg ribbon		
adding HCl to Mg		
grinding $CuSO_4 \cdot 5H_2O$		
heating $CuSO_4 \cdot 5H_2O$		
mixing Fe and S		
heating a mixture of Fe and S		

Name_____ Date_____

6 Mixture Separations

Introduction

The ability to separate and recover pure substances from mixtures is extremely important in scientific research and in industry. Chemists need to work with pure substances, but naturally occurring materials are seldom pure. Often pure substances can be separated from other materials based on differences in physical properties, such as boiling point, freezing point or density. In this laboratory investigation, you will develop, design, and carry out *your own procedure* for separating a mixture. The mixture that you will be given by your teacher consists of poppy seeds, iron filings, sand, and salt.

Introductory Questions

1. You are given a mixture of three different liquids. If each liquid has a different boiling point, explain how you might separate the liquids from each other? _____

_____

_____

_____

_____

2. Identify each of the following as a pure substance or a mixture.

$NaCl_{(aq)}$ _____ Water _____

An alloy _____ $NaCl_{(s)}$ _____

Problem

How can you separate a mixture of poppy seeds, iron filings, sand and salt into individual components?

Equipment

Assorted pieces of lab equipment that you choose

Materials

Sodium Chloride (NaCl)
Sand
A Mixture
Poppy seeds
Iron filings

Procedure

Today, *you* are designing the lab procedure. But, you cannot begin your experiment until you have a *plan of action*. To do that you must first explore the individual characteristics of each substance. As you explore the different characteristics of each sample, record the information in the table provided. At the end of the lab, you will give the teacher four different substances that you have separated form each other.

Characteristics of samples

Properties	Sand	Iron Filings	Salt	Poppy Seeds
Does it dissolve in water?				
Is it magnetic?				
Does it roll?				
Does it float?				

List the steps you will take for separating the mixture before you start to separate the mixture. Have the teacher approve of your method, then follow your procedure.

Conclusions and Questions

1. Were you able to separate the mixture into four separate substances? Explain.

2. If you could do the lab over, what would you do differently?

3. How could you separate each of the following two-part mixtures?

b. sand and gravel

c. salt and sugar (HONORS ONLY)

d. alcohol and water

e. nitrogen and oxygen (HONORS ONLY)

4. Is a solution a heterogeneous or homogeneous mixture?

5. Considering a solution of salt water, which is the solute? Which is the solvent?

6. Why can't you use filter paper to separate salt from water in a salt water solution?

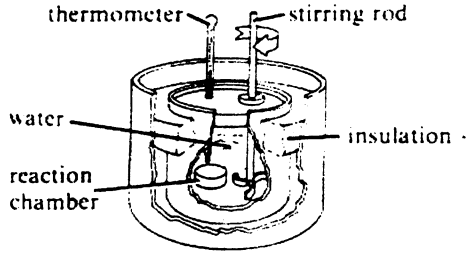
Name_____ Date_____

7 Calorimetry: Heat of Fusion of Ice

Introduction

Heat is a form of energy. Temperature is a measure of the average kinetic energy of particles in a substance.

When there is a physical or chemical change, heat is either released (exothermic) or absorbed (endothermic). It is important to be able to measure this heat. These measurements can be made in a calorimeter.



A simple calorimeter is an insulated container made of two chambers. The outer chamber contains a known mass of water. The material that is to gain or lose heat is placed in the inner chamber. When two substances are in contact, heat will flow from the object with more heat to the object with less heat. The heat exchange will continue until both objects have the same temperature. The heat lost by one object is gained by the second object. In a calorimeter the heat exchange is occurring between the object that is undergoing a physical or chemical change and the water. The change in the water's temperature will be equal to the temperature change of the object. From this information the heat lost or gained can be determined.

The calorimeter we will use will only have one chamber. The ice will be placed directly into the water. The water will supply the energy needed to melt the ice. The temperature change in the water will allow you to determine the ***heat of fusion*** of ice. Heat of fusion is defined as the quantity of heat needed to change one gram of a solid to a liquid at a constant temperature.

The following relationships will be used :

a. Heat lost or gained =
(mass of water) x (change in temperature of water) x (specific heat of water)

b. Heat lost by water = heat absorbed by ice

c. $\frac{\text{Heat needed to melt the ice}}{\text{mass of the melted ice}}$ = heat of fusion of ice

The ***specific heat capacity*** **(Cp)** of a substance is the quantity of heat energy needed to raise the temperature of 1 gram of a substance by 1°C. The specific heat capacity of water is 4.2J/g•K.

Introductory questions

1. Why will it be necessary to know the mass of the water before you begin your experiment?

2. Why is an insulated cup used?

3. Why would putting an insulated lid on the cup improve your results?

4. What is the density of water?

Problem

How can you determine the heat of fusion of ice?

Equipment

Graduated cylinder, 100-mL
Beaker, 250-mL
Styrofoam cup
Thermometer
Spoon or spatula
ring stand
lab burner
wire gauze
safety glasses

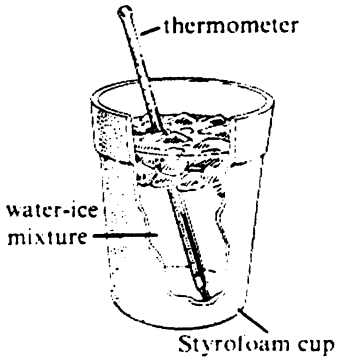
Materials

Water
Ice cubes

Procedure

1. In a 250-mL beaker, heat about 125 mL of water to a temperature of 50°C.
2. Measure exactly 100 mL of the heated water and pour it into the Styrofoam cup. Record the volume of water, V_1.
3. Measure the temperature of the water accurately. Record the temperature, T_1.
4. Immediately add 2-3 ice cubes.

5. Stir the ice-water mixture carefully with the thermometer. **CAUTION:** ***Thermometers break easily.***



6. The cup must contain ice at all times. If the last of the ice is about to melt add another ice cube. Watch the temperature as you stir the mixture.
7. Continue stirring and adding ice as needed until the temperature is no longer changing. Record the final temperature, T_2.
8. Carefully remove the unmelted ice with the spoon and discard the ice in the sink. Try to avoid removing any of the water from the cup.
9. Measure and record the volume of water in the cup, V_2.

Observations and data

V_1 =	mL	T_1 =	°C
V_2 =	mL	T_2 =	°C

Calculations

1. Using the known density of water, find the mass (m_1) of the original volume of water.

2. Find the volume of water due to the melted ice. ($V = V_2 - V_1$)

3. Find the mass (m_2) of this volume of water.

4. Find the change in temperature of the water. ($\Delta T = T_1 - T_2$)

5. Find the heat lost by the original mass of water. ($\Delta Q = m_1 \times \Delta T \times C_p$)

6. Find the heat of fusion of ice. ($\Delta Q \div m_2$ = heat of fusion of ice)

7. Find your percentage of error. (The true value for the heat of fusion of ice is 333.6J/g•K)

Conclusions and Questions

1. List three possible sources of error in this experiment.

2. One source of error is the flow of heat between the water in the cup and the surroundings. Explain how this error is reduced by starting with 50°C water.

3. Define the following terms:

   a. exothermic

   b. endothermic

   c. heat of fusion

   d. specific heat

   e. joule

4. Describe how the Law of Conservation of Energy applies to this experiment.

5. Is the process of melting exothermic or endothermic? Give evidence to support your answer.

6. A solid substance with a mass of 400 g is at its melting point temperature in a calorimeter. While the substance changes from a solid to a liquid at the same temperature, the 400 g mass of water in the calorimeter goes from an initial temperature of 80°C to a final temperature of 30°C.

   a. How much heat did the water lose while the substance cooled?

   b. What is the heat of fusion of the substance that melted?

Name_____ Date_____

8 Calorimetry: Heat of Crystallization of Wax

Introduction

The process of crystallization (solidification) is the exact opposite of the process of fusion (melting). When a substance changes form liquid to solid, heat is released. This is an exothermic change.

When crystallization takes place in a calorimeter, the water will absorb energy. The same relationships that you applied in Lab 7 will be used to calculate the heat of crystallization in this lab.

Introductory questions

1. When does crystallization begin?

2. What is the mass of 1 mL of water?

3. Will the temperature of the water in the calorimeter increase or decrease? Explain.

4. What is the specific heat of water?

Problem

How can you determine the heat of crystallization of wax?

Equipment

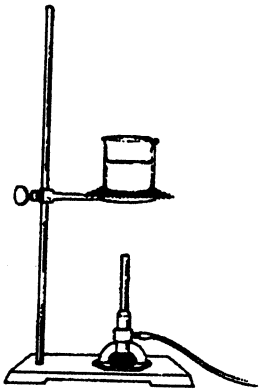
Graduated cylinder, 100-mL
Beaker, 250-mL
Styrofoam cup
Thermometer
Test tube holder
ring stand
lab burner
wire gauze
safety glasses

Materials

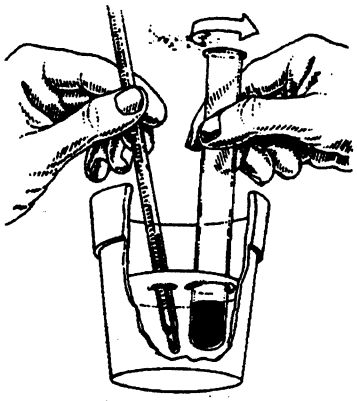
Test tube containing 10.0g of wax

Procedure

1. In a 250-mL beaker heat about 200 mL of water to boiling.



2. Place the test tube containing the 10.0 g sample of wax, into the beaker of boiling water.
3. Gently rotate the tube (be careful of steam) until all of the wax is melted.
4. Measure 100 ml of cold tap water and pour it into the Styrofoam cup.
5. Using a test tube holder remove the test tube containing the wax from the beaker of boiling water. Turn your burner off.
6. Hold the test tube up to the window or light and look for the first sign of cloudiness. This indicates that crystallization has started.
7. Immediately take the temperature of the water in the Styrofoam cup and then place the test tube in the cup with the cold water. Record the temperature. (T_1)
8. Rotate the test tube while the wax is solidifying. Constantly watch the temperature of the water.



9. When the temperature of the water has stopped increasing, Record this temperature. (T_2)

Observations and data

Mass of water in calorimeter (m_1)	g
Mass of wax sample (m_2)	10.0g
T_1	°C
T_2	°C

Calculations

1. Find the change in temperature of the water. ($\Delta T = T_2 - T_1$)

2. Find the heat gained by the water. ($\Delta Q = m_1 \times \Delta T \times C_p$)

3. Find the heat of crystallization of wax. ($\Delta Q \div m_2$)

Conclusions and Questions

1. Define the term of heat of crystallization. What unit is used to express this property?

2. Explain why this experiment could not be conducted using a mixture.

3. How does the heat of fusion compare to the heat of crystallization of the same substance?

Name_____ Date_____

9 Heating and Cooling Curves

Introduction

In this experiment you will take direct measurements on the substance that is undergoing a change. Measurements will be taken during the cooling and heating of this substance. During the cooling part of this experiment, heat will be removed, at a constant rate, from a liquid that is above its melting point, until that liquid has changed to a solid. The temperatures will be taken at specific time intervals to be analyzed later.

The procedure will be reversed for the heating portion of this experiment. Heat will be added at a constant rate. You will start with the same substance in its solid state and heat it until it has changed to a liquid. Again, you will monitor the temperature at specific time intervals.

At the conclusion of the experiment, you will construct a graph to show the two sets of data and how they compare.

Introductory questions

1. As the substance is melting or solidifying, what changes do you expect to see in the temperature of the substance?

2. What do you think the relationship between the freezing point and melting point is?

Problem

How can the freezing and melting points of a substance be determined?

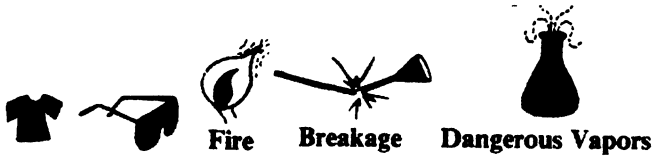
Equipment

Wire gauze	test tube clamp
Beaker, 150-mL	lab burner
Thermometer	safety glasses
Ring stand	timer with second hand
Iron ring	blue pencil
Test tube rack	red pencil

Materials

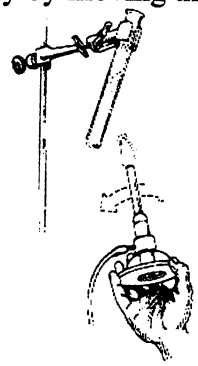
Test tube 18 x 150-mm containing a sample of 1,4-dichlorobenzene, stoppered
Water

Procedure

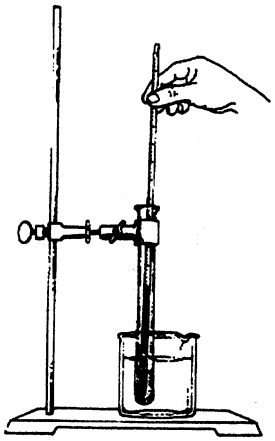


Part I Cooling curve

1. Fill a 250-mL beaker ¾ full with cold tap water. Put it aside until step 8.
2. Clamp the test tube with the 1,4-dichlorobenzene to the ring stand and **REMOVE THE STOPPER.**
3. Heat the sample gently by moving the burner back and forth as shown.



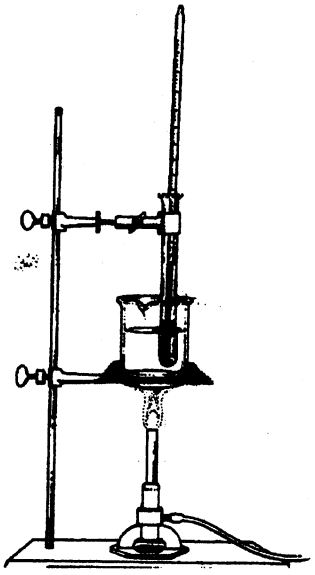
4. As soon as the sample has started to melt, remove the heat and carefully place the thermometer in the sample.
5. Use the thermometer to stir as you resume gently heating the sample.
6. Continue heating until the sample is completely melted. **DO NOT OVERHEAT.**
7. The final temperature should be about 70°C. If it is higher than this, wait for the temperature to drop to this temperature before you go on.
8. When the temperature is 70°C, let the time be 0 minutes. Place the test tube in the beaker of cold water. Make sure the sample is covered by water. **DO NOT LET ANY WATER GET INTO THE TEST TUBE.** See the diagram.



9. One partner should watch and call the times every 30 seconds, as the other partner reads and records the temperatures. The sample must be stirred constantly until it has solidified.
10. Continue this procedure until the sample is at 30°C. Remove the test tube from the beaker of water and place it in the test tube rack.

Part II Heating curve

11. Prepare a water bath as shown in the diagram, **BUT LEAVE OUT THE TEST TUBE.**



12. Heat the water to about80°C. Remove the heat.
13. Let the time equal 0 minutes and attached the test tube to the ring stand as shown in the diagram. Record temperatures as you did before, every 30 seconds.
14. As soon as the thermometer is free to move, gently use it to stir the sample.
15. Continue to stir and record temperatures until the sample has reached 60°C.
16. Remove the thermometer and clean it in the warm water in the beaker.
17. After the test tube has cooled, replace the stopper and return it to where you got it from.

Observations and data

Cooling Curve Data Table

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)
0.0		7.5	
0.5		8.0	
1.0		8.5	
1.5		9.0	
2.0		9.5	
2.5		10.0	
3.0		10.5	
3.5		11.0	
4.0		11.5	
4.5		12.0	
5.0		12.5	
5.5		13.0	
6.0		13.5	
6.5		14.0	
7.0		14.5	

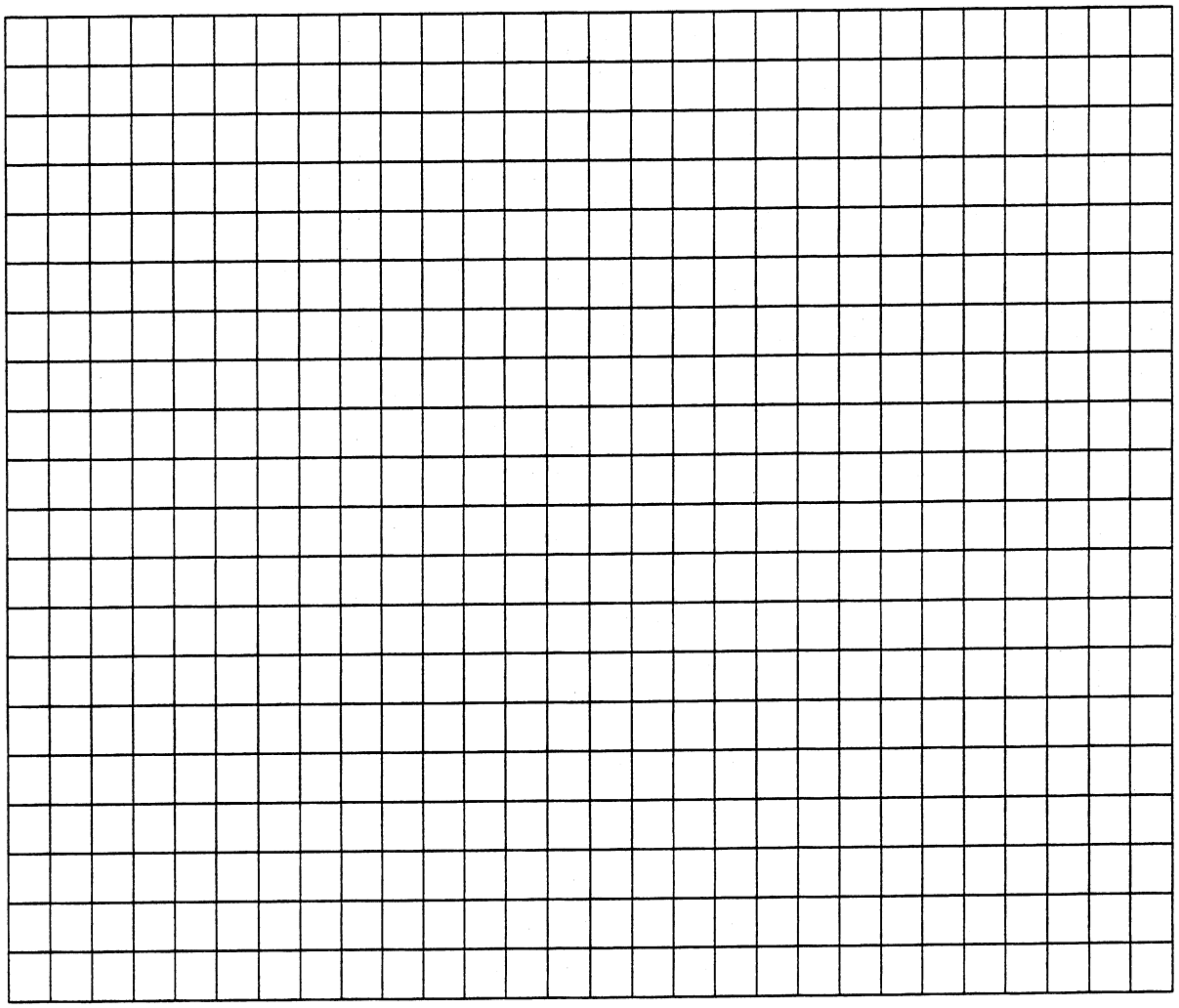
Heating Curve Data Table

Time (min)	Temperature (°C)	Time (min)	Temperature (°C)
0.0		7.5	
0.5		8.0	
1.0		8.5	
1.5		9.0	
2.0		9.5	
2.5		10.0	
3.0		10.5	
3.5		11.0	
4.0		11.5	
4.5		12.0	
5.0		12.5	
5.5		13.0	
6.0		13.5	
6.5		14.0	
7.0		14.5	

Calculations

Plot your data from this experiment on the set of axes below. Label the Y-axis Temperature (°C) and the x-axis Time (min). Use an appropriate scale on each axis. Use a blue pencil to plot cooling data and a red pencil to plot heating.

Heating and Cooling Curves



Conclusions and Questions

1. Referring to your graph, determine the freezing point of 1,4-dichlorobenzene. How does this temperature compare with the melting point temperature of the same substance as indicated on the graph?

2. What phase changes are exothermic? Endothermic?

3. In which phase of a substance do its particles have the greatest kinetic energy?

4. Explain the diagonal parts of the cooling curve in terms of changes in kinetic and potential energy. Do the same for the horizontal portions of the cooling curve.

5. How is this different for the heating curve?

Name_____ Date_____

10 Sub-Atomic Particles

Introduction
In this lab we will be calculating the number and placement of three subatomic particles.

Problem
You will be able to state the numbers of subatomic particles, their charge and mass using the periodic Table.

Introductory Questions

1. What is the charge and mass of the proton, neutron and electron?

2. How do you determine mass?

3. Where are the particles located?

Equipment
None

Materials
Mini Skittles
2 experimental atoms

Procedure
Part I Experimental Atoms
1. Open up experimental atom #1.
2. Count up the different subatomic particles and fill in the information in the Data Table.
3. Use the Reference Tables and the key below to identify the element and complete the Data Table.
4. Repeat steps 1-3 for the other experimental atoms.

Observations

KEY

Green - Proton
Blue - Neutron
Red - Electron

Atom #1

Element Symbol	
Atomic #	
Atomic Mass	
Number of Protons	
Number of Neutrons	
Number of Electrons	
Electron Configuration	
Nuclear Charge	
Atomic Charge	

Atom #2

Element Symbol	
Atomic #	
Atomic Mass	
Number of Protons	
Number of Neutrons	
Number of Electrons	
Electron Configuration	
Nuclear Charge	
Atomic Charge	

Atom # 3

Element Symbol	
Atomic #	
Atomic Mass	
Number of Protons	
Number of Neutrons	
Number of Electrons	
Electron Configuration	
Nuclear Charge	
Atomic Charge	

Atom # 4

Element Symbol	
Atomic #	
Atomic Mass	
Number of Protons	
Number of Neutrons	
Number of Electrons	
Electron Configuration	
Nuclear Charge	
Atomic Charge	

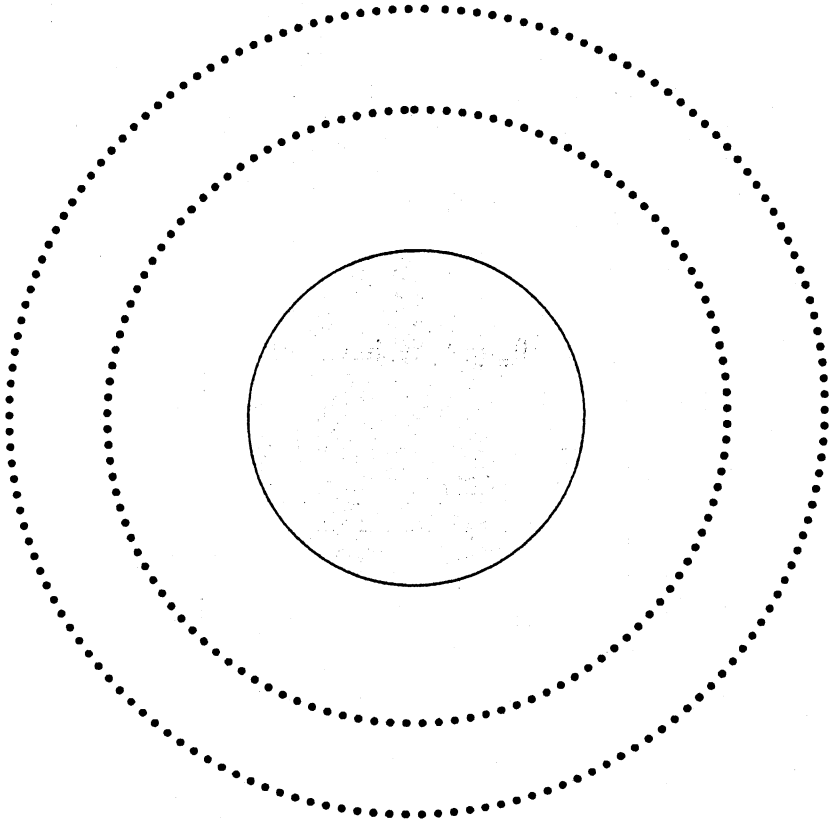
Part II

5. You will be given a "mini" bag of Skittles to build your own atom.
6. Get a "symbol" from your teacher.
7. Build the atom using the orbital model below.
8. When you have successfully built your atom, call your teacher over to check your work.
9. Now you can eat your subatomic particles.

Atom_____

Skittles Key
Red – Protons
Purple – Neutrons
Green – Electrons

Atom:_____



Conclusions and Questions

1. What is the nuclear charge of a sodium ion?

2. Why are all nuclei positive?

3. Why is the charge of an atom neutral?

4. What does an electron configuration show?

11 Periodic Trends in Atomic Radii

Introduction
You will be determining trends between radius and atomic number.

Introductory Questions
1. What is a cation?

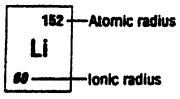
2. What happens is an anion?

Problem
What is the trend between Atomic number and ionic radius?

Equipment
None

Materials
None

Procedure
1. Use the information below to plot ionic radius vs. atomic number.



H^{37}							He^{50}
${}_{60}Li^{152}$	${}_{31}Be^{112}$	${}_{20}B^{80}$	${}_{5}C^{77}$	${}_{171}N^{71}$	${}_{140}O^{66}$	${}_{136}F^{64}$	Ne^{70}
${}_{95}Na^{186}$	${}_{65}Mg^{160}$	${}_{50}Al^{143}$	${}_{41}Si^{118}$	${}_{212}P^{109}$	${}_{184}S^{103}$	${}_{181}Cl^{99}$	Ar^{94}
${}_{133}K^{227}$	${}_{99}Ca^{197}$	${}_{62}Ga^{122}$	${}_{53}Ge^{123}$	${}_{222}As^{121}$	${}_{198}Se^{117}$	${}_{195}Br^{114}$	Kr^{111}
${}_{148}Rb^{244}$	${}_{113}Sr^{215}$	${}_{81}In^{167}$	${}_{71}Sn^{141}$	${}_{62}Sb^{141}$	${}_{221}Te^{138}$	${}_{216}I^{138}$	Xe^{130}
${}_{169}Cs^{262}$	${}_{135}Ba^{222}$	${}_{95}Ti^{170}$	${}_{84}Pb^{175}$	${}_{74}Bi^{151}$	Po^{164}	At^{145}	Rn^{140}

Observations and data

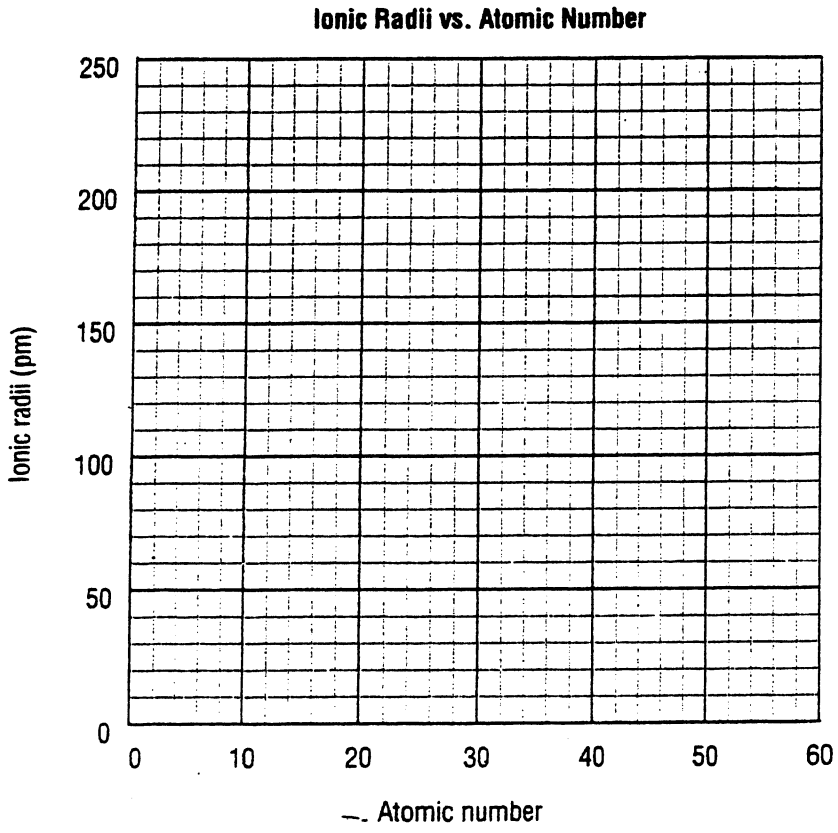


Figure 1

Conclusions and Questions

1. Comment on the sizes of cations compared with the sizes of anions. How do they compare with the size of their atoms?

2. Are the trends shown for periods 2, 3, 4, and 5 similar or different?

3. Describe and explain the shape of each period's portion of the graph.

4. How do the radii for anions and cations change as you go down a group?

Name_____ Date_____

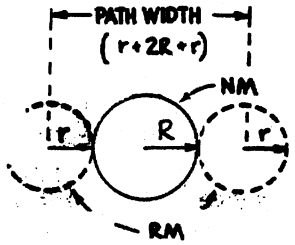
12 Nuclear Marbles

Introduction

Sometimes, people have to resort to something besides theirs senses to determine the shape or size of things, especially if they are for things that are so small we cannot see them. One way to do this is to fire particles at the object you are investigating, and study the path of these particles. This can be done with a particle accelerator. Ernest Rutherford inferred the size of the nuclei of gold atoms by studying how alpha particles were deflected by the nuclei in gold foil. You will try something similar, but with marbles.

You will not measure the diameters of the marbled directly. You will roll a "nuclear" marble toward other target marbles and analyze the data you will collect. The ratio of hits and misses will be used to determine the size of the marble. The fewer hits you have the smaller the object probably is.

When you roll a marble toward a nuclear marble (NM), you have a certain probability of a hit between the rolling marble (RM) and the nuclear marble. One expression of the probability (P) of a hit is the ratio of the path width required for a hit to the width (l) of the target area. The path width is equal to two RM radii plus the diameter of the NM, as shown in the diagram.



The probability that a RM will hit a lone NM in the target area is

$$P = \frac{\text{path width}}{\text{target width}}$$

$$= \frac{2R + 2r}{L} = \frac{2(R + r)}{L}$$

where R = radius of NM
r = radius of RM
R + r = the distance between the centers of an RM and an NM that are touching
L = width of the target area

If the number of nuclear marbles is increases to N, the probability of a hit is increased by a factor of N. Thus, the probability that the RM will hit one of the NM is

$$P = \frac{2N(R + r)}{L}$$

The probability of a hit can also be determined experimentally by the ratio of the number of hits to the number of trials.

$$P = \frac{H}{T}$$

where H = number of hits
T = number of trials

You now have two expressions for the probability of a hit. Assume that both of the expressions are equal to each other. If the radii of the RM and NM are equal, then R + r = d, where d is the diameter of any of the marbles. Combine the two equations for P,

Introductory questions

What is the combined formula for the diameter of a marble?

Problem

How can indirect measurements be used to measure objects that are to small to be seen?

Equipment

3 metersticks
10 marbles

Procedure

1. Set up the 3 metersticks as shown in the diagram. The width between the two sides is 60 cm.

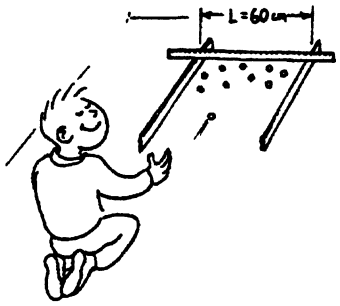


Fig. A

2. Place 9 of the marbles in an area 60 cm wide (L = 60cm), Make sure they are within 20 cm of the back stick and RANDOMLY SCATTERED. None should be up against the meter sticks at any time.
3. One team member will randomly roll the RM toward the 9 NM. The key to this is not to aim or look. Your roll must be completely random.
4. The second person will record the number of hits and misses for 200 trials. You may switch places during the procedure.
5. It is important that you follow the following directions:

a. if more than one marble is hit, it counts as 1 hit.
b. if the RM hits the meter stick first, it is a do over.
c. at no time should a marble be left touching the meter sticks.
d. at all times the marbles should be moved so that they remain RANDOMLY scattered throughout the target area

6. Record the hits and misses in the chart.

Observations

Trial #	Hit	Miss	Trial #	Hit	Miss	Trial #	Hit	Miss	Trial #	Hit	Miss	Trial #	Hit	Miss
1			41			81			121			161		
2			42			82			122			162		
3			43			83			123			163		
4			44			84			124			164		
5			45			85			125			165		
6			46			86			126			166		
7			47			87			127			167		
8			48			88			128			168		
9			49			89			129			169		
10			50			90			130			170		
11			51			91			131			171		
12			52			92			132			172		
13			53			93			133			173		
14			54			94			134			174		
15			55			95			135			175		
16			56			96			136			176		
17			57			97			137			177		
18			58			98			138			178		
19			59			99			139			179		
20			60			100			140			180		
21			61			101			141			181		
22			62			102			142			182		
23			63			103			143			183		
24			64			104			144			184		
25			65			105			145			185		
26			66			106			146			186		
27			67			107			147			187		
28			68			108			148			188		
29			69			109			149			189		
30			70			110			150			190		
31			71			111			151			191		
32			72			112			152			192		
33			73			113			153			193		
34			74			114			154			194		
35			75			115			155			195		
36			76			116			156			196		
37			77			117			157			197		
38			78			118			158			198		
39			79			119			159			199		
40			80			120			160			200		

Calculations

1. Determine the number of hits. H =
2. Determine the number of trials T = Compute the diameter of the marbles using the formula that you derived from the introduction.

Computed diameter d =

4. Measure the marble diameter with a cm ruler.
measured diameter d =

Conclusions and Questions

1. Compare your results for the computed diameter and the measured diameter.

2. What is your percent error?

3. Describe how this experiment is similar to Rutherford's gold foil experiment.

Name_____ Date_____

13 Electrons in Atoms

Introduction

The modern atomic theory of the atom cannot tell you exactly where the electrons in atoms are placed. However, it does define regions in space called orbitals, where there is a 95% probability of finding an electron.

Introductory questions

1. Why is there only a probability of finding an electron in a region of space?

Problem

How does the probability of distribution of electrons in their orbitals relate to the distribution of impacts from your data?

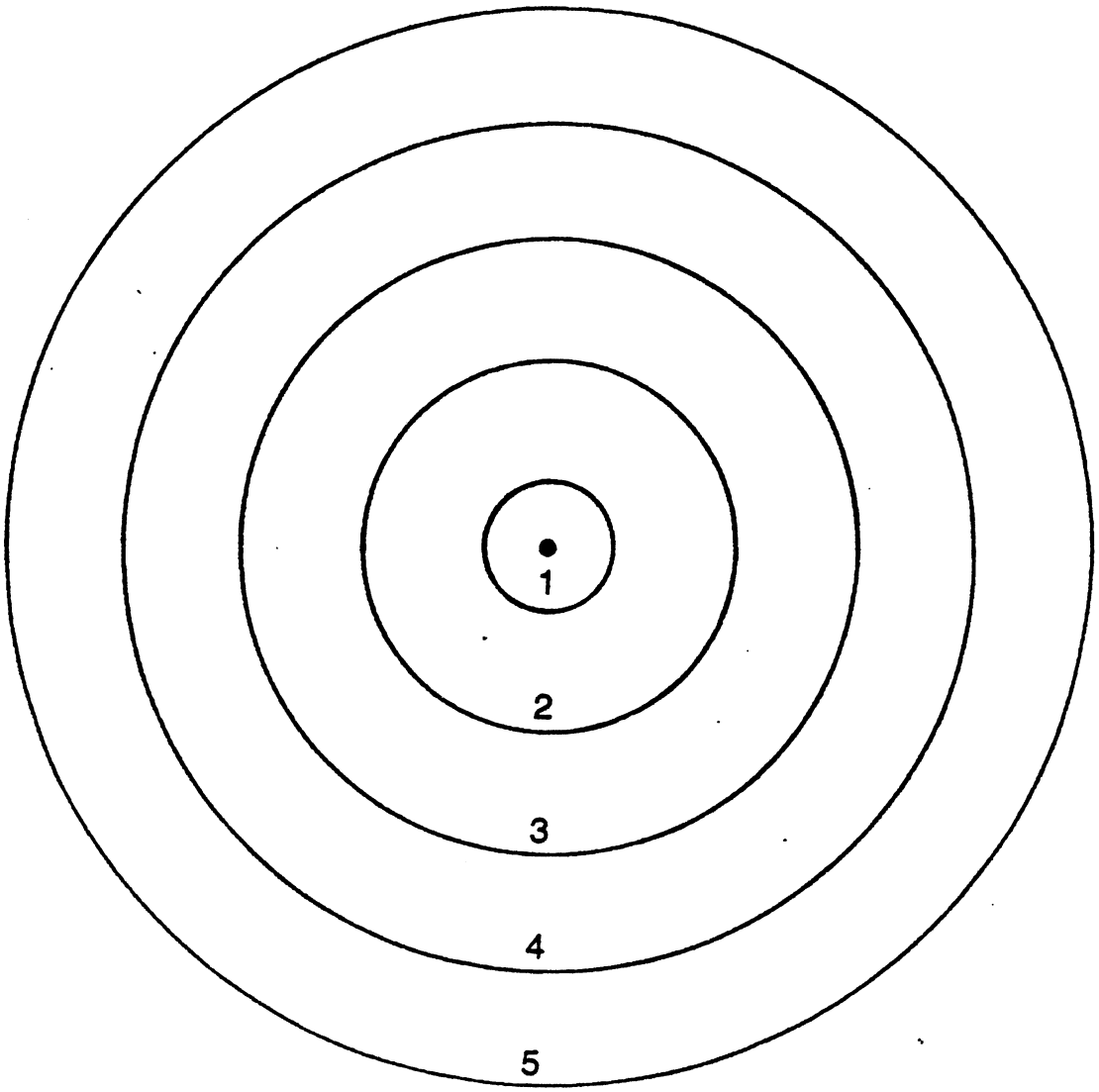
Equipment

Sheet of blank paper
Foam board
tape
dart
safety glasses

Procedure

1. Place the blank sheet of paper over the target sheet.
2. Mark the center "Bull's Eye" on the blank paper.
3. With the 2 centers lined up, tape the 2 sheets to the foam board.
4. Place the foam board on the floor and stand over the target.
5. Hold the dart and place your elbow at your waist.
6. **Without any downward thrusting motion**, aim for the center, and let the dart drop.
7. You may work as a team with your partner. While one person drops the dart, the other can pass the dart back up.
8. Drop the dart 100 times.
9. Trade places and repeat the lab with your partner.

Observations

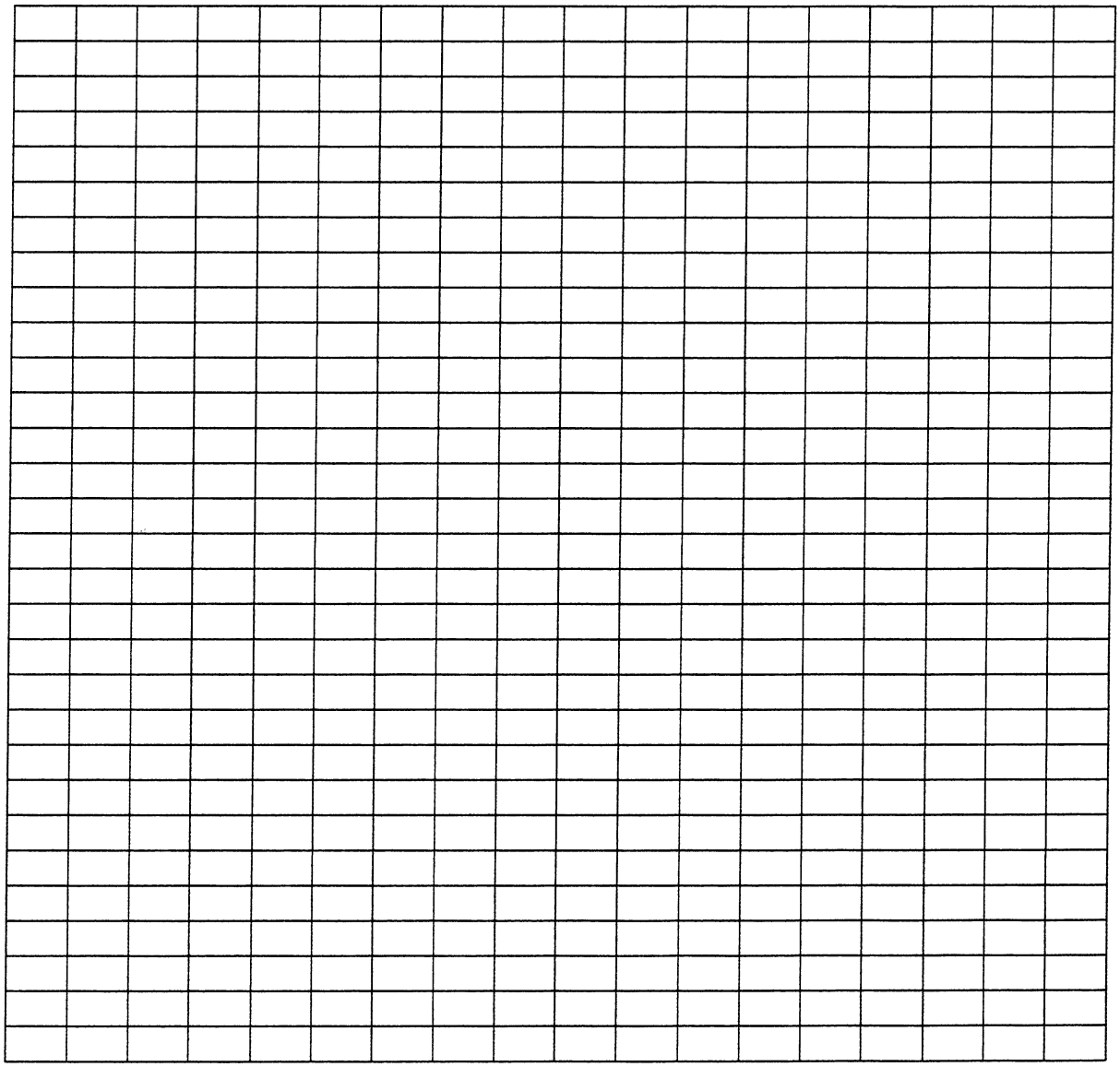


Calculations

1. Count the number of hits in each area.

Region	1	2	3	4	5
# hits					

2. On the graph below, label the y-axis # of hits and the x-axis region. Use an appropriate scale on each axis.



Conclusions and Questions

1. Before you dropped the dart, could you predict where it would hit?

2. If you could not predict the exact spot where the dart would hit, could you predict the area within it would hit?

3. What does the shape of the graph indicate?

4. One dart made all the holes on the target. Could one electron have occupied all of these different locations? Explain.

5. How did the following scientists describe the structure of the atom:
   a. Dalton

   b. Thompson

   c. Rutherford

   d. Bohr

   e. Modern theory

6. How did this lab show that the probability of distribution of electrons in their orbitals relates to the distribution of impacts from your data?

Name_____ Date_____

14 Electron Structure

Introduction

All atoms consist of protons, neutrons, and electrons. Protons and neutrons are clumped together to form a nucleus in the center of the atom. The electrons are very mobile and can be located anywhere from the center of the atom to its outer edge. The electrons cannot be found in the nucleus.

Two of these atomic particles have electrical charge. The proton has a positive charge and the electron has a negative charge. An atom has equal numbers of positively charged protons and electrons. This makes the charge on the atom zero – the atom is said to be neutral. While the protons are only found in the nucleus, the electrons move around so much that they may be thought of as an "electron cloud" that surrounds the nucleus. Different electrons in the electron cloud have different amounts of energy. A model of electron structure would look like a cloud within clouds. Each cloud represents the location of electrons that have a certain energy. The energy levels are numbered from one to seven, with level 1 being closest to the nucleus and having the lowest energy and level seven being farthest from the nucleus and the highest energy level.

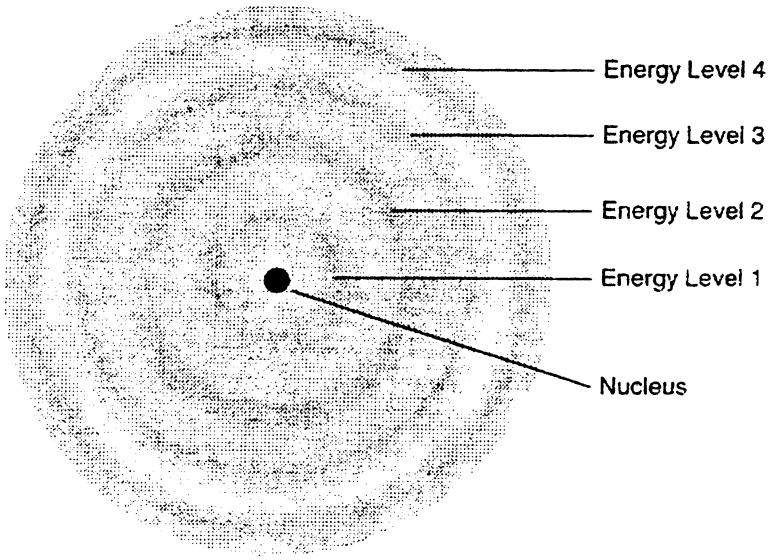


Figure 1. Electron Clouds (Energy Levels 5–7 omitted)

Each energy level can contain more than one electron. However, the maximum number of electrons in each level is not the same, but increases as the energy level number increases. Energy level 1 can contain a maximum of 2 electrons; energy level 2, 8 electrons; energy level 3, 18 electrons; and energy level 4, 32 electrons. These are the maximum number of electrons in each energy level. For levels 5, 6, and 7, the energy

levels and orbital size increases, along with energy levels. The electrons do pair up as they fill the energy levels, but they do not orbit a ring. They move all over the area of the specific "electron cloud" or energy level.

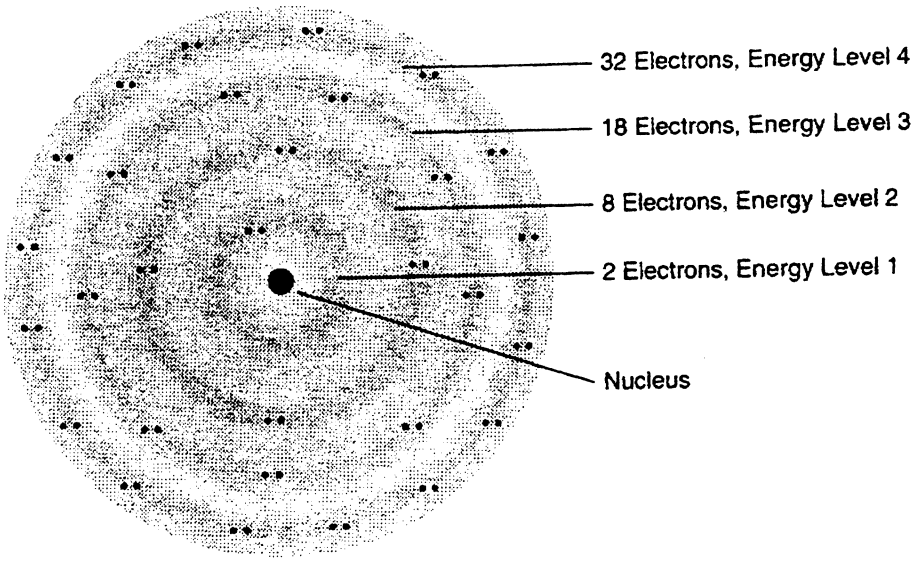


Figure 2. Energy Level Electrons

The elements may be arranged for convenience in a chart called the periodic table. The elements are listed in order of increasing atomic number, which is either the number of protons or the number of electrons in the element. Atomic numbers increase from left to right and from top to bottom. As the atomic numbers of the elements increase, the electrons fill in the energy levels, starting from the lowest energy level to the next higher energy level, and so forth, until the number of electrons in the atom equals the number of protons in the nucleus.

For example, the atomic number of fluorine is nine- a neutral fluorine atom has nine protons and nine electrons. Two electrons are paired and filled energy level one. The remaining seven electrons fill energy level two. Six of the electrons are paired in level two, with one lone unpaired electron.

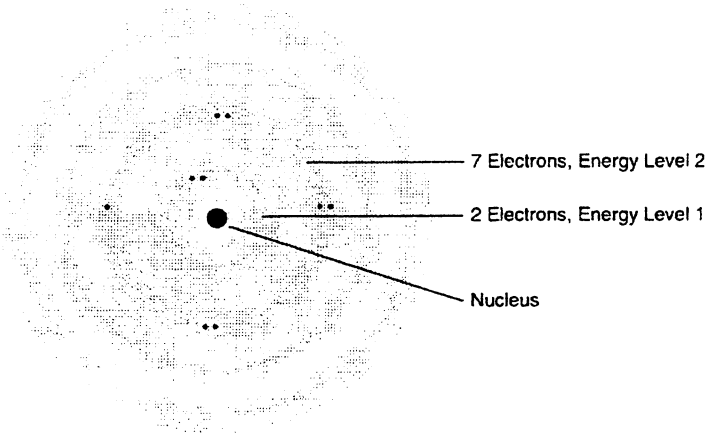


Figure 4. Electron Energy Level Diagram for Fluorine

In chemical reactions, atoms tend to gain or lose electrons. Atoms in which the number of protons is not equal to the number of electrons are called ions. The magnesium ion has 12 protons and 10 electrons.
Ions are written using the element symbol, with the number of extra or deficient electrons at the top and right of the symbol. If the ion has fewer electrons than protons, a + sign is put after the number; if the ion has more electrons than protons, a – sign is used. For every ion, the number of electrons changes. The number of protons remains constant. The number of protons in a neutral atom always equals the number of protons in the ion. If the ion has one extra or deficient electron, the number 1 is omitted from the ions symbol.

The magnesium ion with two less electrons than protons, is written

$$Mg^{2+}$$

If the atom has 9 protons and 10 electrons, then the ion is a fluorine ion and is written;

$$F^-$$

When the electrons are filled in the correct energy level for both Magnesium, Mg, and the magnesium ion, Mg^{2+}, the electron changes are:

	Level 3 — 2 electrons				Level 3 — 0 electrons
Mg	Level 2 — 8 electrons	– 2 electrons	=	Mg^{2+}	Level 2 — 8 electrons
	Level 1 — 2 electrons				Level 1 — 2 electrons

Magnesium loses two electrons from the third energy level, leaving the Mg^{2+} ion with only the first two energy levels filled.

If a fluorine atom gains an extra electron (nine protons and ten electrons), it forms a fluorine ion, F^-. When the electrons are filled in, the extra electron completes energy level two, and the fluorine ion has the first two energy levels completely filled.

F	Energy Level 2 — 7 electrons	+ 1 electron	=	F^-	Energy Level 2 — 8 electrons
	Energy Level 1 — 2 electrons				Energy Level 1 — 2 electrons

In these two examples, both Mg^{2+} and F^- have identical electron structures. However, the nuclei of these two ions contain different numbers of protons and thus ion are still very different elements. One of the reasons that atoms react with each other is for each atom to have completely filled energy levels.

Introductory questions

1. How are the energy levels filled with electrons for a phosphorus atom?

2. How are the energy levels filled with electrons for an oxygen ion, O^{2-}?

Problem

To understand that the basics of electron structure is a fundamental atomic concept.

Materials

Periodic Table
Energy Level Diagrams
Chips; 10 Yellow, 18 Blue, 15 Green

Procedure

1. For each atom or ion, determine the number of protons contained in the atom or the ion. Record this number in the Data Table.
2. Determine the number of electrons and record this number.
3. Use the following color chips to represent the electrons according to the Key below:

s electrons	Yellow
p electrons	Blue
d electrons	Green

4. Count out the appropriate number of chips of each color.
5. Fill the electron energy levels with the correct number of electrons, from lowest energy level to highest, using the correct colors in each level. Record the number of electrons in each level in the Data Table.
6. Repeat this procedure to complete the Data Table.

Observations and data

Atom/Ion	# Protons	# Electrons	# Electrons Level 1	# Electrons Level 2	# Electrons Level 3	# Electrons Level 4
He			**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**
Na+			**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**
K			**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**
B			**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**	**Y-** **B-** **G-**

Cl^-			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
V			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
N			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Kr			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Al^{3+}			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Be			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Br			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
N^{3-}			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Co			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
O			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-
Co^{2+}			Y- B- G-	Y- B- G-	Y- B- G-	Y- B- G-

Conclusions

1. Why do you place the 2 s electrons in energy level 4 before you place the d electrons in energy level 3?

Name_____ Date_____

15 Candium

Introduction
The mass of an element is determined by averaging the mass of naturally occurring isotopes in ratio to their abundance in nature.

Introductory question
1. How do you calculate the average mass of an element?

Problem
Using experimental data, how can the average mass of an "element" be determined?

Equipment
Lab balance

Materials
Sample of Candium

Procedure
1. Obtain a sample of candium.
2. Separate the three isotopes. (m&m's®, Skittles®, Reese's Pieces®)
3. Measure and record the mass of each isotope.
4. Count and record the number of each isotope in your sample.
5. Follow the instructions on p 82 to complete the chart.

Observations and Calculations

	m&m's®	Skittle's®	Reese's Pieces®	Totals
Total mass (grams)				
Number				
Average mass (grams)				
Percent abundance				
Relative abundance				
Relative mass (grams)				

1. Calculate the average mass of each isotope by dividing its total mass by the number of particles of that isotope.

2. Calculate the percent abundance of each isotope by dividing its number of particles by the total numbers of particles and multiplying by 100.

3. Calculate the relative abundance of each isotope by dividing its percent abundance from step 2 by 100.

4. Calculate the relative mass of each isotope by multiplying its relative abundance from step 3 by its average mass.

5. Calculate the average mass of all candium particles by adding the relative masses. This average mass is the atomic mass of candium.

Conclusions and Questions

1. Explain the difference between percent abundance and relative abundance.

2. What is the result when you total the individual percent abundances? The individual relative abundances?

3. The percent abundance of each kind of candy tells you how many of each kind of candy there are in every 100 particles. What does relative abundance tell you?

4. Compare the total values for rows 3 and 6 in the table. Explain why the totals differ and why the value in row 6 best represents the atomic mass?

5. Explain any differences between the atomic mass of your candium sample and that of your neighbor. Explain why the difference would be smaller if larger samples were used.

6. Explain the fact that the atomic masses on the Periodic Table are not whole numbers.

Name_____ Date_____

16 The "Black Box"

Introduction

To learn more about the nature of the atom, Rutherford aimed a beam of alpha particles at a sheet of gold foil surrounded by a fluorescent screen. They found that most of the particles passed straight through the foil. A few particles were significantly deflected. Rutherford concluded that most of the alpha particles pass through the gold foil because the atom is mostly empty space. The mass and positive charge of the nucleus is concentrated in a small region of the atom. Rutherford called this mass the nucleus. Alpha particles coming to close to the nucleus were deflected.
In this investigation you will try to discover the shape of an object, by following the paths of particles passing through the space in which the object is contained.

Introductory questions

1. What does the object represent?

2. What path do you expect most of the particles to follow?

Problem

How can you determine the shape of an object beneath a platform without looking?

Materials

"Black Box"
Paper
Ruler or straight edge
Marble

Procedure

1. Obtain a "black box' from your teacher. ***DO NOT LOOK BENEATH IT AT ANYTIME!***
2. Place it on the floor or table where you will be working.
3. Tape a piece of blank paper to the top platform.
4. From any side, roll a marble beneath the platform, with just enough force to ensure it will roll out the other side. Be careful not to lose the marble into the sinks or beneath heaters or other objects.
5. Using a ruler, or straight edge, trace the path of entrance and exit. If it is not a single straight line, find where the entrance path line and exit path line bisect. Circle this point.
6. Repeat this procedure many times from all sides.
7. When your circled points seemed to show the shape of an object, ask your teacher to check your work.
8. Repeat if necessary.

Observations

Sketch the paths of the marble that showed the shape of the object.

Conclusions and Questions

1. What path did most of the marbles take? Why?

2. Why did some of the marbles deviate from a straight path?

3. Describe how you could determine the shape of your object, even though you couldn't see it.

4. Describe how this experiment is similar to Rutherford's experiment:
   a. What did the object represent?
   b. What did the marble represent?

Name_____ Date_____

17 Flame Tests

Introduction

The normal electron configuration of atoms or ions of an element is known as the "ground state". In this most stable energy state, all electrons are in the lowest energy levels available. When heated, some electrons absorb energy allowing them to move to higher energy levels. When electrons are in this higher energy level it is referred to as the "excited state". This configuration is unstable, and the electrons return to their "ground state", releasing energy in the form of light.

Only metals, with their loosely held electrons, are excited in the flame of a lab burner. Flame tests are useful in determining the presence of metallic ions, by their characteristic color.

Introductory questions

1. What part of a compound exhibits a characteristic color during a flame test?

2. Why can't nonmetals be used in flame tests?

Problem

How can the identity of an known sample with metallic ions be identified?

Equipment

Wooden splints
Lab burner
Colored pencils
Safety glasses
watch glass
spectroscope
matches

Materials

Lithium chloride (LiCl)
Potassium chloride (KCl)
Strontium chloride ($SrCl_2$)
Copper II chloride ($CuCl_2$)
sodium chloride (NaCl)
calcium chloride ($CaCl_2$)
barium chloride ($BaCl_2$)
ethanol

Procedure

Part I Flame test

1. Using your microspatula, place a heaping spatula full of lithium chloride on the watch glass.
2. Add 3 mL of ethanol.
3. Using a lit wooden splint, light the ethanol so that it is burning. **THE FLAME IS LIGHT BLUE AND ALMOST COLORLESS. BE CAREDUL.**
4. Record the color of the flame.
5. Do the same for each powder, cleaning the spatula and watch glass each time. NEVER PUT A WET SPATULA IN THE SAMPLE JAR.

Part II Flame tests using spectroscopes

6. Repeat the procedure for all of the powders, this time looking at the flame through a spectroscope.
7. Sketch the line spectrum that you observe.
8. Obtain an unknown sample and perform these tests, recording your results.

Observations

Part I

Compound	LiCl	NaCl	KCl	$CaCl_2$	$SrCl_2$	$BaCl_2$	$CuCl_2$	unknown
Color								

Part II

Sketch of line spectrum	
LiCl	
V I B G Y O R	

Sketch of line spectrum	
$SrCl_2$	
V I B G Y O R	

Sketch of line spectrum	
NaCl	
V I B G Y O R	

Sketch of line spectrum	
$BaCl_2$	
V I B G Y O R	

Sketch of line spectrum	
KCl	
V I B G Y O R	

Sketch of line spectrum	
$CuCl_2$	
V I B G Y O R	

Sketch of line spectrum	
$CaCl_2$	
V I B G Y O R	

Sketch of line spectrum	
Unknown Number	
V I B G Y O R	

Conclusions and Questions

1. What inaccuracies may be involved in using flame tests for identification purposes?

2. Which pairs of ions produce similar colors in the flame tests?

3. Explain how the colors observed in the flame tests are produced.

4. Define these terms:
   a. quanta

   b. ground state

   c. excited state

5. What is a spectroscope? What is observed if the flame tests are viewed through a spectroscope?

6. What is the identity of your unknown? Give evidence to support your answer.

7. What produces the colors of fireworks?

Name_____ Date_____

17 Emission Spectra and Energy Levels

Introduction

Electrons can be excited by passing an electric current through a sample of an element in its gaseous phase. For this investigation you will use spectrum tubes. A spectrum tube contains a small sample of an element in its gaseous phase. When an electric current is passed trough it, the vapor will glow brightly.

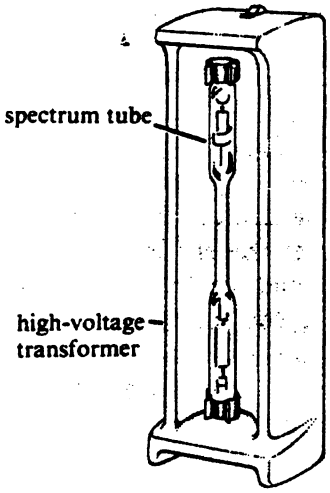


Figure 24-1

When the light from the spectrum tube is passed through a diffraction grating, an emission (bright-line) spectrum is produced. Each element has its own unique emission spectrum. The spectrum consists of a series of bright lines of definite wavelength. Each wavelength is mathematically related to the quantity of energy produced as electrons move from one energy level to another. This is proof that electrons exist at definite, distinctive energy levels in an atom.

Introductory questions

1. What is producing the glow from the spectrum tubes when electricity is passed through it?

2. How can a bright-line spectrum be used to identify elements?

Problem

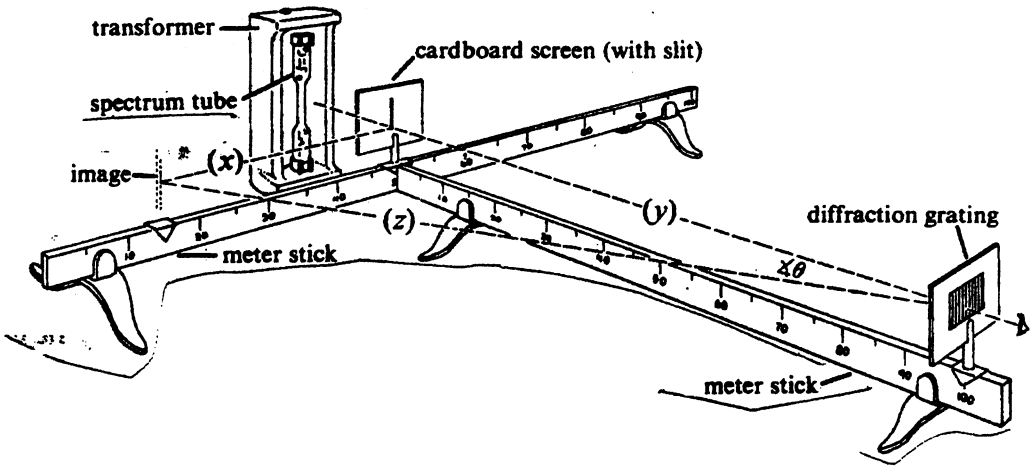
How can you determine the wavelengths of hydrogen and mercury associated with their respective spectral lines?

Equipment

Spectrum tubes, Hg, H_2
Transformer, high voltage
Diffraction grating
Cardboard screen with slit
Colored pencils
2 meter stick supports
2 meter sticks
2 screen supports
safety glasses

Procedure

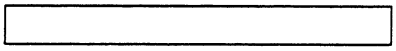
1. Set up the apparatus as shown below.



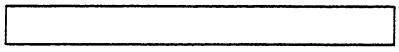
2. The cardboard screen should be placed on the 50 cm mark of one meter stick. The transformer (with the spectrum tube) is placed directly behind the screen so that the glow from the tube is visible through the slit. The diffraction grating should be placed on the second meter stick 100 cm from the tube.
3. One partner will view the emission spectrum by looking through the diffraction grating at the slit in the cardboard screen.
4. The other partner will stand behind the transformer and move a pencil slowly along the meter stick. The student viewing the spectrum should indicate to the other student when the pencil is at the 1st line of the spectrum closest to the transformer.
5. Measure the distance, in centimeters, between the tube and the spectral line. Record this distance.
6. Repeat this procedure for each of the lines of the same gas sample.
7. Turn off the transformer when not in use.
8. Repeat this procedure for the second sample of gas.

Observations

1. Using colored pencils, make qualitative sketches of the spectrum lines.



Hydrogen



mercury

2.

HYDROGEN

Color of line	x (cm)	y (cm)
red		100
blue-green		100
blue		100
violet		100

MERCURY

Color of line	x (cm)	y (cm)
green		100
violet		100

Calculations

Enter all values in the data table at the end of this section.

1. Find the distance (z) from the diffraction grating to the image of the spectral line:

$$z=\sqrt{x^2 + y^2}$$

2. Find the sine of angleθ: $\sin\theta = \frac{X}{Z}$

3 Find the wavelength (λ) in cm: $\lambda = d \times \sin\theta$, where $d = 1.9 \bullet 10^{-4}$cm

4. Find the wavelength in angstroms: $\lambda\text{Å} = \text{cm} \bullet 10^{8}$ Å/cm

color of line	z	sinθ	λ (cm)	λ (Å)
Hydrogen				
red				
blue-green				
blue				
violet				
Mercury				
green				
violet				

Conclusions and Questions

1. Compare your results to the known wavelengths in angstroms listed:
   a. Hydrogen
      i. red 6563 Å
      ii. blue-green 4861 Å
      iii. blue 4342 Å
      iv. violet 4359 Å
   b. Mercury
      i. green 5460 Å
      ii. violet 4359 Å

2. Relate spectral lines to energy levels in an atom and to the term quanta.

3. How might emission spectra be used in studying stars?

Name_____ Date_____

19 Half-Life and Radioactive Decay

Introduction

Many things grow at a fairly steady rate: population, money in the bank, and the thickness of paper that is continually folded over onto itself. Many other things decrease at a steady rate: the value of money in the bank, the amount of vacant area in a place where the population is growing and the amount of material that is undergoing radioactive decay. A useful way to describe the rate of decrease is in the terms of half-life, the time it takes for the quantity to be reduced to half its initial mass. For steady decrease, called exponential decrease, the half-life stays the same.
Radioactive materials are characterized by their rates of decay and are rated in terms of their half-lives. You will explore this idea in this activity. In this activity, you will be using a coin. The head of the coin represents the original substance. The tail represents the product after decay.

Introductory questions

1. What changes are occurring when a radioactive substance decays?

2. What factors can change the half-life of a substance?

3. What do the heads and tails of the coin represent?

Problem

How are radioactive decay and half-life related?

Materials

one coin
red and blue pencils

Procedure

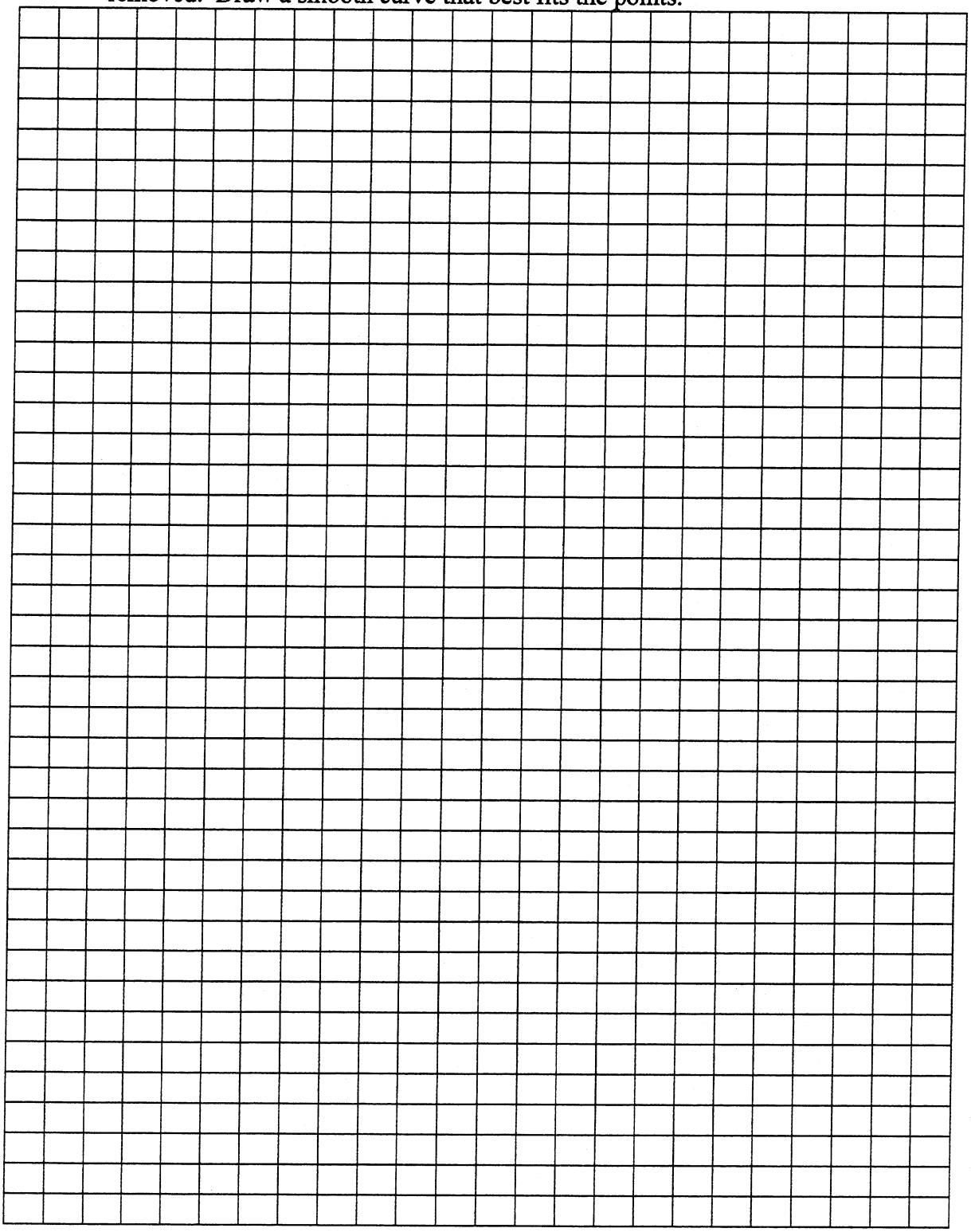
1. Flip your coin 100 times. Keep a tally of all the heads and tails.
2. Record these totals in the chart.
3. Record the number of heads as the number of flips for the next trial.
4. Flip your coin as many times as you just recorded on the chart. Tally the number of heads and tails.
5. Record the number of heads. Add the number of tails recorded in trial one and add it to the number of tails you tallied in trial 2. Record this amount for trial 2.
6. Repeat steps 3-5 until you have one head left. Stop flipping.

Observations

Trial number	Number of flips	Number of heads	Number of tails (cumulative)
0	0	100	0
1			
2			
3			
4			
5			
6			
7			

Calculations

1. Graph the number of heads remaining (vertical axis) vs. the trial number (horizontal axis) in a blue pencil. Draw a smooth curve that best fits the points.
2. On the same graph, using a red pencil, graph the number of cumulative tails removed. Draw a smooth curve that best fits the points.



3. Complete these half-life problems. Show all work
   a. How much of a 100.0 g sample of ^{198}Au is left after 8.10 days if its half-life is 2.70 days?

   b. A 50.0 g sample of ^{16}N decays to 12.5 g in 14.4 seconds. What is the half life?

   c. The half-life of ^{42}K is 12.4 hours. How much of a 750 g sample is left after 62.0 hours?

   d. What is the half-life of ^{99}Tc, if a 500 g sample decays to 62.5 g in 639,000 years?

   e. The half-life of ^{232}Th is $1.4 \cdot 10^{10}$ years. If there are 25.0 g of the sample left after $2.8 \cdot 10^{10}$years, how many grams were in the original sample?

   f. There are 5.0 g of ^{131}I left after 40.35 days. How many grams were in the original sample if its half-life is 8.07 days?

4. Show the complete balanced nuclear equations for the following nuclear reactions. Use Table N.

a. $^{42}_{19}K \rightarrow$

b. $^{239}_{94}Pu \rightarrow$

c. $^{235}_{92}U \rightarrow$

d. $^{1}_{1}H + ^{3}_{1}H \rightarrow$

e. $^{6}_{3}Li + ^{1}_{0}n \rightarrow ^{4}_{2}He +$

f. $^{27}_{13}Al + ^{4}_{2}He \rightarrow ^{30}_{15}P +$

g. $^{9}_{4}Be + ^{1}_{1}H \rightarrow ^{4}_{2}He +$

h. $^{37}_{19}K \rightarrow$

i. $_____ + ^{1}_{0}n \rightarrow ^{142}_{56}Ba + ^{91}_{36}Kr + 3^{1}_{0}n$

j. $^{238}_{92}U + ^{4}_{2}He \rightarrow ^{1}_{0}n +$

Conclusions and Questions

1. Approximately what percent of the pennies remaining were removed on each shake?

2. What is meant by half-life?

3. If each trial represents 100 years, what was the half-life of the pennies in years?

4. What particles can be emitted during decay?

5. What is the meaning of the graph you obtained?

6. If the heads were ^{14}C, what were the tails that were removed? Refer to Table N

Name_____ Date_____

21 Periodic Table Puzzle

Introduction

In 1869 Dmitri Mendeleev developed a periodic table of the known elements. Mendeleev used atomic mass to place the elements in a continuous row and noticed that some properties reoccurred every eight element. He then moved the elements into rows of eight so that similar properties fell into columns. Mendeleev was able to use his periodic table to predict the characteristic properties of undiscovered elements. Although he organized his table by atomic masses, not on atomic numbers as we do today, we give him credit for developing the periodic table.

In this investigation, your instructor has given you 23 of the 24 regular pieces contained in a puzzle. Each puzzle piece has 6 properties. Your goal is to identify these properties and place the pieces into an appropriate pattern or periodic table. Once you have completed this, you will be able to solve today's problem.

Problem

Identify the 6 properties of the missing puzzle piece.

Introductory Questions

1. What property of the elements do we use to organize the modern periodic table?

2. What property of the elements did Mendeleev use to build his periodic table?

3. What was Mendeleev able to do because he developed a properly organized table of elements?

4. What is the name for the rows of elements in the periodic table? _____

5. What is the name for the columns of elements? _____ or _____

Materials

23 of 24 puzzle pieces

Procedure

1. Count to make sure you have 23 puzzle pieces. Turn the pieces face-up and identify six properties of the pieces (such as color) and place them in Observation 1.
2. Place the pieces in a long row in order of one of the six properties you identified. Look carefully for patterns in the remaining properties. If you find that these properties occur along the row at regular intervals, organize the pieces into a chart so that the properties fall into rows and columns.
3. You may find that there are exceptions to the pattern; there are some properties that do not fit exactly into the pattern you perceive. This is normal. It mimics that real periodic table.
4. When you have organized your table, you should be missing one piece. Use the patterns of the properties to figure out the properties of the missing piece. Give that information to your instructor on a piece of paper. Your instructor will verify if you are correct or not.
5. Place the correct properties of the missing puzzle piece in the conclusion.

Observations

Puzzle Number _____

Properties of the puzzle pieces:

1.

2.

3.

4.

5.

6.

Conclusions and Questions

1. List the characteristics of your missing piece.

2. Why do you think is it helpful to organize the elements into a periodic table?

Name_____ Date_____

21 It's in the Cards

Introduction

Dmitri Mendeleev's Discovery of the Periodic Law ranks as one of the greatest achievements in the history of science.

It has survived the test of time and stands to this day as the single most important tool to understand the chemistry of the elements. As we try to understand the essence of this discovery, it is worthwhile to go back in time and look at how it was achieved.
In the years 1868-1870, Dmitri Mendeleev, a professor of chemistry at the University of St. Petersburg in Russia, was writing a new textbook called "Principles of Chemistry". More than 60 individual elements were known, along with a great many facts about their properties and compounds. Mendeleev knew the atomic masses of the elements, their densities, boiling points, and melting points, as well as the formulas of their compounds with hydrogen, oxygen, and chlorine. What was missing was a way to organize these facts, a way to understand how individual facts related to each other – in short, a way to classify the elements. The following quote from Mendeleev reveals his thoughts at the time:

"I wish to establish some sort of system of elements in which their distribution is not guided by chance... but by some sort of definite and exact principle."

Mendeleev decided to arrange the elements according to their atomic mass. He wrote out the exact atomic masses (as they were known at the time) in the margin of a list of the elements, then wrote out separate cards for each of the elements, with their atomic mass and other chemical and physical properties. Using these cards, Mendeleev played "chemical solitaire" for several hours, finally copying to a sheet of paper the arrangement he had worked out with the cards. With slight modification, this became Mendeleev's first *Periodic Table of the Elements.*

Introductory questions

Define each property and give its typical units, if appropriate.

a. ionization energy

b. Atomic radius

c. Melting point

d. Atomic mass

e. Density

f. Electronegativity

Problem

The purpose of this activity is to re-create Mendeleev's discovery of the classification of the elements and the periodic law using a special deck of element cards. The real properties of the elements, but not their names or symbols, are written on these acards. As the cards are arranged and rearranged based on logical trends in some of these properties, the nature of the periodic law should reveal itself.

Materials

Deck of Element Cards

Procedure

1. Spread the cards out on the table.
2. Each card lists the properties of a single element (X) as shown below:

Ionization energy	**Atomic radius**
Atomic mass	**Formula of its oxide X_bO_c**
Formula of its chloride XCl_a	**Melting point**
Density*	**Formula of its† hydride XH_d**
Electronegativity	

*Density values are in units g g/cm^3 for solids and liquids, g/L for gases.
† Dashed lines for a property indicate that no data is available. Some elements, for example, may not form a compound with hydrogen.

3. Workings together, discuss the possibilities for arrangement of the elements cards. Consider the similarities and differences among the elements as well as possible numerical or logical trends in their properties.
4. It is NOT within the rules of this game of chemical solitaire to look up information in a textbook or to use a modern periodic table as a guide! ***Mendeleev's greatest insight in creating the periodic table was in recognizing there were some gaps when the elements arranged in logical order. He had the ingenuity not only to leave blanks in his table for the missing elements, but also to predict their properties.***
5. One of the element cards is also missing in your deck of cards. Decide where the missing element belongs in the arrangement of the elements and rearrange the cards if necessary to accommodate the missing element.
6. On the worksheet, fill in the Table of the Elements to illustrate a logical arrangement of the element cards. To do this, write down only the atomic mass of each element, as shown on its card. Leave a blank space for the missing element. ***Note: The table is 8X8 and contains 32 more squares than are needed – plenty of room to arrange the elements many different ways.***
7. Predict the properties of the missing element by averaging the properties of its nearest neighbors. On the worksheet, complete the card for the missing element by entering its predicted properties alongside the name of each property.

Observations and data

Table of the Elements

Properties of the missing Element

Ionization energy	**Atomic radius**
Atomic mass	**Formula of its oxide**
Formula of its chloride	**Melting point**
Density	**Formula of its hydride**
Electro-negativity	

Conclusions and Questions

1. Mendeleev's Periodic Law can be stated:" *The physical and chemical properties of the elements are periodic functions of their atomic masses*". Looking at your arrangement of the element cards, describe in your own words what the term "periodic function" means.

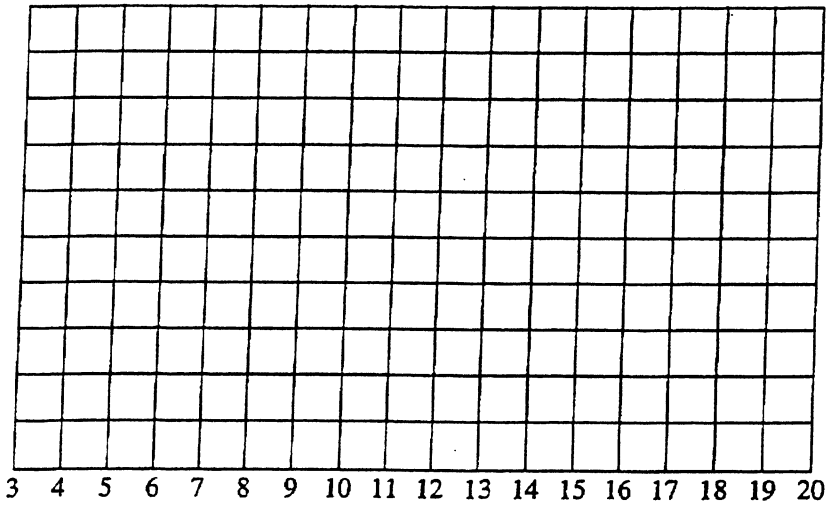
2. Some of the properties listed on each card are periodic properties, others are not. Name one property that is periodic and one that is not.

3. The elements in the modern periodic table are arranged in order of increasing atomic number (instead of increasing atomic mass). Why didn't Mendeleev use atomic number to arrange the elements?

4. From your instructor, obtain a handout showing one possible arrangement of the element cards. Identify each element on the handout with its atomic number and symbol. Use your textbook to obtain this information.

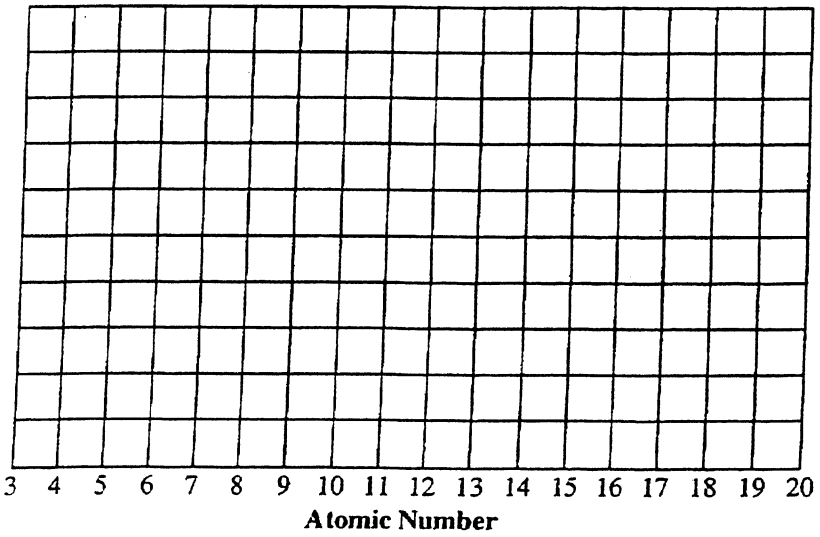
5. Using the possible arrangement of the element cards obtained from your instructor, pick two of the numerical properties of the elements that are periodic and plot their values on the graph below. Give each a descriptive title and label the axes.

Title_____



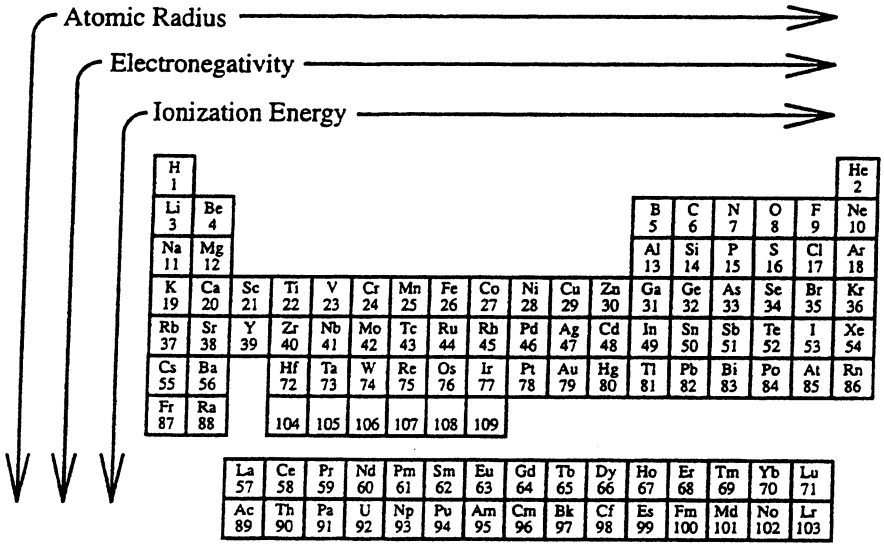
Atomic Number

Title_____



Atomic Number

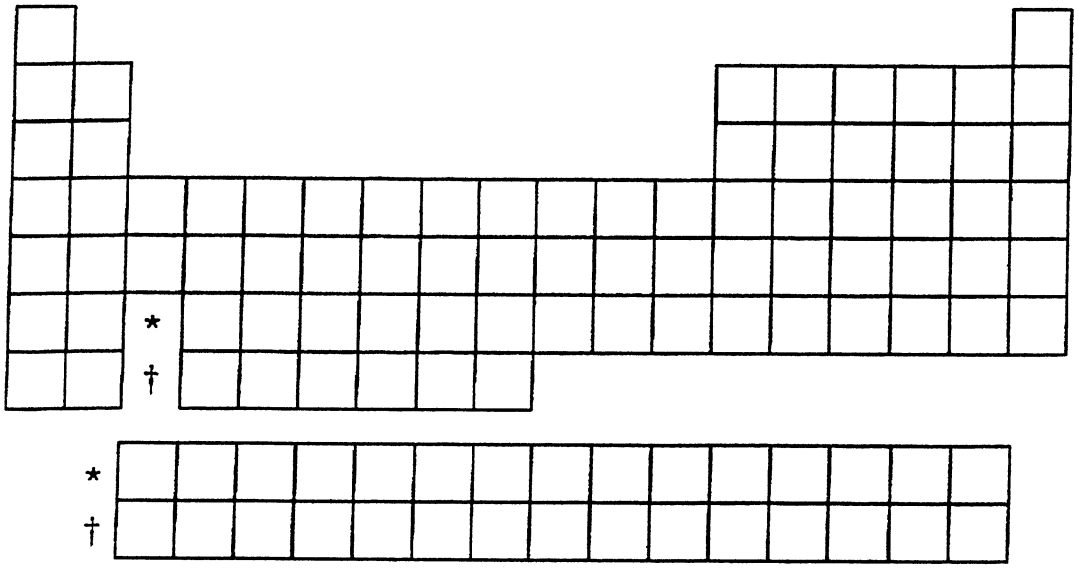
6. There are certain trends in the properties of the elements, both within a column (from top to bottom) and across a row (from left to right) in the periodic table. On the arrow for each property, write the word "increases or decreases" to describe how that property changes.



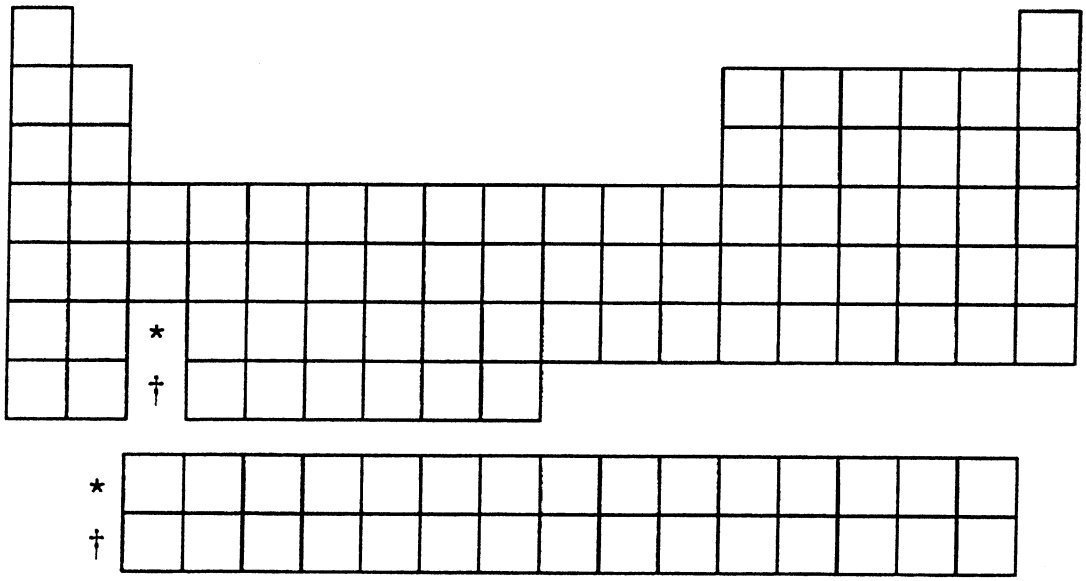
H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56		Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88		104	105	106	107	108	109									

La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

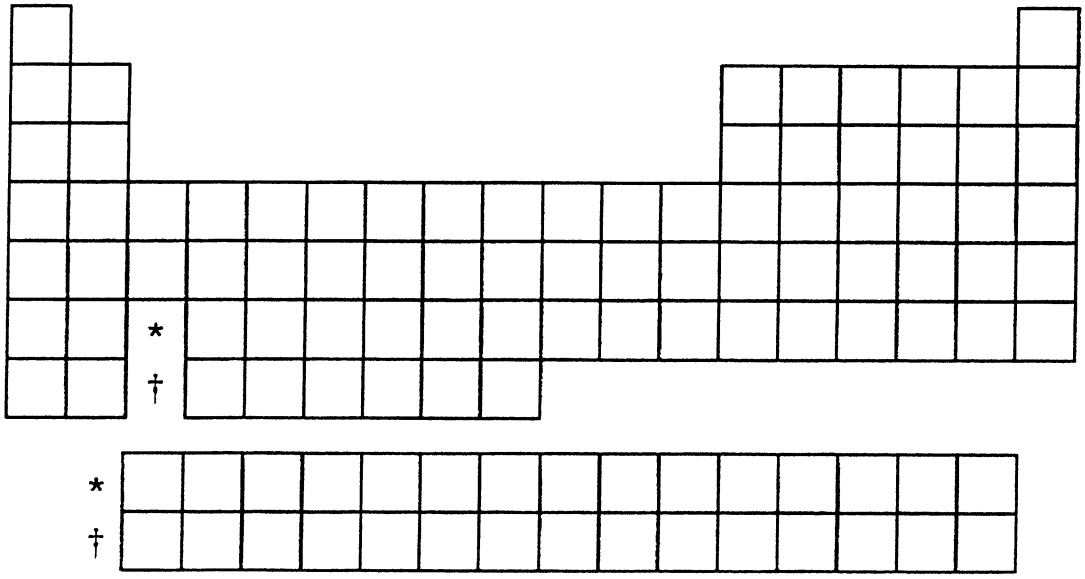
7. On the outline of the periodic table shown below, locate the following: metals, nonmetals, and metalloids.



8. On the outline of the periodic table shown below, locate the following: groups of elements, periods of elements, noble gases, alkali metals, alkaline earth metals, and halogens.



9. On the outline of the periodic table below, locate the following: transition elements, inner transition elements, representative elements.



Element Card Arrangement Handout

520	0.16	899	0.11	801	.083	1086	.077	1402	.070	1314	.066	1681	.062	2081	0.13
6.9	X_2O	9.0	XO	10.8	X_2O_3	12.0	XO_2	14.0	X_2O_5	16.0	XO	19.0	X_2O	20.2	—
XCl	453.7	XCl_2	1560	XCl_3	2300	XCl_4	4100	XCl_3	63.14	XCl_2	50.35	XCl	53.48	—	24.55
0.53	XH	1.85	XH_2	2.46	XH_3	2.27	XH_4	1.25	XH_3	1.4	XH_2	1.7	XH	0.901	—
0.98		1.57		2.04		2.55		3.04		3.44		3.98		—	
496	0.19	738	0.16	578	0.14	786	0.11	1012	0.11	1000	0.10	1251	0.10	1521	0.17
23.0	X_2O	24.3	XO	27.0	X_2O_3	28.1	XO_2	31.0	X_2O_5	32.1	XO_2	35.5	X_2O	39.9	—
XCl	371.0	XCl_2	922	XCl_3	933.3	XCl_4	1685	XCl_3	317.3	XCl_2	388.4	XCl	172.2	—	83.81
0.97	XH	1.74	XH_2	2.7	XH_3	2.33	XH_4	1.8	XH_3	2.07	XH_2	3.17	XH	1.78	—
0.93		1.31		1.61		1.90		2.19		2.58		3.16		—	
419	0.24	590	0.20	579	0.14	762	0.12	947	0.12			1140	0.12	1351	0.19
39.1	X_2O	40.1	XO	69.7	X_2O_3	72.6	XO_2	74.9	X_2O_5			79.9	X_2O	83.8	—
XCl	336.4	XCl_2	1112	XCl_3	302.9	XCl_4	1210	XCl_3	1081	"Missing		XCl	265.9	—	115.8
0.86	XH	1.55	XH_2	5.91	XH_3	5.3	XH_4	5.72	XH_3	Element"		3.1	XH	3.74	—
0.82		1.00		1.81		2.01		2.18				2.96		—	
403	0.25	549	0.22	558	0.17	709	0.16	834	0.14	869	0.14	1008	0.14	1170	0.21
85.5	X_2O	87.6	XO	114.8	X_2O_3	118.7	XO_2	121.8	X_2O_5	127.6	XO_2	126.9	X_2O	131.3	XO_3
XCl	312.0	XCl_2	1041	XCl_3	429.8	XCl_4	505.1	XCl_3	904	XCl_2	722.7	XCl	386.7	—	161.4
1.53	XH	2.6	XH_2	7.31	XH_3	7.30	XH_4	6.68	XH_3	6.24	XH_2	4.66	XH	5.89	—
0.80		0.95		1.78		1.96		2.05		2.10		2.66		—	

Key

Ionization Energy	Atomic Radius
Atomic Mass	X_bO_c
XCl_a	Melting Point
Density†	XH_d
Electro-negativity	

†Density values are in units of g/cm^3 for solids and liquids and g/L for gases.

Name_____ Date_____

22 Group 2A: The Alkaline Earth Metals

Introduction

The elements in group 2 of the periodic table are called the *alkaline earth elements*. Like the elements in group 1 (the alkali metals), the elements in group 2 are chemically active and are never found in nature in the elemental state. Like all members of the *group*, or *family*, the elements in group 2 share certain common characteristics.

The metallic character, the tendency to donate electrons during a chemical reaction, of group 2 elements *increases as you go down the group*. The more metallic of these elements typically react with water to form hydroxides and hydrogen gas. An example of such a reaction would be:

$$Ca_{(s)} + 2HOH \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$$

As metallic character increases (as you go down the group), *the tendency for these elements to form ions increases.* Also, as you go down the group, the solubility of the hydroxides formed by the elements of this group increases. The more active the metal, the more basic its saturated hydroxide solution.

The solubilities of alkaline earth compounds also show some interesting and useful tendencies. For example, the sulfate compounds of alkaline earth metals show decreasing solubilities as you go down the group. This characteristic is used as a means of separating and identifying metallic ions of this group. *Carbonates of all alkaline earth metals are quite soluble.*

In this experiment, you will observe some characteristics of the alkaline earth metals discussed here, and will write balanced equations for all reactions.

Introductory Questions

1. Which elements are the alkaline earth elements?

2. Explain the similarities and differences *in structure* among the elements in this group?

3. What does *metallic character* mean?

4. As you go down group 2, what happens to the metallic character of the elements?

Problem

What are the characteristics of the group 2 elements?

Equipment

Balance
Bunsen burner
Test tubes, 13 x 100-mm (3)
Test tube holder
Weighing paper
nichrome loop
pH paper
stirring rod
test tube rack
spot plate

Materials

calcium turnings (Ca)
magnesium turnings (Mg)
magnesium sulfate crystals ($MgSO_4$)
calcium sulfate crystals ($CaSO_4$)
barium sulfate crystals ($BaSO_4$)
Distilled water
Phenolphthalein

Saturated solutions of:
- calcium hydroxide $Ca(OH)_2$
- magnesium hydroxide $Mg(OH)_2$
- barium hydroxide $Ba(OH)_2$

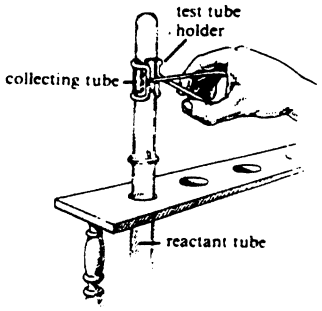
0.1 M solutions of:
- sodium carbonate Na_2CO_3
- magnesium chloride $MgCl_2$
- Calcium chloride CaCl2
- Barium chloride $BaCl_2$

Procedure

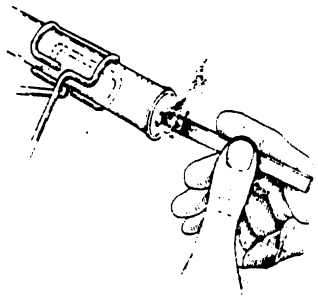
Record all observations and results in the Observations and Data section

Part 1 Test for hydrogen gas (An active metal plus water yields hydrogen gas and a hydroxide.)

1. Pour about 5 mL of distilled water into a clean, dry test tube and stand the tube in the test tube rack. Add several calcium turnings to the water in the tube. To collect the gas being released, invert a clean, dry test tube over the reactant tube, holding the inverted tube with a test tube holder. See the picture below.



2. Test for hydrogen gas by inserting a burning wood splint into the upper part of the inverted tube.



3. Add a few drops of phenolphthalein solution to the reactant tube. After making observations, discard the contents of the tube and clean and dry the tube.
4. Repeat step 1, using several pieces magnesium turnings in place of calcium. If no visible reaction occurs, heat the water to boiling, using a test tube holder over the burner flame. **Caution:** ***Point the mouth of the test tube away form yourself and others.***
5. Once the water is boiling, stand the tube in a test tube rack and, using a test tube holder, invert a collecting tube over the reactant tube. After a few seconds, test for hydrogen gas.
6. Turn off the burner and add a few drops of phenolphthalein to the reactant tube. Record your observations. Discard the contents of the tube, and clean and dry the tube.

Part II Observe the pH of the hydroxides of the group 2 elements.

7. Obtain 5-mL samples of saturated solutions of calcium hydroxide, magnesium hydroxide, and barium hydroxide. Test each solution with pH paper. Record the pH of each solution.

Part III Observe the solubilities of the group 2 elements.

8. Using the balance, measure out a 1-g sample of magnesium sulfate. Place it in a clean, dry test tube.
9. Repeat step 8 for calcium sulfate and barium sulfate.
10. Add 5 mL of distilled water to each tube. Using a glass stirring rod, stir each mixture thoroughly, getting as much of each solid to dissolve as possible. Record your observations of the relative solubilities of each of these compounds.
11. Conduct a flame test for calcium and barium ions. Dip the wire loop into each solution. Test each in the flame and record the color of the flame.

Part IV Observe the solubilities of the carbonates of the group 2 elements.

In this experiment, a double replacement reaction produces a group 2 carbonate. Here are the balanced equations for these reactions.

$$MgCl_2 + Na_2CO_3 \rightarrow 2NaCl + MgCO_3$$

$$CaCl_2 + Na_2CO_3 \rightarrow 2NaCl + CaCO_3$$

Write the last one yourself in Part IV of the Calculations:

12. Using the 0.1 M solutions, add 5 drops of the $MgCl_2$ solution to one spot well, 5 drops of the $CaCl_2$ solution to a second spot, and 5 mL of $BaCl_2$ to the third spot.
13. To each of the solutions in the spot wells, add about 1 drop of the Na_2CO_3 solution. Record your observations. Are these solutions soluble?

Observations and Data

Part I

Ca + HOH: Result of test for H_2 gas

Result of adding phenolphthalein

Mg + HOH: Result of test for hydrogen (before heating)

Result of adding phenolphthalein

Part II pH readings

$Mg(OH)_2$

$Ca(OH)_2$

$Ba(OH)_2$

Part III Apparent solubility

$MgSO_4$

$CaSO_4$

$BaSO_4$

Flame test results

Ca^{2+}

Ba^{2+}

Part IV Observations about solubility of the carbonate of the group 2 elements

Mg^{2+} ($MgCO_3$)

Ca^{2+} ($CaCO_3$)

Ba^{2+} ($BaCO_3$)

Calculations

Write balanced equations for each reaction that occurred in Step 1 and Step4.
Here is the first sample:

Step 1. $Ca_{(s)} + 2H_2O \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$

Step 4.

Conclusions and Questions

1. Describe the reactivity of the metals in group 2 in terms of their location in the group.

2. How does the reactivity of the alkaline earth metals compare with that of the alkali metals (group 1) in the same period?

3. What is the oxidation state for all the alkaline earth metals? Explain why this is so.

4. Explain why the metallic character of the group 2 elements increases as you go down the group.

Name_____ Date_____

23 Hydrogen: Preparation and Properties

Introduction

Hydrogen is the simplest element; it consists of 1 proton and 1 electron. It is unique among the elements in that it does not really belong to any particular group of the Periodic Table. In fact, in some versions of the table, hydrogen is separated from the main body of elements. In some tables, hydrogen appears aligned above groups 1 and 17 because it can both give up and gain one electron depending on what it combines with. Although hydrogen is classified as a *non-metal*, it combines with almost all elements. Hydrogen can form a covalent bond with other non-metals by sharing its one electron. In such compounds, its oxidation number is 1^+. Hydrogen also combines with a group 1 or group 2 metals to form *ionic compounds* called *metal hydrides*. In this case it has an oxidation number of 1^- or 2^- respectively.

In this experiment you will prepare hydrogen gas by reacting an active metal with an acid. The gas will be collected by water displacement, and several properties of hydrogen will be investigated.

Introductory Questions

1. What are some safety techniques that you should be aware of while doing this lab?

2. How do you test for hydrogen gas?

Problem

To prepare and capture hydrogen gas and to investigate some of its properties

Equipment

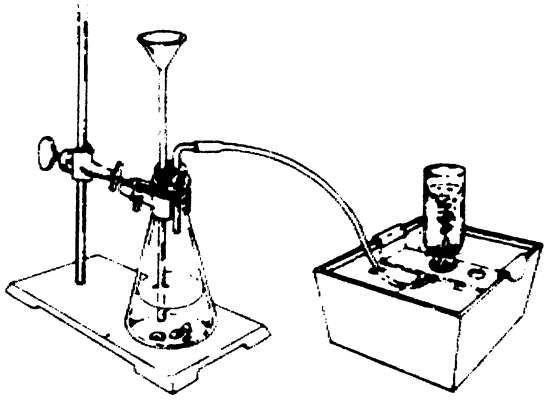
Ring stand
Clamp
Erlenmeyer flask, 250-mL
Thistle tube
Rubber stopper, 2-hole (for flask)
Glass tubing
delivery tube, rubber or plastic
gas-collecting bottles, small (3)
trough
glass squares (2)
safety glasses

Materials

Zinc, mossy (Zn)
6 M Hydrochloric Acid (HCl)
Wood splints
matches

Procedure

1. Place 3 or 4 pieces of mossy zinc in a 250-mL Erlenmeyer flask. Clamp the flask to a ring stand as shown below.



2. Fit the thistle tube and one of the curved pieces of glass tubing into a 2-hole rubber stopper. Place the stopper securely in the flask, making sure that the end of the thistle tube extends below the surface of the water in the flask. This step may have already been done for you.

3. Fill the trough with water. Fill 2 gas-collecting bottles with water and leave them upside down in the trough. Attach the delivery tube.

4. ***Slowly and carefully*** add enough 6M HCL through the thistle tube so that the end of the thistle tube is beneath the acid level. **CAUTION:** ***Handle this acid carefully. It will cause painful burns if it comes in contact with skin.***
5. Allow the bubbles to escape into the trough for about 30 seconds. (If bubbles are not escaping rapidly, inform your instructor.) Then, place a gas-collecting bottle (filled with water) mouth down over the end of the delivery tube. Fill the bottle with hydrogen gas.
6. When the bottle is filled, lift the bottle slightly, and slide a glass square over the mouth of the bottle under water. Holding the glass square in place, lift the bottle out of the water and place it mouth down on the lab table, with the glass square still in place.
7. In the same manner, collect another bottle of hydrogen gas.
8. Disconnect the rubber stopper and remove the tube form the trough. Unclamp the flask and carefully fill the flask with cold tap water. Discard the liquid and any remaining zinc as instructed.

9. Keeping the bottle mouth down, lift one of the bottles of hydrogen gas and insert a burning splint up into the bottle. Record your observations.
10. Keeping the glass square in place over the mouth of the bottle, turn the second bottle of hydrogen gas over. Place an inverted empty gas-collecting bottle on the glass square so that the two bottles are mouth-to-mouth. Slide the glass square away, leaving the two bottle mouths resting against each other. After 30 seconds, test both bottles for hydrogen gas with a burning splint. Record your observations.

Observations and Data

Step 9:

Step 10:

Conclusions and Questions

1. Why is it important to have the thistle tube extend below the surface of the liquid in the flask?

2. Why is the production of gas allowed to operate for 30 seconds before starting to collect hydrogen gas?

3. Write a balanced equation for the reaction used to produce hydrogen gas in this experiment.

4. What product results from the burning of hydrogen (This is what happens when you insert a lighted splint into the bottle of hydrogen gas.)? Write a balanced equation for this reaction.

5. Based on observations from this experiment, describe the solubility of hydrogen gas.

6. Based on your observations from this experiment, describe the density of hydrogen gas compared to the density of air.

Name_____ Date_____

24 Lewis Electron Dot Structures

Introduction

Lewis electron dot structures are a useful and popular way in chemistry classes to visualize the bonding in molecules.

A chemical bond is a strong attractive force between atoms or ions in a compound. Electrons in the outermost shells or orbitals are often the electrons involved in a chemical bond. One type of chemical bond between two atoms is called a covalent bond. The covalent bond consists of a pair of shared electrons from the outermost orbitals. The electrons in this outermost shell, or highest occupied energy level, have a special importance and are referred to as valence electrons. The electrons in the inner orbitals are referred to as core electrons. Since the maximum number of electrons in an orbital is two and since there is one s orbital and three p orbitals for any given shell, the maximum number of valence electrons is 8.

Writing Lewis Electron Dot Structures for Atoms:

A Lewis electron dot structure for an atom includes the element symbol and the "electron" dots. The number of dots is equal to the number of valence electrons which, as explained previously, is equal to the group number of the periodic table where the element is located. To write Lewis dot structures for atoms, the first two dots are placed on the same side, representing the two electrons in the s orbital. Then the dots are placed one at a time on the remaining three sides. These three sides represent the p orbitals. The maximum number of electron dots on any atom is 8. Thus the Lewis electron dot structure for potassium, which is located in Group 1on the periodic table, contains only one dot(below) while the Lewis electron dot structure for bromine, which is located in Group 7 on the periodic table, contains seven dots.

Potassium Bromine

Note: The exact placement of the dots is arbitrary, so the single dot around K can be on any of the four sides.

Writing Lewis Electron Dot Structures for Molecules:

A Lewis electron dot structure for a molecule includes each atom in the molecule connected by bonds. Each bond may be drawn with two electron dots or, a single line. In order to predict the bonding arrangement that occurs between atoms in a molecule, the stable nature of noble gases must be recognized. Noble gases are unreactive and form very few compounds. This stability is due to the outermost s and p orbitals being completely filled and containing two s electrons and six p electrons. Helium has a filled valence orbital with only two s electrons while all of the other noble gases have eight electrons in their valence orbitals, two in the s orbital and six in the p orbital. From these observations of the noble gases, a rule called the octet rule was developed. The octet rule

assumes that atoms form bonds to achieve stability and sharing, each atom in a molecule attains a noble gas configuration.
To write a Lewis dot structure for a molecule, a simple process may be followed. Determine the number of valence electrons supplied by each atom involved. Visualize each atom donating its valence electrons to a large electron pool. When the atoms come together to form a molecule, picture the electrons from the pool being distributed to the molecule so that each atom in the molecule attains a noble gas structure. Some electrons may need to be shared to form bonds, while others may be left unshared.

In General:

1. H atoms want to be surrounded by 2 valence electrons, to attain the stability of He
2. All other nonmetal atoms want to be surrounded by 8 valence electrons.

Step by step procedure:

1. Count the total number of valence electrons available.
2. Draw a "skeleton "structure for the molecule, joining atoms by single bonds.
   a. The skeleton structure for X_2 is: X—X
   b. The skeleton structure for XY_2 is: Y—X—Y

Tips:

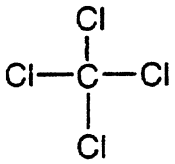
Nature tends toward symmetry, so it is best to make your skeleton structure symmetrical when possible.

Hydrogen is never the central atom; it is always a surrounding atom.

Group 7 halogens tend to be surrounding atoms rather than the central atoms.

Carbon tends to be a central atom.

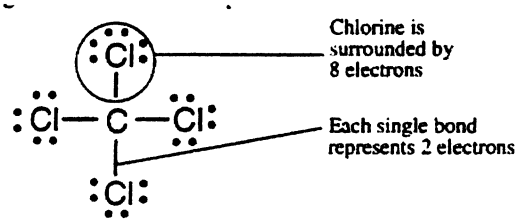
Ex: CCl_4



3. From the total number of valence electrons, subtract two for each single bond in the skeleton structure to determine how many valence electrons are left to distribute.
   a. ex: CCL_4 has 32 total valence electrons. Subtract 2 electrons for each single bond x 4 single bonds. This leaves 32-8 or 24 remaining valence electrons.

4. Distribute the remaining valence electrons as unshared pairs around the atoms in the molecule.

a. Ex:



5. If this point is reached and there are too few valence electrons to give each atom an octet, a multiple bond is most likely needed.

a. Ex: H—C≡N:

Introductory questions

1. Why do atoms become more stable when they have eight valence electrons?

Problem

How can you construct electron-dot diagrams for covalent molecules?

Materials

4 Element Sheets
Scissors
Bingo chips

Pre-Lab:

Working individually or in teams, cut out the 32 element cards from the element sheets.

Procedure

Part I Elements

1. Locate Data Table 1
2. Record the family or group number for each element listed in Data Table 1. Refer to a periodic table to do this step.
3. List the number of valence electrons for each element in the next column of the Data Table.
4. Locate the element card for the first element in the table and set the card on a flat surface. Use bingo chips to place the correct number of valence electrons around the element. Follow the rules for writing Lewis Dot Structures outlined in the background section of this lab. This is the Lewis Dot Structure for this element.
5. Draw the Structure in the last column of Data Table 1.
6. Repeat this procedure for each element in the Table.

Part II

7. Locate Data Table 2.
8. Determine the number of valence electrons for each individual atom in the first compound. Write each separate number in the Data Table.
9. Add up the valence electrons for each atom and record the total number of valence electrons for the compound in the next column.
10. Locate the element cards for each atom in the compound and place them on a flat surface. Gather enough colored bingo chips to equal the total number of valence electrons. Use a different color for each element in the compound.
11. Use the bingo chips to place the correct number of valence electrons in the compound. Follow the rules for atom placement and electron placement.
12. Draw the structure in the Data Table.
13. Repeat this procedure for each compound in the data table.

Observations

Part I

Data Table 1: Elements

ELEMENT	GROUP #	# VALENCE ELECTRONS	LEWIS DOT STRUCTURE
1. H			
2. He			
3. Li			
4. Be			
5. B			
6. C			
7. N			
8. O			
9. F			
10. Ne			
11. Na			
12. Mg			
13. Al			
14. Si			
15. P			
16. S			
17. Cl			
18. Ar			

PART II

DATA TABLE 2: COVALENT COMPOUNDS

Formula	Valence electrons for each atom	Total number of valence electrons	Lewis Electron Dot Structure
1. H_2			
2. HCl			
3. H_2O			
4. NH_3			
5. CH_4			
6. CH_2Cl_2			
7. Cl_2			
8. HOCl			
9. O_2			
10. OH^-			
11. H_3O^+			
12. NH_4^+			

13. SO_4^{2-}			
14. N_2			
15. $CHCl_3$			
16. C_2H_6			
17. CO_2			
18. CH_3OH			
19. H_2O_2			
20. C_2H_4			
21. C_2H_2			
22. SO_2			

Conclusion

1. Why can atoms have no more than 8 valence electrons?

2. Write the Lewis electron dot structures for the following elements:
   a. Barium
   b. Tin
   c. Radon
   d. Gallium
   e. Potassium
   f. Iodine

Name_____ Date_____

25 Ionic Formula Writing

Introduction

In this activity, you will cut cation and anion cards and combine them in the correct ratio to form a neutral ionic compound. The chemical formula and the name of the compound will then be written. This activity leads to the discovery of a systematic method of writing formulas for ionic compounds without the need to experimentally determine each formula.

The *chemical formula* of a compound indicates the kinds of atoms and the relative number of each atom that is combined together to form the compound. *Subscripts* are the numbers written after the chemical element symbols and are used to indicate the number of atoms of each element in the compound. Water, the most abundant chemical on Earth, has the chemical formula H_20 and has a whole-number hydrogen:oxygen ratio of 2 to 1.

Some compounds such as sodium chloride, which is common table salt, are composed of ions. Ions are atoms or groups of atoms that have a net positive or negative charge. *Cations* are atoms or groups of atoms with a net positive charge, such as Ca^{+} of NH_4^{+}, while *anions* or groups of atoms with a net negative charge, such as O^{-2} or NO_3^{-}. Atoms of metallic elements tend to form cations while atoms of nonmetallic elements tend to form anions. Ions join together because of an oppositely charged or electrostatic attraction between a positive cation and a negative anion resulting in an ionic bond. Compounds held together by ionic bonds are ionic compounds.

Ionic compounds are generally composed of an extended network of positively-charges metal ion and negatively-charged nonmetal ions. Although composed of ions, ionic compounds are electrically neutral, i.e., the total positive charge is equal to the total negative charge. The chemical formula indicates the smallest whole number ratio of each element in the network of ions and is called a *formula unit*. For example, magnesium chloride has the chemical formula $MgCl_2$. The overall charge on the resulting ionic compound is zero. Thus the final compound is composed of two Cl^{-} ions for each Mg^{+2} ion.

Some ions consist of a group of covalently bonded atoms that tend to stay together as if they were a single ion. Such ions are called polyatomic ions. An example if the nitrate ion, NO_3^{-}. This *polyatomic anion* contains one nitrogen covalently bonded to three oxygen atoms and has an overall charge of -1. Calcium cautions, Ca^{+2}, in the presence of nitrate anions. NO_3^{-}, will combine in a ratio of two nitrate ions for every calcium ion in order to balance the positive and negative charges and achieve electrical neutrality. Thus the formula for calcium nitrate is written as $Ca(NO_3)_3$, showing that for every calcium ion there are two nitrate ions. Notice that the polyatomic ion is put in parentheses as a unit with the subscript on the outside of the parentheses.

Introductory Questions

1. Define the following:

chemical formula:

_____

subscripts:

_____

anions:

_____

cations:

_____

ionic compounds:

_____

formula unit:

_____

polyatomic anions:

_____

ionic bond: -

_____

Problem

How do you write a formula?

Equipment

Packet of cations and anions
Ion formula chart

Materials
worksheets

Procedure
1. From the cation and anion cards, locate ion cards for the first two combining ions listed in the first column of the data table. The first two, for example, are the aluminum cation and chloride anion. Copy the ionic formula for each ion into the firs column of the data table. Use the Ion Formula Chart to look up the ion formulas.

2. Put the cards for the two combining substances together on a flat surface adding additional cards of the same ion until the heights of the cation and anion column are equal. For example, every aluminum cation needs three chloride anions.

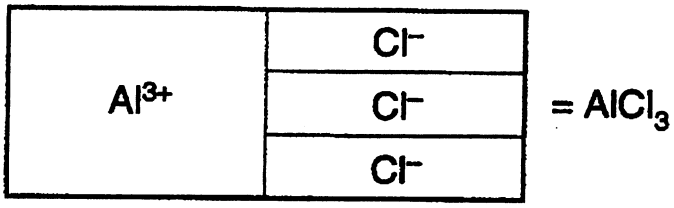


***Figure 1:* Aluminum chloride**

3. Count the number of each ion necessary for the heights to be equal (ie. For the compound to form). For example, one aluminum cation combines with three chloride anions. Record the number of each ion used in Columns 2 and 3 of the data table. Notice that the numbers recorded in Columns 2 and 3 are the subscripts in the chemical formula of the compound.

4. Determine the total positive charge of the cation card(s) and the total negative charge of the anion card(s) used in forming the compound. These two values will be equal yet opposite in value. For example, the total positive charge is +3 (from one aluminum and a +3 charge) and the total negative charge is -3(from three chlorides, each with a -1 charge). Record the ionic charges in columns 4 and 5 of the data table.

5. Write the formula for the ionic compound, using the subscripts to indicate the number of each kind of ion used (from Columns 2 and 3). Follow the rules for writing ionic formulas. Record the formula for the ionic compound in Column 6 of the data table.

6. Write the name of the ionic compound in the last column of the data table.

7. Repeat the steps above for each set of combining ions in the data table.

8. Complete rows 17 – 20 of the data table by randomly choosing four more sets of combining ions (that have not already been combined). Complete each column of the data table.

9. Gather and count the ion cards. There is 3 of each cation(pink) 30 total and 3 of each anion(blue) total 39. Place them in the packet and return the cards.

Conclusions and Questions

1. What did you notice about the height of the cards? What is the relationship between the height and the ion charge?

2. Examining the combing substances and formulas you wrote in your data table, develop a general procedure explaining how to write a formula for an ionic compound.

Data Table

Combining Ions	Number of cations	Number of anions	Total + charge	Total - charge	Formula	Name of compound
1. Aluminum & Chloride	1	3	+3	-3	$AlCl_3$	Aluminum chloride
2. Sodium & Oxide	2	1	+2	-2	Na_2O	Sodium Oxide
3. Iron(II) & Sulfide	1	1	+2	-2	FeS	Iron(II)sulfide
4. Potassium & Sulfate	2	1	+2	-2	K_2SO_4	Potassium sulfate
5. Silver & Nitrate	1	1	+1	-1	$AgNO_3$	Silver nitrate
6. Copper(II) & Hydroxide	1	2	+2	-2	$Cu(OH)_2$	Copper(II)hydroxide
7. Ammonium & Carbonate	2	1	+2	-2	$(NH_4)_2CO_3$	Ammonium carbonate
8. Copper(I) & Phosphate	3	1	+3	-3	Cu_3PO_4	Copper(I)phosphate
9. Magnesium & Bromide	1	2	+2	-2	$MgBr_2$	Magnesium bromide
10. Calcium & Acetate	1	2	+2	-2	$Ca(C_2H_3O_2)_2$	Calcium acetate
11. Iron (III) & Oxide	2	3	+6	-6	Fe_2O_3	Iron(III)oxide
12. Copper(I) & Chromate	2	1	+2	-2	Cu_2CrO_4	Copper(I)chromate
13. Sodium & Bicarbonate	1	1	+1	-1	$NaHCO_3$	Sodium bicarbonate

14. Magnesium & Phosphate	3	2	+6	-6	$Mg_3(PO_4)_2$	Magnesium phosphate
15. Potassium & Cyanide	1	1	+1	-1	KCN	Potassium cyanide
16. Aluminum & Sulfide	2	3	+6	-6	Al_2S_3	Aluminum sulfide
17.						
18.						
19.						
20.						

Name_____ Date_____

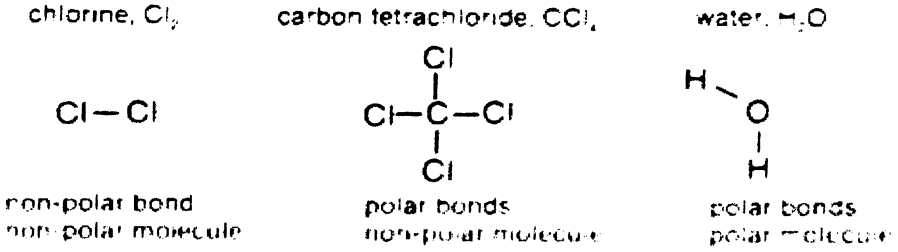
26 Bonds, Polarity, and Solubility

Introduction

"Like dissolves like" is an expression that chemists use when discussing the solubility of substances. It means that, in general, non-polar substances dissolve in non-polar solvents while polar substances dissolve in polar solvents. Ionic substances also dissolve in polar substances.

What determines the polarity of a molecule? First, keep in mind that true molecules contain covalent bonds. (Ionic compounds are not molecular.) The polarity of a molecule is determined by the type of covalent bond (polar or non-polar) and by the arrangement of the bonds. If all the bonds in a molecule are polar, then the arrangement of the bonds determines the polarity of the molecule. If the bonds are arranged symmetrically, the positive and negative charges in the molecule will be evenly distributed and the molecule will be non-polar. If the bonds are not arranged symmetrically, the charges in the molecule will not be evenly distributed. One "end" of the molecule will be slightly negative, and the other "end" will be slightly positive. The entire molecule will be polar.

Examples of the three cases just described are illustrated below.



As indicated, water molecules and carbon tetrachloride both contain *polar bonds*. However, the bonds in the water molecule are not symmetrically arranged. As a result, the oxygen end of the molecule is slightly more negative than the hydrogen end, and the *molecule is polar*. The symmetrical arrangement of polar bonds in the carbon tetrachloride molecule produces an even distribution of charges in the molecule, making it non-polar.

Some substances show an *intermediate polarity* because of complex bonds arrangement in the molecular structures. Such substances show some degree of solubility with both polar and non-polar substances.

When discussing solubility, we usually think in terms of the tendencies of solids to dissolve liquids. However, liquid-liquid solutions are also important in chemistry. Liquids that dissolve in each other are said to be *miscible*. If two liquids dissolve in each completely in all proportions, they are said to be *completely miscible*. If one liquid dissolves in another liquid to a limited extent, the liquids are said to be *partially miscible*. A mixture of two immiscible liquids forms two distinct layers, with the less dense layer on top. A mixture of oil and water is such an example. A mixture of two partially soluble liquids will form two layers only if enough of each liquid is present.

In this experiment, we will investigate the solubilities of an ionic solid and a non-polar solid in liquids of varying properties. We will also investigate the miscibilities of liquids of varying polarities in each other.

Introductory Questions

1. Name the elements that form diatomic molecules.

Which type of covalent bond forms between the atoms in such molecules?

Which polarity do these molecules show, polar or non-polar?

Do you think these molecules would dissolve in water? Explain.

2. What type of covalent bond forms between the oxygen and the hydrogen in a molecule of water?

Is the molecule symmetrical? Is the molecule polar or non-polar?

3. Sketch the structural formula for ethyl alcohol. C_2H_5OH

Do you think it is symmetrical, partly symmetrical of non-symmetrical?
Explain your answer.

Do you think it will be miscible with water? Explain your answer.

Problem

Can you determine if a substance will dissolve in a polar or non-polar solvent?

Equipment

Test tubes, 13x100-mm (4)
Test tube rack
Microspatula
Stoppers or corks (for test tubes)
graduated cylinder, 10-mL
marking pencil
safety glasses

Materials

Solids: Sodium chloride crystals (NaCl)
Iodine crystals (I_2)

Liquids: distilled water (H_2O)
Hexane
Ethyl alcohol

Procedure

Part I Solubility of solids in liquids

1. Place four clean, dry test tubes in a test tube rack. Using a wax marking pencil, number the test tubes 1 through 3.

2.

Measure 5 mL of distilled water in a graduated cylinder and pour it into test tube #1. Using the level of the water in test tube #1 as a guide, add the same amount to the other test tubes as follows:

#2 hexane

- #3ethyl alcohol

3. To each of the test tubes, add 2 crystals of iodine. Stopper and shake each tube. Record your observations in Data Table 1. Clean and dry the test tubes with soap and water.
4. Repeat steps 1 through 3, substituting sodium chloride crystals for iodine.

Part II Miscibility of Liquids

5. Measure 5 mL of distilled water and pour it into a clean, dry test tube. Add 5 mL of hexane to the distilled water. Stopper and shake the tube. Record you observations in Data Table 2.
6. In a clean test tube, repeat step #5, substituting *ethyl alchol* for the hexane. Record your observations in Data Table 2.
7. Measure 5 mL of ethyl alcohol in a clean, dry test tube and add 5 mL of hexane to it. Stopper it and shake vigorously. Record you observations in Data Table 2.

Part III Prediction

9. Before you actually do this next step, make a prediction about what is going to happen. Tell your instructor what your expectation is.

Take a clean, dry test tube and add to it 5 mL of water, 5 mL of hexane and 2-3 crystals of iodine. Stopper it and shake well. Allow it to settle and then observe it.

Observations and Data

Part 1

Describe the degree of solubility by indicating *soluble*, *slightly soluble* or *insoluble*.

Data Table 1

Liquid	I_2 crystals	NaCl crystals
water		
hexane		
ethanol		

Part 2

Describe the degree of mixing by indicating miscible or imiscible.

Data Table 2

	water	hexane	
hexane		XXXXXXXXXX	
ethanol			

Part 3

Sketch and label your observations of the test tube.

Conclusions and Questions

1. In which liquids used in this experiment was iodine soluble? Give a brief explanation of why this is true, based on types of bonding and polarity of molecules.

2. Explain why iodine is not soluble in water.

3. What type of bonding is present in sodium chloride crystals?

In which liquids is sodium chloride insoluble? Explain your answer.

4. In the last procedure where you mix hexane and water, hexane and iodine crystals, explain why you got the results you got.

5. Referring to Procedure Part 1, the solution of alcohol and iodine crystals is light purple. What does this tell you about the solubility of iodine in ethanol?

6. Does this confirm your statement in the Introductory Question #3? Explain.

Name_____ Date_____

27 Names and Formulas of Ionic Compounds

Introduction

Double replacement reactions involve an exchange of positive ions between two reacting compounds. Often these reactions are characterized by the formation of a precipitate, a substance that is not soluble. Table F in your Reference Table gives you information about which substances are soluble and which are insoluble.

Here is a double replacement reaction between two ionic solids that have been dissolved in water.

$$2AgNO_{3(aq)} + Na_2CO_{3(aq)} \rightarrow Ag_2CO_{3(aq)} + 2NaNO_{3(s)}$$

If you study Table F, you will see that Ag_2CO_3 is not soluble in water and will precipitate out of the mixture. It is written as a solid $_{(s)}$. The formation of a precipitate may have a variety of appearances. It may appear as particles at the bottom of the reaction vessel or as a cloudy or grainy appearance of the mixture.

Introductory Questions

1. Using Table F determine which of the following are soluble.

$PbCl_2$ _____ $Ca(NO_3)_2$ _____

2. Name this compound: Na_3PO_4 _____

3. Write the chemical formula for: calcium nitrate _____

Problem

Which chemical reactions will form a precipitate?

Can you name ionic compounds correctly?

Equipment

Pencil
Unlined white paper (2)
Safety glasses
Reaction surface (spot plate)
Ruler –straight edge

Materials

All the chemicals listed in Table A and Table B below.

Procedure

1. On two separate sheets of unlined paper, draw two grids (without the chemical names) similar to Table A. Label one "Table A" and the other one "Table B." You will place these sheets under the spot plates to which you add chemicals. Make sure you include the small letters inside each box which will correlate to your answers in the Conclusion section.

Part I

2. Place a reaction surface (spot plate) over the grid you drew and labeled "Table A" and add a few drops of each chemical as shown in Table A.
3. Use the grid on *this* sheet to record your *observations* for each solution. Describe each precipitate that forms as *milky, grainy, cloudy*, or *gelatinous*. Also, describe the color. If there is no visible reaction, write N.R.

Observations and Data

Table A

	$AgNO_3$ (Ag^+)	$Pb(NO_3)_2$ (Pb^{2+})	$CaCl_2$ (Ca^{2+})
Na_2CO_3 (CO_3^{2-})	a	e	i
Na_3PO_4 (PO_4^{3-})	b	f	j
NaOH (OH^-)	c	g	k
Na_2SO_4 (SO_4^{2-})	d	h	l

Part II

4. Repeat the above experiment steps 2 and 3, using the chemicals in Table B and the second grid you drew labeled "Table B."

Table B

	$FeCl_3$ (Fe^{3+})	$MgSO_4$ (Mg^{2+})	$CuSO_4$ (Cu^{2+})
Na_2CO_3 (CO_3^{2-})	m	q	u
Na_3PO_4 (PO_4^{3-})	n	r	v
NaOH (OH^-)	o	s	w
Na_2SO_4 (SO_4^{2-})	p	t	x

Conclusions and Questions

1. Write a balanced equation for the double replacement reaction that occurred when you mixed the chemicals for "Table A", box "a":

$$AgNO_3 \quad + \quad Na_2CO_3 \rightarrow$$

2. Place all the information below in the Conclusion table.
For each of the 24 reactions you carried out:

- Write the correct chemical formula for each insoluble compound produced by the mixings.
- Name of this chemical compounds produced by the mixings.
- Using Table F from your Reference Table, identify which compounds are *insoluble*

Grid Box	Chemical Formula	Chemical Name	Soluble/Insoluble
a			
b			
c			
d			
e			
f			
g			
h			
i			
j			

k			
l			
m			
n			
o			
p			
q			
r			
s			
t			
u			
v			
w			
x			

Name_____ Date_____

28 Types of Chemical Reactions

Introduction

There are many kinds of chemical reactions and several ways of classifying them. One popular and useful method is to classify reactions into four major types. These are:

1. direct combination or synthesis
2. decomposition or analysis
3. single replacement
4. exchange of ions or double replacement

Not all chemical reactions can be put into one of these categories. However, many can.

In a *synthesis reaction*, two or more substances (elements and compounds) combine to form a more complex substance. Equations for synthesis reactions have the general form $A + B \rightarrow AB$. An example of such a reaction is the formation of water from its constituent elements: $2H_2 + O_2 \rightarrow 2H_2O$.

A *decomposition reaction* is exactly the opposite of a synthesis reaction. In a decomposition reaction, a compound breaks down into two or more simpler substances (elements or simpler compounds). The general form of the equation for a decomposition reaction is: $AB \rightarrow A + B$. The breakdown of water into its elements is an example of such a reaction: $2H_2O \rightarrow 2H_2 + O_2$.

In a *single replacement reaction*, one substance in a compound is replaced by another, more active, substance (an element). Equations for single replacement reactions have two general forms. In reactions in which one metal replaces another metal, the general equation is: $X + YB \rightarrow XB + Y$. In reactions in which one nonmetal replaces another nonmetal, the general form is: $X + AY \rightarrow AX + Y$. The following equations illustrate each of these types of single replacement reaction:

Zinc metal replaces copper (II) ion:

$$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$$

Chlorine (a nonmetal) replaces bromide ions:

$$Cl_{2(g)} + 2KBr_{(aq)} \rightarrow 2KCl_{(aq)} + Br_{2(l)}$$

In a *double replacement reaction*, the metal ions of two different ionic compounds can be thought of as "replacing one another." Equations for this type of reaction have the general form: $AB + CD \rightarrow AD + CB$. Most replacement reactions (both single and double) take place in aqueous solutions containing free ions. In a double replacement reaction, one of the products forms a *precipitate*, an *insoluble gas*, or *water*.

All of the types of reactions discussed here may be represented by balanced molecular equations. Reactions involving ion exchanges may also be represented by ionic equations. In this investigation you will be concerned only with molecular formulas and equations. In a balanced equation, the number of atoms of any given element must be the same on both sides of the equation. Multiplying the coefficients times the subscript of an element must yield the same product on both sides of the balanced equation.

In this experiment, you will observe examples of the four types of chemical reactions described above. You will be expected to balance the equations representing the observed equations.

Introductory Questions

1. Which of the following are substances? a) elements b) mixtures c) suspensions d) matter e) compounds ____

2. Name the four types of chemical reactions.

Problem

Can you classify the type of chemical reaction you observe?

Equipment

Bunsen burner
Crucible tongs
miscrospatula
Test tubes, 15x180mm (7)
Test tube rack
wood splints
sandpaper, fine
evaporating dish
safety glasses

Materials

Zinc, mossy (Zn)
Copper wire, 10cm (Cu)
6 M hydrochloric acid (HCl)
1 M copper (II) sulfate ($CuSO_4$)
Magnesium ribbon
0.1 M zinc acetate ($Zn(C_2H_3O_2)_2$)
0.1 M sodium phosphate (Na_3PO_4)
1. M sodium sulfite (Na_2SO_3)
Copper (II) carbonate ($CuCO_3$)

Procedure

A. Synthesis

Record all your observations in the table in the "Observations" section.

1. Use fine sandpaper to clean a piece of copper wire until the wire is shiny. Examine the wire and record its appearance.
2. Using crucible tongs, hold the wire in the hottest part of the flame for 1-2 minutes. Examine the wire and record any change in its appearance caused by the heat.
3. Place an evaporating dish near the base of the burner. Examine a piece of magnesium ribbon. Using crucible tongs, hold the sample in the flame until the magnesium starts to burn. ***DO NOT LOOK DIRECTLY AT THE FLAME. HOLD THE BURNING METAL AWAY FROM YOU AND DIRECTLY OVER THE EVAPORATING DISH.*** When the ribbon stops burning, put what is left of it in the evaporating dish. Examine this product thoroughly and record your observations.

B. Decomposition

4. Place 2 heaping microspatulas of copper (II) carbonate ($CuCO_3$) in a clean, dry test tube. Record the appearance of the sample.
5. Using a test tube holder, heat the $CuCO_3$ strongly in a burner flame for about 3 minutes. Extinguish the flame and insert a burning wood splint in the test tube to test for the presence of carbon dioxide gas. (CO_2 will put out the flame.) Note and record any change in the appearance of the material in the test tube.

C. Single Replacement

6. Stand a clean, dry test tube in a test tube rack. Add about 5mL of 6 M HCl to the tube. **CAUTION:** ***Handle acids with care. They can cause painful burns.*** Carefully drop a small piece of zinc metal (Zn) into the acid in the test tube. Observe and record what happens.
7. Using a test tube holder, invert a second test tube over the mouth of the test tube in which the reaction is taking place. Remove the inverted tube after about 30 seconds and quickly insert a burning wood splint into the mouth of the tube. (You will hear a "pop" if hydrogen gas is present.) Note and record the appearance of the substance in the test tube.)
8. Add about 5mL of 1 M copper (II) sulfate ($CUSO_4$) solution to a clean, dry test tube. Place a small piece of zinc metal in the solution. Note and record the appearance of the solution and the zinc before and after the reaction.

D. Double Replacement

9. Add about 2 mL of 0.1 M zinc acetate ($Zn(C_2H_3O_2)_2$) to a clean, dry test tube. Next, add about 2 mL of 0.1 M sodium phosphate tribasic (Na_3PO_4) solution to the same test tube. Observe and record what happens and note any changes.
10. Add about 5mL of 1M sodium sulfite (Na_2SO_3) solution to a clean, dry test tube. To this solution, add about 1mL of 6 M HCl. Observe the color odor given off by the mixture by wafting some of the gas toward you nose. **DO NOT SMELL THE GAS DIRECTLY.**

Observations and Data

Sample	Before reaction	After reaction
A. Synthesis 1. Cu 2. Mg		
B. Decomposition 3. $CuCO_3$		
C. Single Replacement 4. Zn + HCl 5. Zn + $CuSO_4$		
D. Double Replacement 6. $Zn(C_2H_3O_2)_2$ + Na_3PO_4 7. Na_2SO_3 + HCl		

Calculations

Balance each of the equations by inserting the proper coefficients where needed. Write the names of the reactant$_{(s)}$ and product$_{(s)}$ in the space provided below the molecular equation for the reaction.

A. Synthesis

1. $Cu_{(s)} + O_{2(g)} \rightarrow CuO_{(s)}$

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

B. Decomposition

3. $CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$

C. Single Replacement

4. $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

5. $Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

D. Double Replacement

6. $Zn(C_2H_3O_2)_{2(aq)} + Na_3PO_{4(aq)} \rightarrow NaC_2H_3O_{2(aq)} + Zn_3(PO_4)_{2(s)}$

7. $Na_2SO_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + SO_{2(g)}$

Conclusions and Questions

1. In this experiment, what method did you use to test for the presence of carbon dioxide?

Name another test for CO_2. Write a balanced equation for this second test.

2. What test was used to check for the presence of hydrogen gas? Write a balanced equation to represent this test.

3. Balance the equation below and identify the type of reaction represented by each equation.

a. $AgNO_{3(aq)} + Cu_{(s)} \rightarrow Cu(NO_3)_2 + Ag_{(s)}$

b. $BaCl_{2(aq)} + Na_2SO_4 \rightarrow BaSO_{4(s)} + NaCl_{(aq)}$

c. $KClO_{3(s)} \rightarrow KCl_{(s)} + O_{2(g)}$

d. $H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$

e. $AlCl_{3(aq)} + NH_4OH_{(aq)} \rightarrow NH_4Cl_{(aq)} + Al(OH)_{3(s)}$

f. $Cl_{2(g)} + NaBr_{(aq)} \rightarrow NaCl_{(aq)} + Br_{2(l)}$

Name_____ Date_____

29 Conservation of Matter

Introduction

Matter cannot be created or destroyed by a chemical change. This principle is known as the Law of Conservation of Matter. (*This law does not apply to nuclear reactions, where matter may be converted into energy.*) During a chemical reaction, the atoms of one or more substances (reactants) undergo rearrangements, which result in the formation of new, different substances (products). Despite the rearrangement, all of the original atoms are present; they are simply in a different form. Because no matter in a chemical reaction has been created nor destroyed, the total mass of all the reactants *must equal* the total mass of all the products. Because of the Law of Conservation of Matter, we can write balanced chemical equations. Such equations make it possible to predict the masses of reactants and products that will be involved in a chemical reaction.

In this experiment, aqueous solutions of three different compounds will be used to produce two separate and distinct chemical reactions. The fact that change occurs during each reaction will be observable.

Balanced chemical equations for the two reactions are:

$$Na_2CO_{3(aq)} + CaCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + CaCO_{3(s)}$$

$$CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_2O + CO_{2(g)}$$

Introductory Questions

1. What does it mean to say a compound is *conserved* in an experiment?

2. A student mixes two chemicals together and they react together. The mass of the products is less than the mass of the reactants. What could be the possible causes for this difference?

Problem

Will the mass of the products in a chemical reaction equal the mass of the reactants?

Equipment

Laboratory balance
Erlenmeyer flask, 125-mL
Rubber stopper (for flask)
Graduated cylinder
test tubes, 13x100-mm (2)
corks (to fit test tubes) (2)
safety glasses
marking pencils

Materials

1 M aqueous solutions of:

Na_2CO_3
$CaCl_2$
H_2SO_4

Procedure

1. In a graduated cylinder, measure exactly 10.0mL of sodium carbonate (Na_2CO_3) solution. Pour into a clean and dry 125-mL Erlenmeyer flask. Stopper the flask. Rinse and dry the graduated cylinder.
2. Measure exactly 3.0mL of 1 M calcium chloride ($CaCl_2$) solution and pour into a clean, dry test tube. Cork and label the tube. Rinse and dry the graduated cylinder.
3. Repeat step 2 with 3.0 mL of 1.0 M sulfuric acid (H_2SO_4) solution. **Caution:** ***Handle this acid with care.***
4. Place the stoppered flask and the corked test tubes together into a beaker. Place all three on the pan of the triple beam balance. Measure and record the combined masses of the test tubes, solutions, stoppers, and beaker. This will be mass *A* in your data table.
5. Remove the flask and the test tube containing the $CaCl_2$ solution from the balance. Pour the $CaCl_2$ into the Na_2CO_3 solution in the flask. Swirl the flask to thoroughly mix the two solutions. Record your observations in the space provided.
6. Replace the stopper and cork in their proper containers. Once again, measure the combined mass of the three containers, stoppers and contents. Record this data as mass *B*.
7. Remove the flask and the test tube containing sulfuric acid from the balance. ***Carefully*** pour the acid into the flask. **With the stopper off**, swirl the flask until the bubbling stops. Record your observations. Allow the flask to cool to room temperature.
8. Replace the stopper and cork. Remeasure the mass of the three containers, stoppers, and contents. Record this as mass *C* in the data table.

Observations and Data

Qualitative Observations

Step 5

Step 7

Quantitative Data

Mass A (before mixing)	grams
Mass B (after mixing Na_2CO_3 and $CaCl_2$)	grams
Mass C (after mixing H_2SO_4 with contents from step 8)	grams

Conclusions and Questions

1. What indications that a chemical reaction was taking place did you observe in step 5? In step 7?

2. Why are you instructed to leave the flask unstoppered after adding the sulfuric acid solution in step 7?

3. Compare masses A, B, and C. Do the masses of the reactants equal the masses of the product?

4. In your opinion, does this investigation verify the Law of Conservation of Matter? Explain.

5. How might this experiment be improved to bring the results in line with the Law of Conservation of Matter?

6. Discuss how the Law of Conservation of Matter relates to the balancing of chemical equations.

7. Here is an equation: $Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$
If the mass of the Zn is 3.50g, the mass of the $CuSO_4$ is 25.0 g, and the mass of the $ZnSO_4$ is 18.5 g, what would you predict the mass of the $Cu_{(s)}$ to be?

8. State the proper procedure when mixing acid and water.

Name_____ Date_____

30 Relating Moles to Coefficients of a Chemical Equation

Introduction

The mole is defined as Avogadro's number (6.02×10^{23}) particles. These particles may be atoms, molecules, formula units, ions, electrons, etc. The concept of the mole is very important, especially when dealing with quantitative studies of chemical reactions. When calculating quantities of solids and liquids, *molar masses* are used. The molar mass of a substance is the mass, in grams, of one mole of particles of that substance. When calculating quantities of gases, molar volumes are used. The *molar volume* is the volume occupied by one mole of a gas at STP.

Chemical reactions are represented by balancing chemical equations. Proper interpretation of an equation provides a great deal of information about the reaction it represents and about the substances involved in the reaction. For example, the coefficients in a balanced equation indicate the number of moles of each substance. Thus, the ratio of moles of a substance to moles of any other substance in the reaction can be determined at a glance.

In this experiment, iron filings will be added to an aqueous solution of copper (II) sulfate. A single replacement reaction will take place; the products will be iron (II) sulfate and copper metal. The balanced equation for this reaction is:

$$Fe_{(s)} + CuSO4_{(aq)} \rightarrow FeSO_{4\,(aq)} + Cu_{(s)}$$

The quantities of iron and copper sulfate used as reactants will be such that the copper sulfate will be in excess. Thus, the iron will be the limiting factor in determining the number of moles of products that will be formed. As the balanced equation indicates, the number of moles of copper produced should be equal to the number of moles of iron reacted.

Introductory Questions

1. Using the letters A, B, and C, write an equation to show a single replacement reaction.

2. You mix two chemicals (A and B) together in a beaker and get a reaction. You get your expected product (C), but some of one of the original chemicals (B) remains in the beaker unreacted. Explain what happened using the terms "in excess" and "limiting factor" and the letters A, B, and C.

3. Here is an equation for a chemical reaction.

$$Na_2CO_{3\,(aq)} + CaCl2_{\,(aq)} \rightarrow 2NaCl_{\,(aq)} + CaCO3_{(s)}$$

If you begin with 2 moles of $Na_2CO_{3\,(aq)}$, (assuming adequate amount of $CaCl_2$), how many moles of $NaCl_{\,(aq)}$ will be produced?

If you begin with 0.5 moles of Na2CO3, (assuming adequate amount of $CaCl_2$), how many moles of $NaCl_{\,(aq)}$ will be produced?

Problem

Find the ratio of moles of reactant to moles of product in a chemical reaction.

Equipment

Balance
Bunsen burner
Beaker, 100-mL
Beaker, 250-mL
Graduated cylinder, 100-mL
ring stand
iron ring
wire gauze
glass stirring rod

Materials

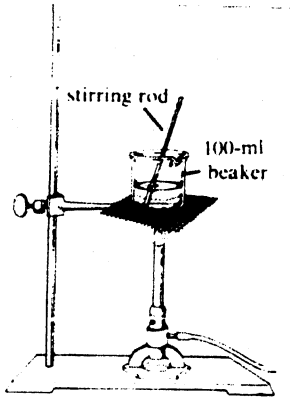
Copper sulfate crystals ($CuSO_4$)
Iron filing (Fe)

Procedure

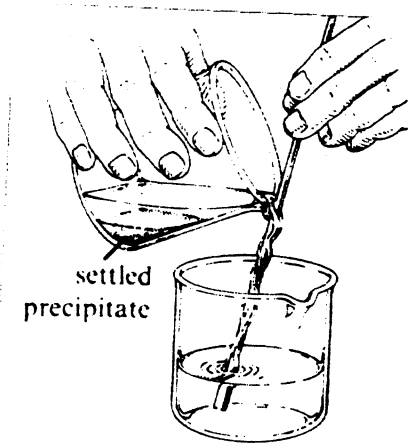
1. Find the mass of a clean, dry 100-mL beaker. Record this as (a) in your data table.
2. Measure out 8.0 grams of copper sulfate crystals ($CuSO_4$) and add these to the beaker.
3. Measure 50.0 mL of water in a graduated cylinder and add it to the crystals in the beaker.

While one partner continues steps 4 and 5, the other partner should carry out the instructions in step 6.

4. Set up the apparatus as shown below. Heat the liquid in the mixture in the beaker to just *below* boiling. **DO NOT ALLOW THE LIQUID TO BOIL.**



5. Continue heating and stir the mixture until the crystals are completely dissolved. Turn off the gas and remove the burner.
6. Using the balance, measure precisely 2.24 grams of iron filings. (*Remember: do not place any reagent directly on the balance pan. Use weighing paper.*) Record the mass as (b) in your data table.
7. Add the iron filings, *small amounts at a time*, to the hot copper sulfate solution. Stir continuously. After all the iron has been added and the mixture stirred, allow the beaker to sit for 10 minutes while the reaction proceeds. Record your observations as (d) in the data table.
8. Decant the liquid into a 250-mL beaker as shown below. Do not disturb the solid as the bottom of the beaker.



9. Now you must wash the solid with water to remove any contaminants. Add about 10 mL of water to the solid in the 100-mL beaker. Stir vigorously. Let the solid settle and decant the liquid. Repeat this washing.
10. Spread the solid out on the bottom of the beaker and place the beaker in a drawer or oven to dry. Complete step 11 and the rest of this experiment at the beginning of the next lab period.
11. Find the mass of the beaker and the dry copper metal. Record this as (c) in the data table.

Observations

(a) Mass of empty beaker g

(b) Mass of iron filings g

(c) Mass of beaker + copper g

(d) Visual observations

Calculations

Use the following formula to convert moles to mass or vice versa.

$$\text{Number of moles} = \frac{\text{Mass}_{(g)}}{\text{Mass per Mole}}$$

1. Find the mass of the copper produced = (c) - (a) g

2. Find the number of moles of copper produced in this experiment. (Use the factor label method.)

3. Find the number of moles of iron reacted in this experiment.

4. Find the whole number ratio of moles of iron to moles of copper in this experiment.

Conclusions and Questions

1. How does the ratio found in calculation 4 above compare with the ratio of coefficients of the same two metals in the balanced equation for this reaction?

2. How many moles of copper sulfate are used to produce the solution in this experiment? Why is this amount of copper sulfate said to be "in excess".

3. Explain why the iron is the limiting factor in this experiment.

4. A general description of the single replacement reaction in this experiment is:

metal + salt in solution yields "new" metal + "new" salt solution.

Give a balanced equation for another example (besides the one in this experiment) of this type of single replacement reaction.

5. Write a balanced equation for this reaction:

Metallic copper plus aqueous silver nitrate yields metallic silver plus aqueous copper nitrate

If 0.25 moles of copper reacts, how many moles of silver metal will be produced?

Name_____ Date_____

31 Mole and mass relationships

Introduction

In a balanced chemical equation, all reactants and products must be represented by symbols or formulas. The total number of atoms of each element must be the same on each side of the equation.

The coefficients of each substance represent the number of moles of each substance. The coefficients for the equation tell the proportions in which these substances must react.

In this experiment, you will investigate the quantitative relationships in the reaction:

$$NaHCO_{3(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + CO_{2(g)} + H_2O_{(g)}$$

A known mass of sodium hydrogen carbonate will be reacted with excess hydrochloric acid. Knowing the mass of $NaHCO_{3(s)}$ that reacts, we can determine from the balanced equation the mass of NaCl that should be produced. We can compare this theoretical value with the actual experimental mass of NaCl produced.

Introductory questions

1. How does balancing an equation satisfy the Law of Conservation of Mass?

Problem

How does the mass of a product compare to the theoretical mass?

Equipment

Balance

Burner	wire gauze
Evaporating dish	ring stand
Watch glass	iron ring
Microspatula	safety glasses

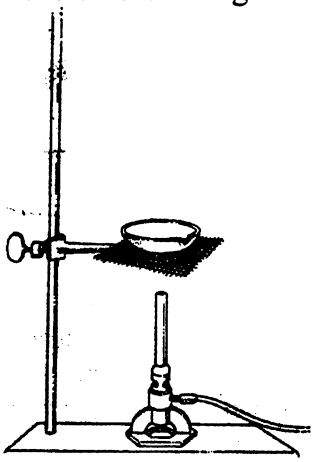
Materials

6 M hydrochloric acid (HCl) in dropper bottle
sodium hydrogen carbonate ($NaHCO_3$)

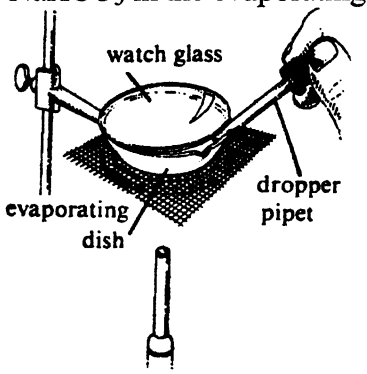
Procedure

Fire **Caustic Substance**

1. Flame dry a clean evaporating dish by heating it in the hot part of the flame for about 5 minutes. Allow the dish to cool.
2. Find the mass of the evaporating dish. Record this mass.
3. Leave the dish on the balance as you move the riders to measure an additional 2.50 g. Add $NaHCO_3$ to the dish, using your microspatula, until the scale is balanced. Record this mass.
4. Set up the ring stand as shown in the diagram.



5. Place the watch glass on top of the evaporating dish and place the dish on the wire gauze.
6. Slowly add HCl to the $NaHCO_3$ in the evaporating dish a few drops at a time.



7. Continue adding HCl, a few drops at a time, until the bubbling stops.
8. Carefully tilt the dish to assure the acid has come in contact with all of the $NaHCO_3$.
9. When you are sure that all of the $NaHCO_3$ has reacted, remove the watch glass and turn it curved side up on the lab bench. Do not put it back on the evaporating dish.
10. With the evaporating dish still on the wire gauze, hold the burner in your hand and gently heat the dish. Move the burner back and forth to avoid splattering.
11. When the substance is dry, allow the dish to cool.
12. Find the mass of the evaporating dish and contents (NaCl). Record this mass.

Observations

	mass (grams)
a) evaporating dish	
b) evaporating dish + $NaHCO_3$	
c) evaporating dish + NaCl	

Calculations

1. Find the mass of the $NaHCO_3$ reactant, (b-a)

2. Find the mass of the NaCl product, (c-a)

Conclusions and Questions

1. In the chemical reaction $CaCO_3 \rightarrow CaO + CO_2$, if 40.0 g of $CaCO_3$ are decomposed:
   a. How many grams of CaO are produced?

   b. How many grams of CO_2 are produced?

2. In the reaction $N_2 + 3H_2 \rightarrow 2NH_3$, if 20.0 grams of hydrogen react:
   a. How many grams of ammonia are produced?

   b. How many grams of nitrogen react?

3. According to the balanced equation for the reaction used in this experiment, what is the ratio of moles of $NaHCO_3$ reacted to moles of NaCl produced?

4. How many moles of $NaHCO_3$ are reacted in this experiment?

5. How many moles of NaCl are produced?

6. What is the ratio of moles $NaHCO_3$ reacted to moles NaCl produced?

7. How does this value compare with the mass attained experimentally?

Name_____ Date_____

32 Mass-Mass Relationships in Reactions

Introduction

As you have learned, given a balanced chemical equation and the mass of one of the substances in the reaction, the mass of any other substance in the reaction can be calculated. Calculations in which a known mass is used to find an unknown mass in a chemical reaction are called mass-mass calculations.

In this experiment, a double replacement reaction will occur when an aqueous solution of a hydrate of zinc acetate is mixed with an aqueous solution of a hydrate of sodium phosphate tribasic. There are two products of this reaction. One is the insoluble solid (zinc phosphate), which will precipitate out of solution. The other is a soluble salt (sodium acetate), which will remain in solution. The insoluble solid will be separated from the liquid, dried, and its mass determined. The value of the experimentally measured mass of the compound will be compared to the theoretical mass of the compound predicted by a mass-mass calculation.

You may recall that the hydrates of ionic substances have water molecules bound into their crystalline structure. However, they look and feel perfectly dry. When doing mass calculations that involve hydrates, you must be careful to include the mass of the water molecules in your calculations. For example, the mass of one mole of $CaBr_2 \cdot 6H_2O$ is 308g/mole.

Ca:	1 x 40. g/mol	= 40 g
2Br:	2 x 80. g/mol	= 160 g
$6H_2O$:	6 x 18.g/mol	= 108 g
	Total =	308 g/mol

Introductory Questions

1. Calculate the mass of one mole of zinc phosphate?

2. Here is an equation for a chemical reaction. $A + 2B \rightarrow C$. If the mass of A is 35.5 grams, and the mass of C is 68.9 grams, calculate the mass of B.

Problem

Can you calculate the theoretical mass of one of the products of a double replacement reaction and yield the same mass in an actual experiment?

Equipment

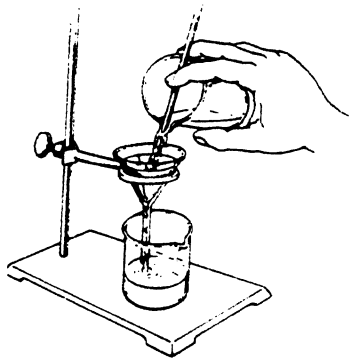
Balance
Graduated cylinder, 100-mL
Beakers, 250-mL (2)
Beaker, 100-mL
glass stirring rod
iron ring
funnel
safety glasses

Materials

Zinc acetate hydrate ($Zn(C_2H_3O_2)_2 \cdot 2H_2O$)
Sodium phosphate tribasic hydrate ($Na_3PO_4 \cdot 12H_2O$)
Distilled water

Procedure

1. Using the balance, measure out exactly 2.19 g of zinc acetate hydrate ($Zn(C_2H_3O_2)_2 \cdot 2H_2O$). Record this mass as (a) in your data table.
2. Place the $Zn(C_2H_3O_2)_2 \cdot 2H_2O$ in a clean 250-mL beaker and add 50 mL of water. Stir thoroughly to make sure all crystals are dissolved. Rinse off the stirring rod.
3. Measure out approximately 2.70 g of sodium phosphate tribasic hydrate ($Na_3PO_4 \cdot 12H_2O$). Place it in a clean 100-mL beaker and add 50.mL of water. Stir until the crystals are dissolved.
4. Pour the $Na_3PO_4 \cdot 12H_2O$ solution into the 250-mL beaker containing the solution of $Zn(C_2H_3O_2)_2 \cdot 2H_2O$. Stir. Record your observations as (d) in your data table.
5. Find the mass of a piece of filter paper. Record this as mass (b). Fold the filter paper and shown by your instructor and place it in the funnel as shown.



6. Pour the mixture from the 250-mL beaker into the funnel as shown. Pour slowly. Do not allow the liquid to rise above the edge of the filter paper in the funnel.
7. Rinse the beaker with about 20mL of water. Pour the rinse water through the filter. Repeat the rinsings and filterings until all the precipitate is out of the beaker.
8. Wash the precipitate by pouring about 10mL of clean water through the filter.
9. Remove the filter paper and precipitate from the funnel and place overnight in an oven to dry at about 45° C.
10. Find the mass of the dry precipitate + filter paper (mass (c) in your data table.)

Observations and Data

(a) mass of $Zn(C_2H_3O_2)_2 \cdot 2H_2O$ g

(b) mass of filter paper g

(c) mass of filter paper + precipitate g

(d) Observations

Calculations

1. Write a balanced equation for the double replacement reaction. (the solid formed by the reaction that you dried in the oven is zinc phosphate hydrate. It's formula is $Zn_3(PO_4)_2 \cdot 4H_2O$.

$$__Na_3PO_4 \cdot 12H_2O_{(aq)} + __ Zn(C_2H_3O_2)_2 \cdot 2H_2O \rightarrow$$
$$Zn_3(PO_4)_2 \cdot 4H_2O_{(s)} + __NaC_2H_3O_{2(aq)} + __ H_2O$$

2. Using mass-mass calculations, find the theoretical mass of the $Zn_3(PO_4)_2 \cdot 4H_2O_{(s)}$ precipitate that should be produced when 2.19 g of $Zn(C_2H_3O_2)_2 \cdot 2H_2O$ react completely.

3. Find the experimental mass of $Zn_3(PO_4)_2 \cdot 4H_2O$ formed, (c) – (a) g

4. Find your experimental error using the formula:

$$\%error = \frac{(observed\ value\ -\ true\ value)}{true\ value} \times 100$$

$$\frac{1.57 - 1.53}{1.53} \times 100 = 2.61\ \%$$

Conclusions and Questions

1. What is another name for a double replacement reaction?

2. Give a brief description of a double replacement reaction. What must one of the products of such a reaction be

3. Suggest some possible sources of error in this experiment.

4. Define the terms filtrate and precipitate.