Grossmont College Chemistry 120 Laboratory Manual

7th Edition

Grossmont College
Periodic Table of the Elements

	1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**
	1A	**2A**											**3A**	**4A**	**5A**	**6A**	**7A**	**8A**
1	**1** H 1.008																	**2** He 4.003
2	**3** Li 6.941	**4** Be 9.012											**5** B 10.811	**6** C 12.011	**7** N 14.007	**8** O 15.999	**9** F 18.998	**10** Ne 20.180
3	**11** Na 22.990	**12** Mg 24.305	**3B**	**4B**	**5B**	**6B**	**7B**	**8B**	**8B**	**8B**	**1B**	**2B**	**13** Al 26.982	**14** Si 28.086	**15** P 30.974	**16** S 32.061	**17** Cl 35.453	**18** Ar 39.948
4	**19** K 39.098	**20** Ca 40.078	**21** Sc 44.956	**22** Ti 47.867	**23** V 50.942	**24** Cr 51.996	**25** Mn 54.938	**26** Fe 55.845	**27** Co 58.933	**28** Ni 58.693	**29** Cu 63.546	**30** Zn 65.382	**31** Ga 69.723	**32** Ge 72.640	**33** As 74.922	**34** Se 78.960	**35** Br 79.904	**36** Kr 83.798
5	**37** Rb 85.468	**38** Sr 87.62	**39** Y 88.906	**40** Zr 91.224	**41** Nb 92.906	**42** Mo 95.962	**43** Tc (98)	**44** Ru 101.07	**45** Rh 102.906	**46** Pd 106.42	**47** Ag 107.868	**48** Cd 112.411	**49** In 114.818	**50** Sn 118.710	**51** Sb 121.760	**52** Te 127.60	**53** I 126.905	**54** Xe 131.293
6	**55** Cs 132.905	**56** Ba 137.327	**57** La 138.906	**72** Hf 178.49	**73** Ta 180.948	**74** W 183.84	**75** Re 186.207	**76** Os 190.23	**77** Ir 192.217	**78** Pt 195.078	**79** Au 196.967	**80** Hg 200.59	**81** Tl 204.383	**82** Pb 207.2	**83** Bi 208.980	**84** Po (209)	**85** At (210)	**86** Rn (222)
7	**87** Fr (223)	**88** Ra (226)	**89** Ac (227)	**104** Rf (261)	**105** Db (262)	**106** Sg (266)	**107** Bh (264)	**108** Hs (269)	**109** Mt (268)	**110** Ds (281)	**111** Rg (272)	**112** Cn (285)	**113** Nh (284)	**114** Fl (289)	**115** Mc (288)	**116** Lv (293)	**117** Ts (294)	**118** Og (294)

6	**58** Ce 140.116	**59** Pr 140.908	**60** Nd 144.24	**61** Pm (145)	**62** Sm 150.36	**63** Eu 151.964	**64** Gd 157.25	**65** Tb 158.925	**66** Dy 162.500	**67** Ho 164.930	**68** Er 167.259	**69** Tm 168.934	**70** Yb 173.04	**71** Lu 174.967
7	**90** Th 232.038	**91** Pa 231.036	**92** U 238.029	**93** Np (237)	**94** Pu (244)	**95** Am (243)	**96** Cm (247)	**97** Bk (247)	**98** Cf (251)	**99** Es (252)	**100** Fm (257)	**101** Md (258)	**102** No (259)	**103** Lr (262)

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Chemistry 120 Laboratory Manual

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Contents

EXPERIMENT 1

Welcome to Lab 1

EXPERIMENT 2

Measuring, Significant Figures, and Glass Working 15

EXPERIMENT 3

Mass, Volume, and Density 33

EXPERIMENT 7

Estimating Avogadro's Number ***127***

EXPERIMENT 8

Double Displacement Reactions ***139***

EXPERIMENT 9

Chemical Reactions ***159***

EXPERIMENT 10

Green Limiting Reagent ***177***

EXPERIMENT 11

Measurement of the Gas Constant and Molar Volume of Oxygen Gas ***187***

EXPERIMENT 12

EXPERIMENT 1

Welcome to Lab

Greetings

Welcome to the wonderful world of chemistry lab.

Purpose

The purpose of lab is to teach you the necessary skills to work in a laboratory and to illustrate the concepts that are covered in lecture. Your laboratory work is the core of your chemistry course. You have a challenging opportunity to observe many of the facts of chemistry under controlled laboratory conditions and to deduce from these observations the basic principles that constitute the foundation of the science of chemistry. This is the scientific method in action.

The purpose of lab reports in preparation for general chemistry is to teach you to organize and analyze data. This will prepare you to write full formal lab reports in general chemistry, organic chemistry, and beyond in your career as a scientist.

Skills

These assignments will help you to practice the following skills that are essential to your success in this course, subsequent courses, and professional life:

- Gain self-confidence by working individually or learn to work collaboratively in a team.
- Learn about time management before (completing both the ***Pre-Lab Preparation*** and the ***Experimental Procedure***), during an experiment (completing ***Report Pages*** for the experiment in the alloted time), and after the experiment (completing ***Data Anaylsis*** and the ***Post Lab Questions).***
- Use qualitative and quantitative observations to gain a hands on knowledge of the concepts being taught in lecture.
- Use your ingenuity and common sense to observe and collect data.
- Critical thinking and experimental interpretation using graphs, calculations, and short answer questions.

Knowledge

You will gain a lot of knowledge by working in the chemistry laboratory. Lab time is limited, so don't waste it! If you don't understand how to do something ask your instructor. Note beforehand any extra equipment required from the stockroom, and obtain all of it at the same time. These assignments will also help you to become familiar with the following important content information.

- Completing a lab experiment.
- Learning laboratory skills such as how to properly work with glassware and equipment.
- Analyzing qualitative and quantitative data
- Completing informal lab reports.

Tasks:

This exercise asks you to collect, organize, analyze, and evaluate data in the lab. You will fill out the lab report pages that include prelab, procedure summary, results and calculations, spreadsheet, and post lab questions.

- Laboratory directions, while quite specific, leave ample opportunity for clear-cut, logical, original, and imaginative thinking. This attitude is a prerequisite in any scientific endeavor.
- Prepare your ***Laboratory Report*** on each experiment with care. If you prepare your own report form, use a permanently bound notebook as prescribed by your instructor, preferably one with 5-mm cross rulings. Always have the data entry portion prepared in advance, and record data directly in your final report as you obtain it. (Data entered on scraps of paper will be confiscated.)
- Where calculations of data are involved, show an orderly calculation for the first set of data, but do not clutter the calculation section with arithmetic details.
- Likewise, think through and answer important questions that are intended to give you an understanding of the principles in which the experimental procedure is based as you perform the experiment. If you don't understand something ask your instructor.

Criteria for Success:

As scientists, we should strive for specificity and accuracy. As such, I encourage you to always come prepared to lab by reading each lab thoroughly before writing your procedure summary and completing the prelab. Make sure that you ask your instructor to clarify anything that you do not understand before, during, and after each experiment.

- Scientists learn much by discussion with one another. You may likewise profit by discussion with your classmates-but not by copying from them. The keystone of all science rests first on integrity. You will also profit by frequent reference to your text while working in the laboratory. (Books are generally even more reliable and complete sources of information than are your classmates!) Your instructor, other chemistry instructors, and peer tutors are also great resources.
- For tabular data on the properties of substances, consult one of your handbooks: the Handbook of Chemistry and Physics (CRC Press, Inc., Boca Raton, Florida) or Lange's Handbook of Chemistry (McGraw-Hill, New York).

Safety Rules

Familiarize yourself with the safety rules given in the ***Introduction***. Obedience to these rules, as modified or added to by your instructor, is essential for the sake of your safety and that of others in the laboratory.

The American Chemical Society (ACS) recommends that you learn the acronym **RAMP** for the four principles of safety:

1. **R**ecognize the hazard.
2. **A**ssess the risks of the hazard.
3. **M**inimize the risk of the hazard.
4. **P**repare for emergencies.

Your instructor will indicate the location and show you the proper use of the fire extinguishers, safety fountains or showers, and first-aid cabinet and supplies. The instructor will also tell you where to obtain safety glasses or goggles. If you are ever unsure of something in the chemistry laboratory be sure to consult your instructor.

Laboratory Regulations

Familiarize yourself with the laboratory regulations in the ***Introduction***. It is essential that these regulations, as modified or added to by your instructor, be followed carefully.

Basic Laboratory Procedures

Check the equipment in the laboratory locker assigned to you, following the procedure described in the ***Introduction*** or as directed by your instructor. If time remains, or if your instructor so directs, read over the procedures for the handling of chemicals, use of laboratory burners, operations with glass tubing, care of laboratory glassware, and volumetric measurements of liquids.

Introduction

The chemistry laboratory can be a place of joy, discovery, and learning. It can also be a place of frustration and danger.

Although every effort has been made to eliminate the use of explosive, very toxic, or carcinogenic substances, there is an unavoidable hazard involved in the use of a variety of chemicals and glass apparatus. In experiments where a potential danger exists, you will find a ***Safety Precautions Section*** in the text of the experimental procedure. Read this section very carefully before you proceed with the experiment.

Your skin and the delicate membranes of your eyes, mouth, and lungs are largely made of protein. We hope that you do not have to experience firsthand the adverse effect that even 6 M solutions, not to mention concentrated solutions of acids and bases, have on protein. The eyes are especially sensitive to acids, bases, and oxidizing agents and must be protected at all times by wearing safety glasses or goggles whenever anyone is working with chemicals or flames in the laboratory. In addition, the open flame of a Bunsen burner presents a constant hazard to clothing and hair. It's likely that you will experience frustration if you come unprepared to the laboratory, neglect to record important data, or frantically try to write up reports an hour before they are due. You can minimize these problems by reading the experiments carefully beforehand, noting the critical data that must be recorded, and thoughtfully considering the data while you collect it in order to avoid careless blunders. We strongly advise you to learn and observe at all times the following laboratory rules and regulations. In doing so you will maximize the joy while minimizing the potential dangers and frustrations of laboratory work, and maintain a safe working environment.

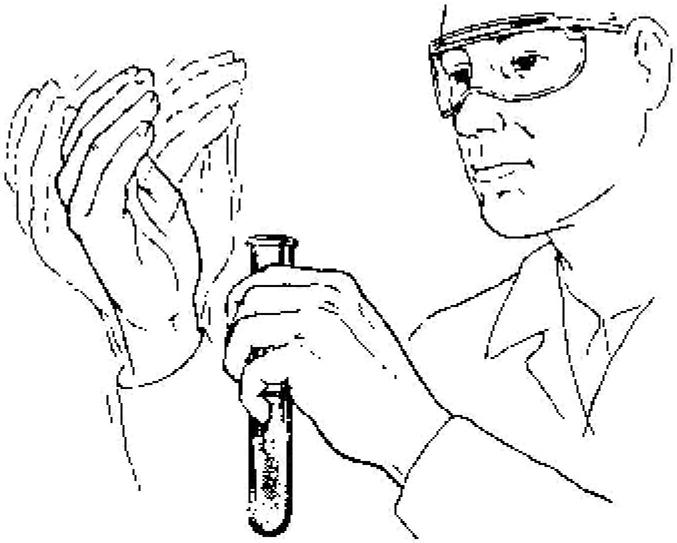
Safety Rules

These rules are designed to ensure that all work done in the laboratory will be safe for you and your fellow students. In addition to the rules listed here, your chemistry department has a set of rules that you will be asked to read and sign as evidence that you have been informed of proper laboratory safety procedures. Your laboratory instructor should also point out the location and demonstrate the use of various pieces of safety equipment such as fire extinguishers, fire blankets, safety showers, eye-wash fountains, and equipment for handling spills. Finally, you should know where first-aid supplies are kept and where a telephone is located for emergencies that require paramedical, fire, or police assistance. The telephone numbers of these on-or off-campus services should be posted in a prominent place.

1. The most important rule is that safety glasses with side panels or goggles must be worn at all times in the laboratory. Goggles provide the greatest safety. Ordinary prescription glasses sometimes cover only parts of the eyes and lack side panels that protect the wearer from chemical splashes that might get into the eyes from the side. For this reason, they should be covered by safety goggles. Contact lenses should not be worn in the laboratory, even under safety goggles. By themselves they offer no protection from splashes and they are considered unsafe even under safety goggles because various fumes for instance, hydrogen chloride gas-may accumulate under the lens and cause serious injuries. If you do not own a pair of safety glasses or goggles, obtain one from the stockroom or your campus bookstore. If any chemical comes in contact with the eye, the most effective first aid is the immediate flushing of the eye with copious amounts of tap water. You are seldom more than a few seconds from a faucet. Continue flushing for at least fifteen minutes and then consult a physician at once. If your laboratory is equipped with eye fountains, familiarize yourself with their use and their location. The same guidelines apply to chemicals that are not reactive with water on the skin.
2. Fire is a constantly present danger. Learn where the nearest fire extinguisher is and how to use it. Your laboratory should also be equipped with a safety shower. If your hair or clothing should catch fire, smother the fire (Stop, Drop and Roll) or douse yourself in the shower.
3. Minor burns, cuts, and scratches are fairly common injuries. However, you must report every such injury to your instructor, who will determine what first aid is appropriate. If you or another student must report to the infirmary or hospital, be certain that someone accompanies the injured person.
4. Bare feet are not allowed in the chemistry laboratory. Closed toe shoes must be worn when laboratory work is being done. Broken glass and spilled chemicals are all too common on the floor of chemistry labs. In addition, we recommend that bare legs, midriffs, and arms be covered with old clothing or, preferably, with a laboratory apron or coat; Clothed from shoulders to knees is advised.
5. Chemical substances should not be inhaled or smelled unless you are instructed to do so. When this is indicated, you must exercise great care in noting the odor of a substance, using your hand to waft its vapor gently toward your face. Whenever possible, avoid breathing fumes of any kind.

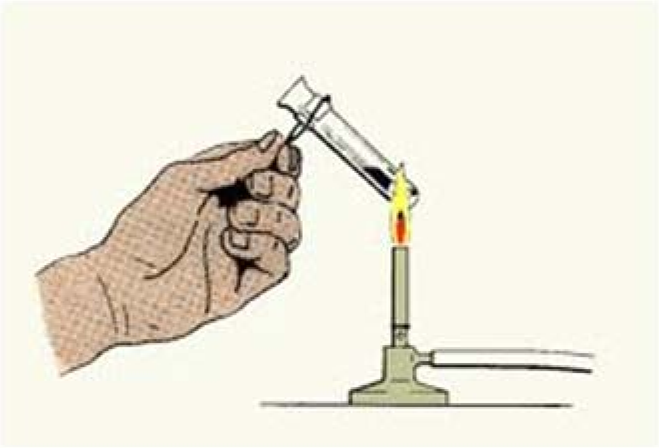
6. The vapors of a number of solutions are quite potent and can irritate or damage the mucous membranes of the nasal passages and the throat. Use the technique displayed in Figure 1.1 on page 5 when you need to note the odor of a gas or vapor.

FIGURE 1.1



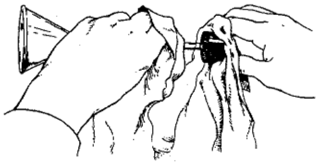
7. In many experiments, it is necessary to heat solutions in test tubes. Never apply heat to the bottom of the tube; always apply it to the point at which the solution is highest in the tube, working downward if necessary. Be extremely careful about the direction in which you point a tube; a suddenly formed bubble of vapor may eject the contents violently (an occurrence called bumping). Indeed, a test tube can become miniature cannon. See Figure 1.2 on page 5 for the proper way to heat contents in a test tube. NEVER point a test tube that is being heated at your neighbor, it may bump and eject its contents. SAFETY GOGGLES worn regularly in the laboratory will protect your eyesight.

FIGURE 1.2 **Do not heat the bottom of the test tube.**



8. Avoid tasting anything in the laboratory. Poisonous substances are not always so labeled in the laboratory. Do not use the laboratory as a play place and do not eat or drink from laboratory glassware.
9. Perform no unauthorized experiments.
10. Never work in the laboratory alone.
11. Beware of hot glass tubing-it looks cool long before it can be handled safely.
12. For reactions involving poisonous gases, use the fume hood, which provides suction to remove such gases or vapors.
13. Neutralize spilled acid or base as follows:
    a. Acid on clothing, use dilute sodium bicarbonate solution.
    b. Base on clothing, use boric acid solution (50 g/L).
    c. Acid or base on the desk, use solid sodium bicarbonate for either, followed by water.
14. To insert glass tubing (including thermometers, long-stemmed funnels, thistle tubes, etc.) through a rubber stopper, first lubricate the tube and stopper with water or *glycerol.* Hold the tubing with a cloth *near the end to be inserted,* and insert with a twisting motion. If you twist a long-stemmed funnel by the funnel end it is easily broken.

FIGURE 1.3



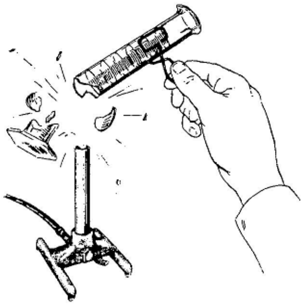
Laboratory Regulations

These regulations are designed to guide you in avoiding unnecessary hazards, and you will also be developing efficient laboratory techniques to obtain reliable, useful data.

1. You must read each experiment thoroughly before entering the lab. If you do not, you will waste a great deal of time (both your own and your instructor’s), and you may expose yourself and others to unnecessary hazards. (You will also routinely fail all pre-lab quizzes if your instructor chooses to use them.)
2. Discard solids into the waste crocks. Never throw solutions, litmus, or any insoluble solids into the sink. When directed wash down liquids into the drain with lots of water; otherwise liquid waste should go into the hazardous waste container. Acids and salts of copper, silver, and mercury are corrosive to lead plumbing.
3. Leave reagent bottles where you found them. Do not take them to your bench. Bring test tubes or beakers to the shelf to obtain chemicals.
4. Read the label twice before taking anything from a bottle.
5. Avoid using excessive amounts of reagent. One to five (1-5) mL is usually ample for test tube reactions.
6. Never return unused chemicals to the stock bottle. You may make a mistake from which other students' experiments will suffer.
7. Do not insert your own pipets, medicine droppers, scoopulas, spatulas, or any utensil into the reagent bottles. Avoid contamination of the stock solution by pouring the solution from the bottle into your clean beaker or other receptacle.

8. Do not lay down the stopper of a bottle. Impurities may be picked up and thus contaminate the reagent when the stopper is returned. Your instructor will demonstrate how to hold reagent bottle caps and stoppers between your fingers.
9. Do not heat thick glassware such as volumetric flasks, graduated cylinders, or bottles; they break easily and heating distorts the glass so that the calibrations are no longer valid. Test tubes may break if they are heated above the liquid level and liquid is then splashed over the hot glass. Evaporating dishes and crucibles may be heated until red hot. Avoid heating any apparatus too suddenly; move the flame intermittently at first.

FIGURE 1.4



Basic Laboratory Equipment

The Laboratory Locker

Check the equipment in the locker assigned you. Refer to Figure 1.5 on page 7 and Figure 1.9 on page 10 for the identification of any unfamiliar items. Ascertain that all items are present, and examine them carefully to be sure they are in an acceptable condition. You are responsible for this equipment and will be charged for any breakage or shortage at the conclusion of the course.

FIGURE 1.5 **Graduated Cylinders, Beakers and Erlenmeyer Flasks:**

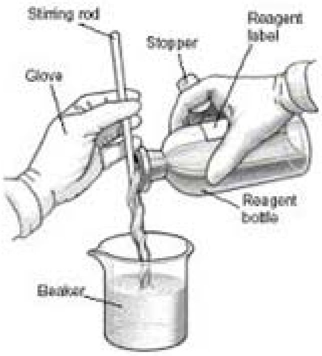
The Handling of Chemicals

Some suggestions bare repeating:

1. Be considerate of others by always bringing your container to the reagent shelf to obtain a chemical. Do not take the bottle to your desk.
2. Maintain the purity of the chemicals in the reagent bottles. Do not withdraw more than you need, and never return any chemical to the bottle. Never contaminate the stopper by laying it down; hold it between your fingers.
3. Do not insert your own medicine dropper into a reagent bottle or the medicine dropper from a reagent bottle down into your own test tube or solutions.
4. If necessary, clean the outside of the reagent bottle of accumulated dust, or other contaminant, rinse the neck and stopper with distilled water, and wipe dry before removing the stopper. Careful observance of these suggestions will prevent the contamination of the stock bottles.
5. If you do spill any chemical, clean it up completely, at once. A dirty laboratory prevents good work from being done.

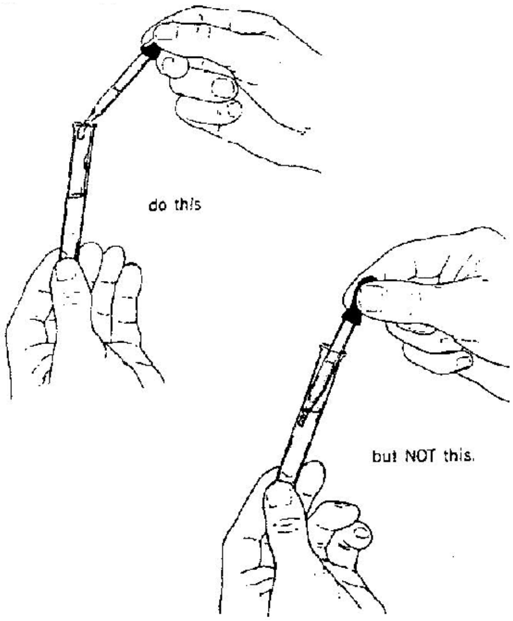
Figure 1.6 on page 8 illustrates one way to dispense of a liquid in the lab.“Glass-to-glass-to-glass” involves pouring an aqueous solution (i.e. a solution where water is the solvent) from a glass (or plastic) container down a glass stirring rod into another container. Doing so will help you to ensure complete transfer of the solution. Your instructor will demonstrate the proper way in which to dispense of reagents.

FIGURE 1.6 **Glass-to-glass-to-glass:**



When putting the contents of a medicine dropper into a test tube, be sure to touch the dropper to the sides of the test tube and allow the liquid to sheet down the test tube. Do not put the medicine dropper into the liquid in the test tube. You do not want to risk contaminating the dropper or liquid.

FIGURE 1.7 **Proper use of a medicine dropper**



Bunsen Burners

The Bunsen burner, used for most laboratory heating, produces a cone-shaped flame, as illustrated in Figure 1.8. Ordinary beakers, crucibles, and other objects to be heated are placed just above the hottest portion of the flame, which is thus allowed to spread about them. If placed down in the cold inner cone of the flame, which consists of unburned gas, the objects are not heated effectively.

FIGURE 1.8 **The various parts of a Bunsen Burner**

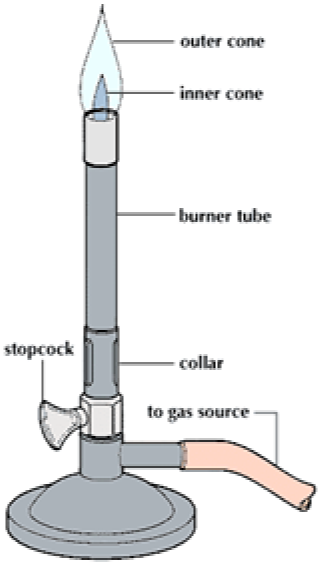
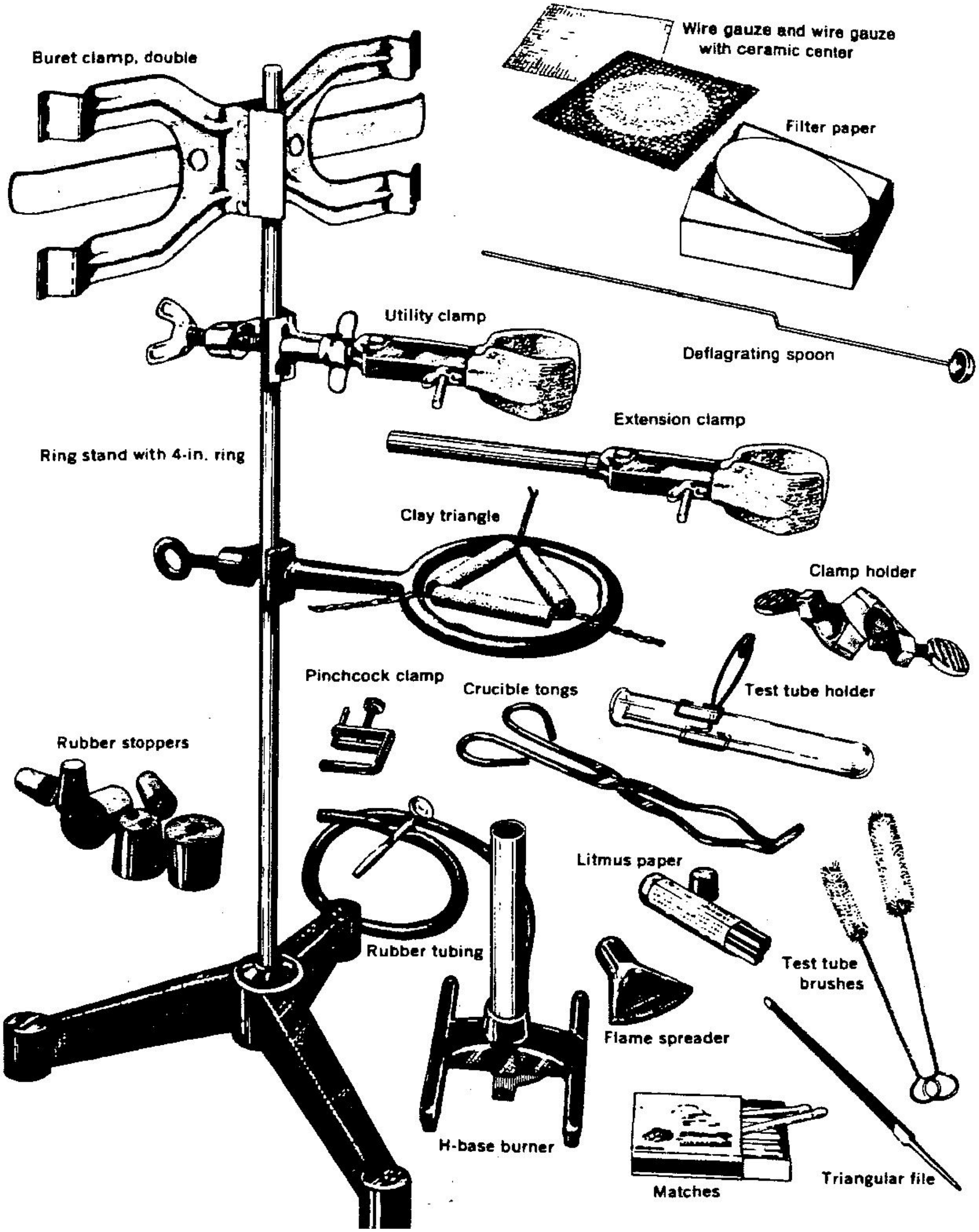
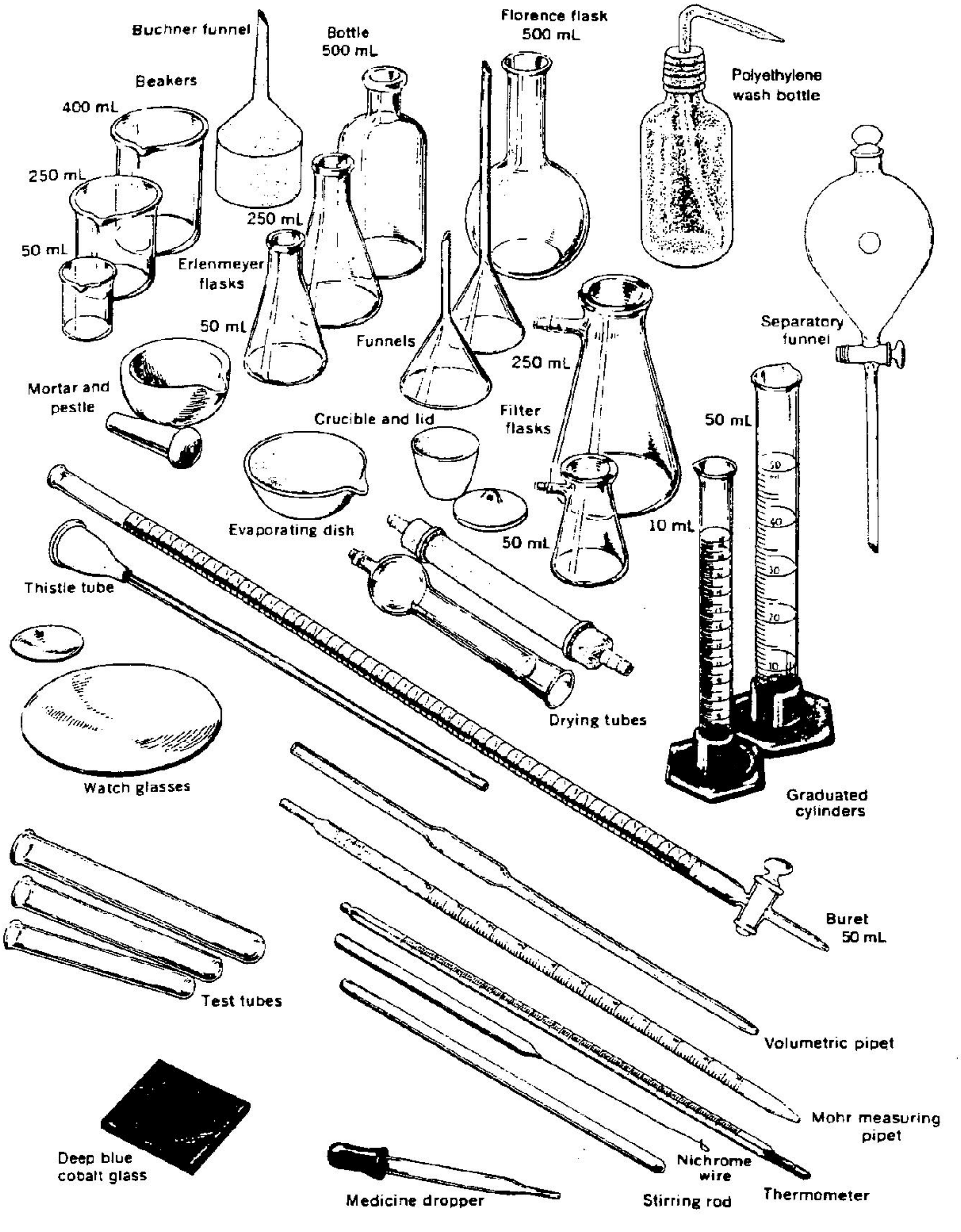


FIGURE 1.9 **Common Items Found in a Chemistry Locker**



[From J. W. Hagen, *Empirical Chemistry*, W. H. Freeman and Company, San Francisco. Copyright 1972.]

FIGURE 1.10 Common Items Found in a Chemistry Locker continued



Operations with Glass Tubing

Glass is not a true crystalline solid and therefore does not have a sharp melting point. In this respect, it more nearly resembles a solid solution or an extremely viscous liquid that gradually softens when heated. It is this property that makes glass working possible.

Soda-lime glass, made by heating a mixture of soda (Na_2CO_3), limestone ($CaCO_3$), and silica (SiO_2), softens readily at about 300 °C to 400 °C in the burner flame. Tubing of this glass is easily bent, but because of its high temperature coefficient of expansion it must be heated and cooled gradually to avoid undue strain or breakage. Annealing by a mild reheating and uniform cooling is often wise. Such glass must not be laid on a cold surface while it is hot, since this introduces strains and causes breakage. Borosilicate glass (such as Pyrex or Kimax) does not soften much below 700 °C to 800 °C and must be worked in an oxygen-natural-gas flame or blowtorch. Because it has a low temperature coefficient of expansion, objects made of it can withstand sudden temperature changes.

Figure 1.11, below, shows the proper way to cut glass tubing and fire polish the ends. Fire polishing smooths the sharp edges so that the glass tubing can be inserted easily into a rubber stopper without cutting your fingers. Figure 1.11 also illustrates the proper technique used to make a bend or a constricted tip. Your instructor will demonstrate making various glass tubing pieces.

FIGURE 1.11

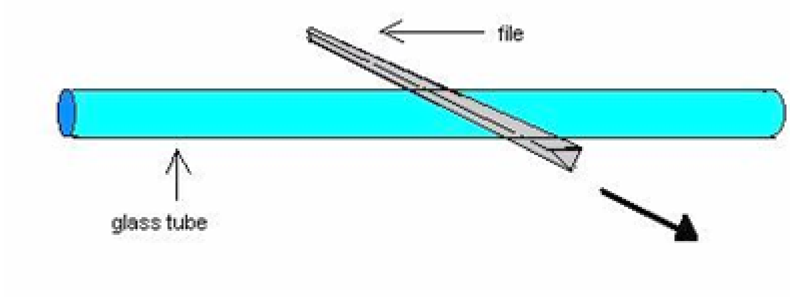
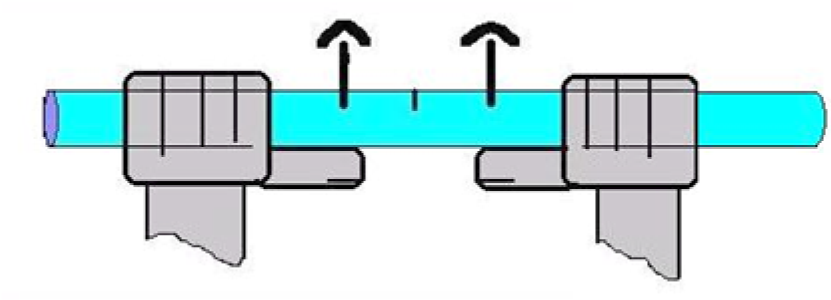


FIGURE 1.12 **Break tube away from you.**



Be careful not to cut yourself! You may use a paper towel to protect your hands.

FIGURE 1.13 Bending glass tubing:

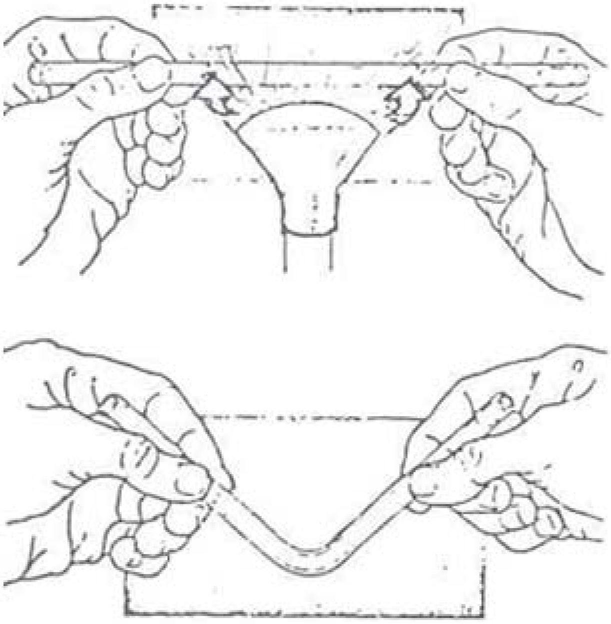


FIGURE 1.14 Bend examples:

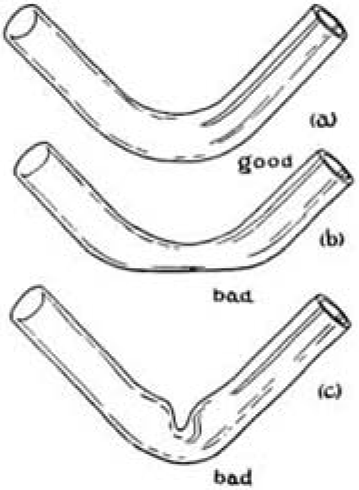
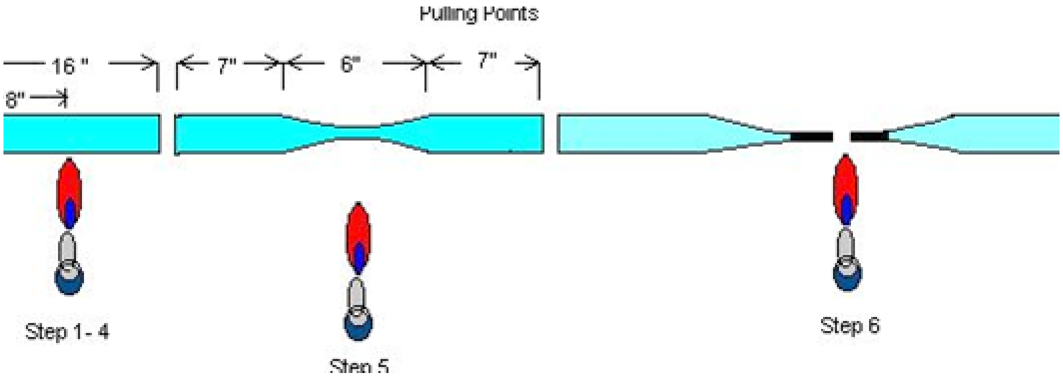


FIGURE 1.15



Care of Laboratory Glassware

Examine all glassware for cracks, stars, and chips. Beakers with cracks may break when heated, which may cause injury. Small chips in borosilicate can sometimes be eliminated by fire polishing. Otherwise, chipped glassware should be discarded to prevent cuts on sharp glass edges.

The recommended procedure for cleaning glassware is to wash the object carefully with a test tube brush, hot water and detergent, then rinse thoroughly with tap water, and finally rinse once again with a small quantity of distilled or deionized water. Then allow the glassware to drain dry overnight in your locker. If you must use a piece of glassware while wet, after cleaning, rinse it with the solution to be used.

Chemical Hoods

The proper use of chemical hoods is an important aspect of laboratory safety. Your instructor will discuss this more during the semester. It is advisable to use the chemical hood whenever you are performing reactions, particularly if the reaction produces noxious vapors. Whenever you are working in the chemical hoods be sure that:

1. The chemical fume hood is turned on (your instructor will do this).
2. The shield is pulled down at least 3/4 of the way down (to the marked sticker, leaving an opening approximately 47 cm or 18 inches) for proper airflow and will protect your head and neck while still allowing your hands and arms inside the hood to work.
3. Air is being properly circulated. Never put papers, lab books, or anything on the chemical hood sashes as this prevents air from being drawn into the hood.
4. The chemical hood is not cluttered with unnecessary materials as this may prevent the hood from working properly.
5. The “racks” of interconnecting horizontal and vertical bars are used to properly secure glassware using clamps whenever appropriate.
6. At the end of the laboratory period make sure that the chemical hood (and any other area that were used during the experiment like the reagent area, waste hood, and balance area) is clean this includes the bench top, the sides, bars, top, glass, sinks, under the sash, and that all clamps are put away.

EXPERIMENT 2

Measuring, Significant Figures, and Glass Working

Background

Significant Figures

In the laboratory many measurements are performed during the course of an experiment. These measurements contain two components: a numerical value and a unit that tells you what was being measured.

For example, if the volume of a liquid is measured using a graduated cylinder, the measurement would be recorded as 14.5 mL. The 14.5 value indicated the amount of liquid in the graduated cylinder. Whenever measurements are made there is a guess factor involved. In this case we are certain that there is 14 mL in the graduated cylinder; however the 0.5 value is estimated based on the precision of the graduated cylinder. As a general rule when a measurement is made, we always record one more number beyond the markings on the instrument (except when using something with a digital readout). For this measurement there are three significant numbers because that is how many numbers were recorded based on the precision of the instrument.

Keeping track of significant figures becomes especially important when measured values are used in calculations.

Rules for determining significant figures:

1. All nonzero digits are considered significant, leading zeros are not significant they are placeholders.

   34,789 L has 5 sf
2. Zeros that are between significant figures are considered significant.

   12.065 inches has 5 sf
3. Zeros that are *after* a significant figure and *after* a decimal point are significant.

   2.340 ft has 4 sf

4. Trailing zeros are ambiguous and should be used in scientific notation. If a trailing zero is followed by a decimal place it is significant.

   25,000 the number of significant figures is unclear 2, 3, 4, or 5 sf?????

   2.5×10^3 shows that there are only **2 sf**

   2.50×10^3 shows that there are **3 sf**

5. Leading zeros are not significant they are only placeholders.

   0.00067 nm has 2 sf

   as can be clearly seen when the value is written in scientific notation: 6.7×10^{-4} nm

6. Exact numbers have an infinite number of significant figures.

   4 pencils 1/2 1 m ≡ 100 cm

7. For addition and subtraction, the number of decimal places in the answer should be the same as the smallest number of decimal places in the data.

   14.562 mL 5 sf, but 3 decimal places
   +6.33 mL 3 sf, but 2 decimal places
   20.892 mL therefore, the answer should have 2 decimal places
   20.89 mL notice that it has only 2 decimal places, but the addition caused it to have 4 sf

   When adding or subtracting numbers with different powers of ten (10^x), one needs to be converted to the other to properly keep track of the significant figures.

   3.4561×10^3 kJ
   $+ 2.43 \times 10^2$ kJ
   becomes:
   34.561×10^2 kJ has 5 sf, but only 3 decimal places
   $+2.43 \times 10^2$ kJ has 3 sf, but 2 decimal places
   36.991×10^2 kJ should have only 2 decimal places
   36.99×10^2 kJ
   3.699×10^3 kJ in proper scientific notation.
   If you convert both numbers to 10^3, the same answer will be obtained.

8. For multiplication and division, the number of significant figures in the answer should be the same as the smallest number of significant figures in the data.

   $2.34 \text{ cm} \times 5.677 \text{ cm} = 13.28418 \text{ cm}^2 = 13.3 \text{ cm}^2$
   3 sf 4 sf 3 sf

9. For logarithms (log), the number of significant figures in the mantissa (original number) is equal to the number of significant figures in the decimal number (i.e. after the decimal point). Rules for natural logarithms, ln, follow the same procedure (recall that $\ln x = 2.303 \log_{10} x$).

   $\log [1.4 \times 10^2] = 2.146128 = 2.15$
   2 sf implies 2 dp in the answer

10. For antilogarithms (10^x), the number of decimal places determines the number of significant figures in the answer. The antilog is the opposite of the log. The exponential function, e, is the opposite of the natural log, ln.

    $10^{-5.67} = 0.000002138 = 0.0000021$ or 2.1×10^{-6}
    2 dp 2 sf

Recording of Data and Observations

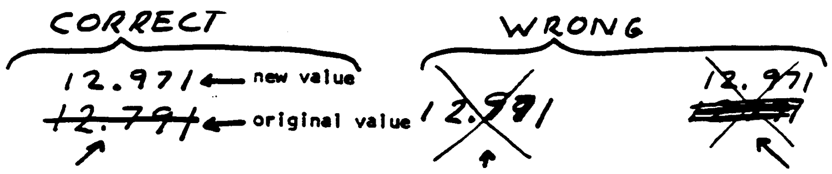
One of the several purposes of the laboratory part of this course is to learn to correctly record laboratory **data** and **observations**. The laboratory notebook of a practicing scientist becomes legal record and has to be kept in a manner that guarantees that it is accurate and not altered at a later date. You should begin to learn in the laboratory how to correctly collect and record relevant data and observations. Hence, the reason for the following rules:

1. All data and observations are to be recorded directly onto the laboratory report sheet in **blue or black non-erasable ink**.
2. Once the data has been recorded, they are to be
    a. **no strikeovers or scratch-outs or white-out.** If an error occurs, put one line through the data so that it can still be read and then write the new observation above it or in an appropriate place so that it can be easily read.
    b. The data section of your laboratory report sheet and all changes in data are to be signed or initialed by your instructor before you leave the laboratory at the close of the period.
3. **No changes** are to be made after this.

These procedures are explained below. Rule 1 should be fairly clear by itself, but note what is intended. Always write data and observations **directly** onto your report sheet in the designated place. Do not use scrap pieces of paper, the palm of your hand or some place on your report sheet not prepared for data or observations. Data found written on scrap pieces of paper, etc., may be confiscated and destroyed by your instructor. This means, for example, that you have to take your report sheet with you when you go to the balance so that you can record the mass directly onto your sheet. It also means to use reasonable care when recording an observation. The recording is to be made with *permanent* blue or black ink so that it cannot be erased or changed. The excuse that you forgot your pen, or do not have one, is **not acceptable**. It is a very good idea to put one or two ballpoint pens in your laboratory locker and leave them there between lab periods so that you will always be properly equipped.

Sometimes you may decide that a particular observation is in error. Rule 2 is the procedure to deal with this problem. As illustrated in Figure 2.1 , draw a line through the discarded observation and write in the new one. Note that the original can still be read. A “strikeover” is where you write the new value on the top of the original value. Thus the original cannot be read. As also illustrated, “strikeovers” or white-out are prohibited. Any method of altering the original recording, except with a single line out is not acceptable. In case you need more space to take down data or observation, use an appropriate blank part of your report sheet on which you have put the necessary labels, etc., to make it clear.

FIGURE 2.1 **Data Correction**



Rule 3 is self-explanatory; however, you should remember that changing data to get “better” answers is **cheating**.

Other parts of your laboratory report DO NOT have to be in ink. The calculations that are made from the data may be done on another sheet of paper first, if you prefer, and then transferred to the report sheet. They may be done in pencil and may be erased and changed if necessary. The same applies to the problems and application of principle sections at the end of the report sheet. Be sure, however, that you always show a complete set-up with correct units and the proper use of significant figures.

Measurements

Mass is determined using a balance. Your instructor will demonstrate the proper use of the quad-beam balance and the more precise electronic balance. Your mass measurement on the quad-beam and electronic balances should be the same to the tenths place. Why might the hundredths and thousandths places be different? Some procedural rules apply to all balances.

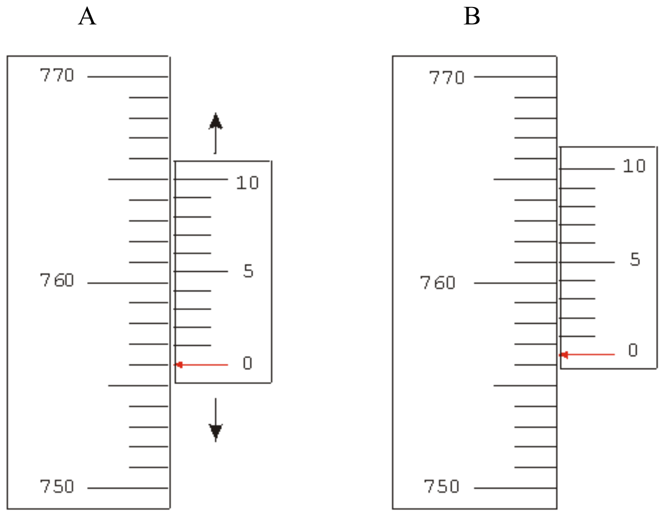
> **NOTE:** We will use the method of mass by difference to minimize any systematic errors in the weighing process. Be sure to use the same balance throughout the experiment!

1. Always keep the balance pans clean and dry. Only clean dry surfaces of glass, metal or weighing paper must ever touch the balance pan. Under no circumstances weigh reagents directly on the balance pan.
2. Always check the zero adjustment before making a mass determination. With the quad-beam balances this means double-checking to see that movable weights are set to the zero position.
3. Objects for mass determination must always be near room temperature. Hot objects cause convection currents that interfere with mass determination and on sensitive balances can produce incorrect readings due to fluctuating effects.
4. Record the mass of the object to the limit of the precision of the balance. Mass determination on the quad-beam balance must always be to the nearest 0.005 grams (three decimal places), and 0.001 g or 0.0001 g (three or four decimal places) on the electronic balances.
5. When done, return movable masses to their zero positions.

The temperature of water will be determined using a thermometer. Record all temperature values to one decimal place.

You will use Vernier calipers with a precision of 0.002 cm to determine the length and width of an object. Vernier allows a precise reading of some value. In Figure 2.2 on page 19, the Vernier moves up and down to measure a position on the main scale. The “pointer” is the line on the Vernier labeled “0”, which is used to determine the “certain digit” for a measurement. Thus the measured position is almost exactly 756 in whatever units the scale has been calibrated, for example, this may be mmHg as on a barometer. The last digit reading, i.e., the “doubtful digit” is read from the Vernier. In the examples below, the Vernier delineates 10 lines, which are used to estimate how far the measurement is between the gap of the lines printed on the scale. Since a single gap between the lines on the scale represent 1 unit increment (750 to 751, 751 to 752, etc.), the 10 lines drawn on the Vernier represent that increment divided by 10, or 1/10 of an increment Therefore, this Vernier allows precision to 0.1 units.

FIGURE 2.2 Barometric Pressure Readings Using Vernier Calipers



Main scale Vernier scale
756.0 mmHg

Main scale Vernier scale
756.5 mmHg

Your instructor will demonstrate how to read the Vernier calipers. Make sure to record all measurements made with the calipers to a precision of 0.002 cm.[1]

If the accepted or actual value for A is 755.87 mmHg, what is the percent error? Calculate the percent error according to the formula:

$$\%\ \text{error} = \frac{\text{observed value} - \text{accepted value}}{\text{accepted value}} \times 100\% \qquad \textbf{(EQ 2.1)}$$

$$\%\ \text{error} = \frac{756.0\ \text{mmHg} - 755.87\ \text{mmHg}}{755.8\ \text{mmHg}} \times 100\% \qquad \textbf{(EQ 2.2)}$$

$$\%\ \text{error} = \frac{0.13\ \text{mmHg}}{755.87\ \text{mmHg}} \times 100\% \qquad \textbf{(EQ 2.3)}$$

Do the subtraction set first and don't round until the end of the calculation. When you subtract 756.0 mmHg from 755.87 mmHg you get 0.13 mmHg, however only the tenths place is significant because your answer can only be as precise as the least precise measurement. The hundredths place is kept until the end to avoid round off error. Notice that the final answer has one significant figure.

$$\%\ \text{error} = 0.017198725\% = 0.02\% \qquad \textbf{(EQ 2.4)}$$

If the accepted value for B is 757.9 mmHg, what is the percent error? Answer: -0.32%.

---

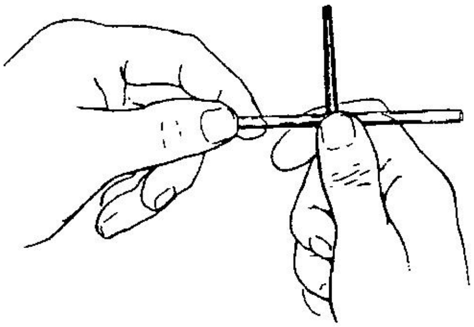
[1] The Vernier explanation uses Copyrighted material and images which can be found at: http://www.upscale.utoronto.ca/PVB/Harrison/Vernier/Vernier.html

Glass Working

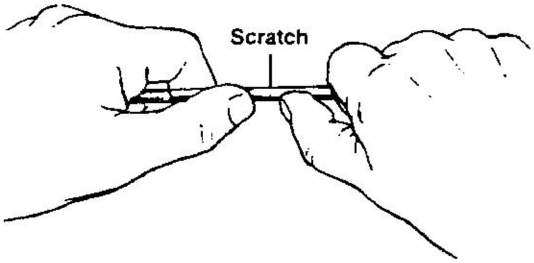
You will be using various pieces of glassware and laboratory apparatus throughout the semester. Many of these pieces are small and relatively easy to fabricate from glass tubing and glass rods; You will be making some of these pieces yourself. Hence, it is important to follow the safety rules and guidelines with respect to working with glass as illustrated in Figure 2.3 on page 21 and Figure 2.4 on page 22. Note the phase change of the glass as you heat it. Why does this phase change occur? **Be sure to fire polish all rough edges!**

FIGURE 2.3 Glass Working

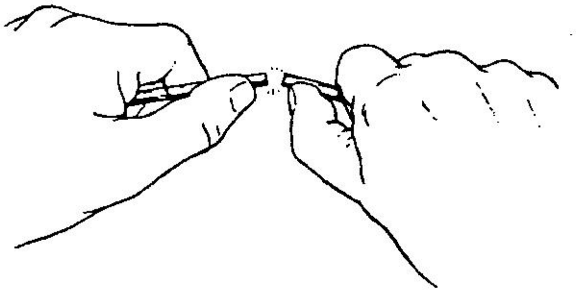
BREAKING A TUBE



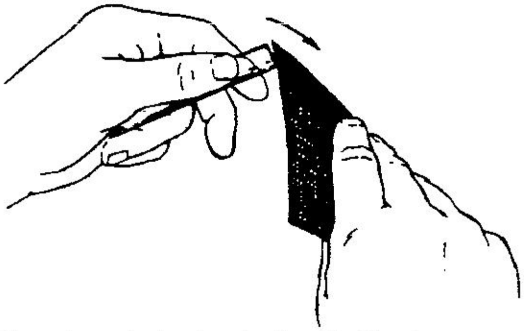
First – Make a single scratch with the triangular file.



Second – Place the thumbs together opposite the scratch.

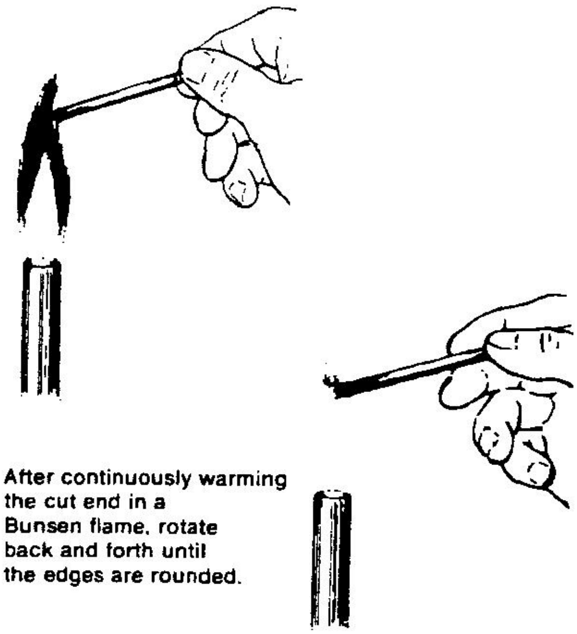


Third – Pull and bend quickly.

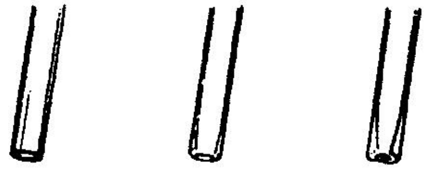


Trim a jagged edge by stroking it with wire screen.

FIRE POLISHING THE END OF A TUBE



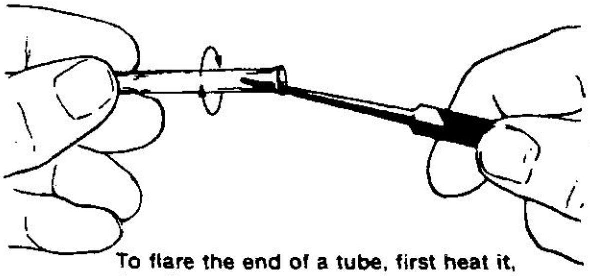
After continuously warming the cut end in a Bunsen flame, rotate back and forth until the edges are rounded.



The cut end

After fire polishing

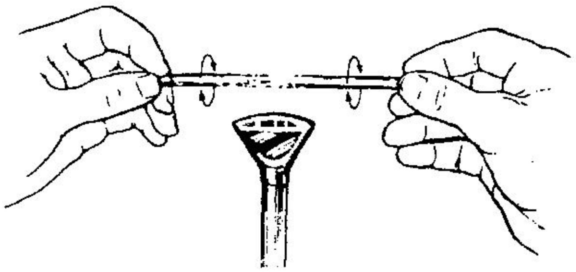
This has been heated too much



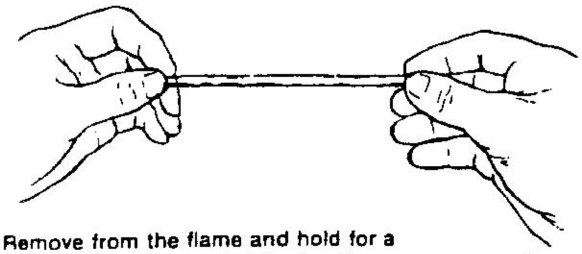
To flare the end of a tube, first heat it, then insert a file and rotate the tube.

FIGURE 2.4 Glass Working

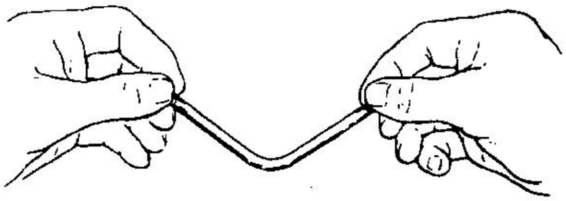
MAKING A BEND



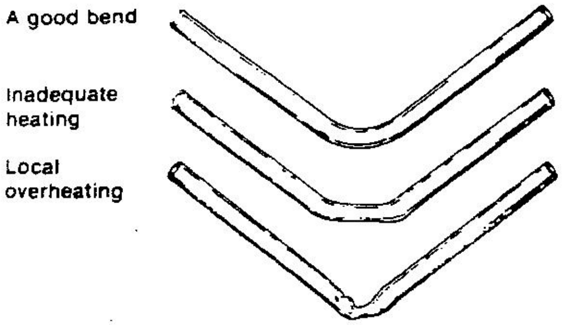
Roll the tube back and forth in the high part of a flat flame until it has become quite soft.



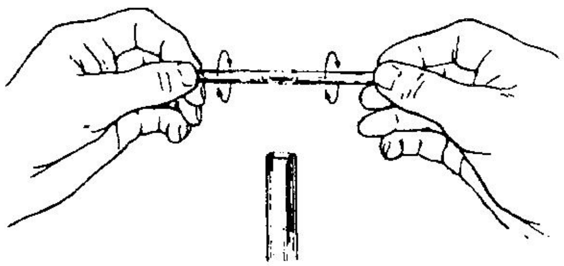
Remove from the flame and hold for a couple of seconds to let the heat become more uniform.



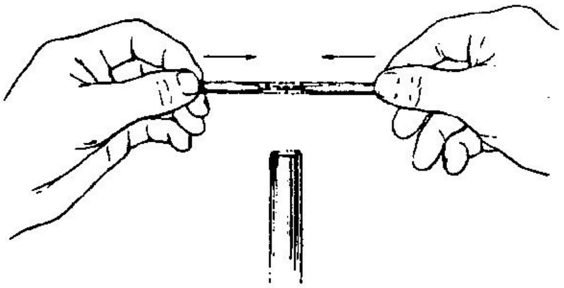
Bend quickly to the desired shape and hold until it hardens.



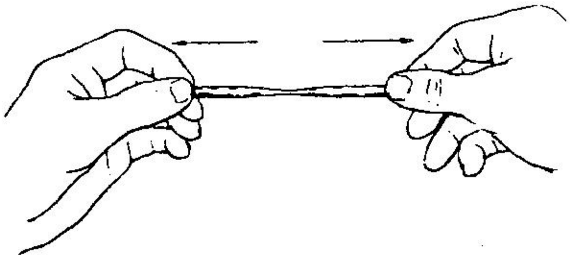
MAKING A CONSTRICTED TIP



Roll the tube in a Bunsen flame until it softens. Don't use a flame spreader.

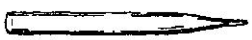


Allow the tube to become shorter as the walls thicken to about twice their original thickness.



Remove from the flame and after a moment pull until the softened region is as small as desired.

Cut to length



Fire polish, or file the tip

Procedure

Part I: Measurements

Mass. You instructor will show you how to correctly operate both the milligram quad-beam (± 0.005 g) and the top loading balance (± 0.001 g or ± 0.0001 g). Each measurement should be recorded to at least 3 decimal places and don't forget to include your units. Your mass of the slug on the quad-beam and electronic balances should be the same to the tenths place. Remember to use the method of mass by difference for the solid compound.

1. First record the number of your metal cylinder.
2. Take the mass a 50 mL beaker.
3. Then add your sample to the container and taking the mass of the container and sample.
4. Finally, you calculate the mass of your sample. Check your measurements with your instructor to ensure accuracy.
5. Calculate the percent error, which can be a positive value (i.e. the observed value is higher than the accepted value) or a negative value (i.e. the observed value is lower than the accepted value).

Length. You will practice using the Verier calipers to determine the diameter and length of your metal cylinder to a precision of 0.002 cm.

1. Measure the length and diameter of your cylinder.
2. Check your measurements with your instructor to ensure accuracy.

Temperature. You will practice taking the temperature of tap water. Be sure to recorder your temperature measurements to 0.1 °C. Be sure to hold the thermometer in the water. It should not touch the bottom or sides of the beaker. Remember you want to take the temperature of the water not the beaker.

1. Obtain about 30 mL of tap water in a 50 mL beaker.
2. Record the temperature of the tap water. Be sure to hold the thermometer in the water, but not touching the bottom or sides of the beaker.

Part II – Mechanics and Use of Bunsen Burners

1. Take apart the Bunsen burner as demonstrated by your instructor. Draw the assembled Bunsen burner and identify each of the following parts of the Bunsen burner: barrel, stand, hose, and valve.
2. Light the reassembled Bunsen burner. Adjust the flame so that
   a. it is yellow and has a smoky flame.
   b. it is noisy and goes out.
   c. it has a lighter blue inner cone and draw the good flame on your lab sheet using colored pencils.
3. Test different parts of the flame for hotness. Insert a wooden splint at the top of the flame and observe the time it takes to ignite. Insert the other end of the splint just above the light blue cone and observe the ignition time.
4. Compare to the previous measurement.

Part III: Glassware Fabrication

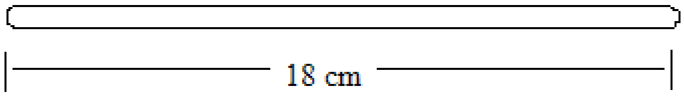
The following list illustrates the several pieces of glassware that you are going to make for this course. Be sure to carefully read through Figure 2.3 on page 21 and Figure 2.4 on page 22. Your instructor will demonstrate the basic operations needed for working glass tubing and rods.

Caution! When you take glass from the glass bin, be sure to check its diameter because the wrong sizes sometimes get put into the various bins. Please recall the safety rule: **all** glass tubing and rod is to be fire polished before it is used!

Stirring Rods

1. Make one stirring rod from the rods that are available. The stir rods should be approximately 18 cm long. (Be sure to use *solid rod* and not tubing!)

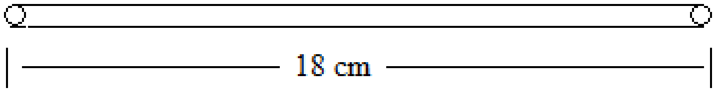
FIGURE 2.5 Stirring Rod



90° Bends

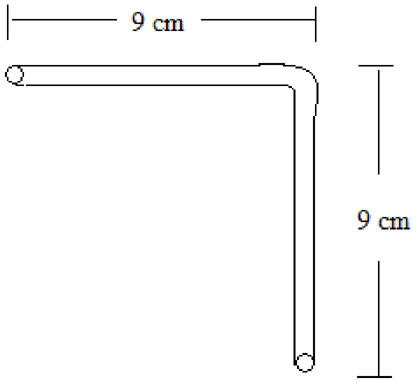
1. Make one straight tube of 6 mm (or of the diameter available) tubing: 18 cm in length as shown in Figure 2.6 on page 24. (Be sure to use *hollow tubing* and not the solid rod!)

FIGURE 2.6 Straight Tubes



2. Using your straight tube make one 90° bend using the 18 cm tubing as diagrammed in Figure 2.7 on page 24. Be sure to fire polish the ends being careful not to heat them too much.

FIGURE 2.7 Bends



Prelaboratory Exercise: Measuring, Significant Figures, and Glass Working

Name: ____________________________

Section:_____________________

Lab Grade

Prelab Questions		20
General Format (Signature, ink, no obliterations, etc.)		11
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		30
Accuracy (%error)		15
Post Lab Questions		24
Total		100

These questions refer to your reading in experiment 1 and experiment 2.

1. What is the proper procedure if you spill a chemical on your hand?

2. Why must you avoid eating, drinking, and gum chewing in the lab?

3. Write the following numbers in proper scientific notation to the correct number of significant figures:

a.	0.000082 m	
b.	0.0000000001420 g	
c.	28419200000 L	

4. Tell the number of significant figures in each of the following measurements:

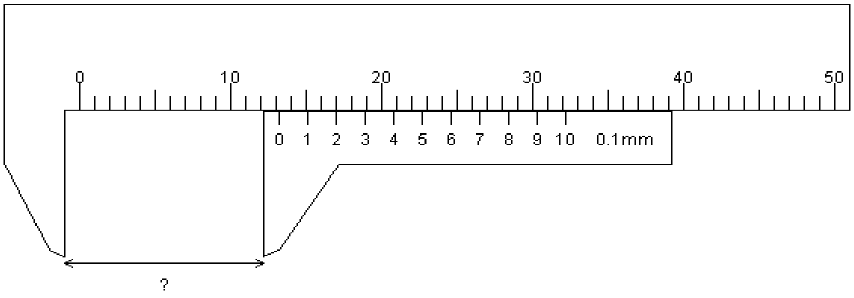
a.	0.025 m	
b.	40.0 g	
c.	0.085 L	
d.	129.50 g	
e.	0.003820 nm	
f.	190.0032 mL	

5. What precision should you record the following measurements?

TABLE 2.1

		Precision
a.	Quad-beam balance	
b.	Electronic balance	
c.	Vernier calipers	
d.	Thermometer	

6. Record the distance measured by the Vernier calipers to the correct precision of the instrument shown:



Data Sheet: *Measuring, Significant Figures, and Glass Working*

Part I: Measurements

Mass. Weigh your assigned object on both types of balances and record your observations and values below. Be sure to make qualitative observations about the slug. What does the slug look like?

TABLE 2.2 **Metal Slug**

Object Description: ____

Object Number: _________

Quad-beam Balance	Mass sample and container	
	Mass container	
	Mass sample	
	Percent error	
Electronic Balance	Mass sample and container	
	Mass container	
	Mass sample	
	Percent error	
	Actual Mass	

Sample Calculations (Show all calculations.)

Did the quad-beam balance or the electronic balance you a more accurate mass value? Use data to support your answer.

Length and Temperature. Record the length and diameter of your metal cylinder using Vernier calipers to a precision of 0.002 cm. Record the temperature of tap water to 0.1 °C.

TABLE 2.3

	Your measurement	Actual Value	Percent Error
Length			
Diameter			
Temperature			

Sample Calculations (Show all calculations).

Part II: Mechanics and Use of Bunsen Burners

Draw a Bunsen burner and identify the following parts of the Bunsen burner: barrel, stand, hose, and valve.

How would you adjust the Bunsen burner so that it has a

a. yellow and smoky flame?

b. is noisy and has a tendency to go out?

Using the good flame, test it with a wooden splint. Record your observations.

Time (seconds) at the top of the flame _________________

Time (seconds) above light blue cone _________________

Compare the two times. Why do you think that one takes longer to light than the other?

Draw the good flame using colored pencils. Use an arrow to identify the hottest portion of the flame.

Part III: Glass Working

Note the phase change of the glass as you heat it.

Why is it necessary to fire polish glass after cutting it?

Why do you need to rotate the glass rod while heating?

Have your instructor check and initial your stir rods and 90° bend. Make sure that all rough edges are fire polished.

TABLE 2.4

	Instructor Check
Stir Rod with smooth ends	
90° Bend with smooth ends	

Post Lab Questions

1. Are the following statements true or false? If false, correct the statement.
   - **a.** The precision of your answer is determined by the largest place of significance when adding and subtracting.
   - **b.** When multiplying and dividing the smallest number of significant figures in the data determines the precision of the answer.
   - **c.** For logarithms the number significant figures in the data determines the number of significant figures in the answer.

2. Perform the following calculations and report the answer, with proper units, to the correct number of significant figures. Use proper scientific notation when appropriate.
   - **a.** $487\ ft \times \frac{12\ \ in}{1\ \ ft} \times \frac{2.54\ \ cm}{1\ \ in} =$
   - **b.** $14\ \ mi \times \frac{1760\ \ yd}{1\ \ mi} \times \frac{3\ \ ft}{1\ \ yd} =$
   - **c.** $5.80\ \ mL \times \frac{2.7\ \ g}{1\ \ mL} =$
   - **d.** $2.37\ \ g + 14.05\ \ g =$
   - **e.** $726.2\ \ kg - 725.89\ \ kg =$

f. $324.55-\left(\frac{6104.5}{22.3}\right)=$

g. $\left[\left(\frac{20.43-1.094}{0.298}\right)+\left(\frac{294.22}{0.3821}\right)\right]+\left[\left(\frac{0.253+15.00}{.0924}\right)\right]=$

h. log (3.423×10^{-5}) =

i. antilog (0.112) =

j. antilog (4.3×10^{-6}) =

k. ln (9.7×10^{2}) =

EXPERIMENT 3

Mass, Volume, and Density

Background

In this exercise you will determine the densities of a liquid and an irregular solid. In the course of performing the experiment you should learn how to correctly use a balance and graduated cylinder. You should also gain a better understanding of the concepts of mass, volume and density.

Chemistry is the study of matter. **Matter** is generally defined as "that which has mass and occupies space". Matter is then is an intrinsic property of matter. An intrinsic property of a substance is independent of the amount of substance present. It is important to distinguish at this time between the concepts of mass and weight as they are not the same. Weight is the effect of the acceleration of gravity on mass. (An astronaut in orbit about the Earth may be effectively weightless, but not massless.) In the English system of weights and measures the equivalent mass of one pound weight is called a slug. Although it is important to understand the difference between the concepts of weight and mass, the common interchangeable usage of the terms is of little practical concern as we generally are "weighing" objects right here on the surface of the Earth where (in metric units) a gram mass exerts a force of a gram weight.

To state that matter occupies space is equivalent to stating that matter has volume. Volume, like mass, is an extensive property of matter. An **extensive property** is one, the value of which depends on the amount of matter. Simply put, the more matter you have the greater is its mass and the more space it occupies. Units of volume may be unique to volume as in a pint, a gallon, or a bushel or may be expressed as cubic length as in a cubic foot or a cubic meter.

Density is an intensive property of matter. **Intensive properties** are independent of the amount of matter involved. Density is defined as the "mass per unit volume" or the quotient of mass and volume of a sample of matter.

$$Density = \frac{mass}{volume} \Rightarrow D = \frac{m}{V} \qquad \textbf{(EQ 3.1)}$$

The density of a substance is one means by which we can identify it. Pure water, for instance has a density at room temperature of 1.00 grams per cubic centimeter. The densities of normal solids range in values from a few tenths of a gram per cubic centimeter to a bit more than 20 grams per cubic centimeter. Liquid water with a density of 1.00 g/cm^3 will float on liquid mercury with a density of 13.6 g/ cm^3 (is mercury the densest substance on Earth?) while vegetable oil at 0.8 g/cm^3 will float on water. Likewise a piece of wood, less dense than water, will float in water, while a stone, more dense than water will sink.

Measurement

Chemistry is an experimental science and central to the experiment is how precise and accurate measurements are. The determination of density requires measurement of both mass and volume. Mass is determined using a balance. Refer to "Measurements" on page 18 in "Measuring, Significant Figures, and Glass Working" for the proper use of the quad-beam balance (used in parts one and two) and the more precise electronic balance (used in part three).

The volume of a liquid sample can be determined using a graduated cylinder. The curved surface of the water in the cylinder is called the **meniscus**. Volume is measured using the very bottom of the meniscus for liquids and solutions. In the laboratory every measurement is made to the limit of the precision of the instrument used. In the case of a 50.0 mL graduated cylinder, you must interpolate the position of the bottom of the meniscus to the nearest one tenth (0.1) mL in every measurement. Your instructor will demonstrate how to use the graduation marks on the cylinder to avoid **parallax error** in reading the meniscus. When transferring an aqueous solutions use the method of "glass-to-glass-to-glass" as described in "Welcome to Lab" Figure 1.6, "Glass-to-glass-to-glass:," on page 8.

Graphing

Scientists experiment. Normally, the goal of an experiment is to determine how the value of one variable, called the **dependent variable**, changes as the experimenter systematically alters the value of another variable, called the **independent variable**. For example, a chemist may be interested in how the volume, V, of a gas sample varies as its pressure, P, is changed; or how the rate of a reaction changes with the concentration of one of the reactants. There are two common ways of presenting the results of experiments like these: tabular form and graphical form. Although both methods are commonly used, and are often used together, the graphical method is generally superior because it gives the scientist a visual feeling for the relationship between the dependent and independent variables.

Graphs are usually plotted on a **Cartesian axis system**. The vertical axis is called the y-axis (or **ordinate**) and the horizontal axis, the x-axis (or **abscissa**). Generally, the value of the dependent variable (or some function of it) is plotted on the y-axis. We assume that you have some familiarity with the Cartesian coordinate system and know how to plot a point represented by the ordered pair (x, y) in such an axis system.

The equation governing a linear relationship between two variables, x and y is given in Equation 3.2:

$$y = mx + b \qquad \textbf{(EQ 3.2)}$$

where y is the dependent variable, x is the independent variable, m is a constant called the slope, and b is a constant called the y-intercept.

Making Graphs

There are several important things to keep in mind when plotting data:

1. Blue or black ink pens should be used throughout, except for lines and circles.
2. All points should be circled as they are marked.
3. A table of all points should be made on the graph paper in an area where it does not interfere with the graph itself.
4. The name of the graph should be printed conspicuously at the top of the paper, as:

   "Density Changes at Various Temperatures:

   The Density of a Liquid (g/mL) vs. its Temperature (°C)"

   Be sure to give a descriptive title and then mention the dependent variable vs. independent variable.
5. Label the axes with the quantities to be plotted.
6. The entire window (for a manual plot, the window is the sheet of graph paper) is used for the plot. Generally, the bigger the graph the better, subject to the constraints below.
7. Refer to Table 3.2 on page 36 and Figure 3.2 on page 37 below. The following ideas should be kept in mind when choosing upper and lower limits:
   a. The lower and upper limits must bracket the range of experimental values being plotted; i.e., the lower limit must be less than or equal to, $\leq$, the smallest experimental value of the variable, and the upper limit must be greater than or equal to, $\geq$, the largest experimental value of the variable. Thus the lower limit of the y-axis in Figure 3.2 on page 37 must be < 0.417 and the upper limit > 0.613. Bracketing should be tight.
   b. The correspondence between the number of units spanned between the lower and upper axis limits and the number of intervals (steps) along the axis must allow for convenient plotting and reading of points. This means that the number of variable units per basic interval must be 2, 2.5, 5, or 10. (These numbers are generic, depending on the order of magnitude of the numbers plotted; the decimal point could be anywhere – e.g. 0.2, 0.25, 0.5, and 1.0).
   c. The lower limit of each axis need not be zero. Although zero is a convenient lower limit for the temperature, T, axis in this case, choice of zero for the lower limit of volume, V, would cause all the points to be cramped into a small area of the graph. The requirement that the plot fill as much as possible of the window takes precedence.
   d. For these reasons the upper and lower axis limits of the variable being plotted for the y-axis in Figure 1 are 0.65 L and 0.40 L.
8. Graphs should be placed on a on separated pieces of paper (not on the back of another).

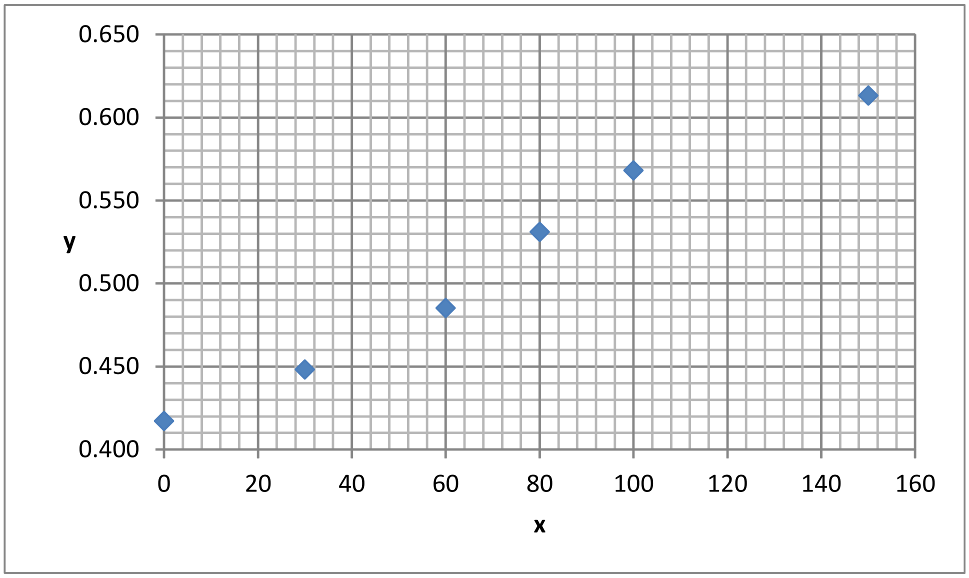
EXAMPLE 3.1 Volume and Temperature of a Gas

The volume of a sample of gas at constant pressure, P, varies with its temperature in ˚C. Varying the temperature, T, on a quantity of gas and measuring the change in volume, V, collected the following data. The volume is obviously the dependent variable in this case.

TABLE 3.2 Volume-Temperature Data

Temperature (˚C)	Volume (L)
0.	0.417
30.	0.448
60.	0.485
80.	0.531
100.	0.568
150.	0.613

FIGURE 3.1 Incomplete graph!

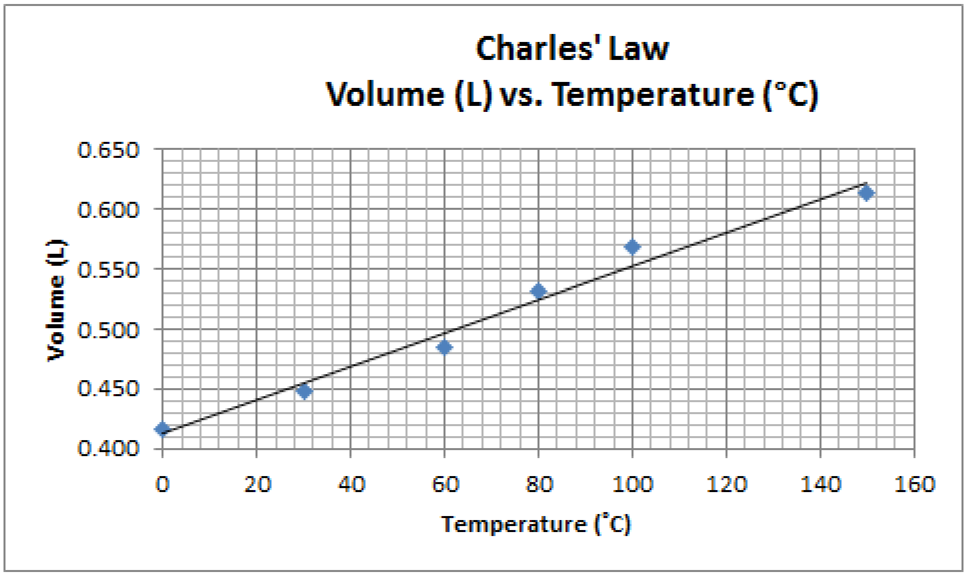


Notice that the x-axis starts at 0, but the y-axis starts at 0.400 to ensure that the graph takes up most of the graph area. The actual quantity plotted on the y-axis is the volume and on the x-axis the temperature. The next step is to draw a line that fits all of the data points. The line need not pass through any of the points and it should be just one line, not a series of lines, i.e. do not "connect the dots".

Clearly V increases as T increases, so a plot of V versus T might be linear. If we put an eye down near the paper and sight along the points, they can be seen to fall fairly closely along a straight line. The best is one that has as many points as possible near it, with roughly an equal number above and below. It is not necessary for any of the points to lie on the line, including the first and the last.

Using a transparent ruler in order to see the points above and below, you can draw this line. It can also be created using a computer plotting software package such as a spreadsheet. Computer programs give the line that best fits all of the data using the method of linear regression analysis, or the method of least squares. The straight line in Figure 3.3 has been drawn this way.

FIGURE 3.2 Better, but still incomplete!



To improve this graph further two valuable bits of information can be determined from this graph: the slope, m, and the y-intercept, b. Add a data table to the graph where is does not interfere with the plot is also a good idea.

The slope of the line can be calculated by choosing two new points (not actual data points), widely separated from each other on the graph and plugging their x and y values into the equation:

$$m = \frac{\Delta y}{\Delta x} = \frac{y_2 - y_1}{x_2 - x_1} \qquad \text{(EQ 3.3)}$$

Actual data points from the data table should not be used since the chances that they will lay on the best-fit line are slim. Widely separating the points chosen for slope calculations is important so that the most significant figures can be obtained. For instance, if we choose points that lay on the line at temperatures of 50. ˚C, and 100. ˚C, the resultant slope could have only two significant figures

$$\Delta x = x_2 - x_1 = 100.\ ^\circ C - 50.\ ^\circ C = 50.\ ^\circ C \qquad \text{(EQ 3.4)}$$

In this case we chose new points that lie on the line at x_1 = 0 ˚C and x_2 = 150. ˚C. We estimate y_1 = 0.415 L and y_2 = 0.623 L giving

$$m = \frac{\Delta V}{\Delta T} = \frac{0.623L - 0.415L}{150.°C - 0°C} = \frac{0.208L}{150.°C} = 0.00139\frac{L}{°C} \qquad \text{(EQ 3.5)}$$

The y-intercept is the value of y where x = 0. Using this graph we can estimate a y-intercept of 0.415 L from the graph, but this is not always possible. To calculate the y-intercept choose a third point off of the line. For example: x_3 = 120 °C and y_3 = 0.582 L giving

$$b = y - mx \qquad \text{(EQ 3.6)}$$

$$b = 0.582\ L - \left(0.00139\frac{L}{°C}\right)(120\ °C) = 0.145\ L \qquad \text{(EQ 3.7)}$$

Substituting these values in Equation 3.5 gives:

$$y = (0.00139\ \mathrm{L/^\circ C})x + 0.415\ \mathrm{L} \qquad \textbf{(EQ 3.8)}$$

Notice that because we are using experimental data for the slope and y-intercept that they both have units. In math the variables x and y are used, however, in chemistry it gives more information to substitute actual variables for x, temperature (T), and y, volume (V):

$$V= (0.00139\ \mathrm{L/^\circ C})T + 0.415\ \mathrm{L} \qquad \textbf{(EQ 3.9)}$$

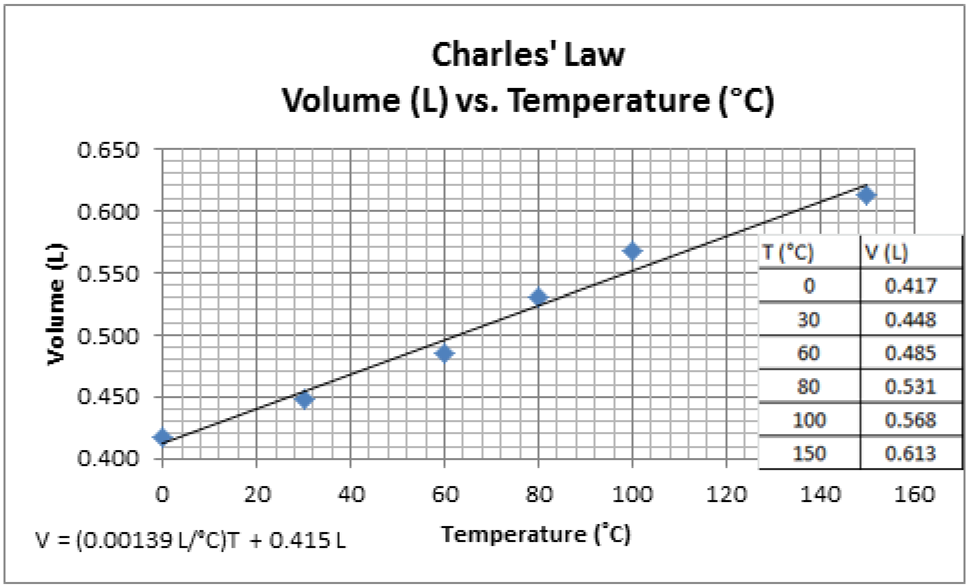
Note the units for V calculate out correctly to (L/˚C)(˚C) + L, or just L.

If the spreadsheet program is used to calculate the slope and the intercept the values that are returned are 0.0013878 and 0.41318 respectively. But of course it would be impossible to obtain so many significant figures using the original data, so they should be rounded off appropriately. The refined equation from the computer program would then be:

$$V = (0.00139\ \mathrm{L/^\circ C})T + 0.415\ \mathrm{L} \qquad \textbf{(EQ 3.10)}$$

which is not so much different from that obtained by graphing by hand or estimating.

FIGURE 3.3 **Best graph**



T (°C)	V (L)
0	0.417
30	0.448
60	0.485
80	0.531
100	0.568
150	0.613

The relationship between volume and temperature for a gas is called Charles' Law, and is usually written as:

$$\frac{V}{T} = k_1 \qquad \textbf{(EQ 3.11)}$$

where k_1 is a constant of proportionality, which depends on the amount and pressure of the gas. You will see Charles' Law again in the chapter on gases.

Procedure

Part I: Determination of the Density of an Irregular Solid

1. Determine which of the five solids you have been assigned as an unknown. Be sure to record your assigned irregular solid on your data sheet.
2. Weigh a clean, dry 50 mL beaker to the nearest 0.005 g on the quad-beam balance.
3. For the first trial fill your beaker with approximately 10 mL of irregular solid (use the markings on the 50 mL beaker to help you estimate how much to obtain).
4. Weigh the beaker and sample on the quad-beam balance and record the mass on your data sheet.
5. Add about 25 mL of water to a graduated cylinder and record the initial volume to the nearest 0.1 mL.
6. Carefully transfer all of the solid into the cylinder containing the water. Avoid splashing.
7. Record the new volume in the cylinder to the nearest 0.1 mL.
8. Decant the water from the cylinder into your largest beaker (use a wire screen to avoid dropping metal into the beaker). The water without any irregular solid can be poured down the drain. Pour the wet sample onto a paper towel. Dry the sample as best you can with the paper towel and deposit it in the proper reclaim container found on the reagent bench. **Do not place the wet metal back into the reagent jar!**

Caution! Safety Notes:

Be very careful not to allow any of the irregular solid to go down the drain.

9. For the second trial, repeat the entire process with a fresh, dry sample of your assigned substance. Obtain approximately 20 mL of irregular solid and repeat the above procedures.
10. For the third trial, repeat the entire process with a fresh, dry sample of your assigned substance. Obtain approximately 30 mL of irregular solid and repeat the above procedures.
11. Calculate the density of your sample from each of the three determinations.
12. Look up the actual density of your sample in the CRC Handbook of Chemistry and Physics (if you are assigned calcium carbonate for your unknown the density is listed on the bottle).
13. Calculate your percent error for each determination.

$$\%\ \text{error} = \frac{\text{observed value} - \text{accepted value}}{\text{accepted value}} \times 100\% \qquad \textbf{(EQ 3.12)}$$

Be sure to show the positive or negative sign (e.g. report your values as +5.67% or -5.67%). What does the sign tell you about the percent error?

14. Create a graph of mass of solid vs. volume solid using the graph paper provided.

NOTE: Be sure that your graph has a proper title, axis labels with units, a data table, circled points, and the equation of the line on it. Remember that your slope and y-intercept will have units and use descriptive variables instead of x and y.

15. Determine the slope of your line to obtain the density of the irregular solid.
16. Calculate the percent error for the density you determined from the slope of your graph.

What is the y-intercept and equation of the line?

Part II: Determination of the Density of an Unknown Solution

1. Obtain a solution sample with your name on it from your instructor. The sample is a salt solution with a density between 1.00 g/mL (the density of pure water) and 1.50 g/mL.
2. For the first trial pour between 45 and 50 mL of the solution into a graduated cylinder and record the volume of the sample to the nearest 0.1 mL. Remember when transferring aqueous solutions to use “glass-to-glass-to-glass.”
3. Determine the mass of a clean, dry 100 or 150 mL beaker to the nearest 0.005 g on a quad-beam balance.

> **NOTE:** We will use the method of mass by difference to minimize any systematic errors in the weighing process. Be sure to use the same balance throughout the experiment!

4. Pour the sample from the cylinder into the pre-weighed beaker and reweigh the beaker.
5. For the second trial: from the beaker pour between 40 and 45 mL of the solution back into the graduated cylinder. Record the new volume of the cylinder (again to the nearest 0.1 mL).
6. Dry the beaker and transfer the solution from the cylinder into the beaker. Record the new mass of the beaker and solution.
7. For the third trial: from the beaker pour between 35and 40 mL of the solution back into the graduated cylinder. Record the new volume of the cylinder (again to the nearest 0.1 mL).
8. Dry the beaker and transfer the solution from the cylinder into the beaker. Record the new mass of the beaker and solution.
9. From the three determinations, calculate the density of the solution and the average density from the three determinations. (watch your significant figures!)
10. Obtain from your instructor the actual density for your unknown.
11. Calculate your percent error in the density.

    Be sure to show the positive or negative sign (e.g. report your values as +5.67% or -5.67%). What does the sign tell you about the percent error?

12. Used salt water solution may be poured down the sink.

Prelaboratory Exercise: Mass, Volume, and Density

Name: ___________________________

Section:____________________

Lab Grade

Prelab Questions		16
General Format (Signature, ink, no obliterations, etc.)		16
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		18
Graph (properly labeled, scaled, overall neatness, etc.)		10
Accuracy (% error)		20
Post Lab Questions		20
Total		100

1. A 50 mL graduated cylinder should be read to the nearest ___________ mL.
2. (True/False) You will get the same weight for an object regardless of whether you weigh it hot or cold.
3. 100 mg is what part of a gram? __________
4. (True/False) Even though the metric system is extremely useful, it is more difficult to add different sized quantities in the metric than in the United States Customary System.
5. With how many different volumes should you determine the density of the liquid in this experiment?

   ____

6. The curved surface at the top of a liquid in a graduated cylinder is called a ____________.
7. (True/False) The densest substance in the world is mercury.
8. The surface of a liquid in a graduated cylinder should be read at eye level to prevent a

   ____________________ error.

9. You can get the true value of the density of the solution you use in this experiment from the

   ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ .

10. A student forgot to zero the balance before weighing his beaker (in “Part II: Determination of the Density of an Unknown Solution”) and got a value that was actually too high. When he weighed the beaker with the liquid in it, however, he remembered to zero and got an accurate value for this weighing. Will the value he gets for the weight of the liquid be high, low or unchanged?

11. If the same student then subsequently measures his liquid volume accurately, will his density result be high or low or unchanged?

12. If a student determines the density of a substance to be 1.80 g/mL, but the actual density is known to be 2.000 g/mL, what is the percent error?

13. A graduated cylinder originally contained 17.2 mL of water. An object weighing 14.1 g was then submerged in the water and the new reading was 21.1 mL. The balance can be read to ± 0.1 g and the cylinder to ± 0.2 mL.

- **a.** Which measurement is the less precise? ____________________________________
- **b.** What is the volume of the object?
- **c.** What is the density of the object?

14. When constructing a graph, the scales should be adjusted so that the graph will cover what part of the page?

Data Sheet: Density Determinations

Part I: Density of an Irregular Solid

Irregular Solid ____________________

TABLE 3.3

	Trial 1	Trial 2	Trial 3
Mass of beaker and solid			
Mass of beaker			
Mass of solid			
Volume of liquid and solid			
Volume of liquid			
Volume of solid			
Literature density of solid			
Calculated density of solid			
Percent error			

Sample Calculations (Remember to show one sample calculation of each type for one trial.)

For the graph:

1. Complete the following table and create a graph of mass of solid vs. volume of solid.

TABLE 3.4

	Volume of solid	Mass of solid
Trial 1		
Trial 2		
Trial 3		

2. What is the slope, m, the graph? Show your calculations here and include the correct units.

3. Remember that the slope of the graph gives the density of the irregular solid. What is the percent error of the graphically determined density of the irregular solid?

4. Did the manual or graphical method give you a more accurate density value?

5. What is the y-intercept, b, of the graph? Show your calculations here and include the correct units.

6. Write the equation of the straight line for your graph. Remember it your line should include descriptive variable and units (i.e. in the form: M = mV +b)

Part II: Density of Unknown Solution

User ID: ________________________________

TABLE 3.5

	Trial 1	Trial 2	Trial 3
Mass of beaker and solution			
Mass of beaker			
Mass of solution			
Volume of solution			
Density of solution			
Actual Value			
Percent error			

Sample Calculations (Remember to show one calculation of each type for one trial only.)

Post Lab Questions

1. What would happen to the determined value of your solid density in “Post Lab Questions” too high, too low, or unchanged in each of the following scenarios? Explain your answer.
    a. There were air bubbles in the water with the metal.

    b. The initial volume of the liquid in the graduated cylinder was recorded too low.

    c. Some water splashed out of the cylinder when the solid was added.

    d. The solid was wet when weighed.

2. Why could you not determine the density of salt in the same way you determined the density of the irregular solid unknown (water displacement) in “Post Lab Questions”?

3. How could the experiment be modified to determine the density of salt?

4. In “Part II: Density of Unknown Solution”, could you increase the number of justifiable significant figures in your answer by using the electronic top-loader balance in place of the quad-beam balance? Explain.

5. If a graduated cylinder initially contains 28.8 mL of water and the addition of 70.42 grams of metal shot causes the water level to rise to 37.3 mL,
   a. what is the volume of the metal shot?
   b. what is the density of the metal shot?
   c. How would the density be affected (too high, too low, unchanged) if a piece of metal shot was lost after it was weighed and before the volume was determined.

EXPERIMENT 4

Physical Properties of Substances

Background

Chemists frequently take advantage of the physical properties of substances to characterize, identify, and isolate them. Examples of physical properties are color, odor, density, melting point, boiling point, crystalline form, hardness, malleability, ductility, solubility, thermal conductivity, and electrical conductivity. Many of these properties can also be explored graphically. As you saw in the experiment, "Mass, Volume, and Density", a plot of mass in grams versus volume is linear. On the other hand, a plot of solubility of a salt such as potassium dichromate, $K_2Cr_2O_7$, versus temperature is curved. Although both linear and curved plots can be useful, the linear plot is generally more so, for several reasons. First, the eye readily sees linearity or deviations from it. On the other hand one cannot tell visually if a curve that resembles a hyperbola is truly hyperbolic. Second, linear plots may be more readily interpolated and extrapolated. **Interpolation** involves reading values between the points of a graph. **Extrapolation** takes us outside all the data points and into regions that may not be experimentally accessible. Extrapolation allows for the predication of the value of the dependent variable under conditions other than those governing the experiment. For these reasons, scientists usually prefer linear plots.In this experiment you will study the physical properties of a variety of substances. You also will learn how to use the Bunsen burners and to dispense reagents properly.

The basis of science is observation. Good scientists make clear and concise observations about the world around them. Scientists sometimes make inferences based upon their observations of substances. In this experiment you will make observations of density, physical state, volatility, color and odor. You should exercise care when smelling any substance because it may have an objectionable odor or be toxic. The proper technique for smelling substances is to waft the vapor towards your nose so that you may smell a very small quantity of it. You should never taste substances you find in the chemical laboratory, as they may be toxic.

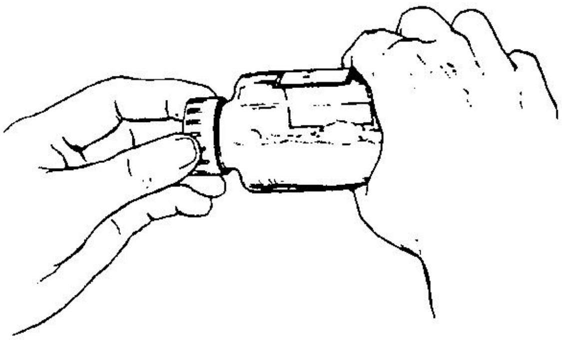
Proper Reagent Dispensing Technique

A chemist should always be confident of what is in a reagent bottle. To avoid contamination, nothing should *ever* be introduced into any reagent bottle. For this reason we have very specific rules for dispensing reagents.

1. For any reagent, the container cap should never be placed on the lab bench. Hold it between your fingers.
2. A reagent should be poured into a beaker or onto a piece of weighing paper, not into a test tube.
3. Spatulas, scoopulas, pipettes or other objects should never be introduced into the reagent bottle.
4. Excess sample should never be poured back into the reagent bottle. If you take more than you need, offer the excess to another student or dispose of it properly.

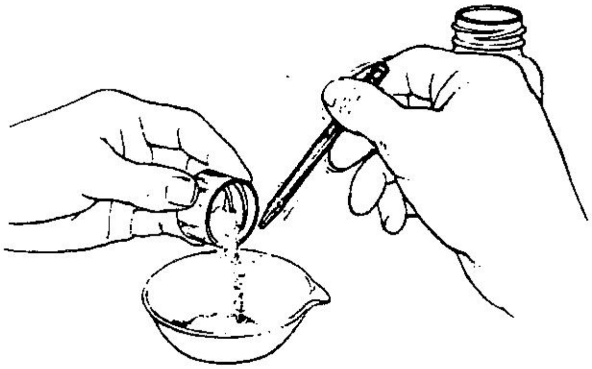
Techniques for transferring solids and liquids are shown in Figure 4.1 on page 55 and Figure 4.2 on page 56. Your instructor will also demonstrate the proper dispensing techniques.

FIGURE 4.1 Method for transferring powders and crystals



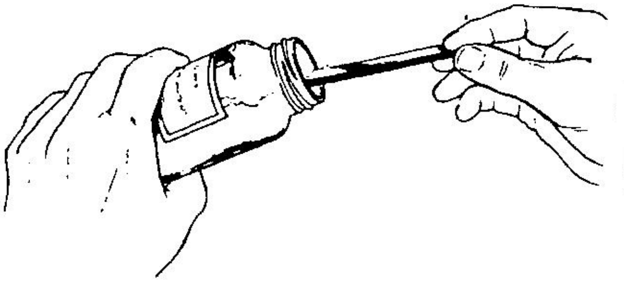
First— Roll and tilt the bottle until some of the contents enters the inside of the plastic cap.

Second— Carefully remove the cap so that some of the contents remains in it.

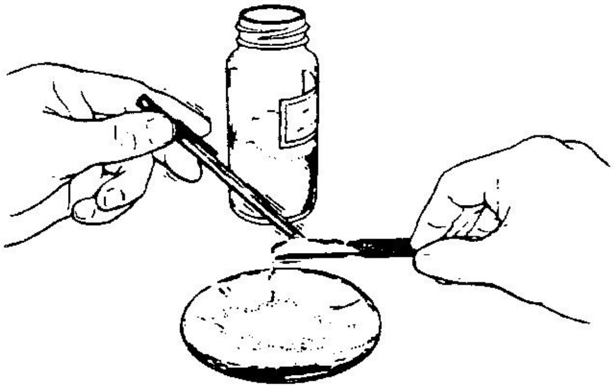


Third— Tap the cap with a pencil until the desired amount falls out.

First Method

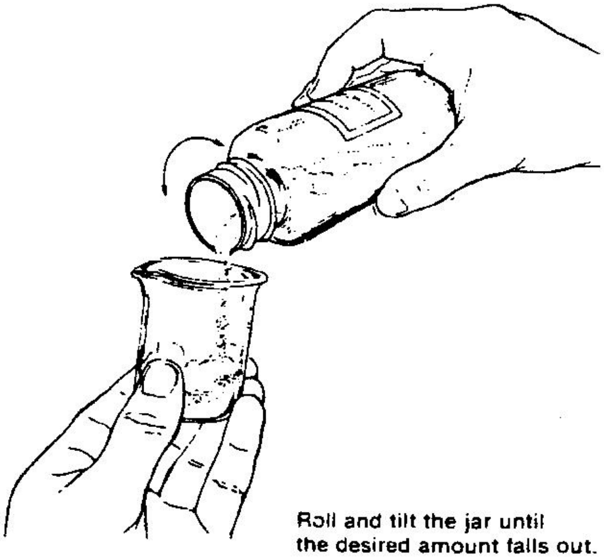


Scoop out a little of the material with the spatula provided.



Tap the spatula until the desired amount falls off.

Second Method

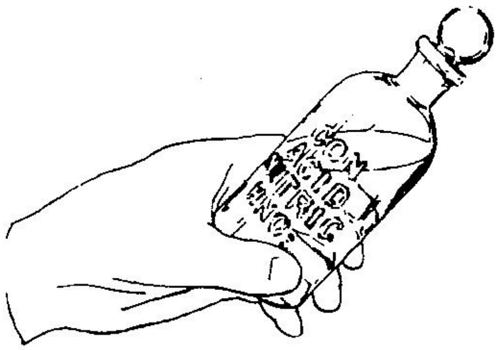


Roll and tilt the jar until the desired amount falls out.

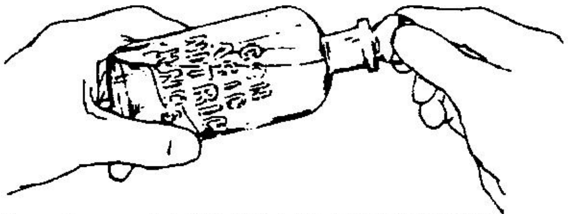
Third Method

FIGURE 4.2 Transferring liquids

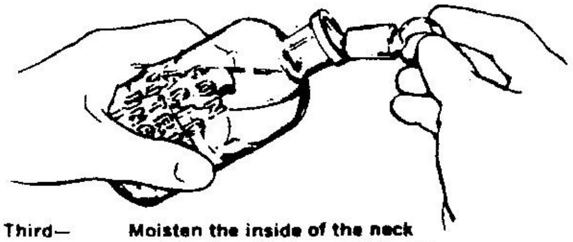
REMOVING A STOPPER



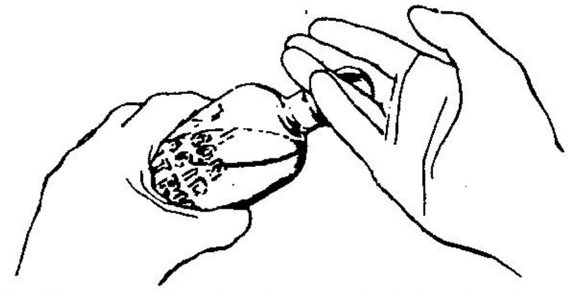
First— Read the label twice.



Second— Hold the stopper in and tilt the bottle until the contents wet the stopper.

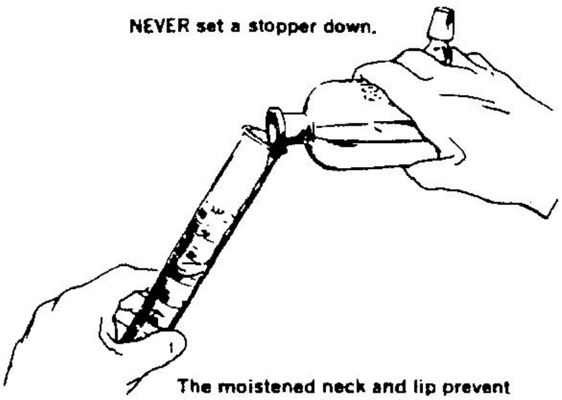


Third— Moisten the inside of the neck and the lip with the wet stopper.



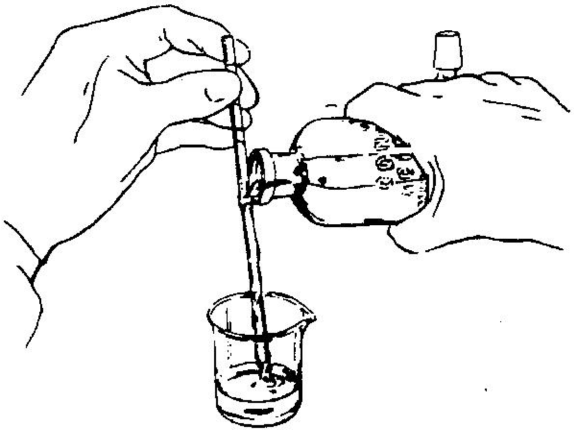
Fourth— Replace the stopper and withdraw it again with the back of the hand.

POURING A LIQUID

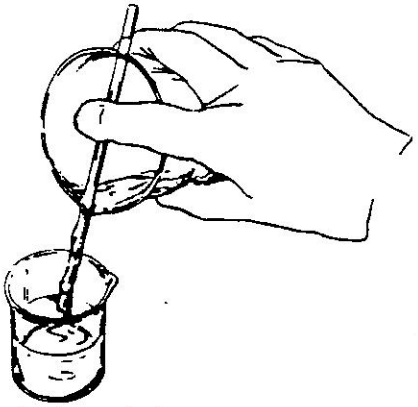


NEVER set a stopper down.

The moistened neck and lip prevent the first drops from gushing out.



Pour down a glass rod when possible.



When pouring from a beaker, the stirring rod can be held in this manner.

Procedure

Part I – Separation of Substances Based on Physical Properties

This part of the experiment involves separating a salt and sand from a mixture of the two. The separation is based on the fact that the salt is soluble in water and sand is not. The stockroom has made a mixture of sand and salt (sodium chloride) for each student. Each mixture has a different mass percent of salt. You will be required to separate this sample into its component substances and will be graded on how carefully you accomplish this separation. (Your grade is determined by comparing your experimental percent salt with the true value. Be careful!) Your mixture with your name on the label is on the instructor's desk.

1. Weigh a clean, dry 150 mL beaker and a clean, dry evaporating dish, each to maximum precision of the balance (0.001 g or 0.0001 g).

> **NOTE:** We will use the method of mass by difference to minimize any systematic errors in the weighing process. Be sure to use the same balance throughout the experiment!

2. Also weigh the container with the unknown sample.
3. Without loss, transfer the entire unknown to the weighed beaker.
4. Weigh the empty sample container, and the beaker plus the unknown. (Note: This double weighing allows a check of the mass of the sample.)
5. Add about 10 mL of deionized water to the unknown in the beaker; gently warm this, while stirring, until the salt dissolves. Then let the mixture settle and carefully decant the clear solution (supernatant liquid) down the stirring rod into the weighed evaporating dish. Remember to use "glass-to-glass-to-glass."
6. Add a second 10 mL portion of deionized water to the sand mixture and repeat the extraction and decantation.
7. Place the evaporating dish and solution on a wire gauze which is on a ring stand, and begin careful evaporation, using a controlled flame.
8. The complete extraction of any remaining salt from the sand will probably require that a third and possibly a fourth 10 mL portion of deionized water be added to the evaporating solution.
9. Finally, decant as much water as possible from the sand (without any loss of sand), and gently heat the beaker of moist sand on a hot plate until it is thoroughly dry (to constant mass).
10. Let this cool completely and weigh it.

> **NOTE:** While the salt is cooling move on to "Part II – Observations of Physical Properties" on page 58. Remember to use your laboratory time as efficiently as possible.

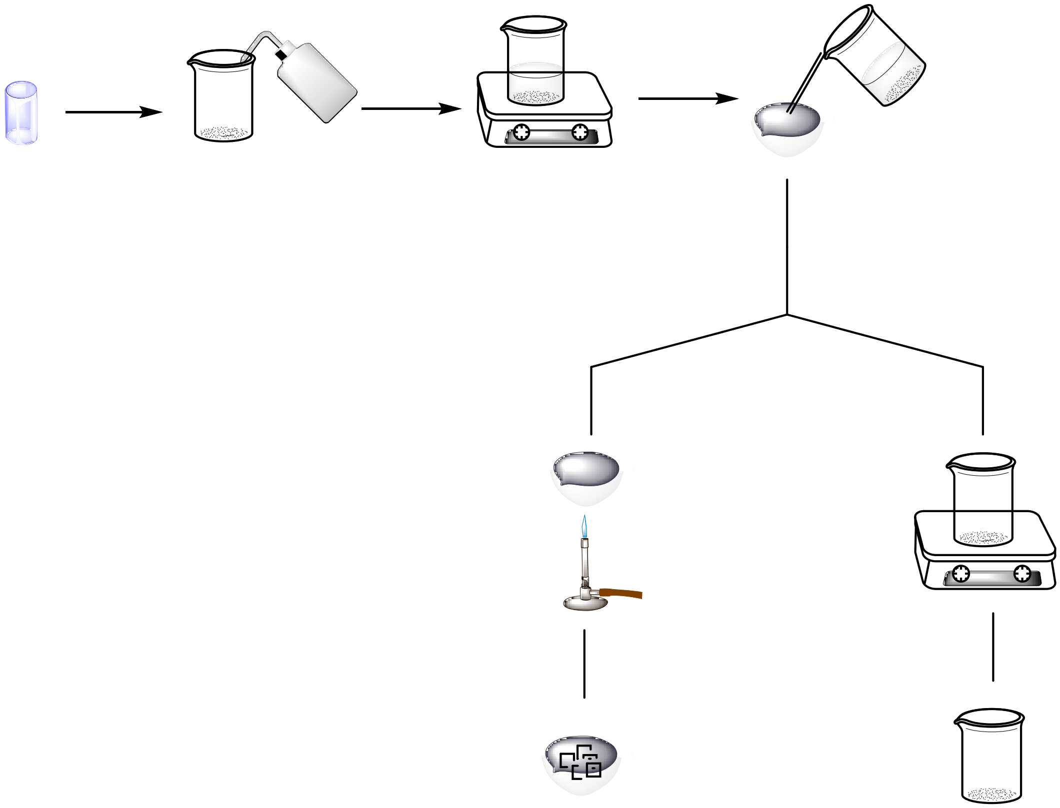
11. As the solution in the evaporating dish becomes concentrated and crystals form, be very careful to avoid splattering.
12. When the salt is completely dry, let the dish cool completely and weigh it.
13. Reheat the dry salt, cool, and reweigh to make sure the sample is completely dry. If the 1st and 2nd weights differ by more than 0.01 g, heat for a 3rd time and reweigh. (This is called drying to a constant mass. Use the last mass in all of your calculations.)
14. From these data you can calculate the percent salt in your unknown mixture. Use the average mass of salt and the average mass of the mixture to calculate the percent salt:

$$\text{percent salt} = \frac{\text{mass salt}}{\text{mass mixture}} \times 100\% \tag{EQ 4.1}$$

15. Observe the sodium chloride that has crystallized in the evaporating dish and compare it to sodium chloride crystals found on the reagent shelf.
16. After recording all data and checking your numbers with your instructor, scrape some of the salt out of the evaporating dish and observe them under a microscope.
17. Do the same with a few crystals of salt from the reagent bottle. Compare your observations.

> ⚠ Dispose of your sand in the trash can and not the sink.
>
> Be sure to return your sample container where you go it.

FIGURE 4.3 **Flowchart: Separation of Salt and Sand**



Part II – Observations of Physical Properties

You will make two sets of observations

1. Observe the odors of the substances provided. Be sure to use the proper technique as demonstrated by your instructor. What can you infer about the relative boiling points of these substances?
2. Observe pre-measured masses of the elements and compounds listed on your data sheet. Record the colors, physical states, and physical appearances of each. What can you infer regarding the relative densities, melting points or boiling points of these substances?

Part III – Using the CRC Handbook

Some properties of Potassium Hexachloroplatinate . Look up potassium hexachloroplatinate in the table called "Physical Constants of Inorganic Compounds." It will be listed under potassium. Look at the many types of information concerning this compound that are listed. Looking across the table you will find several valuable bits of information about this compound including its density and solubility. Answer the questions concerning this part.

Boiling Point of Group VI Hydrides vs. Their Molecular Weight . Look up hydrogen sulfide, hydrogen selenide, and hydrogen telluride (Some may be found under hydrogen; others will be listed under the other elements, as hydrides.) and note the molecular weight (in g/mol or amu) and boiling point (°C) of each of the three compounds. Prepare a hand drawn graph for this part, using the boiling point as the dependent variable. (Start the molecular weight axis at zero amu.) Do not connect the points with a smooth curve, short straight lines between each adjacent pair of points will do (i.e. "connect the dots"). Answer the questions.

Part IV – Observation of Physical Changes and Crystal Structures

Some physical properties can only be observed by a physical change. Examples of these properties are melting point, boiling point, and sublimation point. (Look up the definition of sublimation.) Your instructor may complete this part as a demonstration.

Sulfur

1. Pour some sulfur into a beaker. Observe the crystalline structure of the sulfur under a magnifying glass.
2. Prepare a cone of filter paper and support it either in a funnel or a small beaker.
3. Transfer enough of the sulfur to fill a 15 cm test tube to within three cm of the top.
4. Holding the test tube with a test tube holder, heat it slowly and uniformly in order not to superheat any portion of it. The sulfur will darken if it is superheated; this can be avoided by moving the test tube in and out of the flame.
5. When the sulfur is just melted, it should be an amber yellow, straw-colored, liquid.
6. Pour it into the filter cone previously prepared, and with a wood splint in hand, watch the formation of long needle-shaped crystals.
7. Just as the surface of the liquid begins to solidify, break it open with the wood splint, and quickly pour the remaining molten sulfur into a beaker of tap water.
8. Briefly let the filter cone cool, then break it open and observe the crystals.
9. Compare the newly formed crystals, the sulfur that falls into the water, and to the original sample of sulfur to each other using a magnifying glass or microscope.

Iodine

1. Place a few small crystals of iodine in a dry 250 mL beaker.
2. Cover the beaker with a watch glass and place ice on the watch glass.
3. Support the beaker on a ring stand with wire gauze, and heat the iodine slowly until all the crystals sublime and the vapor deposits on the bottom of the watch glass.
4. Under a magnifying glass observe the crystals that form.

Prelaboratory Exercise Week 1: Physical Properties of Substances

Name: ___________________________

Section:____________________

Lab Grade

Prelab Questions Week 1		10
Prelab Questions Week 2		15
General Format (Signature, ink, no obliterations, etc.)		5
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		40
Accuracy		30
Total		100

1. All data are to be recorded directly onto the data page in ___ ___ ___!

2. You are advised not to taste any chemicals in this experiment. Why not?

3. While doing “Part I – Separation of Substances Based on Physical Properties”, a student weighed the container with the unknown, poured it into a weighed beaker, then washed, dried and reweighed the original container. What did the student do that should not have been done?

4. On the basis of what he saw and recorded, the same student decided the solid was sodium chloride. He was drawing an _**I**_ ___ ___ ___ ___ ___ ___ ___ ___.

5. To pour off a supernatant liquid is to ___ ___ ___ ___ ___ ___ it.

6. A student obtained the following data. Fill in the blank spaces.

TABLE 4.1

a.	mass of sample container and sample	32.345 g
b.	mass of sample container alone	29.972 g
c.	mass of sample	
d.	mass of evaporating dish and salt (1st weighing)	45.676 g
e.	mass of evaporating dish and salt (2nd weighing)	45.671 g
f.	mass of evaporating dish	44.780 g
g.	mass of salt extracted	
h.	percent of salt in sample	

Sample Calculations

7. Suppose the student in question 6 had weighed the empty sample container incorrectly, getting a lower value than it actually weighed. How would this affect the percent salt result (high or low)?

Prelaboratory Exercise Week 2: Physical Properties of Substances

Name: __________________________

Section:____________________

1. What is the conversion of a gas to a solid called?___________________________________

2. A student took a white, crystalline solid and determined its melting point. The

student was making an ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___.

3. Iodine crystals may be easily ___ ___ ___ ___ ___ ___ ___ ___.

4. An experiment calls for 60.0 g of concentrated nitric acid, density 1.42 g/mL. Suppose no balance is available, and you decide to use a graduated cylinder. What volume should you use?

5. The dilute sulfuric acid (an aqueous solution) on the laboratory desk has a density of 1.18 g/mL, and is 25.0% by mass sulfuric acid, the remainder being water.

a. How many grams do 15.0 mL of this dilute sulfuric acid solution weigh?

b. What is the weight of pure sulfuric acid in this 15.0 mL of solution?

6. Calculate the volume of magnesium, density 1.74 g/cm^3, which would be equal in mass to 350.0 cm^3 of lead, density 11.4 g/cm^3.

Data Sheet: Physical Properties of Substances

Part I – Separation of Substances Based on Physical Properties

Unknown number: _____

TABLE 4.2

Mass of sample container and sample		Mass of beaker and sample	
Mass of sample container		Mass of beaker	
Mass of sample (from sample container)		Mass of sample (from beaker weighings)	
Mass of evaporating dish and salt	1st weighing	Mass of beaker and sand (extracted, dry)	1st weighing
	2nd weighing		2nd weighing
Mass of evaporating dish and salt continued (if needed)	3rd weighing	Mass of sand	
	4th weighing	Mass of salt (by subtracting the beaker weighings)	
Mass of evaporating dish			
Mass of salt extracted			
Mass of sand (from subtraction)			

TABLE 4.3

Average mass of salt		Mass percent salt in sample	
Average mass of sand		Actual mass percent salt in sample	
Average mass of sample		Percent Error	

Sample Calculations

TABLE 4.4 **Sketch the sodium chloride crystals**

Sodium chloride from reagent bottle	Sodium chloride from evaporating dish

1. How do the crystals differ?

Part II– Observations of Physical Properties

Record the odors of each of the following substances

TABLE 4.5

Acetone	
Water	
Ethanol	
Para-dichlorobenzene	

2. Can you make any inferences regarding the relative boiling points of each of the substances above based on the odors you detected? Why or why not? Explain your reasoning.

Record your observations regarding each of the following substances. Be sure to note the color, physical appearance, and physical state of each.

TABLE 4.6

Mercury	
Copper	
Magnesium	
Iodine	
Bromine	
Carbon	
Copper carbonate	
Copper nitrate	
Potassium permanganate	
Potassium dichromate	

3. Can you make any inferences regarding the relative densities of copper, magnesium, carbon and mercury?

4. Can you make any inferences regarding the relative melting points of mercury and copper?

Part III– Using the CRC Handbook

Potassium Hexachloroplatinate. Answer the following questions using the CRC Handbook of Chemistry and Physics:

5. What edition of the handbook are you using?

6. According to the table of “Physical Constants of Inorganic Compounds”, what is the formula of potassium hexachloroplatinate?

7. What is its density? (Abbreviations are explained at the beginning of the table.)

8. What happens to it at 250 °C?

9. How soluble is it in water? Remember to include units!

Group VI Hydride. Use the CRC Handbook to complete the table below, create a graph, and answer the questions.

TABLE 4.7 **Complete the table below for the graph of boiling point vs. molecular weight:**

Compound Name	Formula	Molecular Weight (amu)	Boiling Point ( )
Hydrogen Sulfide			
Hydrogen Selenide			
Hydrogen Telluride			

10. What is the name of the first element in the chalcogen group (Group VIA of the Periodic Table)?

11. What is the formula of the compound of this first element that is analogous to the ones graphed?

12. What is the name of the compound?

13. What is the molecular weight of this compound?

14. What is the boiling point of this compound?

15. By extending your graph to the left, what would you expect the boiling point of that compound to be?

16. What is the process of extending the graph out beyond the data points called?

17. Why would interpolation of the boiling point of water be meaningless for this graph?

Part IV – Observations of Physical Changes and Crystal Structures

Characterize and sketch crystals of the various crystals you have prepared. Use colored pencils.

TABLE 4.8 **Sulfur**

	Powdered sulfur from reagent bottle	Sulfur crystals formed in filter cone	Amorphous sulfur in water (Look up this term.)
Sketch			
Characterize			

TABLE 4.9 **Iodine**

	Crystals from reagent bottle	Crystals from watch glass
Sketch		
Characterize		

18. What supporting evidence of Dalton's atomic theory does the formation and growth of crystals suggest? (Base your answer on the observation you noted in "Part IV – Observation of Physical Changes and Crystal Structures" of the report.)

EXPERIMENT 5

Nomenclature

Background

This set of notes is designed to accompany the NOMENCLATURE lectures that your instructor will be giving throughout the semester. The purpose of these notes is to provide a summary of the main points to be covered in each of these lectures and to give you some examples and practice before you do your homework assignments. More detailed information concerning nomenclature can be found in your textbook.

Chemical compounds are designated both by a name and a formula. The name usually gives clues concerning the composition of the substance, but the formula indicates exactly what elements are a part of the compound and in what proportions. Learning how to name the elements and their compounds and to write the corresponding formula is an important part of learning chemistry.

Throughout the several parts below you will find sections labeled for example, “P.G. I-1.” As in the textbook, P. G. stands for performance goal; the I -1 refers to the section and the item number. The P. G.'s indicate what information and skills that you are expected to have upon mastering the given material.

Part I: Naming Binary Ionic Compounds

At the heart of learning to write names and formulas is the need to know the names and symbols of a fair number of the chemical elements.

P. G. I–1

Know and be able to use the **names and symbols** for the elements. Your instructor will indicate which names and symbols you will be expected to know. The names must be learned with the correct spelling. You will find a periodic chart with the elements and their atomic numbers inside your textbook and an alphabetical listing of the elements and their symbols as well. The periodic chart, which has these symbols listed in certain patterns, will always be available during quizzes and examinations. This should be of some help!

EXAMPLE 5.1

The element *carbon* has the symbol of C

The element *copper* has the symbol of Cu

Therefore we can say that formula for carbon is C and that of copper is Cu

Writing Chemical Formulas

All elements are composed of extremely small particles called atoms. An **atom** is the smallest particle that has the properties of the element. All of the atoms in an element are neutral, that is, they have no electrical charge. Each atom is made up of an equal number of positive and negative subatomic particles. The positive particles are called **protons** and the negative particles **electrons**. Because the magnitude of the charge on these subatomic particles is exactly the same, the amount of negative charge just exactly balances the amount of positive charge, and so the atom is electrically neutral. Atoms can either lose or gain a few electrons in chemical reactions. Some atoms, usually metals, lose electrons and thereby become positive in overall charge; others gain electrons and so become negative in charge. An atom which has either lost or gained electrons and is no longer electrically neutral is called an **ion**. Many compounds are formed when positive and negative ions combine; they are classified then as **ionic compounds**. A large percentage of the compounds studied here are ionic compounds.

Binary compounds are those in which only two different elements are combined. The formula for a compound is produced by writing the symbol of each element of the compound next to each other with no spaces in-between. The relative numbers of the various atoms present in the compound are shown with a $_{\text{subscript}}$ following the atom. If only one atom is present, then no subscript is written and the number one is implied.

EXAMPLE 5.2

$CaCl_2$ represents the compound calcium chloride which contains one calcium (Ca) atom and two chlorine (Cl) atoms.

K_3PO_4 represents potassium phosphate which contains three potassium (K) atoms, one phosphorus (P) atom and four oxygen (O) atoms.

$Fe_2(SO_4)_3$ represents iron(III) sulfate which contains two iron (Fe) atoms, three sulfur (S) atoms and 12 oxygen (O) atoms.

P. G. I - 2

Be able to tell the number(s) of each kind of atom in a chemical formula.

Ionic compounds are held together by the electrostatic forces of attraction that occur between positive and negative ions. Ions that consist of only one charged atom are called monatomic ions. The formula of a monatomic ion is written by putting a signed number as a $^{\text{superscript}}$ just to the right of the element symbol which shows the charge of that ion.

EXAMPLE 5.3

K^+ represents the potassium ion with a one positive charge,

Fe^{3+} represents iron with a positive three charge,

O^{2-} represents an oxide ion with a two negative charge, and

F^- represents a fluoride ion with a one negative charge.

The formulas for ionic compounds must be written so the charge of the positive ion, called the ***cation*****,** just balances the charge of the negative ion, called the ***anion***.

EXAMPLE 5.4

Two K^+ cations (total positive charges of +2) combine with *one* O^{2-} anion (charge of -2) to form a neutral compound, K_2O

Two Al^{3+} cations (total charge of +6) combine with *three* O^{2-} anions (total charge of –6) to form Al_2O_3

Observe that the completed formula for the compound does not show any positive or negative signs. It simply shows the elements in the compound and their relative numbers.

In order to write the formulas of the ionic binary compounds, you must know what charge a given atom will have in its compounds. Many elements may have more than one possible charge in their various compounds. In this first part (and in “Part II: Naming Ternary Compounds Containing The “–ate” Anions” on page 79 and “Part III: The “–ite” Anions” on page 81) only those elements which always have the same charge whenever they form compounds will be considered. (Remember, atoms in their elemental form are neutral, it is only when they form compounds that they become charged.) The charge taken by several groups of atoms when they become ions can easily be determined from their position on the Periodic Table. Note that the elements are arranged by **groups or families** that are designated by Roman numerals and by **periods**. (These group numbers correspond to those shown on the Periodic Table given in your textbook and to the one you will be given to use during tests and exams.) The correlation is as follows:

TABLE 5.5

Group Number	Ion Charge	Comments
IA/1A/1	+ 1	hydrogen is -1 when combined with metals
IIA/2A/2	+ 2	
IIIA/3A/13	+ 3	Tl is an exception but not one that you need to know; it can have a +1 or +3 charge.
VA/5A/15	- 3	*only* when combined with metals
VIA/6A/16	- 2	*only* when combined with metals
VIIA/7A/17	- 1	*only* when combined with metals
In addition there are three other cations, silver = +1; zinc = +2 and cadmium = +2.		

P. G. I - 3

Learn the correlations between the elements, their charges as ions and their location on the Periodic Table.

Writing Names From Formulas

To write the name for a binary ionic compound, first determine the name for the two elements in the formula. The name of the first element, always the cation, is unchanged in the compound name; the second element, always the anion will have a new ending, "-ide." The order of each name is the same as in the formula.

TABLE 5.6

$MgBr_2$ magnesium and bromine are the two elements. The name then will be magnesium bromide.

Al_2S_3 aluminum and sulfur are the two elements. The name is aluminum sulfide.

Note that no prefixes such as di-, tri-, etc., are used in the name to indicate the numbers of each kind of atom in the compound. The rule is to **NOT** *use prefixes when naming binary ionic compounds. Non-ionic binary compounds such as CO_2 do make use of prefixes, however these types of compounds will be considered later.*

Writing Formulas From Names

To write the formula from the name, first determine the correct symbol for each atom and what charge that atom will have in the combined form. Then combine the two kinds of ions in the *smallest whole number proportion* that will allow the that total positive charge to cancel out the total negative charge. This then will be the formula of the compound. Remember that the cation always comes first and anion second in both the name and the formula.

EXAMPLE 5.7

Lithium nitride The two ions are Li^{1+} and N^{3-}. Three lithium cations are needed to balance the negative charge of one nitride ion. 3 × (+1) = +3, which just balances the nitride ion.

The formula then is Li_3N

EXAMPLE 5.8

Magnesium fluoride The ions are Mg^{2+} and F^-. Two fluoride ions (2 × (–1) = –2) are required to balance the charge of one magnesium ion.

The formula is MgF_2.

EXAMPLE 5.9

Aluminum oxide The ions are Al^{3+} and O^{2-}. The smallest number that both 2 and 3 will go into evenly is six. Therefore, two aluminum ions (2 × (+3) = +6) and three oxide ions (3 × (–2) = –6) are the least number that can be combined to cancel out all of the charge.

The formula is then Al_2O_3.

P. G. I - 4

Know and be able to use the terms in the Table below. They will be covered in the nomenclature lecture and can be found in the glossary of your textbook.

TABLE 5.10

anion
cation
halogen
halide
metal
nonmetal
alkali metal

Additional information and many more worked examples may be found in your textbook.

Part II: Naming Ternary Compounds Containing The "–ate" Anions

Previously, you learned how to write the names and formulas for binary compounds. The binary compounds are made out of monatomic ions. In this exercise the names and formulas for some ternary compounds containing polyatomic anions will be studied. As the name implies, a polyatomic anion has a negative charge and is made out of two or more atoms. An example is the phosphate ion which consists of one P and four O's and has an overall electrical charge of negative three. The formula is written as PO_4^{3-}

Note that unlike the subscript which refers only to the preceding atom or group, the superscript refers to the whole anion. That is, the *whole ion* consisting of one P and four O's together have a charge of negative three. Table 5.11 on page 79 lists seven common polyatomic ions.

TABLE 5.11

BO_3^{3-}	borate	CO_3^{2-}	carbonate
NO_3^{-}	nitrate	SiO_3^{2-}	metasilicate
PO_4^{3-}	phosphate	SO_4^{2-}	sulfate
ClO_3^{-}	chlorate		

P. G. II - 1

Learn the names of these anions, their formulas and their charges. There are a number of other "–ate" ions that can be named because they have the same pattern or form as three of these seven anions. This relationship to the three anions can be seen by their placement on the periodic table.

The three pattern ions from the list of seven are shown in **bold face type** on Table 5.12 on page 80. There are five additional ions whose names and formulas have the same patterns as the ones above them in period three. Stated a different way, all of the ions from groups VA, VI and VIIA in period three and below have the similar formulas and charge. Note also that for the ions in period three and below that these "-ate" ions have the same charge as the "-ide" ions of that family.

TABLE 5.12

Period Number	Group Number					
	VA	ion	VIA	ion	VIIA	ion
2	NO_3^-	nitrate	-----		-----	
3	**PO_4^{3-}**	**phosphate**	**SO_4^{2-}**	**sulfate**	**ClO_3^-**	**chlorate**
4	AsO_4^{3-}	arsenate	SeO_4^{2-}	selenate	BrO_3^-	bromate
5	-----		TeO_4^{2-}	tellurate	IO_3^-	iodate

P. G. II - 2

Be able to name, write and use the five additional anions shown in the illustration given above.

EXAMPLE 5.13

Give the name for $Ca(ClO_3)_2$.

The cation name is unchanged from that of the element and so is calcium. Notice that the formula for the anion is enclosed in parentheses. This is necessary when the formula contains more than one of the polyatomic ion. The anion name, from the list of seven, is chlorate and so the compound name is calcium chlorate.

EXAMPLE 5.14

Write the formula for zinc arsenate.

The cation zinc is Zn^{2+} *and the anion is* AsO_4^{3-}*. The formula for the cation is determined directly from the Periodic Chart. Arsenate is not one of the seven anions to be memorized but since arsenic is directly below phosphorus in the table, it will have a formula analogous to phosphate.*

In order to make a neutral compound, three zinc cations (total charge = +6) and two arsenate anions (total charge = –6) are needed. Again note the use of the parentheses around the polyatomic anion to show the presence of two anions.

The completed formula then is then $Zn_3(AsO_4)_2$.

P. G. II–3

Learn the names and formulas for the following new anions and add them to your list.

TABLE 5.15

Anion Formula	Anion Name
OCN^-	cyanate
SCN^-	thiocyanate
$Cr_2O_7^{2-}$	dichromate

Compare the first two ions on this list and note that the second one is just like the first except that it has a sulfur atom in place of the oxygen. There are several compounds and ions in which sulfur substitutes for an oxygen atom; when this occurs, the prefix “thio-” is usually added to the name of the compound from which it is derived. Another example would be thiosulfate, $S_2O_3^{2-}$, wherein one oxygen of sulfate has been replaced with a sulfur atom.

P. G. II - 4

New Terms. Know and be able to use the terms in the Table 5.16 on page 81 below which will be explained by your instructor.

TABLE 5.16

polyatomic
alkaline earth
chalcogen
chalcogenide
metalloid
analogous
oxyanion

Part III: The “–ite” Anions

A number of oxyanions exist that contain one less oxygen than the ion with an “-ate” ending. The table shows several examples. The only difference between an anion ending in “-ite” is that it has one less oxygen atom than the corresponding anion ending in “-ate.” As shown with the third example, not all anions ending in “-ate” have corresponding “-ite” anions.

TABLE 5.17

“–ate” ion		“–ite” ion	
SO_4^{2-}	sulfate	SO_3^{2-}	sulfite
NO_3^-	nitrate	NO_2^-	nitrite
CO_3^{2-}	carbonate	no corresponding ion	

P. G. III–1

Learn the rule that anions ending in “-ite” are just like the corresponding anion ending in “-ate” except that they contain one less oxygen atom.

In order to use this rule you must have learned (memorized) the names for the anions ending in “-ate.”

TABLE 5.18

Give the correct name for $KClO_2$.

The cation, K^+ is of course potassium. the anion, ClO_3^-, corresponds to the “-ate” ion, ClO_2^- except that it contains one less oxygen atom, therefore, it must be called chlorite. The compound's name then is potassium chlorite.

TABLE 5.19

Write the formula for sodium arsenite.

The cation, sodium is Na^+. The anion will be just like arsenate with one less oxygen atom. This will be then AsO_3^{3-}. So, the compound formula is Na_3AsO_3.

P. G. III–2

Learn the names and formulas of two new anions. They are: $C_2O_4^{2-}$ called oxalate and $C_2H_3O_2^-$ called acetate.

Part IV: Further Oxyanions of the Halogens

In Part II ions such as bromate, BrO_3^-, and chlorate, ClO_3^- were introduced. The halogens chlorine, bromine and iodine (but not fluorine) make several more oxyanions such as the following ones shown for chlorine.

Note that hypochlorite is *often* written as **OCl^-** in formulas, e.g. **NaOCl** instead of NaClO.

An analogous set of names and ion formulas can be made for bromine and iodine, *although not all of the bromine and iodine analogues actually exist*. In order to name these anions and the compounds in which they are found, you must use what you have learned in Part II and III and then add the following new rules.

TABLE 5.20

Ion Formula	Ion Name	Number of Oxygen Atoms	Anion Pattern
ClO^-	*hypo*chlorite	-2	hypo...ite
ClO_2^-	chlorite	-1	...ite
ClO_3^-	chlorate	-----	...ate
ClO_4^-	*per*chlorate	+1	per...ate

P. G. IV–1

Learn the following rules:

1. If an oxyanion has one less oxygen than an ion that ends in “**-ite**,” then place the prefix “hypo-” directly in front of the “-ite” name.
2. If the oxyanion has one more oxygen than an ion ending in “-ate,” then the “-ate” ion has the prefix “per-” placed in from of the “-ate” name.

ClO_4^- has one more oxygen than chlorate, ClO_3^-, and so it is called perchlorate.

ClO^- has one less oxygen than chlorite, ClO_2^-, and so it is called hypochlorite.

The same method works for the oxyanions of bromine and iodine. Some further examples are: given below:

TABLE 5.21

Name $Ca(IO)_2$.

Iodite is IO_2^- and so this anion must be hypoiodite and the compound is calcium hypoiodite.

Write the formula for iron(III) periodate.

Periodate has one more oxygen than iodate, IO_3^- so the formula would be $Fe(IO_4)_3$.

P. G. IV–2

Learn the names of the following new anions and add then to your current list of nine.

TABLE 5.22

Anion Formula	Anion Name
CrO_4^{2-}	chromate
CN^-	cyanide
OH^-	hydroxide

Two of ions in this group do not follow the rules for naming binary compounds that were given previously. They are the cyanide and hydroxide ions. Although they form ternary compounds their names have the binary form in that the anions end in "-ide." Two examples of their use are:

KCN is potassium cyanide and $Ca(OH)_2$ is calcium hydroxide.

These two anions are the only common exceptions to the rule of ending only binary compounds with "-ide."

P.G. IV–3

Learn the names and formulas of the two new ions and the common name for NH_3.

TABLE 5.23

NH_4^+, ammonium ion, a polyatomic cation	NH_3 ammonia
MnO_4^-, permanganate	

Be sure to keep *ammonia* the compound separate from *ammonium* the ion.

Part V: Cations With Variable Oxidation Numbers

In the first three exercises, only cations which always have the same charge in their compounds have been used. Many cations however, have different charges in their various compounds. Many of these elements, but not all, are the metals of the "B" groups or transition elements. Consider two examples:

(i) **$FeCl_2$**, a light green solid and (ii) **$FeCl_3$**, a brown solid.

Since the chloride ion, Cl^-, has a negative one charge, the iron cation in (i) must have a charge of "+2" and the iron in (ii) must have a "+3" charge. If the method of naming binary compound is applied, both of these distinctly different compounds would be called "iron chloride". Some other method must be used so that each one has a distinctive name.

The Stock System

The problem is solved in a straightforward manner using the **STOCK** system. This system is used for all metals that may have more than one possible *positive oxidation* number in any of their compounds. The term oxidation number refers to the positive charge of the cation in that particular compound. (The term oxidation number has a more general meaning than this; the concept will be expanded at a later time.) To use the Stock system, the oxidation number of the cation (the metal) is written using a Roman Numeral set in a parentheses just behind the metal name. So then, Fe^{2+} is called iron(II) ion, pronounced "iron two ion" and Fe^{3+} is called the iron(III) ion. The Roman Numeral in the parentheses is called the Stock number. (For more detailed information refer to your textbook.)

P. G. V–1

Know and be able to use the following rules:

1. All metals that may have more than one possible positive oxidation number (i.e. charge) are to have their Stock number included as a part of the cation name.
2. All metals that always have a fixed oxidation number either as ions or in compounds are not to a Stock number included with their cation names.
3. All of the metals that are listed in the table in part I are the metals that have fixed charges. All other metals have variable charge and are to have their Stock number included in their names.

Several examples are show in Table 5.24 on page 84:

TABLE 5.24

Formula	Oxidation Number of the Cation	Compound Name Stock System
K_2CO_3	+ 1	potassium carbonate
$FeCl_2$	+ 2	iron(II) chloride
$FeCl_3$	+ 3	iron(III) chloride
$CrPO_4$	+ 3	chromium(III) phosphate
$Al_2(SO_4)_3$	+ 3	aluminum sulfate
CuO	+ 2	copper(II) oxide

As shown in Table 5.24 on page 84, $Al_2(SO_4)_3$ and K_2CO_3 do not have Stock numbers in their names because the cations may have only one possible positive oxidation number. Note that when writing the Stock names that there are no spaces between the name of the cation, the parentheses and the Roman Numerals; this is illustrated below for the compound Fe_2O_3.

EXAMPLE 5.25

Write the name for Fe_2O_3.

To figure out the charge on Fe remember that the charge on O is generally -2, so bring the subscripts up the charges become +3 for the iron and -2 for the oxygen. (This usually works, but not always. Remember that subscripts for ionic compounds should be also low as possible so that the overall charge sums to zero.)

$$(Fe^{3+})_2(O^{2-})_3$$

The name iron oxide does not specify the charge. Thus, is ***incorrect****.*

Iron (III) oxide would be incorrect because no space is needed between the name iron and the parenthesis. Also, ***incorrect****.*

Iron(III) oxide would be ***correct****.*

In order to use the Stock system you must determine the oxidation number of the cation. This is done from the charge on the **anion** and the principle that all compounds have no charge. Be aware that in the Stock system the oxidation number is given **only** for the element with the positive charge and never for the anion.

In a later exercise the stock system will be extended to compounds in which the first element is a nonmetal. The Stock system has the advantage of being clear and easy to use. To see the truth of this statement, compare it to an older system which is still used enough that you should be familiar with it.

An older method

This method, which may be called the "-ous and -ic" method makes use of suffixes to the name of the metal cation to distinguish between two different oxidation states. Additionally, use is often made of the Latin name for the element instead of the English name. For example, Fe^{2+} is called the "ferr*ous*" ion (instead of iron(II) as in the Stock system) and $FeCl_2$ is then called ferrous chloride. The iron(III) ion is called the "ferr*ic*" ion and $FeCl_3$ is called ferric chloride. Some further examples are shown in the next table.

TABLE 5.26

Ion	Ion Name	Compound Name	
Fe^{2+}	ferrous	$FeSO_4$	ferrous sulfate
Fe^{3+}	ferric	FeI_3	ferric iodide
Cu^{+}	cuprous	Cu_2O	cuprous oxide
Cu^{2+}	cupric	$CuCl_2$	cupric chloride
Sn^{2+}	stannous	SnF_2	stannous fluoride
Sn^{4+}	stannic	SnO_2	stannic oxide
Pb^{2+}	plumbous	$PbCl_2$	plumbous chloride
Pb^{4+}	plumbic	PbO_2	plumbic oxide
Hg^{2+}	mercuric	$HgBr_2$	mercuric bromide

Therefore, the compounds in Table 5.24 on page 84 can be named two ways as shown below:

TABLE 5.27

Formula	Oxidation Number of the Cation	Compound Name Stock System	Compound Name Latin System
K_2CO_3	+ 1	potassium carbonate	potassium carbonate
$FeCl_2$	+ 2	iron(II) chloride	ferrous chloride
$FeCl_3$	+ 3	iron(III) chloride	ferric chloride
$CrPO_4$	+ 3	chromium(III) phosphate	chromic phosphate
$Al_2(SO_4)_3$	+ 3	aluminum sulfate	aluminum sulfate
CuO	+ 2	copper(II) oxide	cupric oxide

P. G. V– 2

Learn how to use the Stock system in naming chemical compounds.

P. G. V–3

Learn the names of the six cations in the “archaic system” as shown in the table above and be able to use them in naming compounds.

Any compound which contains the hydroxide anion is called a **base** or sometimes an **alkali**. A base will react with and neutralize an acid to produce a **salt** and water. A salt is a compound made up of a cation and an anion provided that the cation is not H^+ and that the anions are not either OH^- or O^{2-}. Obviously then most of the compounds involved in these nomenclature exercises are classified as salts.

P. G. V–4

Know and be able to use the new terms, *base, alkali* and *salt*.

In this exercise there are two new ions that take a form not yet used. The first is the mercury(I) cation. It's formula is Hg_2^{2+}, not Hg^+ as you might expect. In the mercury(I) cation there are two mercury atoms bonded together with an overall charge of +2, which can be illustrated as:

$$(Hg\text{-}Hg)^{2+} \qquad \text{(EQ 5.1)}$$

Therefore, mercury(I) cation is *always* written as Hg_2^{2+} or as some multiple thereof. Some examples are shown.

TABLE 5.28

Name		Formula
mercury(I) iodide	mercurous iodide	$\mathbf{Hg_2I_2}$ *not* HgI
mercury(I) nitrate	mercurous nitrate	$\mathbf{Hg_2(NO_3)_2}$ *not* $HgNO_3$

Another ion which is similar in some ways is the **peroxide** anion, O_2^{2-}. It has a structure similar to the mercury(I),

$$(O\text{-}O)^{2-} \qquad \text{(EQ 5.2)}$$

in which two oxygens are covalently bonded together but have an overall charge of -2. Some examples are given in Table 5.29 on page 87.

TABLE 5.29

Name	Formula
hydrogen peroxide	$\mathbf{H_2O_{2e}}$ *not HO*
cesium peroxide	$\mathbf{Cs_2O_2}$ *not CsO*
barium peroxide	$\mathbf{BaO_2}$

Peroxides (inorganic) form only with the elements of groups IA and IIA. Compounds such as TiO_2 or PbO_2 are not peroxides because the metals are not in groups IA or IIA. The cations in these compounds must have an oxidation number of 4+ and would be named titanium(IV) oxide and lead(IV) oxide, respectively.

Part VI: Hydrates

Several ionic compounds contain what is called **water of crystallization**. When the substance is crystallized from water, some of the water is brought along into the crystal. An example is copper(II) sulfate which includes five molecules of water in the crystal structure of the compound. This compound is a dry blue crystalline solid. If it is heated to a little over 150°C, the water is expelled and the remaining solid turns white. Its formula is written as:

$$CuSO_4 \bullet 5\ H_2O \tag{EQ 5.3}$$

In order to show that the water is a part of the compound yet has some type of its own independent existence, the number of waters of crystallization are shown by suffixing "• 5 H_2O" to the formula for the compound. The following illustrations show both how the formula is written and how the substances are named.

TABLE 5.30

Formula	Name
$Cr(NO_3)_3 \bullet 9\ H_2O$	chromium(III) nitrate nonahydrate
$Na_2SO_3 \bullet 7\ H_2O$	sodium sulfite heptahydrate
$KF \bullet 2\ H_2O$	potassium fluoride dihydrate

As can be seen, the name for the compound is written as it would be for the **anhydrous compound** and then the number of waters present is suffixed to this name by using the Greek prefix for the number of waters followed by the word hydrate. A **hydrate** is a salt that precipitates with water molecules in the crystal lattice.

Part VII: Binary Nonmetal Compounds

Discusses the prefix method for naming binary nonmetal compounds and couple of new anions.

The Prefix Method for Naming Binary Nonmetal Compounds

This method is simple and easy to use but is restricted to compounds made up of two nonmetals. As in all binary compound names, the name of the first element is unchanged (except for the prefix) and the second element ends in "-ide." The number of atoms of each type is denoted by using the Greek prefix for that number. These nonmetal compounds are bonded together covalently; they do not have ions and so there are no cations and anions. Therefore, you will notice that on the practice sheets and on the computer exercises no provision is made to write out the ion formulas.

P.G. VII–1

The prefixes and their corresponding numbers are listed in Table 5.31 on page 88. Know and be able to use them in naming and writing formulas of compounds.

TABLE 5.31

mono	1	hexa	6
di	2	hepta	7
tri	3	octa	8
tetra	4	nona	9
pent	5	deca	10

Examples are given below:

TABLE 5.32

Formula	Name
N_2O_5	dinitrogen pentoxide
N_2O	dinitrogen monoxide
SF_6	sulfur hexafluoride
CO	carbon monoxide

Note that in compounds which have only one atom of the first element, the “mono-” prefix is not used. This is illustrated by the last two entries in the table. When the prefix ends with an “a” or an “o” and the next syllable starts with a vowel, then the “a” or “o” is usually dropped from the prefix. For example CO is called carbon monoxide not carbon monooxide.

TABLE 5.33 **Some Common Names:**

Formula	Name
$H_2O_{(l)}$	water
$CH_{4\,(g)}$	methane
$C_2H_{4\,(g)}$	ethylene
$C_2H_{2\,(g)}$	acetylene
$NH_{3\,(g)}$	ammonia
$PH_{3\,(g)}$	phosphine
$N_2O_{(g)}$	nitrous oxide (laughing gas)
$N_2H_{4\,(g)}$	hydrazine
$NH_2OH_{(s)}$	hydroxylamine
$NO_{(g)}$	nitric oxide

Part VIII: Acids

Compounds that contain hydrogen in the cation position act are **acids** when dissolved in water. Compounds such as C_2H_6 and $Ca(OH)_2$ contain hydrogen but they are not acidic. This can be seen in that the hydrogens do not occupy the cation position in the formula. There are two main classes of acids, (i) **binary** acids which contain hydrogen and one other element and (ii) **ternary** acids which contain hydrogen and two other elements. (Some acids contain more than three elements but they are named by the rules for the ternary acids). For binary acids, you can easily recognize the formula of the acid as it will have one or more hydrogens in the cation position and the “(aq)” immediately following the formula showing that the substances is dissolved in water. The ternary compounds are almost always named as acids whether or not there is an “(aq)” attached to the formula.

Binary Acids

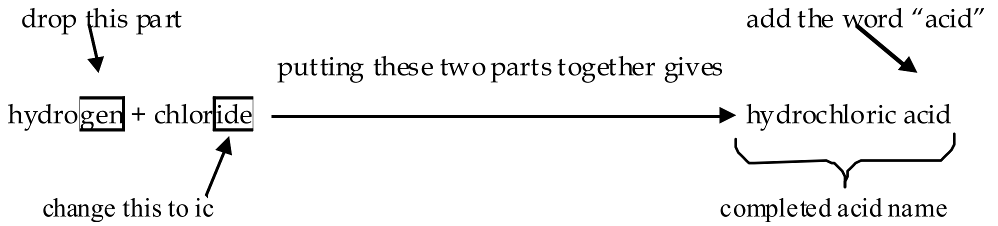
Consider some examples. HCl, HBr and H_2S all have acidic properties when dissolved in water. In their pure state at standard temperature and pressure, STP, (0 °C and 1 atm) they are gasses. The following table shows their names as both the compound and the acid when dissolved in water.

TABLE 5.34

Formula	Compound Name	Acid Name
$HCl_{(aq)}$	hydrogen chloride	hydrochloric acid
$HBr_{(aq)}$	hydrogen bromide	hydrobromic acid
$H_2S_{(aq)}$	hydrogen sulfide	hydrosulfuric acid
$HCN_{(aq)}$	hydrogen cyanide	hydrocyanic acid

To derive the name for a binary acid from its compound name, drop the “-gen” from the hydrogen, change the “-ide” to “-ic” on the anion name, put these two pieces together to form one word and then add the word acid. Perhaps this can be illustrated better with the diagram:

FIGURE 5.1



$H_2S_{(aq)}$ does not quite fit the pattern, however. The “sulfide” becomes “sulfuric” not ~~sulfic~~ as the rule suggests. The acid is then named *hydrosulfuric acid.*

Ternary Acids

Ternary acids are composed of a polyatomic oxyanion such as SO_4^{2-}, and hydrogen ions. They are named somewhat differently than the binary acids. The name of ternary acids is derived from the name of the oxyanion. If the name of the oxyanion ends in “-ate” the acid name is obtained by changing the ending of the anion from “-ate” to “-ic” and adding the word acid. Unlike the binary acid names, there is no “hydro” in the name at all. If the oxyanion ends in “-ite” then the ending is changed to “-ous” and the balance of the procedure remains the same. This process is shown in the following diagram for HNO_3, “hydrogen nitrate.” It should be noted that many of these “compounds” do not exist in the pure state, but only as acids in aqueous solution.

TABLE 5.35

Acid Formula	Oxyanion	Anion Name	Modified Anion Name	Acid Name
$HNO_{3\ (aq)}$	NO_3^-	nitr**ate**	nitr**ic**	nitric acid
$H_2SO_{3\ (aq)}$	SO_3^{2-}	sulf**ite**	sulfur**ous**	sulfurous acid
$HClO_{\ (aq)}$	ClO^-	hypochlor**ite**	hypochlor**ous**	hypochlorous acid
$HIO_{4\ (aq)}$	IO_4^-	period**ate**	period**ic**	periodic acid

Note again that sulfuric acid does not follow the rules exactly, like hydrosulfuric acid did not. This irregularity can soon be learned and should present no problem. As you can also see, the rules work just as well with the “per-” and “hypo-” compounds. Some of the acids such as $HClO_{\ (aq)}$ cannot be obtained in the pure state and exist only in aqueous solutions. Others such as $H_2SO_{4\ (l)}$, and $HC_2H_3O_{2\ (l)}$ are stable and they exist in the pure state.

To summarize the tyes of acids and their names:

TABLE 5.36

Acid Formula	Ion Formula	Ion Name	Anion Pattern	Acid Pattern	Number of Oxygen Atoms	Acid Name
$HCl_{\ (aq)}$	Cl^-	chloride	-ide	hydro...ic	None	hydrochloric acid
$HClO_{\ (aq)}$	ClO^-	*hypo*chlorite	hypo...ite	hypo...ous	-2	hypochlorous acid
$HClO_{2\ (aq)}$	ClO_2^-	chlorite	...ite	...ous	-1	chlorous acid
$HClO_{3\ (aq)}$	ClO_3^-	chlorate	...ate	...ic	-----	chloric acid
$HClO_{4\ (aq)}$	ClO_4^-	*per*chlorate	per...ate	per...ic	+1	perchloric acid

Below are some common inorganic acid:

Most Common Inorganic Acids (* detonates a strong acid)

Formula	Name
$HBrO_{3\ (aq)}$	Bromic Acid
$H_2CO_{3\ (aq)}$	Carbonic Acid
$HClO_{3\ (aq)}$	Chloric Acid*
$HBr_{\ (aq)}$	Hydrobromic Acid*
$HCl_{\ (aq)}$	Hydrochloric Acid*
$HI_{\ (aq)}$	Hydroiodic Acid*
$HNO_{3\ (aq)}$	Nitric Acid*
$HClO_{4\ (aq)}$	Perchloric Acid*
$H_3PO_{4\ (aq)}$	Phosphoric Acid
$H_2SO_{4\ (aq)}$	Sulfuric Acid* (to HSO_4^- + H^+)

Strong acids the first proton dissociates 100% in aqueous solution. In chemical equations where ions are included (total and net ionic equations) write the ionized components, since they are the major species present in the solution.

$$HNO_{3\ (aq)} \rightarrow H^+_{\ (aq)} + NO_3^-{}_{\ (aq)} \qquad \textbf{(EQ 5.4)}$$

Weak acids do not dissociate 100% in aqueous solution, they dissociate anywhere from 1-5%. In chemical equations where ions are included (total and net ionic equations) write the complete weak acid formula, not the ions. The ions are only a minor component in the total solution.

$$HBrO_{3\,(aq)} \rightleftharpoons H^{+}{}_{(aq)} + BrO_3^{-}{}_{(aq)} \qquad \textbf{(EQ 5.5)}$$

Monoprotic acid has one ionizable hydrogen atom.

acetic acid, a weak acid, $HC_2H_3O_{2\,(aq)}$ **(EQ 5.6)**

the first hydrogen is the ionizable hydrogen atom, the other three hydrogen atoms are nonionizable.

$$HC_2H_3O_{2\,(aq)} \rightleftharpoons H^{+}{}_{(aq)} + C_2H_3O_2^{-}{}_{(aq)} \qquad \textbf{(EQ 5.7)}$$

Acetic acid is an organic acid called a carboxylic acid; Often its formula is written to show its structural connectivity of the molecule: $CH_3COOH_{(aq)}$

$$CH_3COOH_{(aq)} \rightleftharpoons H^{+}{}_{(aq)} + CH_3COO^{-}{}_{(aq)} \qquad \textbf{(EQ 5.8)}$$

when written this way the three hydrogens are the nonionizable hydrogen atoms, while the hydrogen atom at the end is the ionizable hydrogen atom.

Hydroiodic acid, a strong acid, $HI_{(aq)}$, and therefore dissociates 100%:

$$HI_{(aq)} \rightarrow H^{+}{}_{(aq)} + I^{-}{}_{(aq)} \qquad \textbf{(EQ 5.9)}$$

Diprotic acids have two ionizable hydrogen atoms.

Sulfuric acid, a strong acid, $H_2SO_{4\,(aq)}$, has two ionizable hydrogen atoms. The first hydrogen atom ionizes 100%:

$$H_2SO_{4\,(aq)} \rightarrow H^{+}{}_{(aq)} + HSO_4^{-}{}_{(aq)} \qquad \textbf{(EQ 5.10)}$$

The second hydrogen atom ionizes as a weak acid:

$$HSO_4^{-}{}_{(aq)} \rightleftharpoons H^{+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \qquad \textbf{(EQ 5.11)}$$

Triprotic acids have three ionizable hydrogen atoms.

Phosphorous acid, a weak acid, $H_3PO_{3\,(aq)}$, has three ionizable hydrogens:

$$H_3PO_{4\,(aq)} \rightleftharpoons H^{+}{}_{(aq)} + H_2PO_3^{-}{}_{(aq)} \qquad \textbf{(EQ 5.12)}$$

$$H_2PO_3^{-}{}_{(aq)} \rightleftharpoons H^{+}{}_{(aq)} + HPO_3^{2-}{}_{(aq)} \qquad \textbf{(EQ 5.13)}$$

$$HPO_3^{2-}{}_{(aq)} \rightleftharpoons H^{+}{}_{(aq)} + PO_3^{3-}{}_{(aq)} \qquad \textbf{(EQ 5.14)}$$

Part IX: Compounds Containing The Acid Anions

Anions which have a charge of -2–or -3 can combine with a hydrogen in the +1 state to form what is being called here an “acid anion.” This produces another anion which can combine with a cation to form a compound. For example the carbonate ion, CO_3^{2-}, can combine with a +1 hydrogen to produce the anion HCO_3^{-}. This ion can then combine with K^{+} to form $KHCO_3$. The product,

$KHCO_3$ is a salt (which when dissolved in water gives the ions K^+ and HCO_3^-) but the H in the HCO_3^- is weakly acidic and $KHCO_3$ can be used to neutralize a base.

These anions are named as shown in the table below. They can be called either by their systematic name or their common name, except for those based on phosphate which must be named as indicated. The ones shown in the table are only examples; there are others which are named in the same way.

TABLE 5.37

Anion	Acid Anion	Systematic Name	Common Name
CO_3^{2-}	HCO_3^-	hydrogen carbonate	bicarbonate
SO_3^{2-}	HSO_3^-	hydrogen sulfite	bisulfite
SO_4^{2-}	HSO_4^-	hydrogen sulfate	bisulfate
PO_4^{3-}	HPO_4^{2-}	hydrogen phosphate	------
$H_2PO_4^-$		dihydrogen phosphate	-----

Compounds made from these ions are named as follows.

TABLE 5.38

Formula	Systematic Name	Common Name
$NaHCO_3$	sodium hydrogen carbonate	sodium bicarbonate
K_2HPO_4	potassium hydrogen phosphate	
KH_2PO_4	potassium dihydrogen phosphate	
$Pb(HSO_4)_2 \bullet H_2O$	lead(II) hydrogen sulfate monohydrate plumbous hydrogen sulfate monohydrate	lead(II) bisulfate monohydrate plumbous bisulfate monohydrate

As shown in Table 5.37 and Table 5.38 above, the phosphate compounds (and other ions with a "-3" charge) cannot be named by the common name system. This is because the phosphate ion can combine with either one or two hydrogen ions and it is necessary to indicate this in the name. The correct method is shown in the tables.

Summary of Inorganic Nomenclature

FIGURE 5.2 Flow Chart for Inorganic Nomenclature

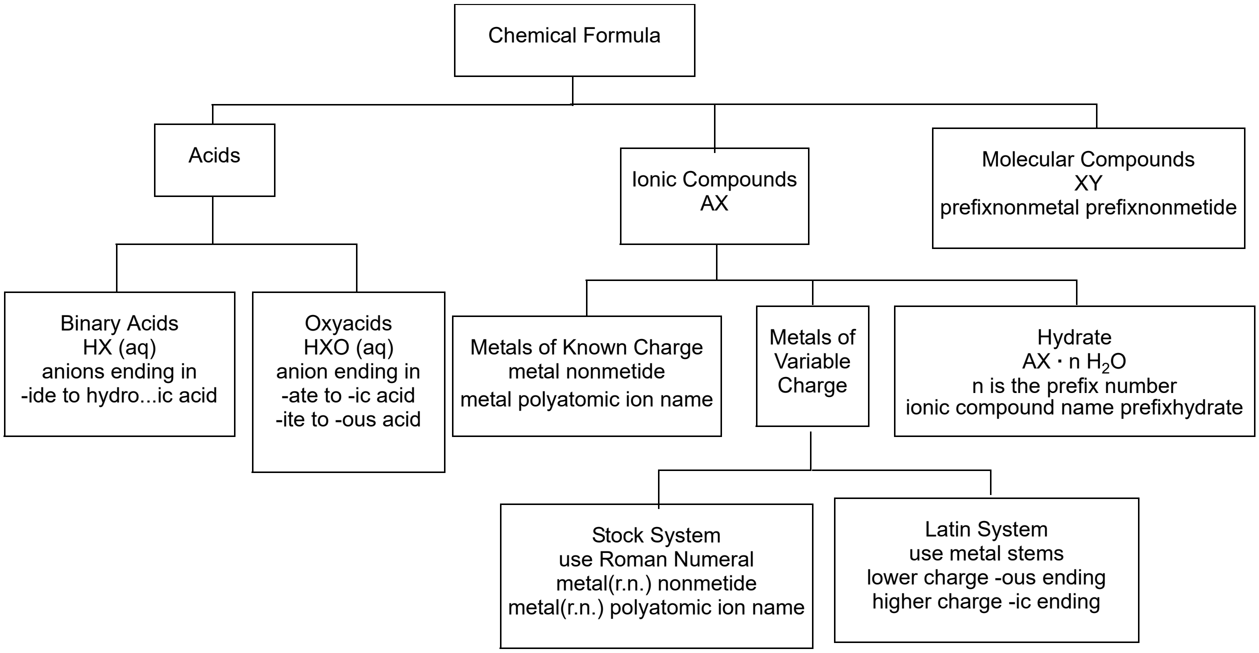


TABLE 5.39

	1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Roman Numeral (r.n.)	I	II	III	IV	V	VI	VII	VIII	IX	X
Greek Prefix	mono	di	tri	tetra	penta	hexa	hepta	octa	nona	deca

In their elemental states metals exist as monoatomic atoms. There are seven nonmetals that exist as diatomic molecules in their elemental state. Mnemonic: **H**ave **N**o **F**ear **O**f **I**ce **C**o**l**d **B**ee**r**

Hydrogen (H_2), Nitrogen (N_2), Fluorine (F_2), Oxygen (O_2), Iodine (I_2), Chlorine (Cl_2), Bromine (Br_2)

TABLE 5.40 Common Monoatomic Anions

Charge	**Ion Name -ide ending**	**Ion**
-1	fluoride	F^-
	chloride	Cl^-
	bromide	Br^-
	iodide	I^-
	hydride	H^-
-2	oxide	O^{2-}
	sulfide	S^{2-}
-3	nitride	N^{3-}
	phosphide	P^{3-}
-4	carbide	C^{4-}

Common Metal Cations of Variable Charge

Element Name	**Latin Root**	**Lower Charge -ous ending**	**Higher Charge -ic ending**
chromium	chrom-	Cr^{2+}	Cr^{3+}
cobalt	cobalt-	Co^{2+}	Co^{3+}
copper	cupr-	Cu^{+}	Cu^{2+}
gold	aur-	Au^{2+}	Au^{3+}
iron	ferr-	Fe^{2+}	Fe^{3+}
lead	plumb-	Pb^{2+}	Pb^{4+}
manganese	mangan-	Mn^{2+}	Mn^{3+}
mercury	mercur-	Hg_2^{2+}	Hg^{2+}
nickel	nickel-	Ni^{2+}	Ni^{3+}
tin	stann-	Sn^{2+}	Sn^{4+}

TABLE 5.41 Common Metal Cations of Known Charge

Charge		Charge		Charge	
+1	Alkali Metals	+2	Alkaline Earth Metals	+3	Aluminum
	Hydrogen		Zinc		Gallium
	Silver		Cadmium		Indium
					Scandium
					Yttrium

TABLE 5.42 Common Polyatomic Ions

	Name	Formula		Name	Formula		Name	Formula
	acetate	$C_2H_3O_2^-$	*	formate	CHO_2^-		phosphate	PO_4^{3-}
	ammonium	NH_4^+		hydride	H^-		hydrogen phosphate	HPO_4^{2-}
	arsenate	AsO_4^{3-}		hydrogen ion/proton	H^+		dihydrogen phosphate	$H_2PO_4^-$
*	azide	N_3^-		hydronium	H_3O^+		selenate	SeO_4^{2-}
	borate	BO_3^{3-}		hydroxide	OH^-		silicate	SiO_4^{4-}
	bromate	BrO_3^-	*	hydrogen sulfide/ bisulfide	HS^-		metasilicate	SiO_3^{2-}
	carbonate	CO_3^{2-}		nitrate	NO_3^-		sulfate	SO_4^{2-}
	bicarbonate or hydrogen carbonate	HCO_3^-		oxalate	$C_2O_4^{2-}$		hydrogen sulfate/ bisulfate	HSO_4^-
	chlorate	ClO_3^-		iodate	IO_3^-	*	superoxide	O_2^-
	chromate	CrO_4^{2-}		mercury(I)/mercurous	Hg_2^{2+}		tellurate	TeO_4^{2-}
	dichromate	$Cr_2O_7^{2-}$		peroxide	O_2^{2-}		thiocyanate	SCN^-
	cyanate	OCN^-		permanganate	MnO_4^-	*	thiosulfate	$S_2O_3^{2-}$
	cyanide	CN^-	*	phosphonium	PH_4^+			

* You do not need to memorize these.

Remember that the charge does not change only the number of oxygen atoms. The -ide ending usually has not oxygen atoms. Some common exceptions are oxide (O^{2-}), hydroxide (OH^-), peroxide (O_2^{2-}) and hyperoxide (O_2^-)

TABLE 5.43 Polyatomic Ion Patterns

prefix...suffix	General Pattern	Number of Oxygen Atoms	Acid Pattern
per...ate	most oxygen atoms	+1	per...ic
...ate	next most oxygen atoms	-----	...ic
...ite	second least oxygen atoms	-1	...ous
hypo...ite	least oxygen atoms	-2	hypo...ous
...ide	usually no oxygen atoms	0 (generally)	hydro...ic

Nomenclature Exercise: IDE IONS

Name: ___________________________________

Section:____________________

Exercise 1a

Give the name of the following compounds:

TABLE 5.44

	Formula	Compound Name
1.	NaCl	
2.	$BaBr_2$	
3.	K_2S	
4.	Al_2O_3	
5.	LiI	
6.	CaS	
7.	BeF_2	
8.	Zn_3N_2	
9.	H_2O	
10.	Na_2S	
11.	Rb_2Se	
12.	Mg_3P_2	
13.	KBr	
14.	Na_2O	
15.	Li_3N	
16.	KH	

Exercise 1b

Give the formulas of the following compounds:

TABLE 5.45

	Name	Cation	Anion	Compound
1.	cesium oxide	Cs^+	O^{2-}	Cs_2O
2.	potassium sulfide			
3.	barium iodide			
4.	zinc chloride			
5.	sodium hydride			
6.	calcium nitride			
7.	gallium oxide			
8.	hydrogen bromide			
9.	lithium phosphide			
10.	beryllium fluoride			
11.	potassium bromide			
12.	magnesium nitride			
13.	potassium selenide			
14.	lithium sulfide			
15.	aluminum iodide			
16.	hydrogen fluoride			

Nomenclature Exercise: ATE IONS

Name: ____________________________________

Section:____________________

Exercise 2a

Give the name of the following compounds:

TABLE 5.46

	Formula	Compound Name
1.	$NaNO_3$	
2.	$CaCO_3$	
3.	Rb_3AsO_4	
4.	$AlPO_4$	
5.	$SrSeO_4$	
6.	Cs_2SO_4	
7.	$Sc(BrO_3)_3$	
8.	$AgIO_3$	
9.	$Ga_2(SO_4)_3$	
10.	$Mg(NO_3)_2$	
11.	$Be_3(AsO_4)_2$	
12.	$Sr(BrO_3)_2$	
13.	$BeCO_3$	
14.	$LiClO_3$	
15.	Ag_2SO_4	
16.	$Zn_3(PO_4)_2$	

Exercise 2b

Give the formulas of the following compounds:

TABLE 5.47

	Name	Cation	Anion	Compound
1.	zinc chlorate			
2.	aluminum nitrate			
3.	potassium borate			
4.	cesium metasilicate			
5.	silver sulfate			
6.	sodium tellurate			
7.	gallium chlorate			
8.	rubidium bromate			
9.	beryllium nitrate			
10.	scandium arsenate			
11.	cadmium iodate			
12.	lithium carbonate			
13.	calcium phosphate			
14.	magnesium selenate			
15.	rubidium sulfate			
16.	strontium arsenate			

Nomenclature Exercise: ITE and ATE IONS

Name: ___________________________________

Section:____________________

Exercise 3a

Give the name of the following compounds:

TABLE 5.48

	Formula	Compound Name
1.	$NaNO_3$	
2.	$Ba(NO_2)_2$	
3.	$CaSO_4$	
4.	$SrSO_3$	
5.	$MgSeO_3$	
6.	$ScAsO_4$	
7.	$GaAsO_3$	
8.	$RbClO_3$	
9.	$CsClO_2$	
10.	$Zn(IO_2)_2$	
11.	KNO_2	
12.	$BeCO_3$	
13.	$KBrO_3$	
14.	$LiClO_2$	
15.	$GaBO_3$	
16.	$AgIO_3$	
17.	$AlPO_3$	

Exercise 3b

Give the formulas of the following compounds:

TABLE 5.49

	Name	Cation	Anion	Compound
1.	aluminum sulfite			
2.	potassium carbonate			
3.	barium nitrite			
4.	rubidium chlorite			
5.	strontium borate			
6.	scandium iodite			
7.	cesium phosphite			
8.	lithium metasilicate			
9.	beryllium selenite			
10.	calcium bromite			
11.	zinc sulfate			
12.	cadmium phosphate			
13.	magnesium chlorate			
14.	sodium selenate			
15.	silver bromate			
16.	potassium arsenite			

Nomenclature Exercise: BINARY COVALENT COMPOUNDS

Name: ____________________________________

Section:____________________

Exercise 4a

Give the name of the following compounds:

TABLE 5.50

	Formula	Compound Name
1.	P_4O_6	
2.	SF_4	
3.	ICl_3	
4.	XeF_6	
5.	CO_2	
6.	P_2Cl_4	
7.	I_4O_9	
8.	Se_2Br_2	
9.	NH_3	
10.	Si_3N_4	
11.	NO_2	
12.	SO_3	
13.	IF_7	
14.	KrF_2	
15.	P_4O_3	
16.	H_2O	
17.	XeO_3	

Exercise 4b

Give the formulas of the following compounds:

TABLE 5.51

	Name	Compound
1.	disulfur decafluoride	
2.	nitrogen triiodide	
3.	trisilicon octahydride	
4.	dinitrogen difluoride	
5.	silicon dioxide	
6.	phosphorus pentachloride	
7.	sulfur hexafluoride	
8.	tetranitrogen tetraselenide	
9.	diiodine pentaoxide	
10.	xenon hexafluoride	
11.	tetraphosphorus heptasulfide	
12.	disilicon hexahydride	
13.	iodine dioxide	
14.	difluorine dioxide	
15.	carbon monoxide	
16.	selenium tetrabromide	

Nomenclature Exercise: MIXED ANIONS

Name: ____________________________________

Section:____________________

Exercise 5a

Give the name of the following compounds:

TABLE 5.52

	Formula	Compound Name
1.	$TiCl_3$	
2.	$ZnCO_3$	
3.	$FeSO_3$	
4.	$Cu_3(AsO_4)_2$	
5.	MnF_2	
6.	Li_3P	
7.	$CaCrO_4$	
8.	KIO_2	
9.	$BeSiO_3$	
10.	$CoSeO_4$	
11.	Ga_2O_3	
12.	$ScBO_3$	
13.	$V(NO_3)_5$	
14.	$CrBr_3$	
15.	$NaBrO_3$	
16.	$FePO_4$	
17.	$Sr_3(AsO_3)_2$	
18.	Cu_2O	

Exercise 5b

Give the formulas of the following compounds:

TABLE 5.53

	Name	Cation	Anion	Compound
1.	cesium nitrite			
2.	nickel(II) chromate			
3.	zinc arsenite			
4.	titanium(IV) borate			
5.	lithium nitride			
6.	sodium metasilicate			
7.	chromium(III) sulfate			
8.	aluminum tellurite			
9.	nickel(II) arsenate			
10.	mercury(II) sulfide			
11.	iron(III) chlorate			
12.	manganese(III) oxide			
13.	vanadium(III) fluoride			
14.	copper(I) carbonate			
15.	scandium nitrate			
16.	beryllium selenide			

Nomenclature Exercise: HYDRATES, SPECIAL ANIONS AND PREFIX METHOD

Name: ___________________________________

Section:____________________

Exercise 6a

Give the name of the following compounds:

TABLE 5.54

	Formula	Compound Name
1.	S_2O_3	
2.	P_2O_5	
3.	$(NH_4)_2S$	
4.	$Cu(NO_3)_2 \cdot 3\ H_2O$	
5.	Cl_2O_7	
6.	$Fe(ClO_4)_2 \cdot 6\ H_2O$	
7.	$KMnO_4$	
8.	$Na_2SO_4 \cdot 7\ H_2O$	
9.	$Mn(C_2H_3O_2)_2 \cdot 4\ H_2O$	
10.	NaOCl	
11.	OF_2	
12.	$(NH_4)_2CrO_4$	
13.	$K_2S \cdot 5\ H_2O$	
14.	CsOI	
15.	$Ca(OH)_2$	
16.	AgCN	
17.	$Ti_2(C_2O_4)_3 \cdot 10\ H_2O$	

Exercise 6b

Give the formulas of the following compounds:

TABLE 5.55

	Name	Cation	Anion	Compound
1.	sulfur trioxide	-------	-------	
2.	tetraphosphorus decoxide	-------	-------	
3.	iodine pentafluoride	-------	-------	
4.	ammonia	-------	-------	
5.	dinitrogen pentoxide	-------	-------	
6.	ammonium phosphate trihydrate			
7.	calcium hypochlorite			
8.	strontium cyanide			
9.	calcium sulfate heptahydrate			
10.	aluminum chloride hexahydrate			
11.	magnesium permanganate			
12.	cobalt(II) perchlorate pentahydrate			
13.	barium hydroxide octahydrate			
14.	lithium acetate			
15.	silver oxalate			
16.	ammonium perchlorate			

Nomenclature Exercise: ACIDS

Name: ___________________________________

Section:___________________

Exercise 7a

Name the following compounds in aqueous solution:

TABLE 5.56

	Formula	Acid Name
1.	$HCl_{(aq)}$	
2.	$HNO_{3\,(aq)}$	
3.	$H_2SO_{4\,(aq)}$	
4.	$H_2SO_{3\,(aq)}$	
5.	$HCN_{(aq)}$	
6.	$H_2S_{(aq)}$	
7.	$HI_{(aq)}$	
8.	$HF_{(aq)}$	
9.	$HC_2H_3O_{2\,(aq)}$	
10.	$H_2C_2O_{4\,(aq)}$	
11.	$HClO_{4\,(aq)}$	
12.	$HClO_{3\,(aq)}$	
13.	$HIO_{(aq)}$	
14.	$HNO_{2\,(aq)}$	
15.	$HBrO_{2\,(aq)}$	
16.	$H_3PO_{4\,(aq)}$	

Exercise 7b

Supply the formula for the following acids:

TABLE 5.57

	Name of Acid	Cation	Anion	Formula of Acid
1.	chlorous acid			
2.	sulfuric acid			
3.	phosphoric acid			
4.	telluric acid			
5.	hypoiodous acid			
6.	perbromic acid			
7.	hydrosulfuric acid			
8.	hydrochloric acid			
9.	nitrous acid			
10.	arsenous acid			
11.	hydroiodic acid			
12.	selenous acid			
13.	chloric acid			
14.	hydrobromic acid			
15.	oxalic acid			
16.	hydrocyanic acid			

Nomenclature Exercise: MOSTLY REVIEW

Name: ___________________________________

Section:___________________

Exercise 8a

Give the name of the following compounds:

TABLE 5.58

	Formula	Compound Name
1.	$(NH_4)_3PO_3$	
2.	$HF_{(aq)}$	
3.	Ba_3P_2	
4.	$Fe(C_2H_3O_2)_2 \cdot 4\ H_2O$	
5.	$LiClO_2$	
6.	$CdSeO_4 \cdot 2\ H_2O$	
7.	$H_2CO_{3\ (aq)}$	
8.	$Ba(HSO_4)_2$	
9.	$MnSO_4 \cdot 7\ H_2O$	
10.	BaH_2	
11.	$HClO_{4\ (aq)}$	
12.	$Be(MnO_4)_2$	
13.	KH_2PO_4	
14.	$Sr(OH)_2$	
15.	$CaSiO_3$	
16.	$PbHPO_4$	

Exercise 8b

Give the formulas of the following compounds:

TABLE 5.59

	Name	Cation	Anion	Compound
1.	chromium(III) phosphate hexahydrate			
2.	nitrous acid			
3.	potassium bisulfite			
4.	lithium hydrogen carbonate			
5.	tin(IV) fluoride			
6.	zinc hydroxide			
7.	arsenous acid			
8.	strontium dihydrogen phosphate			
9.	barium hydride			
10.	potassium bisulfate			
11.	potassium oxalate			
12.	gold(III) cyanide			
13.	hydrosulfuric acid			
14.	scandium iodide			
15.	acetic acid			
16.	dichlorine heptoxide	-------	-------	

Nomenclature Exercise: SPECIAL IONS

Name: ___________________________________

Section:____________________

Exercise 9a

Give the name of the following compounds:

TABLE 5.60

	Formula	Compound Name
1.	$CuO \cdot H_2O$	
2.	$AgSCN$	
3.	$Sn(HSO_4)_2$	
4.	$Al_2(Cr_2O_7)_3$	
5.	$Fe(ClO_4)_2$	
6.	$Pb(C_2H_3O_2)_2$	
7.	BaO_2	
8.	$Hg_2(OH)_2$	
9.	Na_2O_2	
10.	$K_2C_2O_4$	
11.	$AgOCN$	
12.	$Hg(IO_2)_2$	
13.	$W(CN)_3$	
14.	CoC_2O_4	
15.	$Na_2C_2O_4 \cdot 2\ H_2O$	
16.	SrH_2	

Exercise 9b

Give the formulas of the following compounds:

TABLE 5.61

	Name	Cation	Anion	Compound
1.	lead(II) dichromate			
2.	silver oxalate			
3.	ammonium hypoiodite			
4.	iron(III) thiocyanate			
5.	chromium(II) perchlorate			
6.	manganese(III) hydroxide			
7.	aluminum dihydrogen phosphate			
8.	mercury(II) chloride			
9.	tin(II) hydride			
10.	rubidium acetate			
11.	ammonium iodite			
12.	zinc permanganate			
13.	cadmium sulfite			
14.	calcium peroxide			
15.	mercury(I) nitride			
16.	potassium bisulfate			
17.	hydrogen peroxide			

Nomenclature Exercise: REVIEW

Name: ___________________________________

Section:____________________

Exercise 10a

Give the name of the following compounds:

TABLE 5.62

	Formula	Compound Name
1.	$NiF_2 \cdot 6\ H_2O$	
2.	$HClO_{(aq)}$	
3.	$AgClO_3$	
4.	K_3N	
5.	$NaHCO_3$	
6.	$(NH_4)_3PO_3$	
7.	K_2O_2	
8.	BrF_3	
9.	$Cu(HSO_4)_2$	
10.	$Hg_2Cr_2O_7$	
11.	$Cu(CN)_2 \cdot 3\ H_2O$	
12.	$Bi(NO_3)_3$	
13.	ClF_5	
14.	$HIO_{4\ (aq)}$	
15.	PbS	
16.	$Nb_2(SO_3)_5$	

Exercise 10b

Give the formulas of the following compounds:

TABLE 5.63

	Name	Cation	Anion	Compound
1.	bromous acid			
2.	xenon tetrafluoride	----------	----------	
3.	potassium permanganate			
4.	aluminum hydroxide hexahydrate			
5.	potassium bisulfite			
6.	cerium(III) iodite			
7.	potassium peroxide			
8.	palladium(III) tellurate			
9.	boron trihydride	----------	----------	
10.	indium(III) nitrate pentahydrate			
11.	mercury(I) phosphide			
12.	oxalic acid			
13.	neodymium(VIII) carbide			
14.	lithium nitrite monohydrate			
15.	rubidium fluoride			
16.	strontium bisulfide			
17.	water	----------	----------	

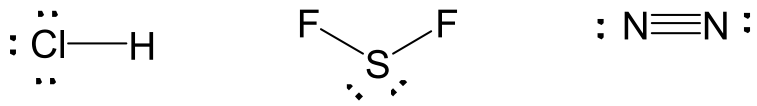
EXPERIMENT 6 *Molecular Models*

Background

Chemists often use models to visualize molecular structures. Molecular geometries can frequently influence chemical and physical properties, thus it is important to begin to recognize the way that atoms bond together and how they orient themselves in a molecule. In this experiment, you will build a variety of molecules using molecular models in order to become familiar with some of the more common geometries.

Lewis Structures. In 1916, G. N. Lewis developed a theory in which he proposed that atoms would bond together to share electrons so that all representative atoms would be surrounded by 8 valence electrons.Hydrogen, H, and helium, He, are exceptions having a maximum of 2 valence electrons surrounding them. Chemists still use that model when drawing **Lewis Electron Dot** structures for molecules. Your textbook outlines the procedure for drawing correct electron dot structures for common compounds. A few examples of correctly drawn Lewis structures are shown in Figure 6.1 on page 115.

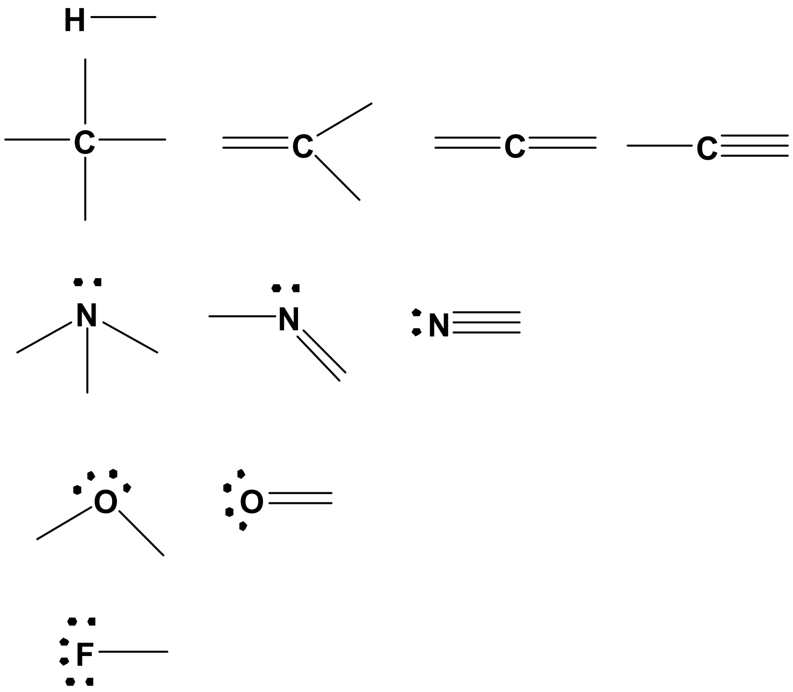
FIGURE 6.1



Remember that all non-bonding electron pairs must be shown in a correct Lewis Structure!

Molecular Geometries. **V**alence **S**hell **E**lectron-**P**air **R**epulsion theory or **VSEPR** is used to predict molecular geometries. VSEPR theory proposes that the structure of a molecule is determined by the repulsive interaction of electron pairs in the valence shell of its central atom. In other words, the bonding pairs and the non-bonding (lone) pairs around a given atom are as far apart as possible.

FIGURE 6.2 **Common Bonding Patterns:**

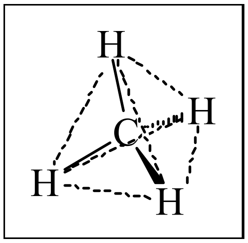


EXAMPLE 6.1 **Methane, CH_4**

has four valence electrons from the carbon atom, and one valence electron from each hydrogen atom.

$$\left(\frac{4VE}{C}\right)(1C)+\left(\frac{1VE}{H}\right)(4H) \;=\; 8VE \qquad \text{(EQ 6.1)}$$

FIGURE 6.3



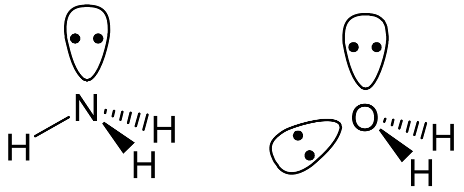
Therefore, methane has four bonding electron pairs around the carbon atom. In order to maximize the distance between these bonding electrons (and minimize the repulsive forces) the hydrogen atoms will orient themselves forming a tetrahedron with 109.5^o bond angles as shown in.Figure 6.3 .

EXAMPLE 6.2 **Ammonia, NH_3, and water, H_2O**

If pairs of non-bonding, lone pairs also surround an atom they will also require space. An example of this is shown in both ammonia and water where the angles between the four sets of electron pairs are still approximately 109.5° (experiment shows the angles are about 107° and 104.5° respectively) as seen in Figure 6.4 .

FIGURE 6.4 **3-D Sketch:**

FIGURE 6.5 **The lone pair electrons are sometimes shown in orbitals:**



Orbital geometry is the geometrical arrangement of all electron groups includes bonding and nonbonding electrons. **Molecular geometry** only looks at the shape of the bonding electrons. An incomplete summary of the orbital geometries, molecular geometries, and approximate bond angles is illustrated in the prelab section in Table 6.5 on page 119. You will complete it.

Molecular Polarity. The polarity of a molecule can be determined by looking at the shape of the molecule and the bonds withing the molecule. A **polar molecule** occurs when there is a net dipole moment on the molecule as determined by electronegativity differences between atoms. A **nonpolar molecule** will not have a net dipole moment. There are some visual indicators that can help determine if a molecule is polar. For example, a molecule can usually be classified as polar if: there is a lone pair (or lone pairs) on the central atom, or there are different types of atoms (or groups of atoms) on the central atom.

EXAMPLE 6.3 **Are methane, water and ammonia polar or nonpolar molecules?**

In methane all 4 bonds are C-H bonds, which would have the same dipole moments. Resulting in a zero net dipole moment.

In water, there are 2 O-H bonds, which would have the same dipole moments. However, there are also 2 lone pairs, which would results in a net dipole moment.

In ammonia, there are 3 N-H bonds, which would have the same dipole moments. However, there is also 1 lone pairs, which result in a net dipole moment.

Double and Triple Bonds. Like single bonds and nonbonding electron pairs, double and triple bonds are considered one “cloud” of electron density.

FIGURE 6.6 **Hydrogen cyanide, HCN**

$$\text{H-C}\equiv\text{N:}$$

One bonding cloud is from the hydrogen-carbon single bond and the other bonding cloud is from the carbon≡nitrogen triple bond.

There are no non-bonding clouds on the carbon atom, which is the central atom.

Therefore, a molecule such as hydrogen cyanide would have two bonding “clouds” and zero non-bonding “clouds” around the central atom.

Hydrogen cyanide would have a net dipole moment and be a polar molecule. Notice that there are different groups with different electronegativities attached to the central atom.

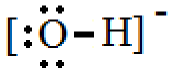
Ions. If you need to draw the Lewis Electron Dot structures for ions, add or subtract valence electrons based upon the charge on the molecule.

EXAMPLE 6.4 Hydroxide ion, OH^-

would have 6 valence electrons from the oxygen, 1 valence electron from the hydrogen, and 1 extra valence electron due to the charge, giving it a total of 8 valence electrons:

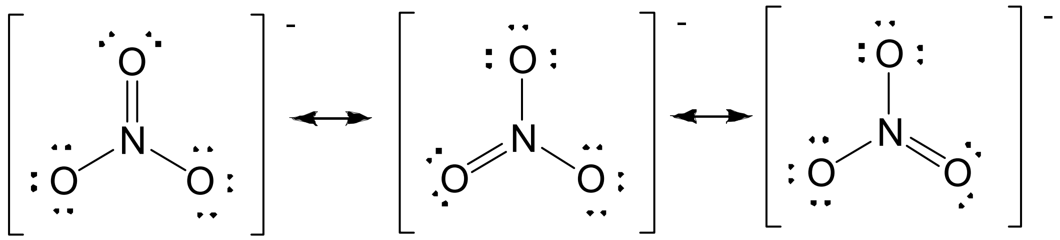
$$\left(\frac{6VE}{O}\right)(1O) + \left(\frac{1VE}{H}\right)(1H) + 1e^{-} \ = \ 8VE \qquad \textbf{(EQ 6.2)}$$

FIGURE 6.7



Another consideration is resonance. Molecules that have **resonance structures** have a double bond that can move around the central atom without changing the structure of the molecule. In other words, atoms are not moving, rather an electron pair is moving, so the octet rule is obeyed.

FIGURE 6.8 Nitrate ion, NO_3^-



Procedure

Build and draw the molecules and ions listed in the tables on the following pages. Be sure to answer the questions regarding these molecules. Your instructor may want to check off each model to ensure that you have made them.

Prelaboratory Exercise: Molecular Models

Lab Grade

Prelab Questions		21
General Format (Signature, ink, no obliterations, etc.)		7
Data and Analysis (molecular models descriptions, etc.)		60
Post Lab Questions		12
Total		100

Name: ___________________________

Section:____________________

1. Complete the following table:

TABLE 6.5

Valence shell electron pairs	Bonding electron "clouds"	Lone pairs	Lewis Dot Formula	Orbital geometry	Approximate bond angles	Molecular geometry
Diatomic molecule	1	No central		Linear	180°	Linear
2	2	0	AX_2	Linear	180°	Linear
3	3	0	AX_3	Trigonal Planar	120°	
	2	1	$\ddot{A}X_2$			Bent
	4	0	AX_4	Tetrahedral	109.5°	Tetrahedral
4	3	1	$\ddot{A}X_3$			Trigonal Pyramidal
	2	2	$:\!\ddot{A}X_2$		<109.5°	Bent

2. How many valence electrons does an atom of nitrogen contain? ______

3. How many valence electrons are in a molecule of $C_4H_6Br_2$? ______

4. How many valence electrons are in an acetate ion, $CH_3CO_2^-$? ______

5. How many valence electrons are in a $NH_3C_2H_5^+$ ion? ______

6. What is the abbreviation for the theory used to determine the geometry of molecules?

7. What is the orbital or electronic geometry of a molecule with 2 nonbonding electron pairs and 2 bonding electron pairs?

8. What is the orbital or electronic geometry of a molecule with 0 nonbonding electron pairs and 2 bonding electron pairs?

9. What is the molecular geometry of a molecule with 1 nonbonding electron pair and 2 bonding electron pairs?

10. What is the molecular geometry of a molecule with 1 nonbonding electron pair and 3 bonding electron pairs?

11. What is the approximate bond angle of a molecule with 1 nonbonding electron pair and 2 bonding electron pairs?

12. What is the approximate bond angle of a molecule with 0 nonbonding electron pair and 4 bonding electron pairs?

Data Sheet: Molecular Models

Fill in the sheet below for the molecules you build. You may use pencil to complete this exercise.

TABLE 6.6

Molecule	CBr_4	PH_3	SiO_2
Total # Valence Electrons			
Lewis Structure			
Number of Bonding clouds of electrons around central			
Number of Non-bonding Electron pairs on central			
Electronic or Orbital Geometry			
Molecular Geometry (Name)			
Molecule Polarity			
3-Dimensional Sketch with Bond Angles Noted			
Bond Angles			

TABLE 6.7

Molecule	C_2H_6	C_2H_4	C_2H_2
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons clouds on one carbon atom			
Number of Nonbonding Electron Pairs on one carbon atom			
Electronic or Orbital Geometry (both carbons)			
Molecular Geometry (Name) (both carbons)			
Molecule Polarity			
3-Dimensional Sketch			
Bond Angle H-C-C			

TABLE 6.8

Molecule	**HF** (Is there a central atom?)	**C_2H_5OH**	**SO_3** (show resonance structures)
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons clouds		C O	
Number of nonbonding Electron Pairs on central atom(s)		C O	
Electronic or Orbital Geometry		C O	
Molecular Geometry (Name)		C O	
Molecule Polarity			
3-Dimensional Sketch			
Bond Angle		H-C-H C-O-H	

TABLE 6.9

Molecule	H_3O^+	CO_3^{2-} (show resonance structures)	NH_2^-
Total # Valence Electrons			
Lewis Structure			
Number of Bonding Electrons clouds on central			
Number of Nonbonding Electron Pairs on central			
Electronic or Orbital Geometry			
Molecular Geometry (Name)			
3-Dimensional Sketch			
Bond Angle			

Post Lab Questions

1. Why are the valence electrons of an atom the only electrons likely to be involved in bonding to other atoms?

2. Why do representative elements tend to form bonds giving them a total of 8 valence electrons?

3. How is the structure around a given atom related to repulsion between valence electron pairs on the atom involved?

4. Why are all diatomic molecules linear, regardless of the number of valence electron pairs on the atoms involved?

EXPERIMENT 7

Estimating Avogadro's Number

Background

In this experiment you will estimate the number of stearic acid molecules in a mole of stearic acid based on an assumption about a thin film of stearic acid on a water surface. A monolayer of stearic acid can be created in which we can estimate the number of molecules. To do this we make assumptions of the size and shape of the stearic acid molecule.

In the International System (SI) of units, the fundamental measure of quantity is the mole. A mole of carbon is 12.0000 grams of carbon-12, and Avogadro's number, N_A, is the number of carbon atoms in this mass. A mole of any substance contains Avogadro's number of particles.

In this experiment you will estimate Avogadro's number by calculating the amount of stearic acid necessary to form a single layer on the surface of water. This single layer is called a monolayer. We will be making some assumptions about how stearic acid molecules arrange themselves to form the monolayer. From these assumptions we can determine the thickness of the layer, and from the thickness we can estimate the number of the stearic acid molecules. If we know the volume occupied by a mole of stearic acid molecules we can use the volume of one stearic acid molecule to estimate Avogadro's number. We simply divide the volume of a mole of stearic acid by the volume of one stearic acid molecule.

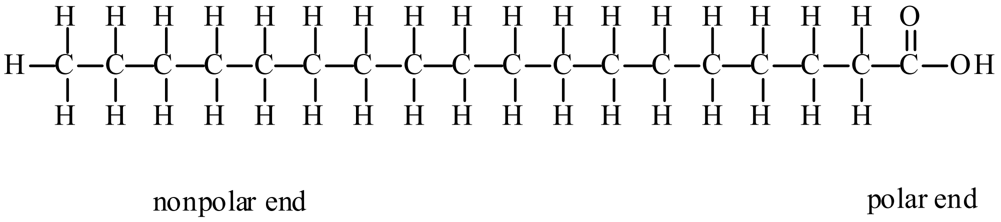
Note – the number calculated will not be accurate, but should be within a power of 10 of the accepted value of **6.02214×10^{23} particles/mol**.

Theory and Basis of the Experiment

One of the properties displayed by water is its polarity. For covalent polar molecules, it is a partial charge separation denoted by a delta, δ. Polar molecules attract each other. The negative end of the dipole moment of one polar molecule is attracted to the positive end of the dipole moment of another polar molecule. Because of this, acetic acid, CH_3COOH, will dissolve in water. Whereas, pentane, $CH_3CH_2CH_2CH_2CH_3$, a non-polar molecule is not water soluble. It stands to reason that a molecule that possesses both a nonpolar and polar end will have the polar end attracted to water and the nonpolar end be repelled by water.

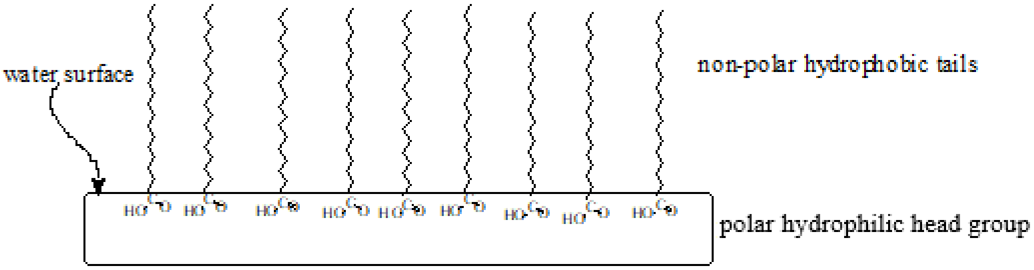
Stearic acid, $CH_3(CH_2)_{16}CO_2H$, has the structure shown below. The acid has a polar portion and a non-polar portion. The large non-polar, hydrophobic portion, of the molecule makes it immiscible with water.

FIGURE 7.1



We will create a layer of stearic acid one molecule thick on the surface of a watch glass filled with water. The polar head group will stay near the water and the non-polar tail will orient themselves away from the water.

FIGURE 7.2



We will create our monolayer by using a solution of the stearic acid in hexane. Hexane is also hydrophobic and will not mix with water. The hexane will evaporate leaving just the thin layer of stearic acid.

If we know the area of the monolayer and we know the volume of the monolayer, we can measure the thickness, *t*, of the monolayer. This thickness will correspond to the length of the stearic acid molecule. If we make the assumption that the stearic acid molecule is a rectangular solid, with a width and length equal to **1/5.44** of the height, where the height equals the thickness of the monolayer, we can calculate the volume of one stearic acid molecule. Knowing the volume of the monolayer and the volume of one molecule we can calculate the number of molecules in the monolayer. By using the density of stearic acid we can determine the number of moles in the monolayer. Finally, we can use the molar volume of stearic acid and divide by the volume of the stearic acid molecule to get Avogadro's number, N_A!

There are two parts to this experiment. In part 1, we will calibrate a dropping pipet so that we can determine the number of drops in 1 mL. This will require a Pasteur pipet for optimum results. In part 2, we will slowly add the stearic acid/hexane solution to a watch glass of water. We will allow the hexane to evaporate leaving just the stearic acid on the water.

EXAMPLE 7.1 Calculation of Avogadro's Number

The following data was collected in an experiment similar to yours and was analyzed as shown below. Remember not to round until the end of the calculation to avoid round off error in your answer.

Drops of hexane/mL 24 drops/mL
Number of drops hexane used for monolayer 7 drops
Diameter of water surface 10.1 cm
Concentration of stearic acid/hexane 0.10 g/L
Molar Mass of stearic acid: 284 g/mol
Density of stearic acid 0.870 g/mL

1. Calculate the volume of the stearic acid/hexane in the monolayer.

$$7 \ drops \times \frac{1 \ mL \ soln}{24 \ drops} = 0.2916666667 \ mL$$

2. Calculate the mass of stearic acid in the monolayer.

$$0.2916666667 \ mL \times \frac{1 \ L}{1000 \ mL} \times \frac{0.10 \ g}{1 \ L} = 2.916666667 \times 10^{-5} \ g$$

3. Calculate the volume of stearic acid in the monolayer in mL

$$2.916666667 \times 10^{-5} \ g \times \frac{1 \ mL}{0.870 \ g} = 3.35249042 \times 10^{-5} \ mL$$

4. Calculate the thickness of the monolayer in cm.

$$A = \pi r^2 = \pi\left(\frac{10.1 \ cm}{2}\right)^2 = 80.11846665 \ cm^2$$

$$t = 3.35249042 \times 10^{-5} mL \times \frac{1 \ cm^3}{1 \ mL} \times \frac{1}{80.11846665 \ cm^2} = 4.184416604 \times 10^{-7} \ cm$$

5. Calculate the volume of a stearic acid molecule in the monolayer.

$$V = l \times w \times h = \left(\frac{t}{5.44}\right)\left(\frac{t}{5.44}\right)(t) = \frac{t^3}{(5.44)^2} = \frac{(4.1844416604 \times 10^{-7} cm)^3}{(5.44)^2 \ molecule} = 2.47575093 \times 10^{-21} \frac{cm^3}{molecule}$$

6. Calculate the number of stearic acid molecules in the surface layer.

$$3.35249042 \times 10^{-5} mL \times \frac{1 \ cm^3}{1 \ mL} \times \frac{1 \ molecule}{2.47575093 \times 10^{-21} \ cm^3} = 1.354130732 \times 10^{16} \ molecules \ stearic \ acid$$

7. Calculate the number of moles of stearic acid in the surface layer.

$$2.916666667 \times 10^{-5} \ g \times \frac{1 \ mol}{284 \ g} = 1.026995305 \times 10^{-7} \ mol$$

8. Calculate the value of Avogadro's number.

$$\frac{1.354130732 \times 10^{16} \ molecules \ stearic \ acid}{1.026995305 \times 10^{-7} \ mol \ stearic \ acid} = 1.318536438 \times 10^{23} \frac{molecules}{mol} = 1.3 \times 10^{23} \frac{molecules}{mol}$$

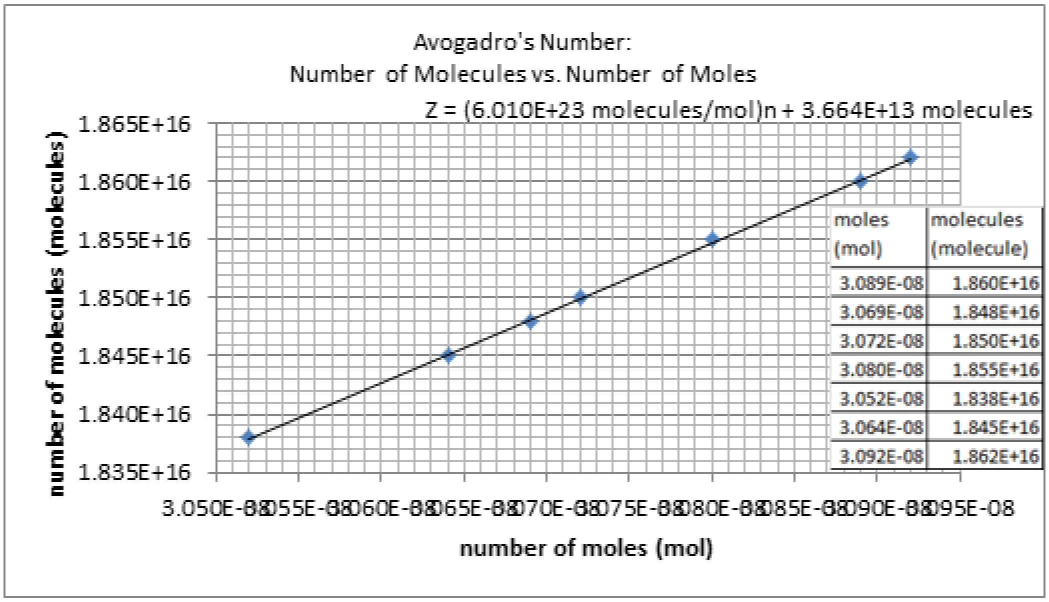
Any impurities will ruin this experiment so be <u>very</u> careful not to put soap, water or dust where it does not belong. **Do not use dirty glassware!**

Another way to analyze this data would be to complete multiple trials of the experiment and graph the number of molecules versus the number of moles. The slope of the line would give you Avogadro's number or an estimation of it.

TABLE 7.2 Avogadro's number data:

moles (mol)	molecules (molecules)	moles x10^{-8} (mol)	molecules x10^{15} (molecules)
3.052E-08	1.838E+15	3.052	18.38
3.064E-08	1.845E+16	3.064	18.45
3.069E-08	1.848E+16	3.069	18.48
3.072E-08	1.850E+16	3.072	18.50
3.080E-08	1.855E+16	3.080	18.55
3.089E-08	1.860E+16	3.089	18.60
3.092E-08	1.862E+16	3.092	18.62

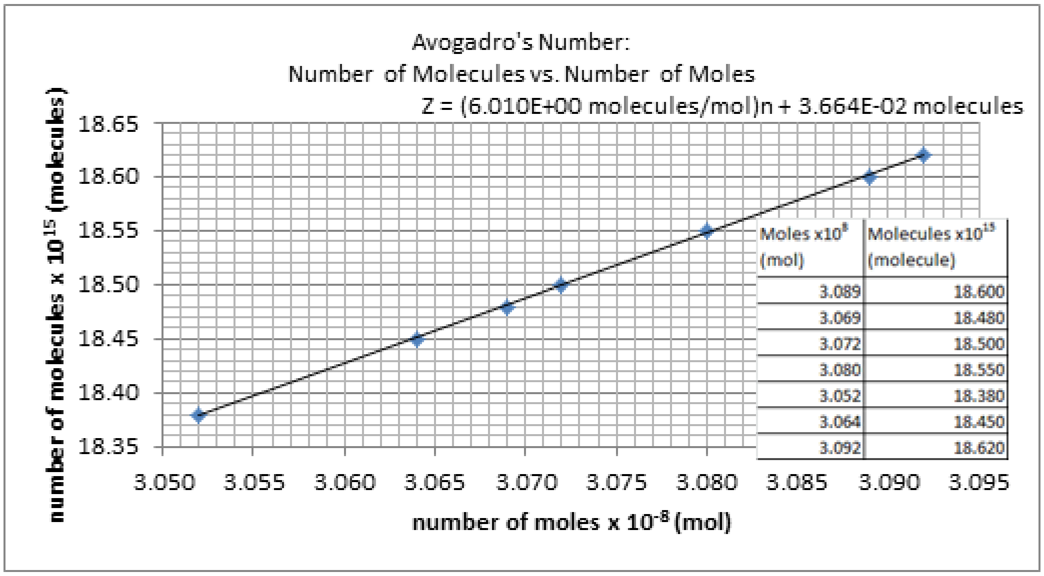
FIGURE 7.3 Unscaled Graph



moles (mol)	molecules (molecule)
3.089E-08	1.860E+16
3.069E-08	1.848E+16
3.072E-08	1.850E+16
3.080E-08	1.855E+16
3.052E-08	1.838E+16
3.064E-08	1.845E+16
3.092E-08	1.862E+16

Often it makes it easier to read the graph if it is scales to more convenient numbers as shown in the third and fourth column in Table 7.3 on page 135 above. The new data can be graphed as shown in Figure 7.4 on page 131.

FIGURE 7.4 Scaled Graph



Moles $x10^{8}$ (mol)	Molecules $x10^{15}$ (molecule)
3.089	18.600
3.069	18.480
3.072	18.500
3.080	18.550
3.052	18.380
3.064	18.450
3.092	18.620

Since the x-axis does not intercept the y-axis at 0 you are not able to read the y-intercept directly off of the graph. However, notice that the graphing program did not scale the units so the equation of the line from Figure 7.3 and Figure 7.4 do not match:

$$Z = \left(6.010 \times 10^{0} \frac{molecules}{mol}\right) n + 3.644 \times 10^{-2}\ molecules \qquad \text{(EQ 7.1)}$$

To fix this, multiply the slope by $10^{15}/10^{-8} = 10^{23}$ and the y-intercept by 10^{15} gives the correct equation of the line of:

$$Z = \left(6.010 \times 10^{23} \frac{molecules}{mol}\right) n + 3.644 \times 10^{13}\ molecules \qquad \text{(EQ 7.2)}$$

Procedure

Part 1: Calibration of the Pipet

1. Obtain a Pasteur pipet from your instructor.
2. Obtain a 10 mL beaker from the stock room. Wash it well. Wash the inside 3 times with approximately 1 mL of aqueous ammonia. Rinse the beaker 3 times with distilled water. Rinse the beaker 1 time with acetone and then dry the beaker thoroughly with a paper towel. Rinse the beaker three times with approximately 1/2-mL of hexane and then add 3 mL of hexane to the beaker.

> **Caution! Safety Notes:**
>
> **ALL WASTE HEXANE MUST BE DISPOSED OF INTO THE ORGANIC WASTE CONTAINERS. DO NOT DUMP HEXANE IN THE SINK.**

NOTE: Use the nitrile gloves provided by the stockroom. Remember that lids are to be kept between your fingers and never put down on counter tops. Do not stick your eye dropper into the reagent bottle!

3. Obtain a clean dry 10 mL graduated cylinder from your drawer. If your graduated cylinder is not clean and dry, clean it and then dry it with acetone and a tissue. Add hexane drop-wise to a volume of 1 mL. Record the number of drops.
4. Repeat this procedure.
5. If the two values are relatively close (within 10% of each other), then find the average. If not, repeat step 3 again.
6. Dispose of excess hexane in the appropriate waste container and save the beaker for part 2.

Part 2: The Monolayer

1. Obtain a watch glass from your drawer. Measure the diameter. Thoroughly clean the watch glass by scrubbing with soap for 2 minutes. Rinse well with tap water and then scrub with aqueous ammonia. Rinse 3 times with tap water and then three times with distilled water. Hold the glass by the edges to avoid putting fingerprints on it.
2. Place the watch glass on a beaker; making sure that the watch glass is parallel to the floor. Fill the watch glass to the brim with **distilled** water.
3. Take your clean 10 mL beaker and pour into it the stearic acid/hexane solution from under the hood.
4. Draw some of the stearic acid/hexane solution into your Pasteur pipet, hold the dropper perpendicular to the surface of the water in the watch glass, and add one drop of the stearic acid solution.
    a. The solution should spread out rapidly across the surface of the water and disappear within a few seconds.
    b. If the watch glass is not properly cleaned then an oily residue may appear after only a few drops of solution. In this case it will be necessary to clean the watch glass again.
    c. If the first few drops disappear rapidly, continue adding the solution drop-wise, counting the drops.
    d. As the monolayer nears completion, the drop of solution forms a circular pattern rather than flowering out.
    e. The circular film of solution contracts as it evaporates and disappears in a relatively short time.
    f. This pattern will be observed for a few drops until finally, one drop strikes the surface and remains as a lens or globule that requires a prolonged period of time to disappear.
5. Record the number of drops. At this point, the surface of the water is covered with a monolayer of stearic acid and one more drop placed at a different point on the water surface forms a second “lens”.
6. Fill out the following table. You must show your calculations for all starred rows. Clearly show these calculations using dimensional analysis.

Prelaboratory Exercise: Estimating Avogadro’s Number

Lab Grade

Prelab Questions		18
General Format (Signature, ink, no obliterations, etc.)		8
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		18
Accuracy		30
Post Lab Questions		26
Total		100

Name: ____________________________

Section:______________________

1. What substance will you use to make a monolayer?

   ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

2. How close should your experimental value be to the actual value of Avogadro’s number?
   - a. Within 1% of actual value
   - b. Within 5% of actual value
   - c. Within a factor of 10 of the actual value
   - d. Within a factor of 100 of the actual value

3. What do polar molecules behave like? ___ ___ ___ ___ ___ ___ ___.

4. Where do you dispose of the waste hexane?

5. What is the procedure for cleaning the watch glass?

6. How do you know when the stearic acid has formed a monolayer covering the entire surface of the water?

Use the data below to answer the following questions:

Drops of hexane/mL 19 drops/mL
Number of drops hexane used for monolayer 3 drops
Diameter of water surface 7.2 cm
Concentration of stearic acid/hexane 0.10 g/L
Molar Mass of stearic acid: 284 g/mol
Density of stearic acid 0.870 g/mL

1. Calculate the volume of the stearic acid/hexane in the monolayer.

2. Calculate the mass of stearic acid in the monolayer.

3. Calculate the volume of stearic acid in the monolayer in mL

4. Calculate the thickness of the monolayer in cm.

5. Calculate the volume of a stearic acid molecule in the monolayer.

6. Calculate the number of stearic acid molecules in the surface layer.

7. Calculate the number of moles of stearic acid in the surface layer.

8. Calculate the value of Avogadro's number.

Data Sheet: Estimating Avogadro's Number

Part 1: Calibration

TABLE 7.3

	Trial 1	**Trial 2**	**Trial 3** (if needed)
Drops of hexane/mL			
Average drops/mL			

Why is it important to do the calibration of the dropper quickly?

Part 2: Estimating Avogadro's Number

TABLE 7.4

	Trial 1	**Trial 2**	**Trial 3** (if needed)
Number of drops of solution used for monolayer			
Diameter of water surface (cm)			

Describe the appearance of the monolayer as you add the final drops of stearic acid.

TABLE 7.5 Calculations

	Trial 1	Trial 2	Sample Calculations
Average drops of hexane/ mL			
Number of drops of solution used for monolayer			
*Volume of solution used (mL)			
Concentration of stearic acid-hexane solution (g/L)	0.10 g/L	0.10 g/L	
*Mass stearic acid in monolayer on water surface (g)			
Density of stearic acid (g/mL)	0.870	0.870	
*Volume of stearic acid in monolayer (mL)			
Diameter of water surface (cm)			
*Area of water surface (of monolayer) ($A = \pi d^2/4$)			
*Thickness of monolayer			
*Volume of stearic acid molecule assuming it is a rectangular solid with a width and length equal to 1/5.44 of the height where the height equals the thickness of the monolayer.			
*Number of stearic acid molecules in the surface layer			
Molecular mass of stearic acid (g/mol)	284	284	
*Number of moles of stearic acid in surface layer			
*Value of Avogadro's number (molecules per mole of stearic acid)			

Post Lab Questions

1. What is the mass of 5.47 mole of stearic acid? (MW = 284 g/mol)

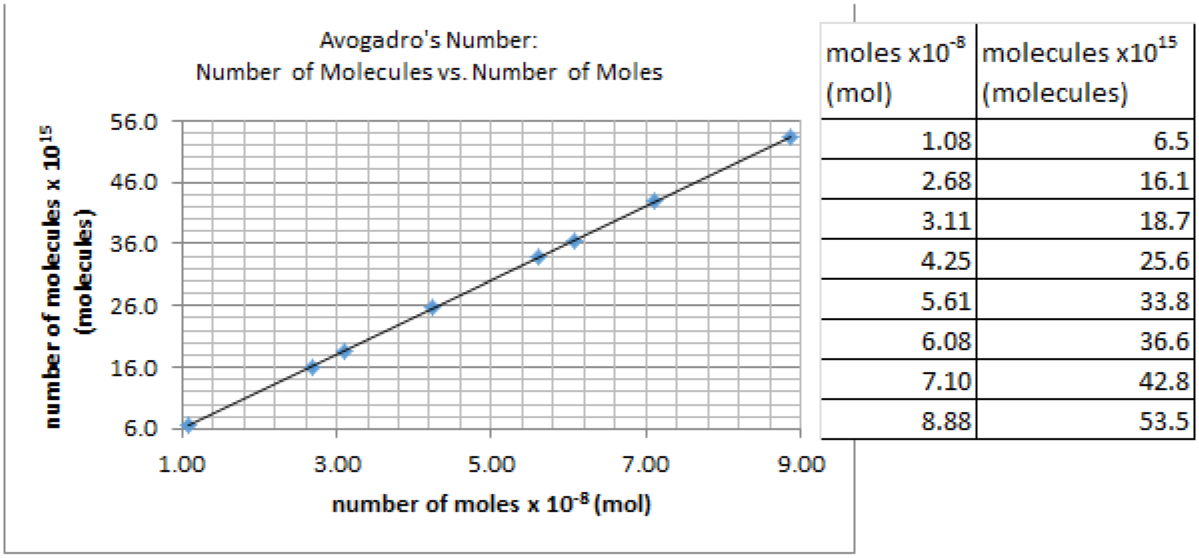
2. The density of stearic acid is 0.870 g/mL. What is the volume of 7.8 x 10^{-6} mole of stearic acid?

3. What assumptions were made in this experiment to arrive at a value for Avogadro's number?

4. What is the mass, in grams, of one molecule of stearic acid?

Answer the following questions about the graph and data below.

FIGURE 7.5



moles x10^{-8} (mol)	molecules x10^{15} (molecules)
1.08	6.5
2.68	16.1
3.11	18.7
4.25	25.6
5.61	33.8
6.08	36.6
7.10	42.8
8.88	53.5

5. Calculate the slope of the line. Remember to use points off of the line not data points and include units.

6. Can the value of the y-intercept, b, be read directly from the graph above? ___________

7. What is the y-intercept?

8. Write the equation for the line in the form Z = mn + b __________________________________

EXPERIMENT 8

Double Displacement Reactions

Introduction

Double displacement reactions are among the most common simple chemical reactions and comparatively easy to study. In each part of this experiment two water solutions, each containing positive and negative ions, will be mixed in a test tube. Consider the hypothetical reaction:

$$AB_{(aq)} + CD_{(aq)} \rightarrow AD_{(?)} + CB_{(?)} \quad \textbf{(EQ 8.1)}$$

where $AB_{(aq)}$ exists as $A^+_{(aq)}$ and $B^-_{(aq)}$ ions in solution and $CD_{(aq)}$ exists as $C^+_{(aq)}$ and $D^-_{(aq)}$ ions in solution. As the ions come in contact with each other, there are six possible combinations that might conceivably cause a chemical reaction. Two of these combinations are the meetings of ions of like charge; that is $A^+_{(aq)} + C^+_{(aq)}$ and $B^-_{(aq)} + D^-_{(aq)}$. Since like charges repel, no reaction will occur. Two other possible combinations are those of the original two compounds; that is, $A^+_{(aq)} + B^-_{(aq)}$ and $C^+_{(aq)} + D^-_{(aq)}$. Since we originally had a solution containing each of these pairs of ions, they can mutually exist in the same solution; therefore they do not recombine. Thus the two possibilities for the combination of each of the positive ions with the negative ion of the other compound is: $A^+_{(aq)} + D^-_{(aq)}$ and $C^+_{(aq)} + B^-_{(aq)}$.

In order for a double displacement reaction to occur one or more of the following must happen:

1. A precipitate is formed.
2. A gas is produced.
3. Water or another slightly ionized substance is formed.

Let us look at some examples:

EXAMPLE 8.1 **Formation of an insoluble salt, criteria 1 met.**

When a solution of sodium chloride and silver nitrate are mixed, the reaction for the hypothetical double displacement reaction is:

$$NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(?)} + AgCl_{(?)} \qquad \textbf{(EQ 8.2)}$$

A white precipitate is produced when these solutions are mixed. This precipitate is definite evidence of a chemical reaction. One of the two products, sodium nitrate ($NaNO_3$) or silver chloride (AgCl) is insoluble. Although the precipitate can be identified by further chemical testing, we can instead look at the **solubility rules Figure 1.1 on page 7** to find that sodium nitrate is soluble, but silver chloride is insoluble. We can then conclude that silver chloride is the precipitate, and the equation can be written in the following manner.

$$\text{Conventional: } NaCl_{(aq)} + AgNO_{3\,(aq)} \rightarrow NaNO_{3\,(aq)} + AgCl_{(s)} \qquad \textbf{(EQ 8.3)}$$

$$\text{Total Ionic: } Na^+_{(aq)} + Cl^-_{(aq)} + Ag^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)} + AgCl_{(s)} \qquad \textbf{(EQ 8.4)}$$

$$\text{Net Ionic: } Cl^-_{(aq)} + Ag^+_{(aq)} \rightarrow AgCl_{(s)} \qquad \textbf{(EQ 8.5)}$$

EXAMPLE 8.2 **No formation of precipitate, gas, or slightly ionizable substance, no reaction**

When solutions of sodium chloride and potassium nitrate are mixed, the equation for the double displacement reaction, hypothetically, is written as:

$$\text{Conventional: } NaCl_{(aq)} + KNO_{3\,(aq)} \rightarrow KCl_{(aq)} + NaNO_{3\,(aq)} \qquad \textbf{(EQ 8.6)}$$

$$\text{Total Ionic: } Na^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + NO_3^-{}_{(aq)} \qquad \textbf{(EQ 8.7)}$$

We get the hypothetical products by simply combining each positive ion with the other negative ion. Has there been a reaction? When we do the experiment we see no evidence of reaction. There is no precipitate formed, no gas evolved, and no obvious evidence of reaction. Thus we must conclude that no reaction occurred. Both of the hypothetical products are soluble salts, so the ions are still present in solution. We can say that we simply have a solution of four kinds of ions, Na^+, Cl^-, K^+, and NO_3^-. This situation is best written as the following:

$$\text{Net Ionic: } NaCl_{(aq)} + KNO_{3\,(aq)} \rightarrow \text{No Reaction} \qquad \textbf{(EQ 8.8)}$$

Acids and Bases

Compounds that contain hydrogen in the cation position act are **acids** when dissolved in water. **Strong acids** the first proton ionizes ~100% in aqueous solution. In chemical equations where ions are included (total and net ionic equations) write the ionized components, since they are the major species present in the solution. Common strong acids are listed in Table 1.8 on page 8.

$$HNO_{3\,(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)} \qquad \textbf{(EQ 8.9)}$$

Weak acids do not dissociate 100% in aqueous solution, they dissociate anywhere from 1-5%, so they are **slightly ionizable substances**. In chemical equations where ions are included (total and net ionic equations) write the complete weak acid formula, not the ions. The ions are only a minor component in the total solution. Some weak acids are listed in Table 1.10 on page 8.

$$HBrO_{3\,(aq)} \rightleftharpoons H^+_{(aq)} + BrO_3^-{}_{(aq)} \qquad \textbf{(EQ 8.10)}$$

Notice in Table 1.8 on page 8, all of the acids are monoprotic except sulfuric acid, which has two ionizable hydrogen atoms. The first hydrogen atom ionizes 100%:

$$H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)} \quad \textbf{(EQ 8.11)}$$

The second hydrogen atom ionizes as a weak acid:

$$HSO_4^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + SO_4^{2-}{}_{(aq)} \quad \textbf{(EQ 8.12)}$$

So in a total ionic or net ionic equation sulfuric acid will ionize into H^+ and HSO_4^-, not 2 H^+ and SO_4^{2-}. Notice that the sulfate ion, SO_4^{2-}, is still formed at the end of the reaction.

Any compound which contains the hydroxide anion is called a **base** or sometimes an **alkali**. **Strong bases** ionize ~100% in water: Some common strong bases are listed in Table 1.9 on page 8.

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} \quad \textbf{(EQ 8.13)}$$

As a result in a unsaturated solution there are only Na^+ ions and OH^- ions, with virtually no NaOH.

Weak bases on the other hand ionize less than 5% depending on the base. Making them **slightly ionizable substances**. As a result in solution most of the base remains in the molecular form and is written as such in total ionic and net ionic equations. Common weak bases include ammonia, NH_3, and methyl amine, CH_3NH_2 are listed in Table 1.10 on page 8.

Examples 1.3 and 1.4, show the reactions of an acid and a base, which is really just another form of a double replacement reaction. It is also called a neutralization reaction. All acid-base reactions result in the formation of a salt and water.

EXAMPLE 8.3Formation of a slightly ionizable substance, criteria 3 met.

When solutions of potassium hydroxide and acetic acid are mixed, the following is the hypothetical chemical equation.

$$\text{Conventional: } KOH_{(aq)} + HC_2H_3O_{2\,(aq)} \rightarrow KC_2H_3O_{2\,(aq)} + H_2O_{(l)} \quad \textbf{(EQ 8.14)}$$

$$\text{Total Ionic: } K^+_{(aq)} + OH^-_{(aq)} + HC_2H_3O_{2\,(aq)} \rightarrow K^+_{(aq)} + C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \quad \textbf{(EQ 8.15)}$$

$$\text{Net Ionic: } OH^-_{(aq)} + HC_2H_3O_{2\,(aq)} \rightarrow C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \quad \textbf{(EQ 8.16)}$$

Notice that potassium hydroxide dissociates into ions, but acetic acid, which is a weak acid, remains as mostly molecules in solution.

The mixture of these solutions produces no visible evidence of reaction, but on touching the test tube we notice that it feels warm. The evolution of heat is evidence of a chemical reaction. This reaction also fulfills the third criteria: the production of water. When a slightly ionized substance is produced the temperature of the mixture rises. More examples of slightly ionized substances are given in Table 1.10 on page 8.

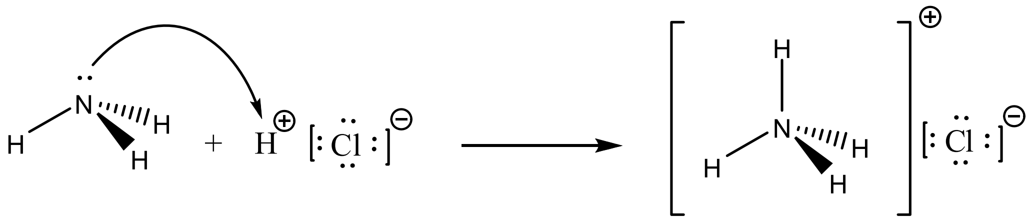
EXAMPLE 8.4 **Formation of a slightly ionizable substance, criteria 3 met.**

When solutions of ammonia and hydrochloric acid are mixed, the equation for the hypothetical double displacement reaction is:

$$NH_{3\,(aq)} + HCl_{\,(aq)} \rightarrow \; ? \qquad \textbf{(EQ 8.17)}$$

There are two ways to think about this reaction of a strong acid, HCl, with a weak base, NH_3.

First, using Lewis structures. Notice that ammonia remains as a molecule in solution, while hydrochloric acid ionized into hydrogen ions, H^+, and chloride ions, Cl^-.



(EQ 8.18)

The lone pair on the nitrogen accepts the hydrogen ion forming the ammonium ion, NH_4^+. So,

$$NH_{3\,(aq)} + HCl_{\,(aq)} \rightarrow NH_4Cl_{\,(aq)} \qquad \textbf{(EQ 8.19)}$$

Second, using the method of ion exchange. $NH_{3\,(aq)}$ may be shown as $NH_3 \cdot H_2O$ or even as NH_4OH, although there is no evidence that an "NH_4OH" molecule exists. However, since aqueous NH_3 reacts as if it was "NH_4OH," it is often written this way in equations.

First, replace $NH_{3\,(aq)}$ with $NH_4OH_{\,(aq)}$ and complete the double displacement reaction:

$$\text{"}NH_4OH\text{"}_{(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)} + H_2O_{(l)} \qquad \textbf{(EQ 8.20)}$$

Then, replace $NH_4OH_{\,(aq)}$ with $NH_{3\,(aq)}$ and $H_2O_{\,(l)}$

$$NH_{3\,(aq)} + H_2O_{(l)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)} + H_2O_{(l)} \qquad \textbf{(EQ 8.21)}$$

Finally, simplify the reaction by canceling $H_2O_{\,(l)}$ from both sides:

$$NH_{3\,(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{\,(aq)} \qquad \textbf{(EQ 8.22)}$$

Notice that both methods give the same overall equation. Now the total and net ionic equations can be written:

$$\text{Conventional: } NH_{3\,(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{\,(aq)} \qquad \textbf{(EQ 8.23)}$$

$$\text{Total Ionic: } NH_{3\,(aq)} + H^+_{\,(aq)} + Cl^-_{\,(aq)} \rightarrow NH_4^+{}_{(aq)} + Cl^-_{\,(aq)} \qquad \textbf{(EQ 8.24)}$$

$$\text{Net Ionic: } NH_{3\,(aq)} + H^+_{\,(aq)} \rightarrow NH_4^+{}_{(aq)} \qquad \textbf{(EQ 8.25)}$$

The mixture of these solutions produces no visible evidence of reaction, but on touching the test tube we notice that it feels warm. The evolution of heat is evidence of a chemical reaction. When a slightly ionized substance is produced the temperature of the mixture rises. More examples of slightly ionized substances are given in Table 1.10 on page 8.

Some reactions produced gases such as H_2, O_2, CO_2, SO_2, NH_3, and H_2S. Many substances form and then decompose such as those in Equation 1.27, Equation 1.31, and Equation 1.32, which you should be familiar with.

EXAMPLE 8.5 **Formation of a gas, criteria 2 met.**

When solutions of sodium carbonate and hydrochloric acid are mixed, the following is the hypothetical chemical equation.

$$Na_2CO_{3\,(aq)} + 2\ HCl_{\,(aq)} \rightarrow 2\ NaCl_{\,(aq)} + H_2CO_{3\,(aq)} \qquad \textbf{(EQ 8.26)}$$

Bubbles of a colorless gas are evolved when these solutions are mixed. Although this gas is evidence of a chemical reaction, neither of the indicated products is a gas. However, carbonic acid, H_2CO_3, decomposes according to the following equation:

$$H_2CO_{3\,(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)} \qquad \textbf{(EQ 8.27)}$$

Therefore, the CO_2 and H_2O are the products that should be written in the equation. The original equation then becomes:

$$\text{Conventional: } Na_2CO_{3\,(aq)} + 2\ HCl_{\,(aq)} \rightarrow 2\ NaCl_{\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)} \qquad \textbf{(EQ 8.28)}$$

$$\text{Total Ionic: } 2\ Na^{+}_{\,(aq)} + CO_3^{2-}{}_{(aq)} + 2\ H^{+}_{\,(aq)} + 2\ Cl^{-}_{\,(aq)} \rightarrow 2\ Na^{+}_{\,(aq)} + 2\ Cl^{-}_{\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)} \qquad \textbf{(EQ 8.29)}$$

$$\text{Net Ionic: } CO_3^{2-}{}_{(aq)} + 2\ H^{+}_{\,(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)} \qquad \textbf{(EQ 8.30)}$$

The following are other substances that decompose into gases as well:

$$H_2SO_{3\,(aq)} \rightarrow H_2O_{\,(l)} + SO_{2\,(g)} \qquad \textbf{(EQ 8.31)}$$

$$\text{“}NH_4OH\text{”}_{\,(aq)} \rightarrow H_2O_{\,(l)} + NH_{3\,(g)} \qquad \textbf{(EQ 8.32)}$$

Some metals ions are able to react with aqueous ammonia to produce a metal hydroxide and ammonium ions.

EXAMPLE 8.6 Formation of a precipitate, criteria 1 met.

When solutions of ammonia and copper(II) chloride are mixed, the following is the hypothetical chemical equation.

$$NH_{3\ (aq)} + CuCl_{2\ (aq)} \rightarrow \quad ? \qquad \textbf{(EQ 8.33)}$$

To complete the double replacement reaction replace $NH_{3\ (aq)}$ with $NH_4OH_{\ (aq)}$ and complete the double displacement reaction:

$$2\ \text{“}NH_4OH\text{”}_{\ (aq)} + CuCl_{2\ (aq)} \rightarrow 2\ NH_4Cl_{\ (aq)} + Cu(OH)_{2\ (s)} \qquad \textbf{(EQ 8.34)}$$

Now replace $NH_4OH_{\ (aq)}$ with $NH_{3\ (aq)}$ and $H_2O_{\ (l)}$

$$\text{Conventional: } 2\ NH_{3\ (aq)} + 2\ H_2O_{\ (l)} + CuCl_{2\ (aq)} \rightarrow 2\ NH_4Cl_{\ (aq)} + Cu(OH)_{2\ (s)} \qquad \textbf{(EQ 8.35)}$$

$$\text{Total Ionic: } 2\ NH_{3\ (aq)} + 2\ H_2O_{\ (l)} + Cu^{2+}_{\ (aq)} + 2\ Cl^-_{\ (aq)} \rightarrow 2\ NH_4^+{}_{\ (aq)} + 2\ Cl^-_{\ (aq)} + Cu(OH)_{2\ (s)} \qquad \textbf{(EQ 8.36)}$$

$$\text{Net Ionic: } 2\ NH_{3\ (aq)} + 2\ H_2O_{\ (l)} + Cu^{2+}_{\ (aq)} \rightarrow 2\ NH_4^+{}_{\ (aq)} + Cu(OH)_{2\ (s)} \qquad \textbf{(EQ 8.37)}$$

During the experiment you will be able to infer solubility based on your experimental observations. At other times solubility rules are useful as shown in Figure 1.1 on page 7.

Summary of Molecules, Formula Units, and Ions in Solution

FIGURE 8.1 **Flow Chart for Separating Compounds into Ions in Aqueous Solution.**

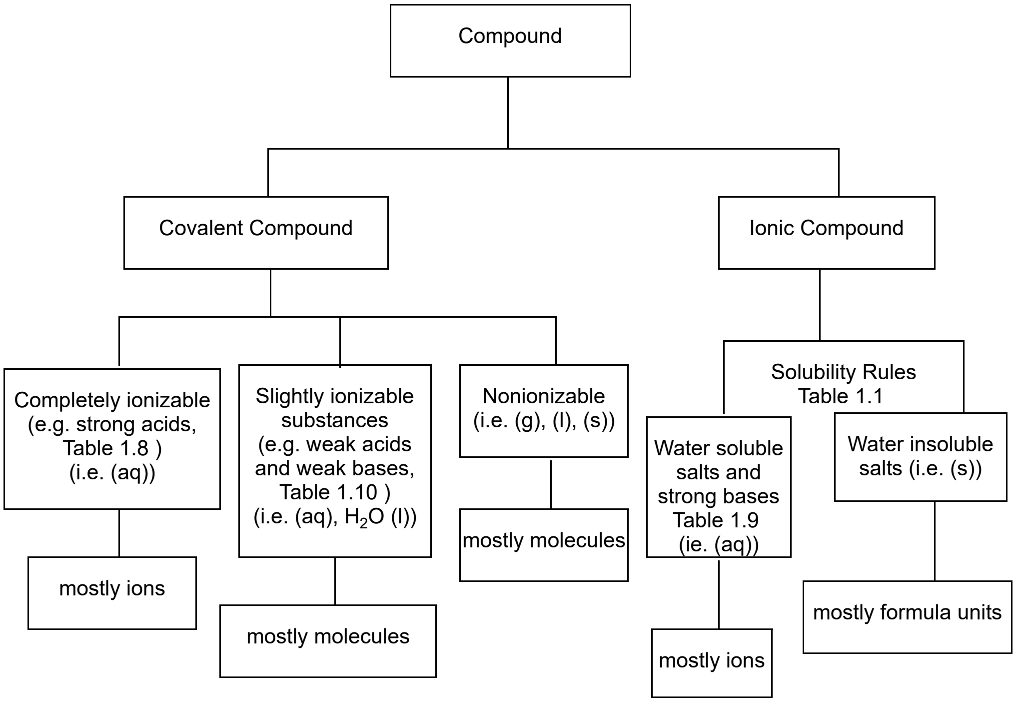


TABLE 8.7 **Solubility Rules**

Soluble Ionic Compound		**Exception**
Compounds Containing	alkali metals, ammonium (NH_4^+),	none
	nitrates (NO_3^-), chlorates (ClO_3^-), perchlorates (ClO_4^-)	none
	acetates ($CH_3CO_2^-$)	aluminum (Al^{3+}) and silver (Ag^+)
	chlorides (Cl^-), bromides (Br^-), iodides (I^-)	Ag^+, mercury(I) (Hg_2^{2+}), and Pb^{2+}
	fluorides (F^-)	alkaline earth metals and lead(II) (Pb^{2+})
	sulfates (SO_4^{2-})	Ca^{2+}, Sr^{2+}, Ba^{2+} insoluble Ag^+, Hg_2^{2+}, and Pb^{2+} slightly soluble
Insoluble Ionic Compound		**Exception**
Compounds Containing	carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), oxalates ($C_2O_4^{2-}$), chromates (CrO_4^{2-}), silicates (SiO_4^{2-})	alkali metals, NH_4^+
	oxides (O^{2-}), hydroxides (OH^-)	alkali metals, NH_4^+ soluble Ca^{2+}, Sr^{2+}, Ba^{2+} slightly soluble
	sulfides (S^{2-})	alkali and alkaline earth metals, NH_4^+

Strong acids and strong bases separate approximately 100% into ions in aqueous solution. They are shown as ions in solution and are broken apart in total ionic and net ionic equations.

TABLE 8.8 **Common Strong Acids**

Name	Dissociation Reaction
Hydrochloric acid	$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
Hydrobromic acid	$HBr_{(aq)} \rightarrow H^+_{(aq)} + Br^-_{(aq)}$
Hydroiodic acid	$HI_{(aq)} \rightarrow H^+_{(aq)} + I^-_{(aq)}$
Perchloric acid	$HClO_{4\,(aq)} \rightarrow H^+_{(aq)} + ClO_{4\,(aq)}^-$
Chloric acid	$HClO_{3\,(aq)} \rightarrow H^+_{(aq)} + ClO_{3\,(aq)}^-$
Nitric acid	$HNO_{3\,(aq)} \rightarrow H^+_{(aq)} + NO_{3\,(aq)}^-$
Sulfuric acid	$H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_{4\,(aq)}^-$
Diprotic acid, both hydrogen react one at a time.	$H_2SO_{4\,(aq)} \rightarrow \rightarrow 2\ H^+_{(aq)} + SO_{4\,(aq)}^{2-}$

TABLE 8.9 **Common Strong Bases**

Name	Ionization Reaction
Lithium hydroxide	$LiOH_{(aq)} \rightarrow Li^+_{(aq)} + OH^-_{(aq)}$
Sodium hydroxide	$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
Potassium hydroxide	$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$
Calcium hydroxide	$Ca(OH)_{2\,(aq)} \rightarrow Ca^{2+}_{(aq)} + 2\ OH^-_{(aq)}$
Strontium hydroxide	$Sr(OH)_{2\,(aq)} \rightarrow Sr^{2+}_{(aq)} + 2\ OH^-_{(aq)}$
Barium hydroxide (slightly soluble)	$Ba(OH)_{2\,(aq)} \rightarrow Ba^{2+}_{(aq)} + 2\ OH^-_{(aq)}$

Notice that the slightly ionizable substances listed are all weak acids or bases. They mostly remain in their molecular form in solution and as a result are not broken apart in total ionic and net ionic equations.

TABLE 8.10 **Slightly Ionized Substances**

	Name	Formula	Decomposes to
	water	$H_2O_{(l)}$	
Acids	acetic acid	$HC_2H_3O_{2\,(aq)}$	
	carbonic acid	$H_2CO_{3\,(aq)}$	$\rightarrow H_2O_{(l)} + CO_{2\,(g)}$
	hydrofluoric acid	$HF_{(aq)}$	
	oxalic acid	$H_2C_2O_{4\,(aq)}$	
	phosphoric acid	$H_3PO_{4\,(aq)}$	
	sulfurous acid	$H_2SO_{3\,(aq)}$	$\rightarrow H_2O_{(l)} + SO_{2\,(g)}$
Bases	aqueous ammonia ($NH_{3\,(aq)} + H_2O_{(l)}$)	$NH_{3\,(aq)}$	
		"NH_4OH" $_{(aq)}$	$\rightarrow H_2O_{(l)} + NH_{3\,(g)}$
	methyl amine	$CH_3NH_{2\,(aq)}$	

Procedure

Each part of this experiment consists of mixing equal volumes of two solutions in a test tube. Use approximately 3 mL of each solution. There is no need to measure this quantity in a graduated cylinder. The important part is to use equal volumes of solution. To determine where the 3 mL mark is on you test tube, fill your 10 mL graduated cylinder with 3 mL of water. Pour this water into the test tube and make a mental note where the fluid level is. This will serve as your 3 mL mark.

Look for evidence of chemical reaction. This may be the formation of a precipitate, the formation of a gas, or the evolution of heat. Make sure that you give the reaction mixtures sufficient time to react. Remember, your observations are to be recorded in ink. However, the interpretation of your observations can be done in pencil.

Formation of a precipitate — Look for the formation of an insoluble compound, the solution may appear cloudy.

Formation of a gas — Look for the formation of bubbles or the solution to effervesce.

Common gases formed during chemical reactions include: hydrogen gas, $H_{2\,(g)}$, oxygen gas, $O_{2\,(g)}$, hydrogen sulfide, $H_2S_{\,(g)}$, ammonia, $NH_{3\,(g)}$, carbon dioxide, $CO_{2\,(g)}$, and sulfur dioxide, $SO_{2\,(g)}$.

When these compounds are “formed” they are unstable and decompose into gases and water:

carbonic acid: $H_2CO_{3\,(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)}$ Equation 1.27;

sulfurous acid: $H_2SO_{3\,(aq)} \rightarrow H_2O_{\,(l)} + SO_{2\,(g)}$ Equation 1.31;

aqueous ammonia, $NH_{3\,(aq)}$ (ammonium hydroxide, “$NH_4OH_{\,(aq)}$”),

“NH_4OH” $_{(aq)} \rightarrow H_2O_{\,(l)} + NH_{3\,(g)}$ Equation 1.32.

Formation of a slightly ionized substance — Heat usually accompanies the formation of water, H_2O, acetic acid, $HC_2H_3O_2$, or any other slightly ionized compound (weak acids, weak bases). Common slightly ionizable substances are shown in Table 1.10 on page 8.

You may also see **turbidity** or waviness in the solution, which also signifies the formation of a slightly ionizable substance.

In each instance where a reaction occurred, write the complete, balanced reaction. Where there is no evidence of reaction write the possible products and show that all ions are spectator ions in the total ionic equation. For the net ionic equation write “No reaction”.

TABLE 8.11 Solutions (Do not record observations here!)

1 Mix 0.1 *M* NaCl and 0.1 *M* $AgNO_3$ solutions.
2 Mix 0.1 *M* NaCl and 0.1 M KNO_3 solutions.
3 Mix 0.1 *M* Na_2CO_3 and 6 M HCl solutions
4 Mix 10% NaOH and dilute (6 *M*) HCl solutions.
5 Mix 0.1 *M* $BaCl_2$ and dilute (3 *M*) H_2SO_4 solutions.
6 Mix dilute (6*M*) NH_3 and dilute (3 *M*) H_2SO_4 solutions.
7 Mix 0.1 *M* $CuSO_4$ and 0.1 *M* $Zn(NO_3)_2$ solutions.
8 Mix 0.1 *M* Na_2CO_3 and 0.1 *M* $CaCl_2$ solutions.
9 Mix 0.1 *M* $CuSO_4$ and 0.1 *M* NH_4Cl solutions.
10 Mix 10% NaOH and dilute (6 *M*) HNO_3 solutions.
11 Mix 0.1 *M* $FeCl_3$ and dilute (6 *M*) NH_3 solutions.
12 Do this part under the hood. Add 1 g of solid Na_2SO_3 to 3 mL of water and shake to dissolve. Add approximately 1 mL of concentrated (12 *M*) HCl solution, drop wise.

Interpretation of observations

In Table 1.12 , Table 1.13 , Table 1.14 , and Table 1.15 on the data sheets:

1. Describe your observations. Remember, your observations are to be recorded in ink.
2. Chemical equations. These may be done in pencil. Make sure you include all phases and charges where necessary. Write:
   a. balanced conventional,
   b. total ionic, and
   c. net ionic equations even if there is no evidence of a reaction to validate that there was indeed no reaction then write NO RXN in the net ionic equation box.

Prelaboratory Exercise: Double Displacement Reactions

Name: ____________________________

Section:_____________________

Lab Grade

Prelab Questions		14
General Format (Signature, ink, no obliterations, etc.)		8
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		48
Post Lab Questions		30
Total		100

1. Give three examples of evidence that a double displacement reaction has occurred.

2. How can you tell if a precipitate forms when you mix two solutions?

3. Balance the following conventional equation:

$$\_\_\_\_\_CaCl_{2\,(aq)} + \_\_\_\_\_K_3PO_{4\,(aq)} \rightarrow \_\_\_\_\_Ca_3(PO_4)_{2\,(s)} + \_\_\_\_\_KCl_{\,(aq)}$$

4. Correct the mistakes in the following chemical formulas:

   a. Na_3Cl_2

   b. $BaOH_2$

   c. $2KSO_4$

5. Would the following compounds be mostly ions, mostly molecules, or mostly formula units in aqueous solution? Figure 1.1 on page 7 may be useful here.

a. oxalic acid, $H_2C_2O_{4\ (aq)}$ ____________________________

b. sodium sulfate, $Na_2SO_{4\ (aq)}$ ____________________________

c. ammonia, $NH_{3\ (aq)}$ ____________________________

d. hydrobromic acid, $HBr_{\ (aq)}$ ____________________________

e. iron(III) hydroxide, $Fe(OH)_{3\ (s)}$ ____________________________

f. potassium hydroxide, $KOH_{\ (aq)}$ ____________________________

Data Sheet: Double Displacement Reactions

TABLE 8.12 Remember to write your phase labels.

1.) Observations NaCl & $AgNO_3$ (See Example 1.1)
Conventional
Total Ionic
Net Ionic
2.) Observations NaCl & KNO_3 (See Example 1.2)
Conventional
Total Ionic
Net Ionic
3.) Observations Na_2CO_3 & HCl (See Example 1.5)
Conventional
Total Ionic
Net Ionic

TABLE 8.13 **Remember to write your phase labels.**

4.) Observations NaOH & HCl
Conventional
Total Ionic
Net Ionic
5.) Observations $BaCl_2$ & H_2SO_4
Conventional
Total Ionic
Net Ionic
6.) Observations NH_3 & H_2SO_4 (See Example 1.4)
Conventional
Total Ionic
Net Ionic

TABLE 8.14 **Remember to write your phase labels.**

7.) Observations $CuSO_4$ & $Zn(NO_3)_2$
Conventional
Total Ionic
Net Ionic
8.) Observations Na_2CO_3 & $CaCl_2$
Conventional
Total Ionic
Net Ionic
9.) Observations $CuSO_4$ & NH_4Cl
Conventional
Total Ionic
Net Ionic

TABLE 8.15 **Remember to write your phase labels.**

10.) Observations NaOH & HNO_3
Conventional
Total Ionic
Net Ionic
11.) Observations $FeCl_3$ & NH_3 (See Example 1.6)
Conventional
Total Ionic
Net Ionic
12.) Observations Na_2SO_3 & HCl (See Example 1.4)
Conventional
Total Ionic
Net Ionic

Post Laboratory Questions

1. Write the equation for the decomposition of sulfurous acid.

2. Using the three criteria for double displacement reactions, and the solubility table, predict whether a double displacement reaction will occur for each of the following. If the reaction will occur, balance the reaction write conventional, total ionic and net ionic equations. **Box the net ionic equation.** Be sure to include the proper phase labels. If no reaction will occur, write "no reaction". All reactants are aqueous solutions.
    - a. Potassium sulfide and copper(II) sulfate →
    - b. Ammonium sulfate and sodium chloride →
    - c. Sodium chromate and lead(II) acetate →

d. Bismuth(III) chloride and sodium hydroxide →

e. Potassium acetate and cobalt(II) sulfate →

f. Zinc bromide and potassium phosphate →

g. Iron(III) chloride and ammonium nitrate →

h. Calcium chlorate and sodium bromide →

i. Sodium chromate and barium nitrate →

j. Sodium carbonate and nitric acid →

k. Potassium hydroxide and ammonium chloride →

l. Sodium acetate and hydrochloric acid →

m. Aqueous ammonia and oxalic acid →

EXPERIMENT 9 Chemical Reactions

Background

In this experiment you will detect and observe a variety of chemical reactions. To experience the five basic types of chemical reactions and to express these reactions using balanced chemical equations.

Chemical change is always associated with a chemical reaction. Atoms, molecules or ions rearrange to form new substances. The substances doing the rearranging are called reactants. The new substances are called products. Keep in mind; the *Law of Conservation of Matter and Energy* is always obeyed. Therefore, all atoms present in the reactants must be accounted for in the products.

There are a few basic observations that indicate a chemical reaction as occurred:

1. A precipitate forms. That is, insoluble solid particles appear in solution.
2. A gas evolves. Bubbles are seen in the reaction solution. The gases we will test for are, oxygen, hydrogen and carbon dioxide.
    a. If the gas evolved is oxygen, O_2, the splint will glow brighter.
    b. If the gas evolved is hydrogen, H_2, you will hear a characteristic “bark” as the hydrogen ignites when an ignited splint is brought to the mouth of the test tube.
    c. Finally, if the gas is carbon dioxide, CO_2, the ignited splint will be extinguished.
3. A color change. If, after mixing the reactants you see a permanent (60 seconds or longer) color change, a reaction took place.
4. Heat is given off (exothermic) or absorbed (endothermic). The reaction vessel becomes warm or cold, or light is given off.
5. Change in pH. An indicator can be used to determine a change in acidity. Remember, blue litmus paper turns red in acidic solution and red litmus turns blue in basic solutions. Phenolphthalein is colorless in its acid form and pink in its base form.

There are five basic types of chemical reactions. They are tabulated as follows:

TABLE 9.1

	Reaction Type	Description	Example
Oxidation-Reduction Reactions (Redox) These are electron transfer reactions.	Combination	Two or more substances combine to form one substance.	$Fe_{(s)} + 3\ Cl_{2\ (g)} \rightarrow 2\ FeCl_{3\ (s)}$ **(EQ 9.1)** $2\ Mg_{(s)} + O_{2\ (g)} \rightarrow 2\ MgO_{(s)}$ **(EQ 9.2)**
	Combustion	Substances combine with oxygen to form one or more oxygen containing compounds, and may also produce light and heat.	Organic compounds react with oxygen to produce carbon dioxide, water, light and heat. $CH_{4\ (g)} + 2\ O_{2\ (g)} \rightarrow CO_{2\ (g)} + 2\ H_2O_{(l)}$ **(EQ 9.3)** A hydrocarbon reacts with oxygen to produce carbon and hydrogen oxides as well as heat and light.
	Decomposition	One substance breaks down into two or more.	**(EQ 9.4)** $2\ KClO_{3\ (s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2\ (g)}$ **(EQ 9.5)**
	Single Replacement	One element reacts with a compound to replace one of the elements of that compound.	$2\ HCl_{(aq)} + Mn_{(s)} \rightarrow MnCl_{2\ (aq)} + H_{2\ (g)}$ **(EQ 9.6)** $Cl_{2\ (g)} + 2\ KI_{(aq)} \rightarrow I_{2\ (s)} + 2\ KCl_{(aq)}$ **(EQ 9.7)**
Double Replacement Reactions These are reactions where ions exchange partners.	Ion Exchange	Cations of both reactants swap places.	$AgNO_{3\ (aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\ (aq)}$ **(EQ 9.8)**
	Acid-base Neutralization	Hydrogen ions are transferred between reactants.	$HNO_{3\ (aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaNO_{3\ (aq)}$ **(EQ 9.9)**

Combination Reactions (Redox Reactions)

In combination reactions reactants combine to form a new substance. Electrons are transferred from one reactant to another so oxidation and reduction take place.

$$A + E \rightarrow AE \qquad \textbf{(EQ 9.10)}$$

Combustion reactions are a specialized type of combination reaction involving oxygen gas as one of the reactants:

$$A + O_{2\ (g)} \rightarrow AO \qquad \textbf{(EQ 9.11)}$$

Single Replacement Reactions (Redox Reactions)

The reactivity of an element is related to its tendency to lose or gain electrons; that is, to be oxidized or reduced. It's possible to arrange nearly all elements into a single series in order of their reactivity. This is known as an **activity series.** A more active element will replace an element from a compound. Generally speaking:

$$A_{(s)} + BC_{(aq)} \rightarrow B_{(s)} + AC_{(aq)} \qquad \textbf{(EQ 9.12)}$$

where A is the more active element and replaces B in the compound. Metals of known charge (alkali metals, alkaline earth metals, aluminum, zinc, silver, etc.) will have their standard oxidation states. Metals of variable charge (transition metals, inner transition metals, heavy metals) can have many different oxidation states depending on the reaction. For the purposes of this experiment metals of variable charge will obtain a +2 oxidation state, except silver which has an oxidation state of +1 being a metal of known charge.

Double Displacement Reactions (aka Ion Exchange Reactions)

In double displacement reactions two aqueous solutions are mixed together to produce a precipitate, slightly ionizable substance, or a gas. For more on this topic review “Double Displacement Reactions” on page 139.

$$AB_{(aq)} + CD_{(aq)} \rightarrow AD_{(?)} + BC_{(?)} \quad \textbf{(EQ 9.13)}$$

Writing Ionic Equations

When you write ionic equations, you need to show the principal species present in the solutions. For a more detailed discussion see “Double Displacement Reactions” on page 139 and your textbook. Some questions to consider:

Are the reactants and products molecular compounds, ionic compounds, or acids?

What are the phase states? Solids, liquids, gases, or aqueous? Remember that solids, liquids, and gases remain unionized in ionic and net ionic equations.

Do the species fully dissociate in solution (i.e. break apart completely into ions in total ionic and net ionic equations)? Species such as water soluble ionic compounds, strong acids, and strong bases.

Partially ionize (i.e. species that are mostly compounds in solution with few ions)? Species such as weak acids and weak bases.

EXAMPLE 9.2 Consider the reaction of copper(II) chloride and potassium phosphate. Once the reactants are mixed a blue precipitate is visible.

a. Write the balanced conventional equation for the reaction.

Remember that the conventional equation shows all species as neutral compounds.

$$3\ CuCl_{2\,(aq)} + 2\ K_3PO_{4\,(aq)} \rightarrow Cu_3(PO_4)_{2\,(s)} + 6\ KCl_{(aq)} \quad \textbf{(EQ 9.14)}$$

b. Write the total ionic equation for the reaction.

Remember that the total ionic equation shows all species as they appear in solution.

$$3\ Cu^{2+}{}_{(aq)} + 6\ Cl^{-}{}_{(aq)} + 6\ K^{+}{}_{(aq)} + 2\ PO_4^{3-}{}_{(aq)} \rightarrow Cu_3(PO_4)_{2\,(s)} + 6\ K^{+}{}_{(aq)} + 6\ Cl^{-}{}_{(aq)} \quad \textbf{(EQ 9.15)}$$

c. Write the net ionic equation for the reaction.

The net ionic equation shows only the species that undergo reaction. No spectator ions. Remember to simplify the coefficients when necessary.

$$3\ Cu^{2+}{}_{(aq)} + 2\ PO_4^{3-}{}_{(aq)} \rightarrow Cu_3(PO_4)_{2\,(s)} \quad \textbf{(EQ 9.16)}$$

EXAMPLE 9.3Consider the reaction of copper(II) nitrate and zinc metal. After several minutes, a deposit forms on the zinc metal and the blue copper(II) nitrate solution fades to colorless. Therefore, zinc is more active than copper and a reaction takes place.

a. Write the balanced conventional equation for the reaction.

$$Cu(NO_3)_{2\,(aq)} + Zn_{(s)} \rightarrow Zn(NO_3)_{2\,(aq)} + Cu_{(s)} \qquad \textbf{(EQ 9.17)}$$

Zinc solid was able to replace copper(II) ions from solution. Therefore, zinc is a more active element than copper.

Notice that the charge on zinc metal, $Zn_{(s)}$, and copper metal, $Cu_{(s)}$, is zero, 0. The zero is often omitted because a charge of zero is understood for substances in their elemental states.

b. Write the total ionic equation for the reaction.

$$Cu^{2+}{}_{(aq)} + 2\ NO_3^{-}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2\ NO_3^{-}{}_{(aq)} + Cu_{(s)} \qquad \textbf{(EQ 9.18)}$$

c. Write the net ionic equation for the reaction.

$$Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)} \qquad \textbf{(EQ 9.19)}$$

A practical approach to understanding redox reactions is to separate the reduction step from the oxidation step by writing a half-reaction omitting the spectator ions.

$$\text{oxidation half-reaction: } Zn^{0}{}_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + 2\ e^{-} \qquad \textbf{(EQ 9.20)}$$

$$[Ar]\ 4s^2\ 3d^{10} \qquad [Ar]\ 3d^{10}$$

Notice that the charge on zinc goes from zero to a charge of plus two. As the electron configuration shows electrons are lost by zinc, which is **oxidation**.

$$\text{reduction half-reaction: } Cu^{2+}{}_{(aq)} + 2\ e^{-} \rightarrow Cu^{0}{}_{(s)} \qquad \textbf{(EQ 9.21)}$$

$$[Ar]\ 3d^{10} \qquad\qquad [Ar]\ 4s^2\ 3d^{10}$$

Notice that the charge on copper goes from two plus to a charge of zero. As the electron configuration shows electrons are gained by copper, which is **reduction**.

$$\text{overall redox reaction: } Cu^{2+}{}_{(aq)} + Zn^{0}{}_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cu^{0}{}_{(s)} \qquad \textbf{(EQ 9.22)}$$

Notice that since the electrons appear as a product in the oxidation step and as a reactant in the reduction step, they “cancel out” in the balanced net ionic overall redox reaction. Of course the electrons do not simply disappear. Fundamentally, at the microscopic level, the electrons are transferred from the zinc atoms to the copper(II) ions. This illustrates the principle that oxidation only occurs in synchromicity with reduction.

EXAMPLE 9.4Consider the reaction of chromium(II) nitrate and copper metal. After a prolonged period of time no color changes, no precipitates, and no formation of gases is observed. Therefore, chromium is more active than copper and no reaction takes place.

a. Write the balanced conventional equation for the reaction.

$$Cr(NO_3)_{2\,(aq)} + Cu_{(s)} \rightarrow \text{no reaction} \qquad \textbf{(EQ 9.23)}$$

Copper solid was not able to replace chromium(II) ions from solution. Therefore, chromium is a more active element than copper. No oxidation or reduction reaction took place. This implies that chromium will lose its electrons to copper, but copper will not lose its electrons to chromium. To test this hypothesis, a strip of metallic chromium could be immersed in a copper(II) nitrate solution.

Using Example 9.3 and Example 9.4 the more active and less active metal can be determined. As well as the element oxidized and reduced.

TABLE 9.5 **Single Replacement Reaction Examples Comparison**

Net Ionic Equation	More Active Element	Less Active Element	Element Oxidized	Element Reduced
$Cu^{2+}{}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$	Zn	Cu	Zn	Cu
$Cr^{2+}{}_{(aq)} + Cu_{(s)} \rightarrow$ no reaction	Cr	Cu	n/a	n/a

So, Zn > Cu, and Cr > Cu, however the reactivity of Zn versus Cr is not known. In order write an activity series for these three elements a third reaction would need to be done between zinc metal and chromium(III) nitrate solution or zinc nitrate solution and chromium metal for example.

Procedure

The experimental instructors are written on the data pages. Read the instructions for each reaction carefully and record your observations in ink in the space provided. Think about what species are present before and after mixing and record those in the appropriate spaces. Finally write the equation for the reaction.

Caution! Safety Notes:

Wear safety goggles! Make sure to dispose of all chemicals according to your instructor's direction!

Solid metal waste you be placed into the solid metal waste container. Solutions and precipitates should be poured into the inorganic waste container.

NOTE: Observations are descriptions of the changes you see occurring. Written notes of observations are considered data. Therefore, record observations in ink on your lab report.

Prelaboratory Exercise: Chemical Reactions

Name: ___________________________

Section:____________________

Lab Grade

Prelab Questions		10
General Format (Signature, ink, no obliterations, etc.)		7
Data and Analysis (describe reactions, identify reactants and products, answer questions)		43
Writing Reactions (complete and correct conventional, ionic and net ion equations)		40
Total		100

1. Looking at the formula for copper(II) carbonate, $CuCO_3$
   a. which two of the three gases would you want to test for upon heating the solid?

   ________________ and __________________.

   b. what tests would you use to confirm the identity of gas?

2. An acid turns the indicator litmus to a ____________ color.

3. We can sometimes tell when a chemical reaction has taken place by:
   a. a change in color of the substance
   b. the evolution of an identifiable gas
   c. either or both of the above

4. You have produced a gas, which you suspect to be hydrogen. Describe how you might confirm your suspicion. State exactly what you would do and what would you expect to observe?

5. Is there any part of this experiment where the top loader balance is used? ________________

6. Identify the type of reactions shown below as a combination reaction, a combustion reaction, a decomposition reaction, a single replacement reaction, a double replacement reaction, or an acid-base neutralization reaction.

a. $C_2H_{2\,(g)} + Br_{2\,(l)} \rightarrow C_2H_2Br_{2\,(g)}$ ______________________________________

b. $C_3H_{8\,(g)} + 5\ O_{2\,(g)} \rightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{\ (l)}$ ______________________________________

c. $2\ NaN_{3\,(s)} \rightarrow 2\ Na_{\ (s)} + 3\ N_{2\,(g)}$ ______________________________________

Data Sheet: Chemical Reactions

Part I. Combination

1. Obtain a 1-inch piece of magnesium ribbon. Hold the piece of Mg ribbon with your crucible tongs and ignite it with your Bunsen burner. **Caution!** ***Do not look directly at the burning ribbon***.

TABLE 9.6 Observations combustion of magnesium ribbon

a. Is the reaction exothermic or endothermic?	
Reactants	**Products**
Balanced Chemical Equation	

Part II. Combustion

2. Pour approximately 1 mL of heptane, C_7H_{16}, into an evaporating dish. (Measure the heptane in a graduated cylinder). Carefully ignite the heptane by holding a burning splint over the heptane until it combusts. Use a test tube holder to invert a test tube over the flame produced by the heptane. Observe the test tube. Disregard the black soot. While the test tube is inverted light a splint from the heptane flame. Quickly move the inverted test tube away from the flame and insert the burning splint into the inverted test tube. What happened to the burning splint? Carefully note your observations.

TABLE 9.7 Observations combustion of heptane

a. After the smoke has cleared, what do you see collect midway up the inverted test tube?	
b. Describe what happened when you held the test tube over the burning heptane and tested the contents. What gas do you think was collected? Why?	
Reactants	**Products**
Balanced Chemical Equation	

Part III. Single Replacement. With the following reactions, evidence of a reaction may not be immediately apparent. Before making a decision be sure to let the reaction stand for approximately 10 minutes. If no reaction occurs, write, “NO RXN” in under balanced chemical equations; Do not write products or other equations.

1. Fill a test tube half full with DI water and place it behind the safety shield in the hood. Sodium metal is stored under a non-reactive solvent such as hexane or oil. Take a small piece of sodium from the stock bottle and place it on a piece of filter paper, being careful not to touch it with your hands. Blot the sodium with the filter paper. As you blot the sodium, apply pressure.

   Using tongs or tweezers pick up the piece of sodium and drop it into the test tube of water, concurrently ignite a splint. While the reaction is proceeding quickly take a burning splint and bring it to the mouth of the test tube as the sodium is reacting. Carefully note your observations.

   Use a stir rod to mix your solution and put a drop on red litmus paper. Phenolphthalein is an acid/base **indicator**, which is colorless in acid solution (contains H^+ ions) and is bright pink in base solution (contains OH^- ions). Put a few drops of phenolphthalein into the test tube.

TABLE 9.8 **Observations sodium metal and water**	
a. Is the sodium malleable?	
b. What happened when the sodium was placed in water?	
c. What gas do you think was evolved? Explain your answer using your observations.	
d. What color does red litmus paper turn? Is the solution acidic or basic?	
e. What color is the reaction solution once the phenolphthalein is added?	
f. Does the solution contain hydrogen ions, H^+, or hydroxide ions, OH^-?	
Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	

2. Obtain and clean with sandpaper to expose the surface: 3 pieces of zinc, 2 pieces of copper and 1 piece of lead. Place 6 test tubes in a rack and label each appropriately. Observe the contents of each of the test tubes for evidence of a reaction. Record your observations in the tables below.

TABLE 9.9 **Observations Tube 1: Copper strip and approximately 4 mL silver nitrate**

Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

TABLE 9.10 **Observations Tube 2: Lead strip and about 4 mL copper(II) nitrate**

Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

TABLE 9.11 **Observations Tube 3: Zinc strip and about 4 mL lead(II) nitrate**

Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

TABLE 9.12 **Observations Tube 4: Zinc strip and about 4 mL magnesium sulfate**

Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

TABLE 9.13 **Observations Tube 5: Copper strip and about 4 mL 3 M sulfuric acid**	
Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

TABLE 9.14 **Observations Tube 6: Zinc strip and about 4 mL of 3 M sulfuric acid**	
Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	
Which element is oxidized?	Which element is reduced?
Which element is more active?	Which element is less active?

a. From your observations, arrange Cu, Ag, Pb, Zn, and Mg in order of their activities, listing the most active first. Explain how you determined this order **using your experimental data**.

b. Where would H^+ fit in the activity series you developed in question 1 (**use your experimental data**)? Can you conclusively determine its position? Why or why not? If not, what experiment could you perform to determine its position?

c. Would magnesium react with dilute sulfuric acid? Why or why not? If so, write a balanced chemical reaction?

Part IV. Decomposition

Into a test tube, place about 2 mL of 3% hydrogen peroxide along with a pea size amount of MnO_2. Manganese dioxide, the common name for manganese(IV) oxide, is used as a catalyst. Remember, that a **catalyst** is used to lower the activation energy of a reaction, thereby speeding up the reaction. It does not get consumed in the reaction and therefore is not part of your balanced chemical equation. It is written above the reaction arrow in a chemical equation.

Note any evidence of the reaction. Has a gas been evolved? Light a wood splint and blow out the flame. Put the glowing splint into the test tube. What happens?

Remember, when writing your balanced equation, H_2 gas **was not** produced.

TABLE 9.15 **Observations decomposition of hydrogen peroxide**	
a. If a gas was evolved, what was the identity of the gas? Explain how you came to this conclusion using experimental evidence.	
Reactants	**Products**
Balanced Chemical Equation	

Part V. Double Replacement

For the following reaction be sure to include a total ionic equation and a net ionic equation.

1. In a small test tube, place approximately 3 mL of 0.1 M $BaCl_2$ and add 3 mL of 0.1 M K_2SO_4. Carefully note your observations.

TABLE 9.16 **Observations reaction of barium chloride and potassium sulfate**

One way to determine which substance precipitated out is to do a little bit of investigation. Place 3 mL each of DI water into two test tubes. Add a pea size amount of potassium chloride into one test tube and the same amount of barium sulfate into the other. Stopper and shake the two tubes. Note your observations.	
a. What happened to the potassium chloride solid when mixed with water?	
b. What happened to the barium sulfate solid when mixed with water?	
c. What does this tell you about the precipitate that formed?	
Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	

2. In a *clean, dry* test tube, place a small amount of calcium carbonate. Carefully and slowly add about 3 mL of 6 M HCl solution. Remember that carbonic acid is NOT a gas, but an aqueous acid that decomposes into carbon dioxide and water.

   Allow the reaction to proceed for 60 seconds, the insert a lit splint into the test tube. What did you observe? From your observation, determine what gas is evolved.

TABLE 9.17 **Observations reaction of calcium carbonate and hydrochloric acid**

a. What did you observe after you allowed the reaction to proceed for 60 seconds? Explain how you used this information to determine the identity of the gas.	
Reactants	**Products**
Balanced Chemical Equation	
Total Ionic Equation	
Net Ionic Equation	

EXPERIMENT 10 *Green Limiting Reagent*

Background

In this experiment you will learn more about the concept of a limiting reactant in a chemical equation. When two or more reactants react chemically they do so in a certain mole ratio (the stoichiometric ratio). If one of the reactants is present in an amount less than the stoichiometric proportion to that of the other (or those of the others), it will limit the amount of products produced. One reactant is present in **excess**, and the one that limiting the amount of product that can be formed is called the **limiting reactant** (or limiting reagent). A method for determine the excess and limiting reactant is shown below.

Read the experiment and review the material on limiting reagent stoichiometry in your lecture textbook.The initial change end (or equilibrium) (ICE) table method will be used as illustrated below.

1. Suppose that the initial mass of lead(II) acetate trihydrate is 3.000 g and the initial mass of potassium iodate is 3.000 g. To find the actual moles of both reagents first convert grams of both reagents to moles:

$$3.000\ g\ Pb(C_2H_3O_2)_2 \bullet 3H_2O \times \frac{1\ mol\ Pb(C_2H_3O_2)_2 \bullet 3H_2O}{379.3\ g\ Pb(C_2H_3O_2)_2 \bullet 3H_2O} = 0.007909\ mol\ Pb(C_2H_3O_2)_2 \bullet 3H_2O$$

   When a hydrate is dissolved in water the water of hydration becomes part of the solution. So,

$$0.007909\ mol\ Pb(C_2H_3O_2)_2 \bullet 3H_2O \times \frac{1\ mol\ Pb(C_2H_3O_2)_2}{1\ mol\ Pb(C_2H_3O_2)_2 \bullet 3H_2O} = 0.007909\ mol\ Pb(C_2H_3O_2)_2$$

2. Using the reactants a balanced chemical equation can be written. The theoretical stoichiometric mole ratio can be seen from the equation:

$$\mathrm{Pb(C_2H_3O_3)_{2\,(aq\,)} + 2\ KIO_{3\,(aq\,)} \rightarrow Pb(IO_3)_{2\,(s)} + 2\ KC_2H_3O_{2\,(aq\,)}} \qquad \textbf{(EQ 10.1)}$$

   where one mole of lead(II) acetate react with two moles of potassium iodate forming one mole of lead(II) iodate and two moles of potassium acetate. Often you will be required to determine the balanced chemical equation.

3. An ICE table will be used. Each part will be discussed separately, but in practice will be combined.

a. To determine the limiting reagent, compare your actual mole ratio to the theoretical mole ratio obtained from the balanced chemical equation.

Theoretical mole ratio	Actual mole ratio	Limiting Reagent is KIO_3
$\frac{2\ mol\ KIO_3}{1\ mol\ Pb(C_2H_3O_2)_2}=2$	$\frac{0.01402\ mol\ KIO_3}{0.007909\ mol\ Pb(C_2H_3O_2)_2}=1.773$	

The actual ratio of 1.773 is lower than the theoretical mole ratio, which means that the numerator of the ratio, namely “mol KIO_3” is too low, and so KIO_3 is the limiting reagent.

b. To use the ICE table fill in the initial moles of reactants and products will be filled in. Notice that grams is not used here!

	$Pb(C_2H_3O_3)_{2\ (aq)}$	$+ 2\ KIO_{3\ (aq)}$	$\rightarrow Pb(IO_3)_{2\ (s)}$	$+ 2\ KC_2H_3O_{2\ (aq)}$
Initial moles	0.007909 mol	0.01402 mol	0.00000 mol	0.0000 mol

c. Next the change of moles will be filled in. At this point the change in moles is not known so “x” is used. For $Pb(C_2H_3O_2)_2$ the change is -1x because we are losing reactant, where the 1 comes from the coefficient of the balanced chemical equation. The change for $KC_2H_3O_2$ is +2x because we are gaining product, where the 2 comes from the balanced equation.

Change moles	-x	-2x	+x	+2x

d. Then the end moles is found:

End moles	0.007909 moles -x	0.01402 moles -2x	x	2x

e. Now “x” can be found by setting the end moles of the limiting reagent equal zero and solving for “x.” In this case potassium iodate was found to be the limiting reagent:

0.01402 moles - 2x = 0 moles

x = 0.007010 moles

f. Complete the end moles by plugging in “x” to find the end moles of all reactants and products::

End moles numerical	0.007909 mol -x = 0.007909 mol - 0.007010 mol = 0.000899 mol left	0.01402 mol -2x 0.01402 mol - 2(0.007010 mol) = 0.00000 mol left	x = 0.007010 mol	2x = 2(0.00710 mol) = 0.01402 mol

4. Finally, convert moles of your precipitate into grams to calculate the theoretical yield:

$$0.007010\ \text{mol}\ Pb(IO_3)_2 \times \frac{557.00\ \text{g}\ Pb(IO_3)_2}{1\ \text{mol}\ Pb(IO_3)_2} = 3.905\ \text{g}\ Pb(IO_3)_2$$

5. Lastly, calculate the percent yield:

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \qquad \textbf{(EQ 10.2)}$$

In the past various reagents have been used for this experiment including lead(II) nitrate and potassium iodate. In an effort to increase awareness of sustainability in the next generation of the California workforce this protocol was developed with California SB 70 funding. In other words these protocols “green” our labs and “green” our students. Therefore lead(II) nitrate and potassium iodate have been replaced by calcium chloride dihydrate and sodium carbonate which produce calcium carbonate, which is essentially chalk and can be disposed of in the regular trash can, and aqueous table salt (sodium chloride), which can be poured down the sink.

There are twelve principles of green chemistry two of which are addressed in this revision:

3) Less Hazardous Chemical Syntheses: Wherever practicable, synthetic methods should be designed to use and generate substances that possess little or no toxicity to human health and the environment.

4) Designing Safer Chemicals: Chemical products should be designed to affect their desired function while minimizing their toxicity.

Method. For this experiment you will work with the mathematical reaction between calcium chloride dihydrate and anhydrous sodium carbonate, which produces insoluble calcium carbonate. From the balanced reaction equation you will determine the correct stoichiometric ratio of reactants. You will then combine different weights of reactants, collect and weigh the insoluble product, and compare the results to those predicted by the equation.

Procedure

1. A letter, corresponding to a certain amount of each of the two reactants, will be assigned to you:

TABLE 10.1

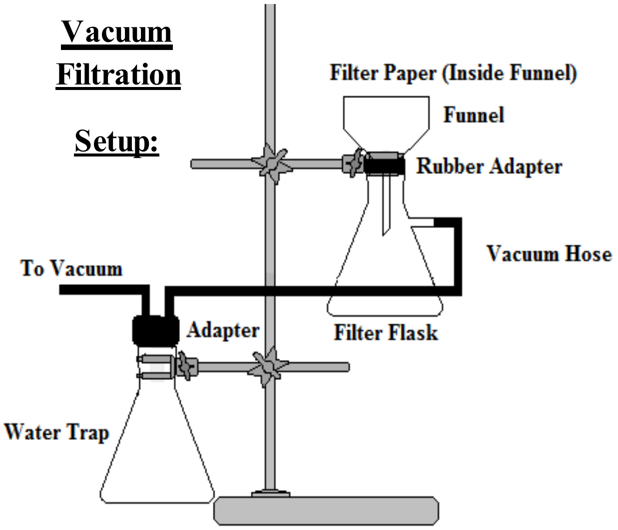
Letter	mass Na_2CO_3 (grams)	mass $CaCl_2 \cdot 2\ H_2O$ (grams)
A	1.000	1.650
B	1.000	1.000
C	1.100	1.750
D	1.050	1.500
E	1.200	1.300
F	1.250	1.250
G	1.100	1.750
H	1.250	1.100

2. Obtain from the reagent shelf—in two separate containers, such as two 50 mL beakers—quantities of those reagents approximating the values assigned. (Sample sizes will be on display on the instructor’s desk.)
3. Weigh out the proper amounts to the nearest milligram. Try to get within ± 0.050 grams of your assigned quantity. Remember it is more important to know exactly how much of each reactant you have than to get exactly the amount assigned.

> **NOTE:** We will use the method of mass by difference to minimize any systematic errors in the weighing process. Be sure to use the same balance throughout the experiment!

4. Do this by transferring reagent with your spatula from your beaker to a pre-weighed 250 mL beaker, until the assigned amount is reached.
5. To each 250 mL beaker add 50 mL of DI water and dissolve both reagents.
6. With a stirring rod, gently pour and mix the calcium chloride dihydrate solution into the sodium carbonate solution. Remember “glass-to-glass-to-glass.”
7. Rinse the beaker with DI water to assure a quantitative transfer of calcium chloride dihydrate. Gently stirring the mixture and set it aside.
8. Meanwhile set up a Büchner funnel and vacuum train. The equipment may be checked out from the stockroom and the instructor will demonstrate the set-up.

FIGURE 10.1



9. Weigh a piece of fiberglass 7.0 cm filter paper to the nearest milligram and record the weight.
10. Prepare your watch glass by labeling it on the bottom so as not to interfere with your sample. Then weigh the watch glass to the nearest milligram and record the weight.
11. Place the filter paper into the Büchner funnel and dampen it to make it stick. Slowly turn on the vacuum line. (Too great a vacuum will cause the paper to tear.)
12. Carefully pour the supernatant into the funnel without transferring the solid. Don't forget to use the technique of glass-to-glass-to-glass here.
13. Slowly pour the rest of the contents into the funnel so that the suspended precipitate goes onto the center part of the filter. If you can keep the precipitate back from the edge, it will be much easier to remove the filter paper, dry it and weigh it without loss of any solid.
14. Remove the remaining precipitate from the beaker by using three 10 mL portions of water and your rubber policeman to scrub the solid calcium carbonate, $CaCO_3$, from the sides of the beaker.
15. Allow the suction to dry the precipitate for a minute or so; then turn off the vacuum and carefully remove the filter paper and precipitate, losing as little as possible. Dry the paper and reaction product on a labeled and weighed watch glass for 30 minutes in an oven set at 110°C.
16. Weigh the cooled filter paper, watch glass and precipitate to the nearest milligram and record.
17. If at all possible, reheat your filter paper and precipitate for another 30 minutes. Cool and reweigh. The second weighing should agree with the first one to within 0.004 g.

Calculations

Determine the expected weight of product from the reaction equation and your <u>actual</u> initial weights. Compare the expected weight with your result and calculate your percent yield.

Prelaboratory Exercise: Limiting Reagent

Lab Grade

Prelab Questions		26
General Format (Signature, ink, no obliterations, etc.)		8
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		20
Accuracy (% error)		20
Post Lab Questions		26
Total		100

Name: ___________________________

Section:____________________

1. A student weighs out 1.100 grams of copper(II) chloride dihydrate and 3.500 grams of silver nitrate:

a. Calculate the molar masses for the reactants:

	$CuCl_2 \cdot 2\ H_2O$	$AgNO_3$
Molar Mass (g/mol)		

b. Calculate the moles of copper(II) chloride dihydrate.

c. Calculate the moles of copper(II) chloride from the moles of copper(II) chloride dihydrate.

d. What happens to the water of hydration when the solid dissolves?

e. Calculate the moles of silver nitrate.

f. Write the balanced chemical equation for the reaction of copper(II) chloride and silver nitrate.

g. Determine the limiting reagent by comparing the theoretical mole ratio to the actual mole ratio:

Theoretical mole ratio	Actual mole ratio	Limiting Reagent is
$\frac{CuCl_2}{AgNO_3} =$	$\frac{CuCl_2}{AgNO_3} =$	

h. Write the balanced chemical equation for the reaction and complete the table below. Be sure to use the correct number of significant figures. Remember that leading zeros are not considered significant

Reaction				
Initial moles				
Change moles				
End moles				
End moles numerical				

i. What does x equal?

j. Calculate the molar masses for the products:

	$Cu(NO_3)_2$	AgCl
Molar Mass (g/mol)		

k. How many grams of precipitate can be collected?

l. If 0.8525 g of silver chloride is collected what is the percent yield?

m. How many grams of the excess reagent are left at the end of the reaction?

2. Complete the table below:

	Na_2CO_3	$CaCl_2 \cdot 2\ H_2O$	$CaCO_3$	NaCl
Molar Mass (g/mol)				

Data Sheet: Limiting Reagent

TABLE 10.2

Data	
Assigned Letter	
Spreadsheet ID	
Actual weight of Na_2CO_3 used mass Na_2CO_3 + beaker =	
mass beaker =	
mass Na_2CO_3 =	
Describe the Na_2CO_3	
Actual moles Na_2CO_3 used	
Actual weight of $CaCl_2 \cdot 2\ H_2O$ used mass $CaCl_2 \cdot 2\ H_2O$ + beaker =	
mass beaker=	
mass $CaCl_2 \cdot 2\ H_2O$ =	
Describe the $CaCl_2 \cdot 2\ H_2O$	
Actual moles $CaCl_2 \cdot 2\ H_2O$ used	
Actual moles $CaCl_2$ used	
Weight of filter paper	
Weight of watch glass	
Describe the wet $CaCO_3$	
Weight of filter paper, watch glass and precipitate 1st drying	
Weight of filter paper, watch glass and precipitate 2nd drying	

Sample Calculations

TABLE 10.3

	$Na_2CO_{3\,(\;\;)}$	+ $CaCl_{2\,(\;\;)}$	→ $CaCO_{3\,(\;\;)}$	+ $NaCl_{(\;\;)}$
Initial moles				
Change moles				
End moles				
End moles numerical				

TABLE 10.4

Theoretical mole ratio $\frac{CaCl_2}{Na_2CO_3} =$	Actual mole ratio $\frac{CaCl_2}{Na_2CO_3} =$	Limiting Reagent is

Sample Calculations

TABLE 10.5

Results	
Theoretical weight of $CaCO_3$	
Actual weight of $CaCO_3$ collected	
Describe the dried $CaCO_3$	
Percent Yield	

Sample Calculations

Post Lab Questions

1. Describe the effect on the actual yield and percent yield (increase, decrease, or no change), if the following errors occurred. Be sure to explain your reasoning.
    a. The filter paper was still wet when the final product was weighed.
    b. Some product fell through the filter paper and was seen in the filter flask during the vacuum filtration.
    c. Some product fell on the floor and was scraped up with some dust particles, and then weighed for the final mass.
2. What are you to do with unused chemical in the lab?

3. In the reaction below 25.00 g of $MnC1_2$ is reacted with 100.0 g of PbO_2, excess KCl, and excess HCl.

$2\ KC1_{\ (aq)} + 2\ MnCl_{2\ (aq)} + 5\ PbO_{2\ (s)} + 4\ HC1_{\ (aq)} \rightarrow 2\ KMnO_{4\ (aq)} + 5\ PbC1_{2\ (s)} + 2\ H_2O_{\ (l)}$ **(EQ 10.3)**

a. How many grams of $KMnO_4$ can be produced by this reaction? Use an ICE table.

b. If 21.42 g of $KMnO_4$ is actually produced, what is the percent yield?

c. How many grams of what starting substances will be left over after the reaction?

d. How many grams of what starting substance (i.e., $MnCl_2$ or PbO_2) must be added to the original quantities of reactants so that there will be neither PbO_2 nor $MnC1_2$ left over after the reaction?

EXPERIMENT 11

Measurement of the Gas Constant and Molar Volume of Oxygen Gas

Background

This experiment will allow you to gain practical experience in the collection and measurement of the properties of gases. From the data collected you will experimentally determine the values of the universal gas constant, R, and the molar volume of a gas, V_m, under standard conditions (STP) based upon your experimentally derived value of the gas constant.

In this experiment, two values will be determined, the gas constant, R, and molar volume, V_m. The calculations are based on the ideal gas law:

$$PV = nRT \qquad \textbf{(EQ 11.1)}$$

where P = pressure of the gas; V = volume of the gas; n = moles of the gas; T = temperature of the gas with units of Kelvin (K); and R = the gas constant. There are two ways that can be used to determine R. First, mass is converted to moles then Equation 11.1 can be rearranged and used:

$$R = \frac{PV}{nT} \qquad \textbf{(EQ 11.2)}$$

The second way is derived as follows. Moles are difficult to measure directly, however the number of moles of a substance is equal to the mass of the substance divided by its molecular mass so the equation above may be rearranged as below to replace moles by mass:

$$MM = \frac{m}{n} \Rightarrow n = \frac{m}{MM} \qquad \textbf{(EQ 11.3)}$$

where m = mass and MM = molar mass.

Plugging Equation 11.3 into Equation 11.1 the new expression is:

$$PV = \frac{mRT}{MM} \qquad \textbf{(EQ 11.4)}$$

Equation 11.4 can now be rearranged to solve for R:

$$PV = \frac{mRT}{MM} \Rightarrow R = \frac{PV(MM)}{mT} \tag{EQ 11.5}$$

Rearranging Equation 11.1 an expression for molar volume, V_m, can be found:

$$PV = nRT \Rightarrow V_m = \frac{V}{n} = \frac{RT}{P} \tag{EQ 11.6}$$

In this experiment you will measure the mass, volume, temperature, and pressure of a gas to experimentally determine values for R and V_n. The literature values are listed below under STP conditions (1 atm and 0 °C):

$$R = 0.08206\frac{L\ atm}{mol\ K} = 62.365\frac{L\ torr}{mol\ K} \tag{EQ 11.7}$$

$$V_m = 22.414\frac{L}{mol} \tag{EQ 11.8}$$

In this experiment, oxygen gas will be produced by the thermal decomposition of potassium chlorate in the presence of the catalyst manganese dioxide. Catalysts are substances, which increase the rate of a reaction but are not used up by the reaction. Catalysts are not used up in the reaction, so they appear above the reaction arrow. Manganese dioxide (common name for manganese(IV) oxide) functions as a catalyst by lowering the activation energy required to decompose potassium chlorate. The reaction for the decomposition of potassium chlorate is shown below:

$$2\ KClO_{3\ (s)} \xrightarrow[heat]{MnO_2} 2\ KCl_{(s)} + 3\ O_{2\ (g)} \tag{EQ 11.9}$$

The oxygen gas will be collected through water displacing a volume of water equal to the volume of oxygen gas produced during the reaction. Therefore, the oxygen gas will be saturated with water vapor and be considered “wet”. The pressure inside the system will be equilibrated to atmospheric pressure. So Dalton’s Law of Partial Pressures can be used to determine the pressure of the “dry” oxygen gas. Dalton’s Law of Partial Pressures says that the total pressure is equal to the sum of the individual partial pressures, p_i, of the gases:

$$P_{total} = \sum p_i \tag{EQ 11.10}$$

Since the pressure inside the system is equilibrated to the atmospheric pressure in this case the barometric pressure, P_{bar}. P_{bar} is the sum of the partial pressure from oxygen gas and the vapor pressure of water. **Vapor pressure** is the pressure exerted when a solid or a liquid is in dynamic equilibrium with its vapor:

$$P_{bar} = P_{O_2} + P_{H_2O} \tag{EQ 11.11}$$

Vapor pressure values are temperature dependent therefore the temperature of the water is taken. The solubility of oxygen gas is also temperature dependent, although considered negligible in this experiment. The temperature of the water and oxygen gas should be about the same.

The graph of vapor pressure versus temperature is nonlinear. Linear graphs are often preferred for many reasons. It is easy to observe deviations from linearity, whereas it can be difficult to determine if a curve resembles a hyperbola or is truly hyperbolic. Linear plots may be more readily interpolated and extrapolated. In the following pages the creation of linear graphs from curved plots will be explored.

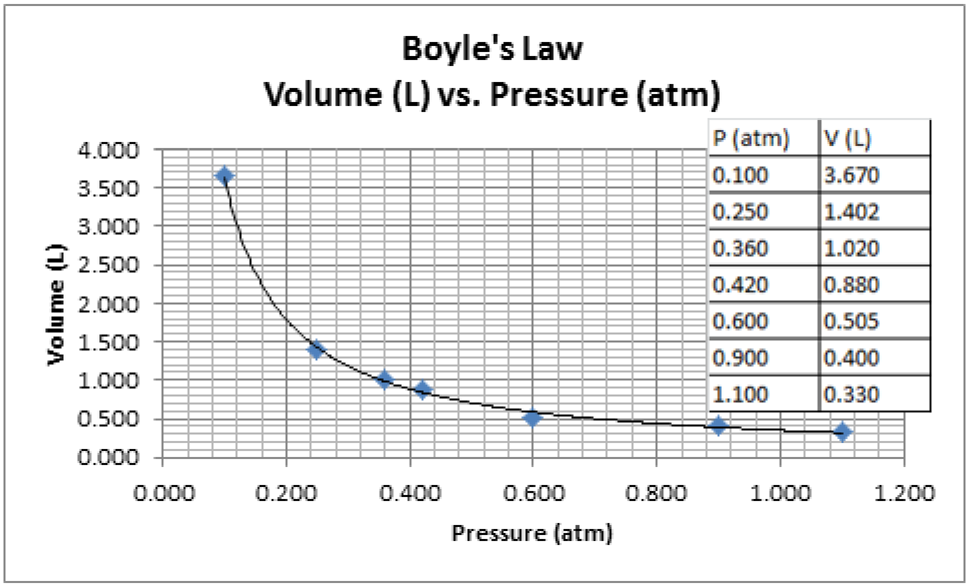
EXAMPLE 11.1 **Volume and Pressure of a Gas**

The following data was obtained by changing the pressure on a gas at constant temperature in Table 11.2:

TABLE 11.2 **Volume-Pressure Data**

Pressure (atm)	Volume (L)
0.100	3.670
0.250	1.402
0.360	1.020
0.420	0.880
0.600	0.505
0.900	0.400
1.10	0.330

FIGURE 11.1



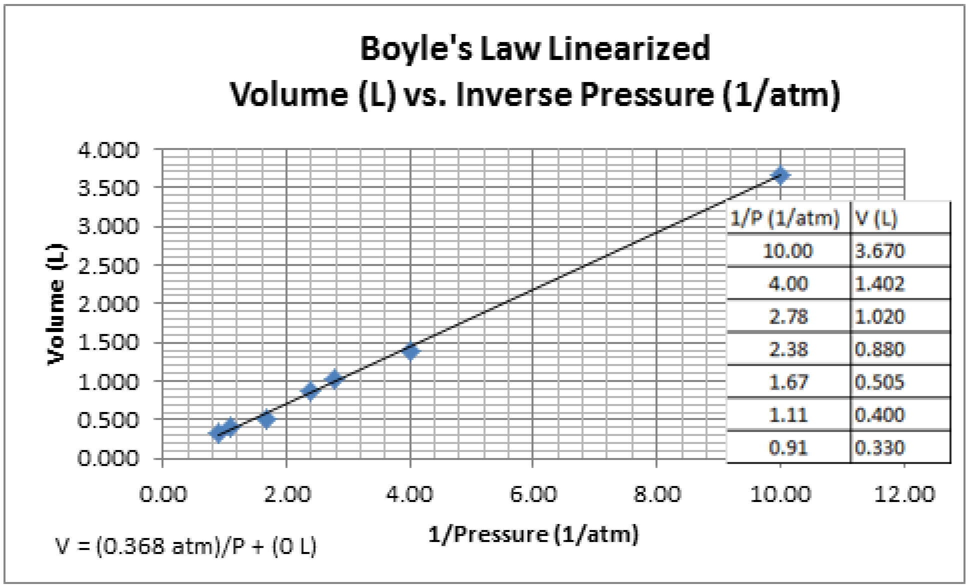
As the pressure is increased, the volume decreases, suggesting an inverse relationship. The third column of the table shows calculated values of inverse of P or 1/P. The trick to constructing a linear plot is to cast the function relationship between dependent and independent variables into the form of y = mx + b; therefore a linear plot might be obtained by plotting V vs. 1/P.

TABLE 11.3 **Volume-Pressure Data**

Pressure (atm)	Volume (L)	1/P (atm^{-1})
0.100	3.670	10.0
0.250	1.402	4.00
0.360	1.020	2.78
0.420	0.880	2.38
0.600	0.505	1.67
0.900	0.400	1.11
1.10	0.330	0.909

The actual quantity plotted on the y-axis is the volume and on the x-axis the inverse pressure. The result is Figure 11.2:

FIGURE 11.2



1/P (1/atm)	V (L)
10.00	3.670
4.00	1.402
2.78	1.020
2.38	0.880
1.67	0.505
1.11	0.400
0.91	0.330

Upper and lower variable limits of the axes have been chosen by the guidelines above. The plot is clearly linear. The slope, calculated by regression analysis, is 0.368 L atm. The intercept is zero, within experimental error, so the equation relating the variables is:

$$V = (0.368\ L\ atm)\left(\frac{1}{P}\right) + (0L) \qquad \textbf{(EQ 11.12)}$$

This is a form of Boyle's Law, usually stated as:

$$PV = k \qquad \textbf{(EQ 11.13)}$$

where k depends on quantities held constant, that is, the amount and temperature of the gas.

EXAMPLE 11.4 Problem for Solving Boyle's Law Problems:

Using an equation like Equation 11.3 various pressure and volume conditions can be calculated. For example, what will the pressure be at a volume of 5.721 L?

Rearranging Equation 11.12 for P:

$$V = (0.368\ L\ atm)\left(\frac{1}{P}\right) \quad \text{becomes} \quad P = (0.368\ L\ atm)\left(\frac{1}{V}\right) \qquad \textbf{(EQ 11.14)}$$

Putting the volume value into the equation the pressure can be calculated:

$$P = (0.368\ L\ atm)\left(\frac{1}{5.721\ L}\right) = 0.0643\ atm \qquad \textbf{(EQ 11.15)}$$

EXAMPLE 11.5 Consider the radioactive decay of barium-128 by electron capture.

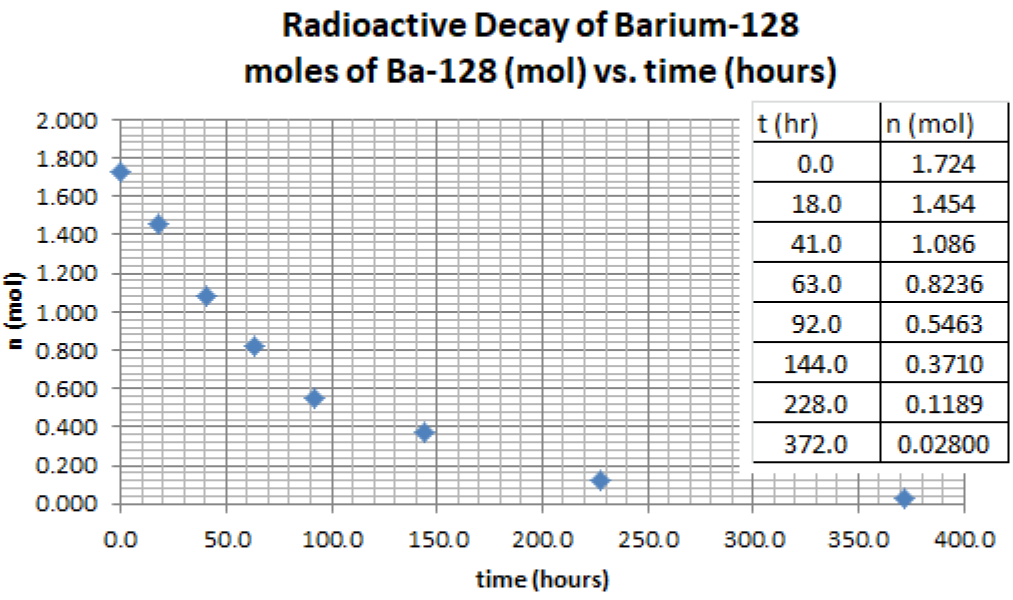
An initial amount of 220.7 grams (1.724 moles) is allowed to decay over a period of about two weeks. The number of moles of the initial isotope remaining is recorded periodically:

TABLE 11.6 Radioactive Decay of Ba-128

Time (hours)	Moles, n, (mol)
0.000	1.724
18.0	1.454
41.0	1.086
63.0	0.8236
92.0	0.5463
144.0	0.3710
228.0	0.1189
372.0	0.02800

A graph of these data appears as Figure 11.3:

FIGURE 11.3



t (hr)	n (mol)
0.0	1.724
18.0	1.454
41.0	1.086
63.0	0.8236
92.0	0.5463
144.0	0.3710
228.0	0.1189
372.0	0.02800

The actual quantities plotted on the x and y-axises are time and moles respectively. Figure 11.3, though not in the form of a straight line, this is a useful graph from several points of view. The trend line indicates a logarithmic function. The bottom of the line approaches zero as a limit, meaning it will take a very long time for the last bit of Ba-128 to decay. It may also be noted that the time it takes for half of the Ba-128 to decay (down to 0.862 moles) seems to be about 55 hours. The time for half again to decay (down to 0.431 moles) is about 110 hours, or another 55 hours. This time is in fact a constant, and is called the half-life or $t_{1/2}$.

In order to obtain a straight line from this data it is necessary to plot the logarithm of the number of moles vs. the time. The logarithmic function can be to any base, but usually base e (the natural logarithm, abbreviated ln) and base 10 (abbreviated log) are used. The relationship between the two is:

$$\ln x = 2.303 \log_{10} x \qquad \textbf{(EQ 11.16)}$$

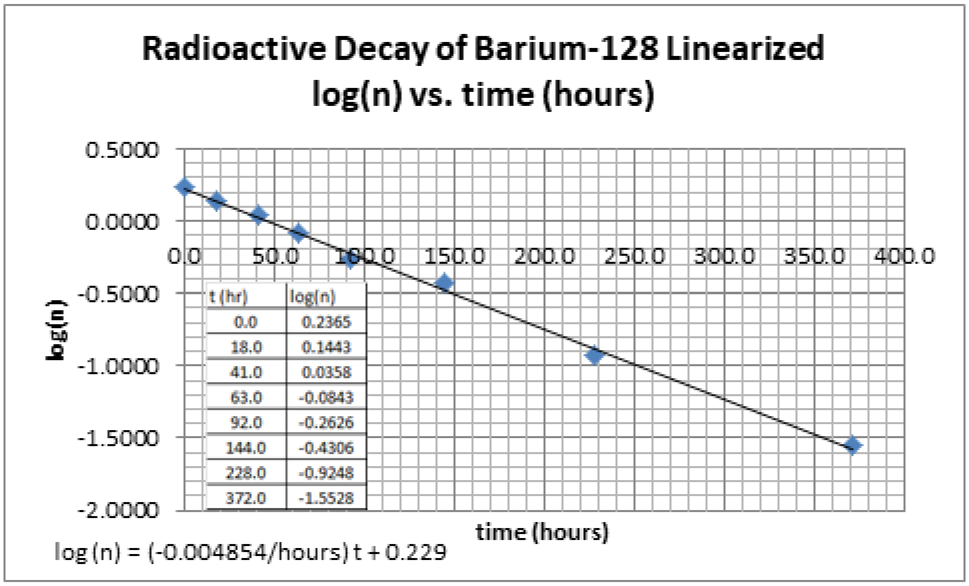
Using base 10 logarithm the data can be recalculated

TABLE 11.7

Time (hours)	Moles, n, (mol)	log(n)
0.0	1.724	0.2365
18.0	1.454	0.1443
41.0	1.086	0.0358
63.0	0.8236	-0.0843
92.0	0.5463	-0.2626
144.0	0.3710	-0.4306
228.0	0.1189	-0.9248
372.0	0.02800	-1.5528

Notice that logarithms have no units. Also using the significant figure rules for logs and that because there are 4 significant figures for the moles there are four decimal places for the log(n).

FIGURE 11.4 **Graphing log n vs. time**



t (hr)	log(n)
0.0	0.2365
18.0	0.1443
41.0	0.0358
63.0	-0.0843
92.0	-0.2626
144.0	-0.4306
228.0	-0.9248
372.0	-1.5528

log (n) = (-0.004854/hours) t + 0.229

The actual quantity plotted on the x-axis the time, while on the y-axis it is the logarithm of the moles, log(n). The slope of the line is given by the computer spreadsheet program as -0.004854, but don't forget to add the units of hours^{-1}. The y-intercept is 0.229 with no units because logarithms are unit-less values. The negative sign for the slope indicates that it tilts down from left to right. The resulting equation for the line is:

$$\log n = (-0.04854\ \text{hours}^{-1})(t) + 0.229 \qquad \textbf{(EQ 11.17)}$$

EXAMPLE 11.8 Problem Solving for Moles

How many moles of barium-128 will remain after 425.0 hours?

Use Equation 11.17 and solve for the number of moles, n:

$$\log n = (-0.04854\ \text{hours}^{-1})(t) + 0.229 \qquad \textbf{(EQ 11.18)}$$

$$\log n = (-0.04854\ \text{hours}^{-1})(425.0\ \text{hours}) + 0.229 \qquad \textbf{(EQ 11.19)}$$

$$\log n = -1.829 \qquad \textbf{(EQ 11.20)}$$

$$n = 10^{-1.829} = 0.0148\ \text{moles} \qquad \textbf{(EQ 11.21)}$$

Don't forget that n has units of moles even though logarithms are unitless!

Procedure

Caution! Safety Notes:

The $KClO_3/MnO_2$ is a powerful oxidizing mixture that may explode on excessive heating if the mixture is contaminated. Be very careful not to get any glycerin in the mixture.

Wear your safety goggles at all times! Especially when anyone in the laboratory is heating the $KClO_3/MnO_2$ mixture!

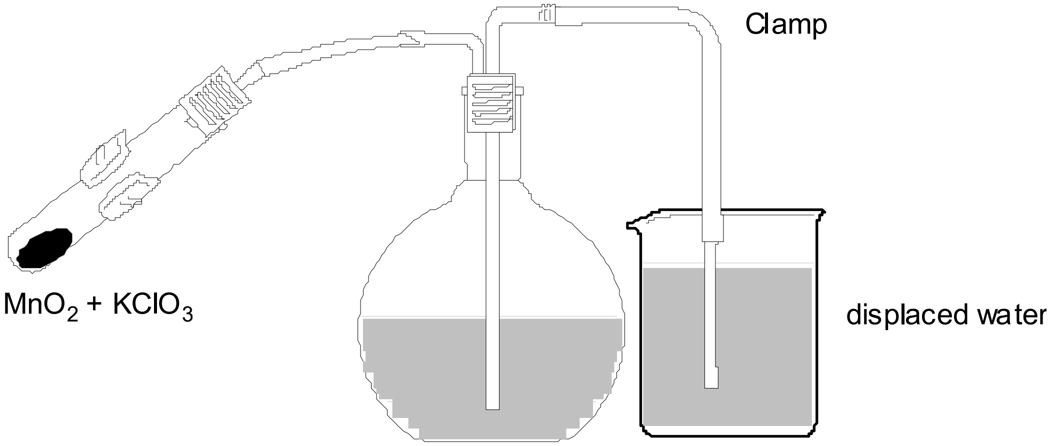
The $KClO_3/MnO_2$ waste should be disposed by dissolving the $KClO_3$ mixture in water and pouring it into the inorganic substances hazardous waste container. If solid $KClO_3$ is put in the wastebasket it may start a fire.

Clean and dry a large (8-inch size) test tube. Weigh the test tube. Transfer 2.1-2.3 g of the $KClO_3/MnO_2$ mixture into the test tube and reweigh it. Make sure to record your mass to the precision of the balance used. Remember to use a milligram balance and record all masses to at least the nearest 0.001 g.

NOTE: We will use the method of mass by difference to minimize any systematic errors in the weighing process. Be sure to use the same balance throughout the experiment!

Preliminary assembly

Assemble the apparatus shown in Figure on page 195. A correctly assembled apparatus will be on display at the instructor's station as well



NOTE: Because you will be collecting a gas sample, all of the tubing and stopper connections of the apparatus must be gas tight. If any part of the apparatus leaks during the experiment, your data will not give accurate results. (Also, note that a leaking apparatus constitutes a blunder and is not just an experimental error!)

1. Fill a Florence flask to the beginning of the neck and large beaker with water with about 300 mL of water. Place the two holed stopper in the mouth of the Florence flask and the long tubing into the beaker. Make sure that the pinch clap is open.
2. Before attaching the test tube containing the weighed reaction mixture, use a pipet bulb to blow gently into the connecting tubing (from the end that is to be attached to the test tube) to start a siphon action between the one-liter flask and the 1000 mL beaker.
3. Siphon the water back and forth several times until you have removed all of the air bubbles. Throughout the siphoning process, take care that no water enters the short piece of tubing in the two-holed stopper that will ultimately be connected to the test tube.
4. With the water level midway in the neck of the 1 liter flask you are ready to begin the **equalization process**.

Equalization Process

Notes: The system must be equilibrated to the atmospheric pressure before starting your decomposition reaction.

You should complete the **equalization process** before starting your decomposition reaction. In this step you are going to adjust the pressure inside of the 1 liter flask and the test tube (i.e., the system) so that it will be the same as the pressure on the outside, that is, the pressure in the room, which you will determine by reading the barometer. The pressure of the system is to be equalized before the oxygen is produced and again at the end of the experiment just before the volume of the displaced water is measured.

1. To equalize the system:
    a. Raise or lower the beaker till the level or height of the water in the beaker is the same as the level or height of the water in the flask.
    b. Close the clamp while maintaining the same water level in the flask and beaker. Again, the purpose of this process is to ensure that the pressure in the flask is the same as the pressure in the room, the atmospheric pressure
    c. Now attach the short piece of rubber tubing that connects to the stopper and cap the large test tube containing the $KClO_3/MnO_2$ mixture. Be sure that all stopper connections are airtight so that no gas will be lost to leaks in the system.
    d. Then close the pinch clamp.
    e. Have your instructor check to make sure that the system is air tight BEFORE you continue.
2. Take the glass delivery tube out of the beaker and empty the water out of the beaker into one of the cup sinks in the hood or a sink. Handle the delivery tube from the flask very carefully so as to not shake any water out of it. Set the beaker down but do not dry it out with anything. Return the beaker to its position under the delivery tube and reopen the pinch clamp. From this point forward, **DO NOT** take the end of the delivery tube out of the water that will soon flow into the beaker.
3. Only a few drops of water should flow into the beaker. **If water continues to flow into the beaker, there is a leak in the system and you must stop, fix the leak, and repeat the equalization process before you continue.** If you have a leak, make sure that the stopper is tightly seated in the flask. Check that all of the rubber tubing-glass tubing connections are tight. When you are sure they are all tight, repeat the procedure – flush the system and equalize the pressure in the system.

Decomposition Reaction

1. After the instructor gives you permission, make certain that the pinch clamp is open and start slowly heating the $KClO_3/MnO_2$ mixture in the tube. After 1-2 minutes of gentle heating, turn up the heat level from the burner so that the mixture melts and starts to decompose. Carefully watch the progress of the decomposition and regulate the heat level so that there is a slow but continuous evolution of oxygen gas.

> **NOTE:** Be sure to heat slowly or the oxygen gas will be generated faster than it can escape through the glass tube and it will shoot the stoppers out of the test tube or the flask, meaning you must start over! If the water level in the flask gets close to the bottom of the glass tubing that leads to the beaker, immediately stop heating the test tube. If the water level lowers past the bottom of the glass tubing you must start over!

2. As the reaction progresses, the mixture will begin to solidify and you will have to supply more heat. When little gas seems to be evolving, heat the tube from all sides and from the end strongly to ensure that the last bit of $KClO_3$ has been decomposed. At this point, you should have 600 to 700 mL of water in the beaker. The end of the tube that extends almost to the bottom of the flask must still be submerged under the water. If the water level falls below the end of the tube, gas will escape and your results will be invalid.
3. When no more oxygen gas is evolved, let the system cool to room temperature. Make sure that the water delivery tube remains in the beaker of water and that the pinch clamp remains open. During the cooling, some water flows back into the flask to compensate for the contraction of the gas upon cooling.
4. While you are waiting for the cooling, measure the barometric pressure in the room. This measurement should have four significant figures. Your instructor will show you how to do this.
5. When the system has cooled to room temperature, **equalize the pressure** and while the water levels are the same in the beaker and the flask, close the pinch clamp. **Be sure you have the pinch clamp securely in place before you do this next step!**
6. Record the temperature of the gas in the flask, to one decimal place (± 0.1 °C), by releasing the stopper slightly and inserting a thermometer into just the gas. After the temperature comes to equilibrium, measure the temperature of the water remaining in the flask to one decimal place.
7. Record the temperature of the water in the flask, to one decimal place (± 0.1 °C).
    a. Use the CRC Handbook to determine the vapor pressure of water between 20 °C and 120 °C, including one point at the temperature of your water in the flask.
8. When it is cool, weigh the test tube and KCl/MnO_2, residue.
9. Measure the volume of water displaced into the beaker with a 1000 mL (± 5 mL) graduated cylinder.

Calculation Hints

1. The water vapor pressure at a variety of temperatures can be found in the CRC Handbook.
2. Calculate the value of the gas constant, R, by substituting the known quantities from you measurements into the ideal gas equation and solving for R.
3. Evaluate the molar volume, V_m, by using the value of R just calculated and the standard temperature and pressure (STP) conditions in the ideal gas equation.

EXAMPLE 11.9

The following data was collected in an experiment similar to yours and was analyzed as shown below. Remember not to round until the end of the calculation to avoid round off error in your answer.

Mass of empty test tube 25.783 g
Mass of test tube and $KClO_3/MnO_2$ 27.921 g

a. Calculate the mass of $KClO_3/MnO_2$.

27.921 g - 25.783 g = 2.138 g

The test tube and $KClO_3/MnO_2$ is heated and decomposed as shown in Figure on page 195.
Mass of test tube and residue 27.057 g

b. Calculate the mass of oxygen gas liberated.

27.921 g - 27.057 g = 0.864 g

c. Calculate the moles of oxygen gas.

$$0.864\ g\ O_2 \times \frac{1\ mol\ O_2}{31.998\ g\ O_2} = 0.0270\ mol\ O_2$$

Volume of H_2O displaced 642 mL
The volume of O_2 gas liberated is assumed to equal the volume of water displaced, so 642 mL or 0.642 L.
Temperature of oxygen in flask 22.1 °C + 273.15 = 295.3 K
Temperature of water left in flask 22.0 °C
Barometric Pressure 758.3 mm Hg
From the CRC Handbook the Vapor Pressure of Water at 22.0 °C 19.837 mm Hg

d. Using Dalton's Law of Partial pressures, Equation 11.11, calculate the pressure of oxygen gas.

$$P_{O_2} = P_{bar} - P_{H_2O} = 758.3\ mmHg - 19.837\ mmHg = 738.5\ mmHg$$

e. Calculate the experimental value of R in L mmHg/mol K using Equation 11.2:

$$R = \frac{PV}{nT} = \frac{(738.5\ mmHg)(0.642\ L)}{(0.0270\ mol)(295.3\ K)} = 59.5\frac{mmHg\ L}{mol\ K}$$

f. Calculate the Molar Volume, V_m, at STP (0 °C = 273 K and 760 mmHg) conditions using Equation 11.6.

$$V_m = \frac{V}{n} = \frac{RT}{P} = \frac{\left(59.5\frac{mmHg\ L}{mol\ K}\right)(273\ K)}{760\ mmHg} = 21.4\frac{L}{mol}$$

4. Compare both your value for R and for V_m to the accepted values. This comparison should consist of percentage error calculated by:

$$\%\ \text{error} = \frac{\text{observed value} - \text{accepted value}}{\text{accepted value}} \times 100\% \qquad \textbf{(EQ 11.22)}$$

Be sure to show the positive or negative sign. What does the sign tell you about your percent error?

5. Calculate the moles of $KClO_3$ in the original mixture from the moles of oxygen gas produced using stoichiometry and the balanced chemical equation, Equation 11.9.
6. Then calculate the mass of $KClO_3$ from the moles of $KClO_3$ using the molar mass of $KClO_3$.
7. Next calculate the percent of $KClO_3$ in the original mixture from the mass of $KClO_3$ and the mass of the mixture:

$$\%\ KClO_3 = \frac{m_{KClO_3}}{m_{mixture}} \times 100\% \qquad \textbf{(EQ 11.23)}$$

8. Prepare two separate graphs using data obtained from the CRC Handbook. Be sure to make a data table before you start making the graph using a reasonable number of data points. This is very important for the second plot since you must calculate the data points from the information given in the handbook.
    a. Vapor pressure of water vs. the Celsius temperature (Use a French curve to draw a smooth curve between the data points).
    b. Base ten logarithm of the vapor pressure vs. the reciprocal of the Kelvin temperature.
9. Determine the slope, y-intercept, and the equation of the line for graph b. Don't forget to put the equation for the straight line on the graph as well as a data table.

Prelaboratory Exercise: Gas Constant and Molar Volume

Lab Grade

Prelab Questions		10
General Format (Signature, ink, no obliterations, etc.)		10
Data, Analysis, Graph (observations, questions, units, significant figures, sample calculations, etc.)		36
Accuracy (% error)		30
Post Lab Questions		14
Total		100

Name: ___________________________

Section:_____________________

1. Write the balanced equation for the production of the oxygen gas in this experiment. Don't forget the catalyst.

2. How much of the $KClO_3$ mixture are you to use in you experimental work?

Between ______________ and ______________ grams

3. In this experiment MnO_2 is used as a catalyst. What is the function of the MnO_2 in this reaction?

4. After you have set up your apparatus as shown in Figure on page 195 and you think that the system is air tight, what are you supposed to do next (i.e., before actually starting the experiment)?

5. Before you start displacing the water with the oxygen and again when you are all done, you must equalize the pressure inside of the system so that it is the same as that in the room.

- **a.** Why is this necessary?
- **b.** How is it done?

6. A sample of oxygen gas is collected over water using the same procedure you will use in this experiment. The system is equalized and the temperature of both the water and the gas is 22°C. The atmospheric pressure is 745 mm Hg, what is the pressure of the O_2 gas in the flask if the vapor pressure of water at 22 °C is 19.827 mm Hg?

7. Use the data below to answer the following:

TABLE 11.10

a.	Mass of empty test tube	18.767 g
b.	Mass of test tube + sample	20.989 g
c.	Mass of test tube assembly + residue	20.164 g
d.	Mass of O_2, liberated	
e.	Volume of H_2O displaced	675 mL
f.	Volume of O_2 evolved	
g.	Temperature of O_2 in flask	24.0 °C
h.	Temperature of H_2O left in the flask	23.9 °C
i.	Barometric pressure	742.0 torr
j.	Vapor pressure of water	22.377 torr
k.	Calculated pressure of O_2 gas	
l.	Calculated Value of R in L atm/mol K	

Sample Calculations:

Data Sheet: Gas Constant and Molar Volume

TABLE 11.11 Experimental Data

Mass of empty test tube	
Mass of test tube and $KClO_3/MnO_2$	
Mass of $KClO_3/MnO_2$ mixture	
Mass of test tube assembly and residue	
Mass of O_2 liberated	
Moles of O_2 liberated	
Volume of H_2O displaced	
Volume of O_2 liberated	
Temperature of O_2 in flask	
Temperature of H_2O left in the flask	
Barometric pressure	
Vapor Pressure of H_2O	
Calculated pressure of O_2 gas	

Sample Calculations

TABLE 11.12 Results

Experimental Value of R in L torr/mol K	
Literature Value of R	62.365 L torr/mol K
Percentage error in R	
Experimental Value of R in L atm/mol K	
Literature Value of R	0.08206 L atm/mol K
Percentage error in R	
Molar volume, V_m under STP conditions	
Literature Value of V_m	22.414 L/mol
Percentage error in V_m	
Calculated moles of $KClO_3$	
Calculated mass of $KClO_3$	
Percentage of $KClO_3$ in the original mixture	

Sample Calculations

Temperature dependence of the vapor pressure of water

1. What is vapor pressure?

2. Use the CRC Handbook to determine the vapor pressure of water between 20 °C and 120 °C, including one point at the temperature of your water in the flask.Then complete the rest of the table and draw the graphs
    a. Graph 1: Vapor pressure vs. temperature (Don't forget to use a French curve.)
    b. Graph 2: Log of vapor pressure vs. reciprocal Kelvin temperature .

TABLE 11.13 **Complete the following table to help you make your graphs.**

Temperature (°C)	Temperature (K)	1/T (1/K)	Vapor Pressure (P_{vap}) ( )	log(P_{vap})

3. Using graph 1, estimate the vapor pressure of water at 80 °C? ______________

4. What is the slope, m, of graph 2? Show your calculations here and include the correct units. Remember that logarithms do not have units!

5. What is the y-intercept, b, of graph 2? Show your calculations here and include the correct units.

6. Write the equation of the straight line of graph 2. Refer to Figure 11.2 on page 190 and Figure 11.4 on page 192 and the accompanying explanations in the prelaboratory information sheet before you attempt to write this equation. Remember it should be in the form: $\log(P) = m\left(\frac{1}{T}\right) + b$

7. What should the vapor pressure of water at 160 °C? Use your equation to calculate your answer. Show the set up.

Post Lab Questions

1. How would your result (i.e., the R value) be affected if you did not wait for the test tube to cool before measuring the volume of water? Explain your answer.

2. A mixture of gases at 760 mm Hg pressure contains 65.0% nitrogen, 15.0% oxygen, and 20.0% carbon dioxide, by volume. What is the partial pressure of each gas in mm of Hg?

3. Assume that 1.40 g of a $KClO_3$ /MnO_2 mixture (65% $KClO_3$, by weight) was decomposed in an experiment similar to yours and all final measurements were made at 22°C on a day when the atmospheric pressure was 716 torr.
    - **a.** What is the pressure of the O_2 if the vapor pressure of water at 22 °C is 19.827 mm Hg?
    - **b.** How many moles of O_2 should have been evolved?
    - **c.** What volume of O_2 would be evolved under the experimental condition?
    - **d.** What is the STP volume of this O_2?

EXPERIMENT 12

Volumetric Analysis: The Titration of Acids and Bases

Background

This experiment will allow you to gain practical experience in the preparation of standard solutions, use of a pipette and a buret, and performing standard titrations. You will use this experience to experimentally determine the concentration of acetic acid in an unknown concentration solution. There are three parts to this experiment, they are

1. Preparation of a standard solution of oxalic acid (~0.07 M);
2. Preparation of a sodium hydroxide solution (~0.1 M) which will be standardized using the standard oxalic acid solution;
3. Determination of the concentration of acetic acid in an unknown vinegar solution. (0.1-0.2 M).

Standard solutions are solutions with known concentrations, generally to four significant figures. There are two different ways to make a standard solution. We can make a primary or a secondary standard. A **primary standard** is prepared directly by dissolving a known mass of sample to make a known volume of solution. Dissolving an approximate amount of sample into a volume of solvent and determining its exact concentration through titration experiments prepares a secondary standard. Primary standards are prepared from compounds that are at least 99.9% pure, have a definite composition, are water soluble, are easily weighed, and do not change composition on contact with air. Oxalic acid dihydrate ($H_2C_2O_4 \cdot 2\ H_2O$) fits these criteria and therefore may be used as a primary standard. Sodium hydroxide absorbs water when it comes into contact with air and therefore it is difficult to obtain a pure, dry sample to weigh. For this reason the sodium hydroxide solution will be titrated with the oxalic acid standard to become a **secondary standard**.

In the first part of this experiment you will prepare a solution of known concentration (to four significant figures) of oxalic acid. The oxalic acid crystallizes with two water molecules per oxalic acid in the crystalline network. For this reason, we will weigh out an appropriate amount of oxalic acid dihydrate to dissolve in water. The water molecules in the crystal network will become part of the water of solution once it is dissolved. For this reason the molar concentration of oxalic acid dihydrate will be the same as the molar concentration of oxalic acid.

EXAMPLE 12.1 Determining the concentration of oxalic acid solution:

Data:

Mass oxalic acid dihydrate = 5.364 g
Total volume solution = 500.0 mL
Moles oxalic acid dihydrate =

$$5.364\text{ g } H_2C_2O_4 \bullet 2H_2O \times \frac{1\text{ mol } H_2C_2O_4 \bullet 2H_2O}{126.07\text{ g } H_2C_2O_4 \bullet 2H_2O} = 0.04255\text{ mol } H_2C_2O_4 \bullet 2H_2O \quad \textbf{(EQ 12.1)}$$

Molarity of oxalic acid =

$$\frac{0.04255\text{ mol } H_2C_2O_4 \bullet 2H_2O}{500.0\text{ mL } solution} \times \frac{1\text{ mol } H_2C_2O_4}{1\text{ mol } H_2C_2O_4 \bullet 2H_2O} \times \frac{1000\text{ mL solution}}{1\text{ L } solution} = 0.08510\text{ M } H_2C_2O_4 \quad \textbf{(EQ 12.2)}$$

Often reagents will be bought and stored in their concentrated form and you need to preform a dilution before using them.

EXAMPLE 12.2 Dilution of a stock solution of sodium hydroxide:

Data:

For this approximately 500 mL of a 0.2 M sodium hydroxide solution is needed. The stockroom has a stock solution of 16 M sodium hydroxide. How much concentrated sodium hydroxide solution do you need to dilute?

Answer:

There are two ways to solve this type of problem using stoichiometry and algebra:

Using stoichiometry:

$$V_{concentrated} = (0.500\ L_{dilute})\left(\frac{0.2\ mol\ NaOH}{L_{dilute}}\right)\left(\frac{1\ L_{concentrated}}{16\ mol\ NaOH}\right) \quad \textbf{(EQ 12.3)}$$

$$V_{concentrated} = 0.00625\ L_{concentrated} = 6\ mL_{concentrated} \quad \textbf{(EQ 12.4)}$$

Using algebra:

M_1 = 0.2 M
V_1 = 500 mL= 0.500 L
M_2 = 16 M
V_2 = ?

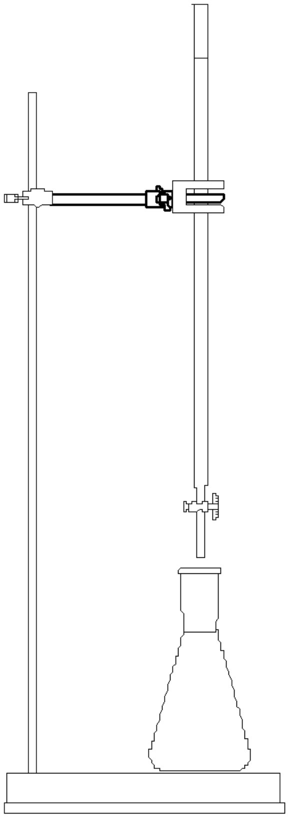
$$M_1V_1 = M_2V_2 \quad \textbf{(EQ 12.5)}$$

$$V_2 = V_1\frac{M_1}{M_2} = (500\ mL)\left(\frac{0.2\ M}{16\ M}\right) = 6.25\ mL = 6\ mL \quad \textbf{(EQ 12.6)}$$

To make about 500 mL of a 0.2 M sodium hydroxide solution, start with about 6 mL of 16 M sodium hydroxide solution and add DI water until you get about 500 mL.

In order to standardize the sodium hydroxide solution you will perform a titration.

FIGURE 12.1 **Titration set-up showing buret, ring stand and Erlenmeyer flask**



A **buret** is a hollow, graduated cylindrical tube equipped with a regulating device for controlling the flow of a liquid. Since it has a narrow, uniform diameter, it is possible to deliver small quantities of liquid with a high degree of accuracy. The graduations on a buret begin with zero and increase in number from the top down, but do not extend to the bottom.

Sodium hydroxide reacts with oxalic acid according to the reaction below:

$$H_2C_2O_{4\,(aq)} + 2\ NaOH_{\,(aq)} \rightarrow 2\ H_2O_{\,(l)} + Na_2C_2O_{4\,(aq)} \quad \textbf{(EQ 12.7)}$$

You will measure a 25.00 mL aliquot of the oxalic acid solution into a flask and add an indicator.

> An aliquot is an exact quantity of a substance or solution.

An acid-base **indicator** is a substance that changes color when a solution changes pH. The common indicator used for acid-base titrations is phenolphthalein. Phenolphthalein is colorless in a solution that is acidic and bright pink in a solution that is basic. In this titration the oxalic acid solution is acidic and therefore phenolphthalein will be colorless. The sodium hydroxide solution will be added drop wise from a buret into the flask containing the oxalic acid and indicator. As the sodium hydroxide is added to the flask it will react with the oxalic acid and be neutralized. (The set up is shown in Figure 12.1 on page 219) At the point where all of the oxalic acid is reacted, the next drop of sodium hydroxide will make the entire solution basic and it will turn pink. At this point you have completed the titration. Your end point color should be a faint pink, not bright pink!

EXAMPLE 12.3 **Determining the molarity of sodium hydroxide standard.**

Data:

Volume oxalic acid solution = 25.00 mL
Molarity oxalic acid solution = 0.08510 M $H_2C_2O_4$ from Equation 12.2
Volume sodium hydroxide to titrate = 19.43 mL

Solve this type of problem using stoichiometry:

moles sodium hydroxide reacted =

$$25.00\ \mathrm{mL}\ \mathrm{H}_2C_2O_4 \times \frac{1\ \mathrm{L}\ \mathrm{H}_2C_2O_4}{1000\ \mathrm{mL}\ \mathrm{H}_2C_2O_4} \times \frac{0.08510\ \mathrm{mol}\ \mathrm{H}_2C_2O_4}{1\ \mathrm{L}\ \ \mathrm{H}_2C_2O_4} \times \frac{2\ \mathrm{mol}\ NaOH}{1\ \mathrm{mol}\ \mathrm{H}_2C_2O_4} = 0.004225\ \mathrm{mol\ NaOH} \quad \textbf{(EQ 12.8)}$$

molarity of sodium hydroxide =

$$\frac{0.004225\ \mathrm{mol\ NaOH}}{19.43\ \ \mathrm{mL}\ solution} \times \frac{1000\ \mathrm{mL\ solution}}{1\ \mathrm{L}\ solution} = 0.2190\ \mathrm{M}\ \ NaOH \qquad \textbf{(EQ 12.9)}$$

In order to determine the concentration of acetic acid in the vinegar solution, you will titrate it with the standardized sodium hydroxide solution. The equation for this reaction is:

$$HC_2H_3O_{2\,(aq)} + NaOH_{\,(aq)} \rightarrow H_2O_{\,(l)} + NaC_2H_3O_{2\,(aq)} \qquad \textbf{(EQ 12.10)}$$

EXAMPLE 12.4 **Determining the molarity of vinegar solution**

Data:

Volume vinegar solution = 25.00 mL
Volume sodium hydroxide to titrate = 22.84 mL
Molarity of sodium hydroxide solution = 0.2190 M NaOH from Equation 12.9

Solve this type of problem using stoichiometry:

moles acetic acid reacted =

$$22.84\ \mathrm{mL}\ NaOH \times \frac{1\ \mathrm{L}\ NaOH}{1000\ \mathrm{mL}\ NaOH} \times \frac{0.2190\ \mathrm{mol}\ NaOH}{1\ \ \mathrm{L\ NaOH}} \times \frac{1\ \mathrm{mol}\ \mathrm{HC}_2H_3O_2}{1\ \mathrm{mol}\ NaOH} = 0.005002\ \mathrm{mol}\ \mathrm{HC}_2H_3O_2 \textbf{(EQ 12.11)}$$

molarity of acetic acid =

$$\frac{0.005002\ \mathrm{mol}\ \mathrm{HC}_2H_3O_2}{25.00\ \ \mathrm{mL}\ solution} \times \frac{1000\ \mathrm{mL\ solution}}{1\ \mathrm{L}\ solution} = 0.2001\ \mathrm{M}\ \ \mathrm{HC}_2H_3O_2 \qquad \textbf{(EQ 12.12)}$$

In order to get the best precision possible, you should repeat each titration until you get 3 trials that are within 1% of each other.

Procedure

Caution! Safety Notes:

- Wear safety glasses at all times.
- All solution can go down the drain because your acid and base solutions will be neutralized by each other during the titrations.

Part I. Preparation of Standard Oxalic Acid Solution

1. Carefully weigh a 100 or 150 mL beaker and record its mass. Measure between 2.1 and 2.3 g of pure oxalic acid dihydrate crystals into the beaker and weigh again.

> Remember to use the milligram balances and record all masses to at least the nearest 0.001 g.

2. Add 30-60 mL of deionized water to the beaker and dissolve the crystals. You may gently heat the solution to speed up this process.
3. Transfer the solution quantitatively into a clean 250 mL volumetric flask. Rinse the beaker with 15-20 mL of deionized water and pour this solution into the volumetric flask and repeat. This will ensure that all of the oxalic acid is transferred into the volumetric flask.
4. Fill the volumetric flask to within about 2 cm of the mark and allow it to sit for a minute. This will allow any water clinging to the edges of the neck to drain into the flask. Using an eyedropper, fill the flask to the mark with water.
5. Stopper the flask and mix the solution by repeated inversion and swirling. This requires about 30 inversions and takes close to 1 minute.
6. If necessary, you will transfer this solution into one of the **clean** 500 mL bottles underneath the chemical hoods.
7. To transfer, first pour a small amount (~10 mL) of the standard solution into the bottle and rinse the inside walls of the bottle. Discard this wash solution.
8. Do this two times to prevent the standard solution from being diluted with any pure water that might have remained in the bottle after washing.
9. Then transfer your solution to the bottle.
10. Calculate the molarity of the oxalic acid solution using the molarity of the oxalic acid solution using the mass of acid used and the volume of the volumetric flask and record the concentration of your solution on the bottle.
11. Your label should look like the one pictured in Figure 12.2 .

FIGURE 12.2

> 0.5318 M Oxalic Acid
> 7/27/04 *Tina Titration*
> Chem 120 Section 1234

Part II. Preparation of Sodium Hydroxide Solution by Dilution

1. Calculate the volume of concentrated sodium hydroxide solution you must use to prepare 500 mL of approximately 0.1-0.12 M sodium hydroxide solution. (Check the label on the reagent bottle to determine its approximate concentration. The concentration will be 6 M)

> **Hint:** See “ Dilution of a stock solution of sodium hydroxide:” on page 218. or “Prelaboratory Exercise Week 1: Volumetric Analysis”question 4.

2. Measure an appropriate amount using your graduated cylinder. Try to measure to within 1 mL of the desired amount of reagent. The exact amount is not important because you will be standardizing this solution later.
3. Pour the concentrated base solution into a clean (need not be dry) 500 mL bottle and fill the bottle up to the shoulder with deionized water.
4. Shake the bottle well and label it as above recording the concentration to 1 significant figure.

Glass Handling Technique:

Pipetting a Liquid

Volumetric pipets are useful for quick delivery of liquids with greater accuracy and precision than a graduated cylinder. The volumetric pipet delivers a single specific aliquot or volume of liquid.

Preparation of the Pipet –

- a. Obtain a pipet from the stockroom.
- b. Inspect the pipet to ensure it is free of chips or cracks.
- c. Wash the pipet with soapy tap water three times, tap water three times, and then rinse with DI water three times. A clean pipet should have no water droplets adhering to its inner wall.
- d. Transfer the liquid that you intend to pipet from the reagent bottle into a clean, dry beaker; remember never to put a pipet directly into a reagent bottle. Be sure to label the beaker with the correct chemical formula so that you don't mix up your reagents later!
- e. Dry the tip of the pipet with a paper towel.
- f. Draw 2-5 mL of solution into the pipet using a pipet bulb.
- g. Then swish and rotate the solution around the pipet and discard the rinse liquid. This will be used to rinse the inner walls of the pipet with the liquid and ensure that the water adhering to the pipet walls does not dilute it.
- h. Repeat this rinse step one or two more times.

Filling the Pipet –

- a. Place the pipet tip well below the surface of the liquid.
- b. Using a pipet bulb, draw the liquid into the pipet until the level is 2-3 cm above the "mark" on the pipet.
- c. Remove the bulb and quickly cover the top of the pipet with your finger.
- d. Allow some liquid to drain out of the pipet until the meniscus is at the mark in the pipet. (You may wish to practice this step with deionized water until you are able to control the flow easily with your finger.) Wipe any liquid clinging to the outside of the pipet with a dry towel.

Delivery of the Liquid –

- a. Deliver the liquid to the receiving vessel by releasing the finger from the top of the pipet.
- b. Hold the pipet tip to the wall of the receiving vessel to avoid splashing.
- c. Do not blow or shake the last remaining bit of liquid out of the pipet.
- d. Do not blow any remaining solution out with the pipette bulb, as these pipettes are designed to retain a small amount of liquid in the tip when they have dispensed the appropriate amount of reagent.

Clean up –

- a. Once you are finished with the pipet, rinse it several times with deionized water and return it to the stockroom or the appropriate container placing the tip up.

Glass Handling Technique:

Titrating – Use of a Buret

Preparation of the Buret –

- **a.** Obtain a clean buret from the stockroom. Close the stopcock and fill with deionized water. Place a beaker beneath the tip and allow the water to drain. Check to see that the water flows freely and that there are no leaks at the stopcock.
- **b.** The walls of the buret should be free of water droplets; this indicates a clean buret. If your buret is not clean, you should wash it thee times with soapy water, rinse three times with tap water, then three times with DI water, finally rinse three times with your titrant as outlined below.
- **c.** Pour the some of the reagent you will use for the titration or titrant from the reagent bottle into a beaker. Be sure to label this beaker as NaOH.
- **d.** Close the stopcock and add about 5 mL of the titrant to the buret using a funnel. Tilt and roll the buret so that the titrant rinses all of the interior surfaces of the buret and be sure to run the titrant through the spout.
- **e.** Dispose of this rinse solution and repeat 1 or 2 more times to ensure that all interior surfaces of your buret are wet with the titrant solution.

Filling the buret -

- **a.** Clamp the buret into the buret clamp, close the stopcock, and fill it with titrant using a funnel. Allow the titrant to drain into a waste container until there is a solid column of titrant extending down to the tip, completely free of air bubbles.
- **b.** Allow 10-15 seconds for the titrant to drain from the walls and record the initial volume in the buret to the nearest 0.01 mL.

Performing the Titration –

- **a.** Pipet an aliquot of the solution to be titrated in an Erlenmeyer flask, add indicator and place it on a white surface underneath the buret. A color change is much easier to detect on a white background.
- **b.** Begin adding titrant to the reaction vessel. You will swirl the flask with the one hand and manipulate the stopcock with the other. (Generally, left hand will control the stopcock for right-handed individuals.)
- **c.** As you titrate you will see the indicator begin to change color near the point of addition of the titrant. This color will fade quickly as you swirl the solution. Near the endpoint, the color will take longer to fade away.
- **d.** As you reach this point, rinse the walls of the titration flask with deionized water and slow the rate of titrant addition until a drop (or less) makes the color change of the indicator persist for 30 seconds.
- **e.** Stop, allow 10-15 seconds for the titrant to drain from the buret walls, read the final volume of the buret to the nearest 0.01 mL, and record. Often it is helpful to hold a white card with a wide black mark on it behind the buret to see the meniscus better.
- **f.** To add less than a drop of titrant to the receiving flask, suspend a drop from the buret tip, touch it to the side of the receiving flask, and wash the wall of the Erlenmeyer flask with deionized water.

Clean up –

- **a.** After completing a series of titrations, drain the titrant from the buret, rinse the buret with several portions of deionized water, and drain each rinse through the tip. Place it in the appropriate container tip up.

Part III. Titration: Sodium Hydroxide Standardization

1. Pipette 25.00 mL of oxalic acid solution into a clean but not necessarily dry Erlenmeyer flask. (You may do three samples as you will be doing at least 3 titrations.). Refer to “Pipetting a Liquid ” on page 223 for detailed directions.
2. Add 2-3 drops of phenolphthalein to the Erlenmeyer flask. (Do not forget this step or you will not see any endpoint.)
3. Set up a buret using your sodium hydroxide solution as the titrant.
4. Titrate your oxalic acid solution to a pink phenolphthalein endpoint as described in “Titrating – Use of a Buret” on page 224.
5. Repeat until you have three titrations that differ from the average by no more than 0.05 mL of titrant.
6. Calculate the molarity of your standard sodium hydroxide solution. You should be able to determine the concentrations of the standardized sodium hydroxide to 4 significant figures.
7. Complete the excel spreadsheet for this experiment (the spreadsheet can be found on the chemistry department home page, under lab archives) if requested by your instructor.

Part IV. Titration: Acetic Acid Molarity

1. Clean your pipet by rinsing it with DI water and three aliquots of your vinegar solution.
2. Pipette 25.00 mL of vinegar solution into a clean but not necessarily dry Erlenmeyer flask and titrate this sample.
3. Add 2-3 drops of phenolphthalein and repeat until you get three titrations that differ from the average by no more than 0.05 mL of titrant.
4. Calculate the molarity of your acetic acid (vinegar) solution of unknown concentration.
5. Complete the excel spreadsheet for this experiment (the spreadsheet can be found on the chemistry department home page, under lab archives) if requested by your instructor. You should be able to determine the concentrations of the vinegar solution to 4 significant figures.

Prelaboratory Exercise Week 1: Volumetric Analysis

Lab Grade

Prelab Questions Week 1		15
Prelab Questions Week 2		13
General Format (Signature, ink, no obliterations, etc.)		6
Data and Analysis (observations, questions, units, significant figures, sample calculations, etc.)		18
Accuracy (% error)		30
Post Lab Questions		18
Total		100

Name: ____________________________

Section:______________________

1. What is a primary standard?

2. You have weighed out precisely 2.471 g of oxalic acid dihydrate, and diluted it to 250.0 mL. What is its molarity? (Hint: Example 12.1)

3. What is a secondary standard?

4. You have diluted 9 mL of roughly 6 M NaOH to 500 mL. What is the approximate molarity of the base? (Pay attention to significant figures) (Hint: Example 12.2)

5. What is the name of the indicator you are to use and what color does it turn at the end point?

6. What is an equivalence point?

7. What is the meaning of the term “end point” in a titration and what is true about the system at the end point?

8. Write the equation for the reaction of oxalic acid and sodium hydroxide.

9. A 20.00 mL sample of the standard oxalic acid solution (use molarity from question 2 above) requires 24.16 mL of your base to reach the end point. Calculate the molarity of the base. (Hint: Example 12.3).

10. Write the equation for the reaction of acetic acid and sodium hydroxide.

11. 25.00 mL of an unknown concentration of acetic acid is titrated with 40.62 mL of the standard base (use molarity from 9 above). What is the molarity of this acid? (Hint: Example 12.4).

Prelaboratory Exercise Week 2: Volumetric Analysis

Name: ___________________________

Section:____________________

1. What are volumetric pipets and burets and what are they used for?

2. To what volume are you to record the values from the buret? __________

3. 0.3811 g sample of KOH will just neutralize what volume of 0.2000 M H_2SO_4?
   - a. Write the balanced chemical equation.
   - b. What volume of 0.2000 M H_2SO_4 will just neutralize the KOH sample?

4. In a experiment similar to yours, a solution is prepared by adding 4.512 g of slaked lime also known as calcium hydroxide to 500.00 mL of water.
   - a. What is the molarity of the solution?
   - b. What is the molarity of the ions in the solution?

c. Write the balanced chemical reaction between calcium hydroxide and perchloric acid.

d. You are given a solution of perchloric acid of unknown concentration. If it took 25.77 mL of calcium hydroxide solution to completely neutralize 15.00 mL of perchloric acid, what is the molarity of the perchloric acid?

e. The actual molarity of the perchloric acid 0.4237 M, what is the percent error?

Data Sheet: Volumetric Analysis

Record all of your original data on this sheet in ink and you may transfer it to the lab book pages in pencil to do the work-up.

Part I. Preparation of Oxalic Acid Standard

TABLE 12.5

mass of oxalic acid dihyrate and beaker	
mass of beaker	
mass of oxalic acid dihydrate	
volume of volumetric flask	
concentration of oxalic acid	

Sample Calculations

Part II. Preparation of Sodium Hydroxide Standard

TABLE 12.6

molarity of concentrated NaOH stock solution	
Volume of NaOH stock solution used	
Final volume of dilute NaOH solution	
Approximate concentration of NaOH Solution	

Sample Calculations

Part III. Standardization of NaOH Solution

Pipet volume _______________

TABLE 12.7

	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**
Molarity of oxalic acid solution (from Table 12.5)				
Volume of oxalic acid solution used				
NaOH buret, final volume reading				
NaOH buret, initial volume NaOH added (should be about 0.00 mL)				
Volume of NaOH solution used				
Molarity of NaOH				

Average Molarity of NaOH solution______________________

Write the balanced chemical equation for the reaction that takes place during the standardization of sodium hydroxide.

Sample Calculations (Remember to show one calculation of each type for one trial.)

Part IV. Standardization of Unknown Acid

Pipet volume ________________ unknown identification_____________________________

TABLE 12.8

	Trial 1	Trial 2	Trial 3	Trial 4
Volume of unknown acid solution used				
Molarity of base (fromTable 12.7)				
NaOH buret, final volume reading				
NaOH buret, initial volume of NaOH added				
Volume of NaOH solution used				
Molarity of unknown acid				

Average Molarity of vinegar solution______________________

Write the balanced chemical equation for the reaction that takes place during the standardization of the acetic acid solution of unknown concentration.

Sample Calculations (Remember to show one calculation of each type for one trial.)

Post Lab Problems

1. Consider the following and briefly explain what would be the effect on the calculated analyte concentration; i.e. why would the calculated analyte concentration be higher, lower, or unchanged than the actual value if:
   - a. the last drop of base was not rinsed from the buret into the reaction flask?
   - b. the reaction flask contained traces of water in it before the acid was added with the volumetric pipet?
   - c. the volumetric pipet used to transfer the acid solution contained traces of water?

2. Commercial lye is an alkaline mixture that contains mainly sodium hydroxide, but major impurities include sodium chloride, sodium carbonate, sodium chlorate, sodium sulfate, potassium and heavy metals such as iron and nickel. A 0.8500 g sample of commercial lye is dissolved in water and titrated with 35.00 mL of 0.5377 M HCl.
   - a. Write the balanced chemical equation.
   - b. What mass of sodium hydroxide was in the lye sample?
   - c. What is the percent purity of the lye sample? (i.e., what is the mass percent of NaOH in the lye?)

3. What is the molarity of a 12.6% solution of sodium carbonate with a density of 1.31 g/mL?

4. What is the mass percent of potassium iodide in a 2.89 M solution of KI with a density of 1.27 g/mL?

5. A 30.0 mL aliquot of 0.300 M H_3PO_4 is mixed with 90.0 mL of 0.200 M KOH, and the mixture is evaporated. Will the salt that crystallizes out be K_3PO_4, K_2HPO_4, or KH_2PO_4? Show your calculations by using an ICE table (or two or three).

Hint: when a polyprotic acid such as H_3PO_4 is mixed with a base, the protons will react one at a time. For example, if the one mole of OH^- was mixed with one mole of H_3PO_4, it would produce one mole of $H_2PO_4^-$ and one mole of water:

$$H_3PO_{4\,(aq)} + KOH_{\,(aq)} \rightarrow KH_2PO_{4\,(aq)} + H_2O_{\,(l)} \quad \textbf{(EQ 12.13)}$$

The next proton on the $H_2PO_4^-$ would start to react only when more OH^- was added, over and above the first mole:

$$KH_2PO_{4\,(aq)} + KOH_{\,(aq)} \rightarrow K_2HPO_{4\,(aq)} + H_2O_{\,(l)} \quad \textbf{(EQ 12.14)}$$

The final proton on the HPO_4^{2-} would start to react only when more OH^- was added, over and above the second mole:

$$K_2HPO_{4\,(aq)} + KOH_{\,(aq)} \rightarrow K_3PO_{4\,(aq)} + H_2O_{\,(l)} \quad \textbf{(EQ 12.15)}$$