

NO RETURNS

Chemistry 60

LABORATORY EXPERIMENTS FOR

BEGINNING CHEMISTRY

Proceeds of this manual are used for Pierce student scholarships

Department of Chemistry



Fifth
Edition

Periodic Table of the Elements

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< learn these elements

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INTRODUCTION

A laboratory is a place where an investigator can examine some small part of the world of nature. The scientific investigator is looking for answers to specific questions about the composition and behavior of matter under various conditions. To find the answers to the questions the investigator creates situations in which certain substances and processes can be observed and measured. The purposeful creation of situations to observe is called **experimentation**.

Observations that can be made with your own senses, without using a measuring instrument, are qualitative observations. The fact that an insoluble substance forms when two solutions are mixed is a **qualitative** observation. The observations that hydrogen sulfide has a foul odor and that copper sulfate forms a blue solution in water are qualitative in nature.

How much insoluble material forms when measured amounts of two solutions are mixed is a **quantitative** question. To answer it, a measurement must be made with an appropriate measuring instrument. In general, qualitative observations lead to answers about what is happening; quantitative measurements help answer how much.

EXAMPLE 1:

Does an insoluble precipitate form when a silver nitrate solution and a hydrochloric acid solution are mixed?

Answer (by direct observation): Yes.

In many cases the measurement or direct observation you make is not the answer to the question posed. It is often necessary to deduce (figure out) from other known facts and relationships just what the observation means with respect to what you want to find.

EXAMPLE 2:

What is the insoluble substance formed when silver nitrate and hydrochloric acid solutions are mixed?

You cannot answer this question by direct observation because one white powder looks pretty much like any other white powder. You must deduce the answer from known facts about the solubility of the possible products of the reaction. This means you must know what species are present and which ones can react with each other. By examining the properties of the possible products, you can arrive at a conclusion about the probable identity of the white powder.

EXAMPLE 3:

Is aluminum denser than water?

From the direct observation that a solid piece of aluminum sinks in water, one could deduce that aluminum is denser than water.

EXAMPLE 4:

What is the density of aluminum?

This is a quantitative question, so we must make some measurements. However, there is no instrument that measures the density of solids directly. We must measure something that is related to density and calculate what the density must be. To know what measurements to make, we must know what the relationship is.

Each experiment in this course will be proposed to you as a problem, either qualitative or quantitative, for you to solve by making the appropriate observations or measurements and deductions. The background facts and relationships will be reviewed. Untested assumptions about the experiment will be examined. Then we will decide what measurements we can make with the available apparatus to find the answer to the problem.

Knowledge of the characteristics of the materials, substances, and apparatus to be used in a particular experiment can be obtained by reading the laboratory textbook and by prior planning of the experiment and your laboratory notebook. You will be amazed at how this will speed and smooth your journey through this laboratory course.

LABORATORY MANNERS

Your laboratory experience will be pleasant if everyone in the laboratory observes courteous behavior and common sense. Common sense is easier to follow if you know what you are doing and why you are doing it. The following points are offered as reminders of good laboratory manners.

1. **Protect your eyes with approved safety eyewear at all times.** You, of course, will be careful and never have an accident, but you never know about your neighbor!
2. Dress for lab: **always wear closed toed shoes** that protect the tops and soles of your feet while you are in the laboratory. People do spill things, and concentrated acid between the toes is a thrill you will never forget. **Tie long hair back and do not wear loose flowing garments.**
3. **Let nothing touch your lips** during a laboratory session. The laboratory is not a kitchen. Food, drinks, and gum can too easily become contaminated by toxic (poisonous) substances. Avoid touching your face with your hands- if you must, wash them first. **Wash your hands before you leave lab** to avoid contaminating your lunch.
4. Find out where fire extinguishers, safety showers, and eye-wash fountains are and know when and how to use them.
5. In case of chemical spills on skin or clothing, wash the affected area immediately with plenty of water; then tell the instructor.
6. Follow the directions in your lab manual. If you think you have a better way of doing an experiment, discuss it with the instructor. Please do not experiment on your own yet.
7. **Keep your work area clean and dry.** You will discover that spills and breakage occur less often if you keep your glassware at the back of the bench. Look before you reach.
8. **Take only as much chemical as you need** from the stock bottles. **Never** return unused chemicals to a stock bottle; you might contaminate the rest of the contents of the bottle.
9. Use the fume hood or lab snorkels for reactions which release toxic or noxious gases- your lab manual will direct you when these are to be used.
10. For your health and that of the environment, **dispose of chemicals properly**- your lab manual will indicate where to dispose of used chemicals.
11. Focus and awareness is important in the laboratory- so **do not use any electronic devices**- cell phones, texters, ipods, etc.- while in lab.
12. When using common equipment, like balances or shared reagent bottles, always **leave them clean when you are done** and return them to their original location.
13. Electronic balances are expensive and delicate instruments, read the posted instructions before using them. **Never add chemicals to a container on a balance**, remove it first, add the chemical and then return it for reading.
14. Return all of your glassware and lock your lockers at the end of lab. Return all common equipment (rings, test tube clamps, wire gauze, burets, etc.) to where they belong.
15. **Report all accidents to your instructor.**
16. **Place all backpacks, jackets and books not being used in a cubby** on the back wall.
17. **Read the experiment carefully and prepare your lab notebook before each lab experiment.** Use common sense and ask your instructor whenever you are unsure.

TIPS ON TECHNIQUE

1. **Never heat a beaker or flask directly with a flame.** Use a wire gauze support to spread the heat. A direct flame may cause the glass to crack.
2. Never heat soft (non-Pyrex) glass or plastic apparatus. Graduated cylinders, funnels, vials, etc., will break or melt every time. Glassware labeled Kimex or Pyrex can be heated.
3. Never heat volumetric glassware (those used to accurately measure volumes) such as volumetric flasks, pipets, or burets. Even if they do not break the heat will distort them and destroy their accuracy.
4. When dispensing a liquid from a reagent drop bottle, do not touch the tip to your test tube, spot plate or flask. This can contaminate the drop bottle. Always drop the reagent holding the dropper bottle above your container.
5. Keep books, papers, and personal gear off the lab bench- store them in the cubbies in the lab. The only exception is your laboratory notebook. Soggy textbooks are hard to read and impossible to sell.
6. **ALWAYS** follow the instructions in your lab book and if you are not sure about anything ask your instructor.
7. Do not leave a burner or an apparatus with running water unattended. Beakers boil over or boil dry when you are not looking, and hoses invariably disconnect when you turn your back.
8. Always heat the contents of a test tube by placing the test tube in a hot water bath. Heating a test tube directly in a flame can be dangerous to those around you. The heated liquid can be ejected from the tube and burn someone else in the lab.
9. Do not waste water- it is a precious commodity in California.



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

MEASUREMENTS AND ERROR

SIGNIFICANT FIGURES IN MEASUREMENTS AND CALCULATIONS

Measurement is the operation of comparing a property of a substance to the same property of a standard. We find the length of an object by counting the number of standard lengths laid end to end that correspond to that dimension. We find the mass of an object by counting the units of standard mass which experience the same force (weight) in the earth's gravity. Since every measurement amounts to a physical comparison, there is always a certain amount of inexactness or uncertainty in the value of that measurement. This error or uncertainty may come about in several ways:

1. The measuring instrument used may not reproduce the standard accurately. One does not use a ruler to construct an airplane, for example.
2. The measuring instrument may not be sensitive to small differences in the value of a measurement. A truck scale is not used to measure the weight of the truck driver.
3. The measurement taken with a certain instrument may not be reproducible in repeated trials (precision). A precise instrument reproduces a measurement within a very small interval.
4. Different methods of measuring may result in different values.

Notice that carelessness and mistakes are not included in the list. The errors and uncertainties that we are discussing are those that a conscientious worker cannot avoid with the available measuring devices. You should be aware of these limitations so that you do not place more reliance on your answers than the data warrants.

ACCURACY

The accuracy of a measuring device can be determined by measuring a primary standard. If a mass balance reports an average value of 1 kilogram for a standard kilogram, it is accurate, at least at that one value. If a thermometer reads zero degrees Celsius in a mixture of pure water and ice, it is accurate at that temperature, because the freezing point of pure water is defined as 0°C.

If the instrument is accurate at one value it probably gives true values at other measurements not too far removed from the test value. In general, however, a number of different independent measurements must be made to test for accuracy. If the measurements all agree, the value probably is true.

For example, we will determine the volume of a metal object by three independent methods. We will determine the volume (1) from its dimensions, (2) from the indicated liquid level change on immersion in a graduate, and (3) by application of the buoyancy law. If the values all agree within a certain interval, we can be confident about the accuracy of the value.

SENSITIVITY

The sensitivity of a measuring device is the smallest unit that can be read or estimated of the property it is measuring. An analytical balance (used in Chemistry 101) will give a measurable response to the weight of the fingerprint of the operator (± 0.0001 gram). The milligram top loading balance we will use in this class will not be affected by fingerprints but can be affected by air currents in the lab. It has a sensitivity of ± 0.001 gram and is ten 10 times less sensitive than the analytical balance. The capacity and sensitivity of an instrument are usually inversely

proportional to each other. If a measuring device has a large capacity or range, it is generally not very sensitive. This makes it difficult to detect small differences in large measurements.

PRECISION

The precision of a measurement is related to the reproducibility of the value. If an archer puts five arrows within a two-inch circle he is shooting precisely. If the circle happens to be twelve inches from the bull's-eye, he is precise but not very accurate.

The precision of an instrument can be determined only by repeated measurements of the same property of an object. The precision is related to the sensitivity of the measuring device and is usually less than the last decimal unit that can be read (or estimated) on the scale of the instrument.

EXAMPLE:

Suppose an object is weighed repeatedly on several balances.

Balance 1. The scale is in increments of 0.1 gram and the hundredths figure is estimated.

	MEASURED WEIGHT	DEVIATION
Trial 1.	10.19 g	0.01 g
2.	10.23 g	0.05 g
3.	10.14 g	0.04 g
4.	10.21 g	0.03 g
5.	<u>10.11 g</u>	<u>0.07 g</u>
Sum	50.88 g	0.20 g

The average weight value is $50.88 / 5 = 10.18$ grams.

The deviation listed to the right is the amount each value differs from the average. The average deviation is the sum of the absolute value of each deviation, divided by the number of trials. In this case, the average deviation is $0.20 / 5 = 0.04$ g. That is, on the average, there is an uncertainty of ± 0.04 in our answer. We can say that the probable weight of the object is 10.18 ± 0.04 grams. The value has four significant figures, even though the last one is somewhat uncertain. The digits 1, 0, 1, and 8 are all significant.

If we divide the average deviation by the average value, we obtain a measure of the relative uncertainty of the measurement. In this measurement the relative uncertainty is $0.04\text{g}/10.18\text{g} = 1/254$ or one part in 254.

Balance 2: The scale on this balance can be read without estimate to 0.01g, and the thousandths unit estimated. On this balance the values for the mass of the same object might be:

	MEASURED WEIGHT	DEVIATION
Trial 1.	10.193 g	0.011 g
2.	10.181 g	0.001 g
3.	10.176 g	0.006 g
4.	10.169 g	0.013 g
5.	<u>10.189 g</u>	<u>0.007 g</u>
Sum	50.908 g	0.038 g
Average	10.182 g	0.008 g

The value can be reported to five significant figures: 10.182 $\pm$ 0.008 grams. The relative uncertainty of the balance is 0.008/10.182, or one part in 1270. This balance is five times as precise as Balance 1.

SIGNIFICANT FIGURES (or DIGITS)

The number of significant figures in a measurement is an indication of the precision of the measurement. The last digit in a value is presumed to be uncertain. If you report that you obtained 1.0 gram of product, the reader should presume that you mean something between 0.5 and 1.5 grams. If you report 1.00 gram, the reader knows that you measured it to the nearest tenth of a gram and there is some uncertainty about the third significant figure. If not specified, the uncertainty in the last significant figure is presumed to lie within 5 units of the reported digit. For example, 1.283 means 1.283 $\pm$.005, or "at least 1.278 and not more than 1.288."

The rules for counting significant figures are as follows;

- All non-zero digits are significant
- Zeros between non-zero digits are significant
- Leading zeros are not significant
- Trailing zeros are significant if a decimal point is shown; otherwise they are assumed to be insignificant.
- Exact numbers have infinite significant figures.
- Defined numbers have infinite significant figures.

If the measurement has no significant figures to the right of the decimal point, it is best to write the number in exponential form to indicate the number of significant figures. For example, you cannot be sure how many significant figures there are in 2500. If there are four, you should write 2.500 $\times 10^3$. If only two significant figures are justified by the measurement, record it as 2.5 $\times 10^3$.

If we have had no prior experience with an instrument, a single measurement leaves us in doubt about the precision of the recorded value. We should take enough measurements to treat the values statistically. However, time is also important, and after sufficient experience with one instrument, you will know what its inherent precision is. Then you will know how much confidence to place on a single measurement.

USE OF LABORATORY EQUIPMENT TO MEASURE MASS AND VOLUME

Measuring the Mass of Solids

Using the electronic balance

In this lab, we will use the electronic balance to measure the mass of an object. Electronic balances have many advantages over other types of balances mainly the ease with which measurements are obtained. To measure the mass of an object follow the rules below.

- Make sure the electronic balance is placed on a flat, stable surface. Wind, shaky surfaces, or any kind of force can affect the reading accuracy of the balance.
- Press the "ON" button and wait for the balance to show zeroes on the digital screen. If the balance was already on and not reading 0.000 g, press the "zero/tare" button and wait for the display to read zero. The symbol "⚖" will appear on the left side of the display when the balance has stabilized.
- Place the solid object on the balance and wait for the mass to stabilize. When the symbol "⚖" appears, record the mass in your notebook. Always report all the digits in the mass that the balance indicates. For our electronic balances, it'll be to 3 places after decimal.
- If you are measuring the mass of a liquid or a solid chemical, you must weigh the substance in a beaker or a flask. To automatically deduct the weight of the container, press the "Tare" or "zero" button while the empty container is on the balance and wait for "⚖" to appear. Carefully add the substance to the container. Ideally this is done while the container is on the balance but if it's easier, you may remove the container from the balance, add the appropriate amount of chemical, then return the container to the balance to determine the mass of the chemical. Record the mass as indicated on the display.
- While determining the mass of an object, use the same balance for the entire experiment.

Measuring the Volume of Liquids

The pieces of laboratory glassware we will use in this semester for measuring volume include the beaker, graduated cylinder, volumetric flask, Erlenmeyer flask, pipet, and buret.

To Deliver and To Contain

The glassware used for measuring liquid can be divided into two categories: those that contain and those that deliver a specified amount of liquid. Glassware that contain include graduated cylinders and volumetric flasks (labeled with "TC" meaning "to contain"); for these, the inscribed markings increase in value from bottom up. Glassware that deliver include pipets and burets (labeled with "TD" meaning "to deliver"); for these, the inscribed markings increase in value from top to bottom, thus indicating how much liquid has been delivered not how much remains in the glassware.

Glassware Used for Measuring Volume

Some of the glassware in the chemistry lab is inscribed with markings for measuring the volume of liquids. This includes beakers, Erlenmeyer flasks, graduated cylinders, and burets. Even though beakers and Erlenmeyer flasks have markings on them they are generally not accurate enough for measuring volume. Thus, these pieces of glassware are primarily used to hold liquids during experiments. When measuring the volume of a liquid using the graduated cylinder or buret, the rule is to estimate **one more digit past the smallest division** on the measuring device. Other glassware, such as the volumetric flask and the pipet only have a single line that indicates volume since they are used to measure only one specific volume and thus are much more accurate than the other measuring devices.

The Meniscus

When liquid, such as water, is placed in a glass or plastic container, surface tension causes the surface of the liquid to take on a curved shape, called a meniscus. Lab glassware used for measuring liquid is calibrated so that an accurate result is obtained by reading the bottom of the meniscus, when it is viewed at eye level. Viewing the meniscus at any other angle will give incorrect results. The accuracy of the markings on the different glassware vary significantly. The 50 mL graduated cylinder we will use in this lab have a precision of ± 0.5 mL, the 10 mL graduated cylinder has a precision of 0.1 mL, the buret with a precision of 0.05 mL, the pipet 0.01 mL.

Note that the buret and pipet are more precise than the graduated cylinder.

YOUR LABORATORY NOTEBOOK

A laboratory notebook is a permanent record of experiments carried out in the laboratory. Everyone who may be occupied in scientific or technological research must learn to keep a proper record of their experimental conditions and observations. There are a number of reasons for this practice. Nobody's memory for details is infallible, and it is often necessary to repeat an experiment or report the findings in a communication. The notebook serves as an original source for the data. For people doing original research, the laboratory notebook is a legal record that establishes their claim of discovery.

You will benefit in several ways by keeping a careful record of your laboratory experience. The experiments in this text are illustrative of most of the major lecture topics, so your notebook will help you review for examinations. Careful observing and recording habits will give you an advantage in subsequent courses where notebooks are required. Communication skills are valuable in employment applications, too.

The laboratory notebook is your diary of your laboratory experience. As such, it may contain anything that you feel is pertinent. The minimum content of the notebook must show:

1. What was done
2. When it was done
3. What was observed
4. Who performed the experiment

To be of the greatest value, all notebook entries of observations and data should be made at the time of the experiment. Notes made on stray scraps of paper are worthless. Entries should be made **in ink** in a permanently bound notebook with sequentially numbered pages, dated and signed at the bottom of each page. In research laboratories, it is also required that each page be witnessed by a competent person who is familiar with the work.

A laboratory worker sets up or contrives a situation to observe in such a way that the observations or measurements answer a question. Each experiment therefore starts with a question to be answered.

Before you start any laboratory work, you should write a general outline of what you propose to do and how the work relates to the problem you seek to solve. In your outline, note the known facts and assumptions which bear on the problem.

As you complete each phase of the experiment, describe your actions in sufficient detail so that someone else could repeat the experiment if necessary. Name or describe the apparatus used, and give the amounts of all substances employed. Record the measurements and observations in ink at the time that you make them. List repetitive measurements in parallel columns when repeated trials are made.

USE OF BOUND LABORATORY NOTEBOOK

Information obtained from scientific work is of interest to scientists from all over the world, especially if they are investigating similar things. When a set of successful experiments is completed and analyzed the results are typically published in a scientific journal allowing anyone to see what was done. These results can now be carefully scrutinized and added to, or corrected, by other scientists to increase our body of knowledge. Eventually this information may also be of interest to nonscientists as well, as it will help us to better understand the natural world around us and make more informed decisions about how to live our lives.

The information obtained from experiments in its simplest form is called **data** and it serves as a record of what occurred during an experiment. Without the data there is nothing to support the results or conclusions that are made. It is necessary that an accurate and permanent record of the data be made at the time it is collected. Without this record the experimental work cannot be published. This record is also needed should the experimental work be challenged by other scientists or in a court of law. All experimenters, whether working in academics or industry, must keep a notebook with these records. It is a legal requirement. To help you to follow this regulation you will be keeping a notebook of all of your experimental data during this semester. The notebook should be clearly organized and labeled using neat columns or tables whenever possible. It takes some experience to learn how to construct these tables properly so to help you initially sample data tables will be given in this lab book for the first few experiments. Copy these tables neatly into your notebook **using ink** before coming to lab. When in lab, write all of your data directly into these tables.

Please follow these guidelines when using your notebook for lab reports.

Use black or blue ball point pen only and do not at any point use whiteout; if you make a mistake, simply put a line through it and write the corrected information next to it.

NEVER erase any work which is incorrectly recorded. Draw a neat line through the incorrect number and make the new record nearby:

~~21.08g~~ 21.48 g

1. Write your name, your instructor's name, course number, and lab time on the front label. Write your last name across the cut bottom edges of the pages. Write your locker number and lock combination on the inside of the front cover.
2. Your notebook has duplicate numbered pages. Insert the cardboard backing behind each set of pages before beginning to write. Ensure that the writing appears clearly on the duplicate copy. These are the copies you will turn in to your instructor. If these are not legible, turn in the original.
3. Use the first two pages to make a Table of Contents:

Date	Experiment #	Title of Experiment	Page #	Grade
------	--------------	---------------------	--------	-------
4. Start every experiment on a fresh page, using only right-hand pages. Put the date, experiment # and title of experiment at the top. Use clear sub-headings where several experiments are done under one main heading. Write going across the page rather than down the column.
5. In writing up an experiment, organize it in the following way, unless directed otherwise:

HOW TO WRITE YOUR LAB REPORT

Prelab (to be completed before experiment and class)

- a. **Objective:** State concisely what the purpose or objective of the experiment is.
- b. **Background:** Write the definition of the terms asked for in the manual.
- c. **Procedure:** Write as briefly as possible what is to be done and the correct order of doing it. It is acceptable to use an outline and number the steps.
- d. **Prelab Assignment:** Answer all prelab questions when applicable.

In-lab (to be completed while doing the experiment before leaving lab)

- e. **Data:** Every measurement or reading must be recorded in ink when it is made. A clear label or statement should show exactly what has been measured. Digits should be arranged in columns vertically under each other. The value of a neat, permanent record made at the time of observation cannot be overemphasized. NEVER record laboratory data on any piece of paper other than your notebook.
- f. **Calculations:** Arrange each calculation neatly and in detail on the page (it is best to first do calculations on scratch paper and then copy them into your notebook). Always include units and show a prerounded and then final rounded answer. This page may also be used for any omissions from the write-up, for drawing diagrams or for any required graphs.

Post-lab (to be completed after doing the experiment before next class)

- g. **Results:** List the results of the calculations. When appropriate, give a table of summarized results. Whenever possible express the results in tabular form. Clearly label all results and insure that all numbers are accompanied by appropriate units.
- h. **Discussion:** Write a statement summarizing your results and observations. Explain what the results mean! (What is the "take-home lesson"? What significant conclusions can be reached on the basis of the results?) Comment on the probable sources of error and limitations of any analytical methods employed. Briefly comment on the precision and accuracy of your own work and any unusual observations made.

IMPORTANT: A well thought out and well written discussion shows your understanding of the experiment and your grade will depend on this aspect of the write-up.

- i. **Answers** to all Post lab Exercises and/or Questions when applicable.

Come to lab PREPARED to do the assigned experiment. This requires that you will have read and studied the experiment at home and **prepared your notebook by writing the problem, theory, outlined procedure, answers prelab questions, and set up tables for recording your data into.** A sample lab report appears below. Examples of how to properly prepare your notebook will be given for the first few experiments in this book.

The best time to complete the write-up of an experiment is as soon as possible. When there is time left, start working on your calculations right in lab- this way you can obtain help. Your instructor will tell you when the completed report will be due.



*Wear shoes in the laboratory. People do spill things,
and concentrated acid between the toes
is a thrill you'll never forget!*

SAMPLE NOTEBOOK ENTRY

OBJECTIVE:

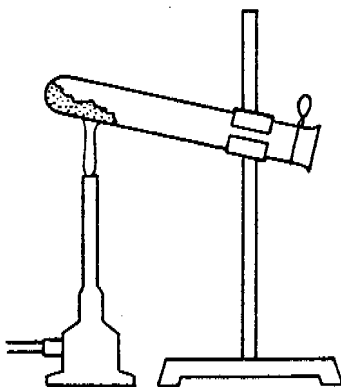
To determine the percent water in copper sulfate pentahydrate by heating.

BACKGROUND:

Empirical formula- the smallest whole number mole ratio of a compound/molecule.

PROCEDURE:

A quantity of copper sulfate hydrate was measured by weighing a test tube with and without the sample. The sample was then heated with a flame of a Bunsen burner in the manner pictured until no more weight loss could be detected. A second sample was treated in the same way.



EXPERIMENTAL DATA:

	TRIAL 1	TRIAL 2
1. Weight of test tube + sample before heating	<u>23.85 g</u>	<u>23.39 g</u>
2. Weight of test tube	<u>19.10 g</u>	<u>19.50 g</u>
3. Weight of sample before heating (by difference)	<u>4.75 g</u>	<u>3.89 g</u>
4. Weight of sample + t.t. after 1st heating	<u>22.35 g</u>	<u>22.25 g</u>
5. Weight of sample + t.t. after 2nd heating	<u>22.17 g</u>	<u>22.02 g</u>
6. Weight of sample + t.t. after 3rd heating	<u>22.14 g</u>	<u>21.99 g</u>
7. Weight loss of sample (#1 – #6)	<u>1.71 g</u>	<u>1.40 g</u>

CALCULATIONS: The percent water is the weight loss (#7) divided by the weight of the sample before heating (#3) times 100.

TRIAL 1

$$\frac{1.71}{4.75} \times 100 = 36.0\%$$

TRIAL 2

$$\frac{1.40}{3.89} \times 100 = 36.0\%$$

$$\text{Average} = 36.0\%$$

RESULTS:

	Trial 1	Trial 2	Average
% water in salt hydrate	36.0%	36.0%	36.0%

DISCUSSION: The weight percent water in copper sulfate pentahydrate averages 36.0% to three significant figures. The two trial results were very precise, in fact exactly the same, suggesting that the % mass of water in a particular salt hydrate is constant and independent of how much salt hydrate is used. Sources of error include not evaporating all the water, which would increase the calculated percent of water in the sample.

What you need to know for Quiz Before Beginning Lab

- Significant figures
- Common laboratory equipment (page A13)
- Safety in the lab (safety quiz) (page A10)
- Basic parts of the lab report

Sample Quiz 1

1. Name the lab glassware shown to the right is:



2. How many significant figures are in each of the following numbers?

47 students _____ 0.004506800 _____

3. Calculate the following to the correct number of significant figures:

$2.34 + 1.235 \times 3.7 =$ _____

4. Answer True/false

Long hair need not be tied up if you are not using a burner _____

Small amounts of acids or base can be poured down the sink _____

5. Where should the following information be recorded in a students' notebook? Choose from the following;

objective, background, procedure, data, calculations, results, discussion

Mass of sodium chloride in test tube: 2.342 g _____

EXPERIMENT 1

MEASUREMENT, VOLUME, AND DENSITY

PRE-LAB

Objective: Write short statements describing the purpose of each part of the experiment. Since there are four parts to this experiment, write a separate statement for each part.

Background: Define the terms density, buoyancy force, extensive property, intensive property, precision, and accuracy.

Procedure: Briefly write/ outline what you will be doing in each part of the experiment. Write a separate procedure for each part.

Data: Write separate data tables for parts a, b, c, and d. Use the lab manual as your guide for parts a, b, and c. Put all data tables following each other without leaving space for calculations. Create a data table for part d by reading through the procedures and noting all measurements you are asked to record. Create clear labels for each measurement.

OBJECTIVE OF EXPERIMENT 1

To calculate the density of an aluminum cylinder from its mass and volume. The mass will be measured using an electronic balance (Part A), and the volume will be found using three different methods: by measuring dimensions (part B), by displacement (part C), and by buoyancy force (part D).

BACKGROUND OF EXPERIMENT 1

DENSITY

Matter has mass and it occupies space (has a volume). Mass is the amount of matter present in an object, while volume is the amount of space a substance occupies. These two properties, mass and volume, are called **extensive properties**. An extensive property is one that depends upon the amount of substance present. The more we have of any sample of a substance the larger will be its mass and volume.

A relationship exists between these two extensive properties and it is called the density. Density is the ratio of the mass to the volume of a given amount of matter.

$$\text{Density} = \text{mass/volume}$$

Density is an **intensive property**. An intensive property is a property independent of the amount of substance present. It is a fundamental characteristic of the substance itself. Any object made

of a particular pure substance, no matter what its size or shape, has the same density. For example any aluminum object, a piece of foil or a cylinder, has a density of 2.702 g/cm^3 (or g/mL) and any amount of pure water has a density of 0.997 g/mL , at room temperature. It is only because the volume of an object changes slightly as temperature changes that the density values will also change. The volume of matter almost always increases with temperature, so density decreases as the temperature rises and increases as temperature falls. It is for this reason that any density value must be reported with a temperature. (Since the change of density with temperature is small for solids, it may be ignored in this experiment.)

PART A: MEASURING MASS

OBJECTIVE A

To measure the mass of the aluminum cylinder.

BACKGROUND A

Mass is defined as the amount of matter present in an object. The mass of an object is always constant. Mass is determined using an instrument known as a **balance**. There are many different types of balances. In this class you will be using an electronic top loading balance. Like any mechanical device, the balance may be out of adjustment and so each semester all of the balances are calibrated (adjusted) with samples of known masses to ensure that they are reading properly. This calibration process need only be done periodically. An important adjustment that must be made before **every** use of the balance is called **taring or zeroing**. This is done by simply touching the zero button when the balance is clean and there is nothing positioned on it, and then waiting until it reads 0.000 g .

Error in Experiment
All measurements have some error associated with them. The extent of this error depends upon both the measuring devices being used and the measurer herself. To minimize error the device should be working properly and be used properly. Calculating deviation and average deviation are one way to determine error when related to a repeating series of a measurement. They quantitate the precision (see page 11) or the closeness (reproducibility) of a series of measurements of the same dimension. Another way is to calculate error and percent error . These measurements can only be calculated when the quantity being determined experimentally is known, such as the density of aluminum in this experiment. When a value such as this is known it is called the true value . In this case it is 2.702 g/cm^3 . Error and percent error quantitate the accuracy (see page 10) or closeness of a measured (experimental) value to its true value. Formulas for error and percent error appear on page 27.

PROCEDURE A

Make three determinations of the mass of your cylinder on the electronic balance and record them in ink into the data table you have constructed in your notebook (see below). **Important:** **always** wait until the stabilization symbol “**⏏**” appears in the **lower left** of the display before recording the mass or adding or removing anything from the balance (read page 14, “Measuring the mass of solids” before using the balance).

IN-LAB

Data A: Complete the data table for part A.

Calculations A: perform the calculations asked for in part A on the next page.

Repeat this format (data followed by calculations) for the remaining parts of this experiment.

Sample Data Table for part 1A:
Copy this format into your notebook

Data: Mass Determination

	Measurement	Deviation from average
Trial 1	_____	_____
Trial 2	_____	_____
Trial 3	_____	_____
Sum	_____	_____
Average (in g)	_____	_____

CALCULATIONS A

Calculations: *(Do these on scratch paper first and then transfer into your notebook under the calculations section.*

- 1) Calculate the average mass and the average deviation for the set of measurements and transfer them into your notebook in the data section (as shown above).

The average is the result obtained when the sum of all measured quantities is divided by the total number of quantities.

$$\text{Average} = \frac{\text{Sum of quantities}}{\text{Number of quantities}}$$

The deviation is the absolute value of the difference between each value and the average value.

$$\text{Deviation} = | \text{trial value} - \text{average value} |$$

PART B: CALCULATING VOLUME BY MEASURING EXTERNAL DIMENSIONS

OBJECTIVE B

To measure the external dimensions of an aluminum cylinder and calculate its volume.

BACKGROUND B

A vernier caliper is a device used for measuring the distance between two calibration marks. As opposed to a ruler where the measurer positions the object and reads the measurements, a caliper fixes the position of an object between its “jaws”. Therefore the caliper is a more precise measuring device than the centimeter ruler. After calibrating the caliper to read zero when the jaws are completely closed, the jaws of the caliper, once tightened around an object, will accurately define the distance to be measured. The digital read out makes reading the measurement simple. The digital caliper you will use provides millimeter readings to the nearest hundredth place (all non-zero numbers on a digital instrument are considered significant).

VOLUME

The volume of a regularly shaped solid object may be calculated from its dimensions if the geometry is not too complicated. It is known that the volume of a cube is the cube of the length of an edge, and the volume of a rectangular parallelepiped (a box) is the length x width x height.

These two volumes are illustrated in Figure 1.2, along with two other common shapes, the right cylinder and the sphere. The abbreviations used in the formulas are s, side; l, length; w, width; h, height; d, diameter. π (the Greek letter pi) is the mathematical constant equal to approximately 3.1416.

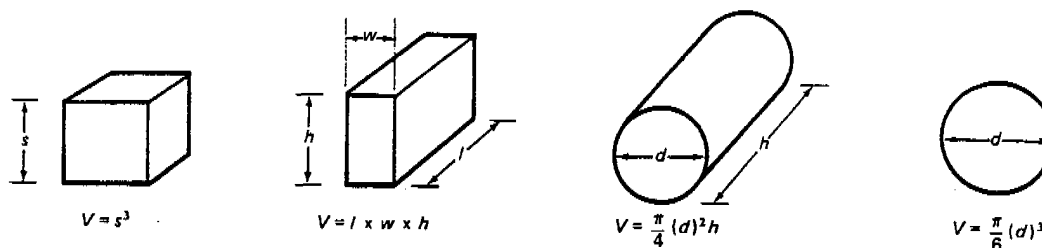


Figure 1.2

PROCEDURE B

Remove the aluminum cylinder in your locker and ensure that it is clean and free of tape. What dimensions must you measure in order to calculate its volume?

In this part of the experiment, you will measure the dimensions of the cylinder with a vernier caliper. After calibrating your caliper to read zero with the jaws of the device completely closed, insert your cylinder between the set of jaws on the bottom of the caliper and tighten. Make three measurements of each dimension of your object with the vernier caliper and record the data in your data table. Be sure to label the dimensions properly.

Sample Data Table for part 1B:

Copy this format into your notebook

Dimensions to be measured:	<u>height (h)</u>	<u>Diameter (d)</u>
	Measurement	Measurement
Trial 1	_____	_____
Trial 2	_____	_____
Trial 3	_____	_____
Sum	_____	_____
Average (in mm)	_____	_____
Average (in cm)	_____	_____

CALCULATIONS B

- 1) Calculate the average value, in units of mm, for each dimension measured for the caliper (height and diameter)
- 2) Convert the average values for height and diameter from units of mm to cm and copy them into your data table.
- 3) Calculate the volume of your cylinder using the average of the values (in cm) you determined from the caliper and highlight your final answer. Use the formula:

$$\text{Volume} = \frac{\pi}{4} d^2 h$$

and show all of your work. Transfer all of these calculations neatly into your notebook.

Here is a sample calculation.

$$V = \frac{\pi}{4} d^2 h = \frac{3.1416}{4} (1.23 \text{ cm})^2 (2.65 \text{ cm}) = 0.7854 (1.5129 \text{ cm}^2) (2.65 \text{ cm}) = 3.1488 \text{ cm}^3 = \underline{\underline{3.15 \text{ cm}^3}}$$

*(Show any formulas you are using and all of your steps here including units next to each measurement. Always round off just your final answer using appropriate significant figures that rely on your **original data**- in this case 3 digits!)*

- 4) Calculate the density of your object using the average volume you calculated above and the average mass you found in part a.
- 5) The true density for aluminum is **2.702 g/cm³**. Calculate the error and percent error for your experimentally determined densities using the formulas below.

$$\text{Error} = | \text{Experimental Value} - \text{True Value} |$$

$$\% \text{ Error} = \frac{\text{Error}}{\text{True Value}} \times 100$$

PART C: MEASURING VOLUME BY DISPLACEMENT

OBJECTIVE C

To measure the volume of a cylinder by direct displacement and by Archimedes' principle

BACKGROUND C

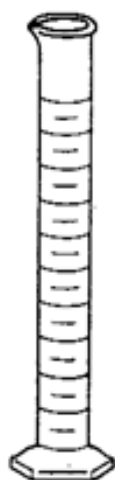
A solid object submerged in a liquid will displace a volume of liquid equal to the volume of the object. Displacement is a very common sense concept and explains why you wouldn't fill your bathtub up to the very top- it would obviously spill over when you got in it. In fact if you measured the amount of water that spilled out of your full tub when you carefully submerged yourself in it, that volume would be equal to the volume of your body! So, this is a very easy way to determine the volume of an irregular shaped object- like your body or a regular shaped object like a cylinder. If some liquid is in a graduated cylinder and you drop an object into the cylinder, the increase in the height of the liquid surface is a direct measure of the volume of the object.

$$\text{Volume of cylinder} = \text{Volume}_{\text{after submerging cylinder}} - \text{Volume}_{\text{before submerging cylinder}}$$

The method is only as precise as the precision of the graduated cylinder, and a 50 mL cylinder is usually not precise to more than ± 0.5 milliliter. This method, therefore, chiefly serves to get an estimate of the volume of irregular objects, or to check the arithmetic on calculations of volume based on the dimensional measurements.

PROCEDURE C

Add some deionized water to a clean 50 mL graduated cylinder and record the initial volume to the tenth place. Read the volume with your eye level with the meniscus, or curved surface of the water and record it in ink in your notebook. It is typical to read the bottom surface of the meniscus (see Figure 1C). Note that the error in the graduated cylinder is ± 0.5 mL (which means that you would record a 0 or a 5 in the tenths place for it to be a proper measurement). Take your cylinder and drop it gently into the graduated cylinder. Shake any bubbles out and record the volume of the water with the object submerged. The difference between the final and initial volumes is the volume of liquid displaced. This is a measure of the volume of the object.



Graduated Cylinder

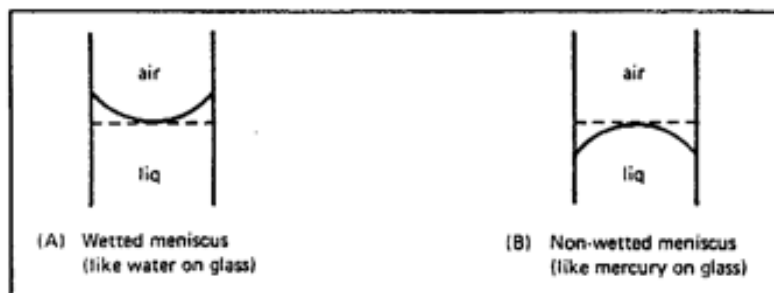


Figure 1C

Sample Data Table for part 1C:

Copy this format into your notebook

Initial volume in the graduated cylinder: _____

Final volume in the graduated cylinder: _____

CALCULATIONS C

- 1) Calculate the volume of the object to the correct number of significant figures by finding the amount of water displaced.
- 2) Calculate the density of your object using the volume you found above and the average mass of the object from part A.
- 3) Calculate the error and percent error for your experimentally determined density just like you did in part B.

PART D: MEASURING VOLUME BY ARCHIMEDES' PRINCIPLE

OBJECTIVE D

To measure the volume of a cylinder by Archimedes' (buoyancy) principle.

BACKGROUND D

A more precise determination of volume can be made using a balance and Archimedes' principle. Archimedes' principle states that when an object is immersed in a liquid, the liquid exerts an upward force known as the buoyancy force. This buoyancy force is equal to the weight of the liquid displaced (Figure 1D).

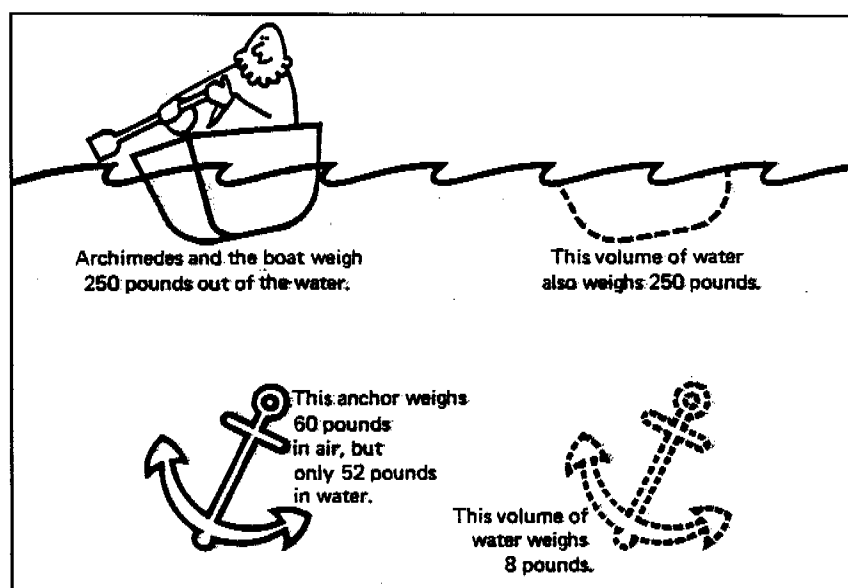


Figure 1D

Imagine that the anchor in the figure above is your cylinder. The buoyancy force can be calculated by taking the difference between the weight of the cylinder in air and its weight when suspended fully submerged in the liquid ($\text{weight}_{\text{in air}} - \text{weight}_{\text{in water}}$). If the buoyancy force on the cylinder is measured, that will tell us the **mass** of the amount of water displaced by the cylinder. If the liquid is pure water, its **density** can be looked up (at the relevant temperature). This allows us to solve for the cylinder's volume using the formula $d = m/V$, rearranged to $V = m/d$.

$$\text{Mass of water displaced} = \text{buoyancy force} = \text{weight}_{\text{in air}} - \text{weight}_{\text{in water}}$$

$$\text{Volume of water displaced} = \text{Volume}_{\text{cylinder}} \quad (\text{The water is displaced by the cylinder, therefore the volume of the water displaced is the same as the volume of the cylinder.})$$

$$\text{Density of water displaced (density}_{\text{water}}) = \frac{\text{Mass of water displaced (buoyancy force)}}{\text{Volume of water displaced (volume}_{\text{cylinder}})}$$



$$\text{density}_{\text{water}} = \frac{\text{buoyancy force}}{\text{volume}_{\text{cylinder}}}$$



$\text{volume}_{\text{cylinder}} = \frac{\text{buoyancy force}}{\text{density}_{\text{water}}}$

Equation to find
volume Part D

PROCEDURE D

To measure the buoyancy force on your cylinder you will need to use the coiled wire hanging from the bottom of the balance. (Do not remove it or pull on it. See your instructor if you are having difficulties.) Tare (zero) the balance with the wire coil hanging freely. Prepare your 150 mL beaker so that it is about 2/3 filled with deionized water and place it on the jack directly below the coil on the balance. Place your cylinder in the coil and completely submerge the coil-cylinder pair in the beaker of water. Be careful to insure that the coil/cylinder is not touching the bottom or sides of the beaker. You can use the jack to adjust the height of your beaker. (Is it necessary to know the weight of the beaker of water for this problem?) Measure the apparent mass of the submerged object (How should this compare to the same object's mass in air?) Record the measurement in your notebook. Return with your beaker and the cylinder to your bench. Obtain a thermometer and determine the temperature of the water in your beaker to the nearest tenth °C (one place after decimal) and record it in your data table. Using table 1 below, look up the density of water at this temperature and record it.

Data Table for part 1D

For this part of the experiment and most future ones you will be creating your own data table. Use the examples given in the previous parts to help guide you.

CALCULATIONS D

- 1) Calculate the volume of the cylinder by using the mass of the water displaced (buoyancy effect) and the density of the water from Table 1. Hint: Use the “equation to find volume” on the previous page.
- 2) Using the volume determined above and the average mass in air from part A, calculate the density of the cylinder.
- 3) Calculate error and percent error for your experimentally determined density of your cylinder.

Table 1: Density of Water (in g/cm³) at Temperatures 18 to 24°C in 0.1°C increments.

⁰ C	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9
18	0.99860	0.99858	0.99856	0.99854	0.99852	0.99850	0.99848	0.99846	0.99844	0.99842
19	0.99840	0.99838	0.99836	0.99834	0.99833	0.99831	0.99829	0.99827	0.99824	0.99822
20	0.99820	0.99818	0.99816	0.99814	0.99812	0.99810	0.99808	0.99806	0.99804	0.99801
21	0.99800	0.99797	0.99795	0.99793	0.99790	0.99788	0.99786	0.99784	0.99782	0.99779
22	0.99777	0.99775	0.99772	0.99770	0.99768	0.99766	0.99763	0.99761	0.99758	0.99756
23	0.99754	0.99751	0.99749	0.99747	0.99744	0.99742	0.99739	0.99737	0.99734	0.99732
24	0.99730	0.99727	0.99725	0.99722	0.99720	0.99717	0.99715	0.99712	0.99710	0.99707

POST-LAB

Results: Prepare a summary results table (see below) containing your results from all four parts of the experiment.

	Mass	Volume	Density	Error	% Error
Part A		-----	-----	-----	-----
Part B	-----				
Part C	-----				
Part D	-----				

Discussion: Write a thoughtful discussion that analyzes the results and addresses these questions.

- a. How do the different values for the volume of your cylinder compare? Which do you think is the most accurate? Why?
- b. How do the different values for density compare? Which do you think is the most accurate and why?
- c. What were some experimental sources of error that could have led to an increase or a decrease in the calculated density? **Note that carelessness and mistakes are not sources of error.** Errors are those that a conscientious worker cannot avoid with the available measuring devices.

(This section should show whether or not you understood the experiment.)

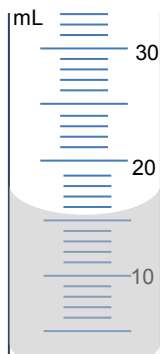
What you need to know for Experiment 1 Quiz

- Reading the volume of liquid in a graduated cylinder
- Extensive/intensive properties
- Calculating density/ error/% error/ average
- How different sources of error would affect calculated density

Sample Quiz 2

1. The following is a graduated cylinder which has an error of ± 0.5 mL:

What is the volume of liquid shown? _____



2. State whether each of the following is an extensive or intensive property:

Temperature _____ mass _____

3. Find the average of the three trials: 5.683 g, 5.681 g, 5.688 g

4. What is the error in density for a sample if the experimental density was 2.13 g/cm^3 and the true value is 2.437 g/cm^3 ? _____

5. A student sets out to determine the density of a metal sample just like you did in experiment 1. To measure the volume, he immersed his sample in a graduated cylinder containing water and measured the volume of the displaced liquid. If some water splashed out when he placed his sample in the cylinder, how would his calculated density compare to the true density of his sample. Choose from higher, lower, no change, can't tell.

Experiment 2

Preparation and Measurement of the Density of Solutions

PRE-LAB

Objective: Write a short statement describing the purpose of the experiment.

Background: Define the terms solution, solvent, solute, and aqueous solution.

Procedure: Briefly write/outline what you will be doing in the experiment.

Prelab Calculations: Perform the two calculations asked for on page 40.

Data: Set up the data table for all the information you will be collecting in this experiment.

OBJECTIVE

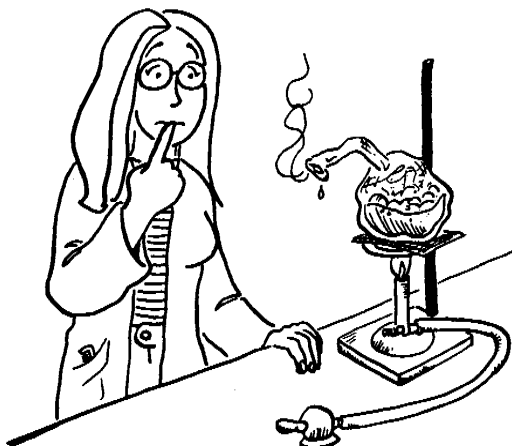
To prepare and measure the density of a 10.0 percent composition (by mass) salt solution and a 40.0 percent composition (by volume) alcohol solution.

BACKGROUND

A **solution** is a homogeneous mixture made up of two or more pure substances that has all of these components thoroughly mixed at the molecular level. The component present in the largest amount is called the **solvent** and the component(s) present in the smaller amount(s) is known as the **solute(s)**. When solids are dispersed in a liquid the solids are usually the solutes and the liquid is the solvent. Many of the solutions we will work with use water as the solvent and are called **aqueous** solutions. They can be represented by writing the formula of the solute followed by (*aq*). For example, sodium chloride dissolved in water would be written $\text{NaCl}(\text{aq})$. When the solvent of a solution is not specified, the solvent is presumed to be water.

Another property of a solution is that its components can be mixed in varying amounts. This creates solutions of different concentration. Addition of a lot of solute to a fixed amount of solvent gives a more concentrated solution while a small amount of solute gives a more dilute solution. The composition of the mixture may be quantitated by specifying the ratio of the mass of one component to the total mass of the solution (% composition by mass or % m/m). The composition may also be quantitated by the ratio of the volume of one of the components to the sum of the volumes (% composition by volume or % v/v).

As we learned in experiment 1, the density of a pure substance is a physical property that is independent of the amount of the substance we have. Each pure substance has a characteristic density value. For example, we saw that the density of water is approximately 1.0 g/mL (remember that density varies slightly with temperature). When substances with a density higher than 1.0 g/mL are added to water the density of the resultant solution is greater than 1.0 g/mL. Conversely, adding substances with a density lower than 1.0 g/mL to water decreases the density of the resultant solution to less than 1.0 g/mL. In this experiment we will prepare aqueous solutions containing either sodium chloride or ethanol as the solute and then measure their density using the buoyancy concept of Archimedes' principle.



Volumetric flasks are NEVER heated!

To prepare a given solution we must know the pure substances to be used and how much of the solution to prepare. For example let's say we want to make 100.0 g of a 10.0% (by mass) salt solution [NaCl(aq)]. Masses are always additive, so the sum of the masses of the components is the mass of the solution. Since % represents parts out of 100, in the case of a 10.0% solution, 10.0 mass parts out of every 100.0 mass parts of solution is solute (salt) and the rest, 90.0 mass parts, is water. The following statement can be made for a 10.0% NaCl(aq):

10.0 g salt $\equiv$ 90.0 g water $\equiv$ 100.0 g solution,

where $\equiv$ means equivalent to. Conversion factors can be constructed from any combination of two of these terms (see example 1 on the next page).

In general we cannot say that volumes of different liquids are additive. Therefore volumes of unlike substances do not give a volume of solution which equals the sum of the volumes of the components. A 40.0 % (by volume) alcohol solution means 40.0 volume parts alcohol and the rest, 60.0 volume parts, is water. The volume of the solution is not necessarily 100 volume parts. The following statement can be made for a 40.0% alcohol solution:

40.0 mL alcohol $\equiv$ 60.0 mL water

Conversion factors can be constructed from any combination of these terms (see example 2 on the next page).

Percent composition, like density, is an intensive property since the amount of the solution doesn't affect its % composition.

In order to prepare a solution one must first decide what quantity of solution is needed, and then compute the quantities of components to be measured. Below are examples of how calculations could be made for preparing solutions of each type of % composition- mass and volume.

EXAMPLE 1: % composition by mass

Prepare 80.0 g of a 10.0% (by mass) salt solution ($\text{NaCl}(aq)$). How much salt and how much water are needed?

Given: g of salt solution
 Want: g of salt
 Road map: g of salt solution $\rightarrow$ g of salt
 Use conversion factor:
 10.0 g salt/100 g solution

$$\text{Calculation: } 80.0 \text{ g of solution} \times \frac{10.0 \text{ g salt}}{100 \text{ g solution}} = 8.00 \text{ g salt}$$

The remainder is water: $80.0 \text{ g solution} - 8.00 \text{ g salt} = 72.0 \text{ g water}$

There is no way to prepare a specified volume of a solution by adding volumes of unlike substances, like we did above with masses, unless the densities of the solution and the components are known. The volume of one of the components is specified and the volume of the other is computed from its relationship to the one specified.

EXAMPLE 2: % composition by volume

Prepare a 40.0% (by volume) alcohol solution ($\text{C}_2\text{H}_6\text{O}(aq)$) containing 25.0 mL of alcohol. How much alcohol and how much water are needed?

Given: mL of alcohol solution
 Want: mL of water
 Road map: mL of alcohol solution $\rightarrow$ mL of water
 Use conversion factor:
 60.0 mL water/40.0 mL alcohol

$$\text{Calculation: } 25.0 \text{ mL of alcohol solution} \times \frac{60.0 \text{ mL water}}{40.0 \text{ mL alcohol solution}} = 37.5 \text{ mL water}$$

After measuring out **37.5 mL of water** in your cylinder, you would add **25.0 mL of alcohol** to it.

The densities of the solutions could be found by measuring the mass and the volume of each solution. If by chance an accurate measuring device for volume is not available as is the case here, we can use the buoyancy law to calculate the density. Remember that the buoyancy law says that the difference between an object's weight in air and its weight when it is submerged in a liquid is the weight of the liquid displaced.

We have an aluminum cylinder whose accurate volume can be calculated from its actual density value of 2.702 g/cm^3 and its mass in air. If we know the mass of solution that the cylinder displaces (which is the buoyancy force) we can calculate the solution's density.

$$d_{\text{soln}} = \frac{\text{weight of the solution displaced by cylinder (buoyancy force)}}{\text{volume of solution displaced by cylinder (volume of cylinder)}}$$

Note from experiment 1: buoyancy force = mass in air – mass in solution

PROCEDURE

Salt Solution: Prepare 100.0 grams of a 10.0 percent, by mass, salt solution. Put on your protective eyewear. (You and your lab partner may prepare one solution to share between you.) Tare a 150 mL beaker on the balance- this means putting your dry beaker on the balance and pressing the tare knob. Wait for the balance to read 0.000 g. Remove the beaker and add sodium chloride to the beaker until you reach the calculated amount (remember to remove the beaker from the balance when adding any chemical- the balance will retain its tared memory). Once you get close to 10.0 g on the balance, (its not essential to get 10.000 g since you only need 3 significant figures), return with your beaker and salt to your bench. So that you don't have to weigh the mass of the water on the balance also, assume that water has a density of 1.0 g/mL , and measure the calculated mass of deionized water using your graduated cylinder. Pour it into the same beaker. Stir the mixture until the salt is completely dissolved and this will give you a 10.0% by mass salt solution Label the beaker as NaCl solution and go on to prepare the solution in Part B.

Alcohol Solution: Prepare a 40.0 percent, by volume, alcohol solution containing 30.0 milliliters of alcohol. (You and your lab partner may prepare one solution to share between you). The volumes of alcohol and deionized water may be measured separately in a graduated cylinder. (The alcohol solution can be found in a hood.) Mix the liquids in your 150 mL beaker and label it as alcohol solution.

Once you and your partner have made the solutions, take both solutions, each of your aluminum cylinders, each of your lab notebooks, and pens to a free balance. Tare the balance and place your aluminum cylinder in the coiled wire below the balance (gently dry the wire with a paper towel if it is wet). Wait for the balance to stabilize and record its mass in your notebook. This is the mass of your cylinder in air. Now place your beaker with NaCl solution under the balance and completely submerge the coil-cylinder pair in the beaker of NaCl solution. Be careful to insure that the coil is not touching the bottom or sides of the beaker. After the balance reading stabilizes, record it in your notebook. This is the mass of your cylinder in NaCl solution. Replace the beaker containing NaCl solution with the beaker containing the alcohol solution and determine the submerged mass of your cylinder again. Record it in your notebook. Allow your lab partner to repeat the same three measurements using their own cylinder. Using the determined masses, complete calculations 1-5 below. You may use your notes from the calculations in experiment 1 part D as a guide to help you calculate the buoyancy effect on your cylinder when submerged in the solution.

Disposal: The salt solution can be poured down the drain and the alcohol solution should be poured into the waste container in the hood marked "alcohol solution".

PRELAB CALCULATIONS

Carefully read the background section of this experiment and use the examples to guide you through the following two calculations.

- 1) What mass of salt and what mass of water are needed to prepare 100.0 grams of 10.0 percent by mass salt solution?

Hint: remember that a 10.0 percent by mass salt solution = 10.0 g salt / 100.0 grams of solution. Use this mass percent as a conversion factor for the following solution map:

mass solution $\rightarrow$ mass salt

Remember, the solution contains both water *and* salt!

- 2) What volume of water must be added to 30.0 milliliters of alcohol to obtain a 40.0 percent (by volume) solution of alcohol?

To get started, write the percent by volume as a conversion factor between volume of alcohol and volume of water, as shown in the background section.

IN-LAB

Data: Complete the data table you prepared.

Calculations: Perform the calculations asked for below.

CALCULATIONS

- 1) Calculate the volume of the aluminum cylinder you used from the actual density of aluminum (2.702 g/mL) and its mass in air.
- 2) a) Calculate the buoyancy force on the cylinder due to the salt solution.
b) Determine the experimental density of the salt solution from the buoyancy force and the volume of the aluminum cylinder (determined in 1) above).
- 3) According to the CRC Handbook of Chemistry and Physics the actual density of a 10.0% salt solution at 25°C is 1.0726 g/mL. Calculate the error and % error of your calculated density for the salt solution.
- 4) a) Calculate the buoyancy force on the cylinder due to the alcohol solution.
b) Determine the experimental density of the alcohol solution from the buoyancy force and the volume of the aluminum cylinder (determined in 1) above).
- 5) According to the CRC Handbook of Chemistry and Physics the actual density of 40.0% (v/v) alcohol solution at 25°C is 0.926 g/mL. Calculate the error and % error of your calculated density for the alcohol solution.

Note: You should be able to get errors of less than 10% for both solutions in this experiment. If your values are higher, repeat the measurement(s) for the mass of the cylinder in the solution(s) and recalculate.

POST-LAB

Results: Create a summary results table

Conclusion: Analyze the results and errors in a written discussion. Include any sources of error that could have led to an increase or a decrease in the calculated density values for the solutions.

What you need to know for Experiment 2 Quiz

- Definition of the terms solution, solvent, solute, and aqueous solution.
- Calculating mass (or volume) of solute/solvent needed given percent composition.
- Determining density from buoyancy force and volume of cylinder

Sample Quiz 3

1. Fill in the blank with one of the following terms: solution, solvent, solute, aqueous solution.

The component present in the smaller amount is the _____

2. How many mL of alcohol and how many mL of water are needed to prepare a 35.0% alcohol solution containing 15.0 mL alcohol.

3. Use the following data to answer the questions below: A salt solution was prepared and a 2.45 cm^3 aluminum cylinder was submerged in the solution. The mass of the cylinder in air was 7.325 g and its submerged mass was 3.768 g.

a. What is the buoyancy force on the cylinder? _____

b. What is the density of a salt solution? Include units _____

Experiment 3

Solubility and Chemical Change

PRE-LAB

For this experiment, prepare your notebook in the following manner:

On the first page:

Objective: Write a brief summary of the three types of problems that you will address in this experiment.

Background: Write brief definitions for each of the following terms that we will be using in this experiment: **solution, miscible, immiscible, soluble, insoluble, homogeneous mixture, heterogeneous mixture, reaction, precipitate, exothermic, and endothermic.**

On the next two pages:

Turn your notebook horizontally and divide it into three equally sized columns. Label the columns with the following headings: **procedures, observations and conclusions.** Write a **brief summary** of the procedure in your lab notebook for each experimental part in the procedure column **before you come to lab.** You will enter observations and conclusions in the other columns when you do the experiment. Spread this table over **two pages** so that you have plenty of room to write. The terms you defined in the background section will be used in the conclusion section.

OBJECTIVE

1. Determine the miscibility of different liquids and their relative densities.
2. Determine whether or not a solid will dissolve in water.
3. Determine whether a chemical reaction occurs when two solutions are mixed.

BACKGROUND

Part 1: Mixing two different liquids.

When two different liquids are mixed they may form a homogeneous mixture or **solution**. These liquids are said to be **miscible** resulting in one visible layer or **phase**. Some liquids are miscible in all proportions while others are only partly soluble in each other. When two different liquids do not mix (or mix very little) with each other they will form two distinct layers or phases and are said to be **immiscible**. The phase having the lower density will always float on top of the other. When the two immiscible liquids are both colorless, it may be difficult to determine which phase is which liquid. One can add a few drops more of one liquid and see which phase gets larger in size.

Part 2: Mixing a solid in water.

Some ionic compounds will dissolve in water and others will not. When a small amount of a solid is added to and dissolves in a particular liquid it forms a homogeneous single phase and is said to be **soluble** in the liquid. If the solid is not soluble, or **insoluble**, it will form a cloudy

heterogeneous mixture from which solid particles will separate out, forming two phases. All mixtures should be mixed well before determining whether all of the solid can indeed dissolve. Solubility tables for ionic compounds in water can be used to confirm your results.

Part 3: Is it a reaction or not?

If two different aqueous solutions are mixed several different observations can be made. A number of these observations are evidence that a new chemical(s) has been formed. This is called a chemical change or **reaction**. One way in which a reaction can be confirmed is when the mixing of two solutions produces an insoluble solid product. This results in a cloudy, heterogeneous mixture and the solid formed is called a **precipitate**. A solubility chart (see Table 3.1 on page 44) lists ionic compounds that are soluble (dissolve in water) and insoluble (do not dissolve in water).

Other chemical reactions result in a gaseous product being formed that escapes as bubbles from the solution. There are also some reactions in which dissolved components combine to produce colored substances. These reactions can be recognized by a color change. Sometimes no observed change takes place, generally indicating that no reaction has occurred.

If there is a temperature change, you can describe the processes or reactions that occur in any of the parts above using the terms exothermic or endothermic. The term **exothermic** describes a process or reaction that releases energy, usually in the form of heat; while the term **endothermic** describes processes or reactions that absorb heat energy.

IN-LAB

Fill in the observations and conclusions columns for each part by answering the questions asked for under the procedures section. Remember to use the terms you defined in the background section. The answers to the prompting questions in the procedure should go in the **conclusions** section.

Observations vs. Conclusions:

Part of what you will learn in this experiment is how to distinguish between writing observations and conclusions. **Observations** are descriptions of what you observe, that is see, feel, smell, or hear. Your observations should be written in simple, non-scientific language. **Conclusions** are statements describing what you are deducing or inferring from your observations. Here you should **use the appropriate scientific language**, including the new terms you are introduced to in this experiment.

For example:

Observation 1: after mixing the chemicals the test tube became warm

Conclusion 1: the process occurring gave off heat and was therefore **exothermic**

Observation 2: after mixing the 2 liquids well, two separate layers (phases) were visible

Conclusion 2: these two liquids formed a heterogeneous mixture and are therefore **immiscible**

PROCEDURE

Wear your protective eyewear throughout the experiment.

EXPERIMENTAL TIPS

Use of Reagent Droppers Bottles: To avoid contamination when dispensing chemicals using droppers, never touch the end of the dropper to a container. This may contaminate the end of the dropper.

Quantities: This is a qualitative experiment, meaning that the amounts of the chemicals you use are approximate and need **not** be measured carefully. For all parts of this experiment, make sure that you **mix the chemicals well** before making your observations

Disposal of Chemicals: You will be using a variety of different chemicals in this experiment. Years ago chemicals of all sorts were disposed of down the drain or in the trash, without much concern about their affect on us or the environment. Today this is not the case- there is much more public consciousness of our safety and the hazards of chemicals in the environment are regulated by many laws and statutes. Most chemicals must be collected after use and disposed of in the proper means. Learning about the proper disposal of chemicals is an important part of this laboratory experience. This lab book will indicate how to dispose of used chemicals so follow them carefully. If you still have questions, ask your instructor.

Part 1. Mixing Different Liquids:

a. Is hexane (C_6H_{14}) miscible (does it mix) with water?

Place about a **half** inch of deionized water from your wash bottle into a small test tube and add about a **quarter** of an inch of hexane (found in the hood). Mix well. What do you observe? Did the water end up on the top or the bottom? (Since you added more water than hexane, the layer with the greater volume is the water.) If you can't tell, add some more water and see into which layer it goes.

Empty the liquids into the red container in the hood marked “organic waste”.

b. Is methanol (CH_3OH) miscible with water?

Place about a half inch of deionized water in a test tube and add about a quarter of an inch of methanol. Mix as before. What do you observe? Is there any way of telling which has the greater density? Try dripping some water down the side of the tube and see if it layers on top or sinks to the bottom to determine relative densities.

Empty the liquids into the red container in the hood marked “organic waste”.

c. Is methanol (CH_3OH) miscible with hexane (C_6H_{14})?

Place about a half inch of methanol in a test tube and add about a quarter of an inch of hexane. Mix as before. Let it stand for a few minutes. Are there two phases? Which phase is which chemical?

Empty the liquids into the red container in the hood marked “organic waste”.

Part 2: Dissolving a Solid in Water:

a. Is solid barium chloride (BaCl_2) soluble in water?

Place a **very small** amount (about the size of a match head) of barium chloride dihydrate in a small, **dry** test tube and add about an inch of deionized water from your wash bottle. Mix well. Did the solid dissolve or not? **Save this for later use.**

b. Is sodium sulfate (Na_2SO_4) solid soluble in water?

Place a **very small** amount (about the size of a match head) of sodium sulfate in a small, **dry** test tube and add about an inch of deionized water. Mix as before. Did the solid dissolve or not? **Save this for later use.**

c. Is barium sulfate (BaSO_4) solid soluble in water?

Place a **small** amount of barium sulfate in a small, **dry** test tube and add about an inch of deionized water. Mix as before. Is the mixture cloudy? Did the barium sulfate dissolve?

d. Is there a reaction between barium chloride (BaCl_2) and sodium sulfate (Na_2SO_4)?

Now mix the barium chloride solution from (a) and the sodium sulfate solution from (b) together. Let it stand for a few minutes before making observations. Does this mixture look like any of the mixtures of solid and water that you just prepared? What do you conclude? Did a chemical change occur? If so, can you suggest what one of the products might be? **Empty all solutions from part 2 into the container in the hood marked "aqueous waste."**

Part 3: Mixing Aqueous Solutions: (Carry out these procedures in a clean spot plate.)

a. Is there a reaction between potassium nitrate and sodium chloride solutions?

Add a few drops of potassium nitrate (KNO_3) solution into a clean well of your spot plate. Add an equal amount of sodium chloride (NaCl) solution and mix using your stirring rod. Does any visible reaction occur? Record your observations.

b. Is there a reaction between iron(III) chloride and potassium thiocyanate solutions?

Into another clean well on your spot plate, add a few drops of iron(III) chloride (FeCl_3) solution and potassium thiocyanate (KSCN) solution and mix. What happens?

c. Is there a reaction between iron(III) chloride and sodium hydroxide solutions?

Place several drops of iron(III) chloride solution into a clean well and then add sodium hydroxide solution drop by drop. Mix and observe what happens?

d. Is there a reaction between sodium carbonate solution and hydrochloric acid?

Place several drops of sodium carbonate (Na_2CO_3) solution into a clean well and then add dilute hydrochloric acid [$\text{HCl}(aq)$] found on the bench top drop by drop, with mixing. Watch the solution carefully as each drop of acid is added and as you mix. What happens? Is there a chemical reaction? What type of product (physical state) is formed?

e. Is there a reaction between copper(II) sulfate and ammonia solutions?

Add a small amount of the dilute ammonia (NH_3) solution found on the bench top drop by drop, while mixing, to a well containing copper(II) sulfate (CuSO_4) solution. What happens? Is there a chemical reaction?

Repeat any reactions that you want to observe again.

Empty all solutions from your spot plate into the container in the hood marked "aqueous waste."

POST-LAB

Write a discussion analyzing the results you obtained in each part. Include the following:

For Part 1:

- Indicate the relative densities of the 3 liquids: water, hexane and methanol as determined from the experiment.
- Look up the actual densities of these 3 liquids (in g/mL) by “googling” them on the Internet and write them in your report (write the reference site’s URL in each case). Use them to **confirm or deny** the results you obtained and discuss this.

For Part 2:

- Compare the solubility results you obtained in this experiment with solubility table 3.1. Your instructor will review this with you.

<u>Compounds containing these ions are soluble</u>	<u>Exceptions making compound insoluble</u>
$\text{Li}^+, \text{Na}^+, \text{K}^+, \text{NH}_4^+, \text{NO}_3^-, \text{C}_2\text{H}_3\text{O}_2^-$	No exceptions
$\text{Cl}^-, \text{Br}^-, \text{I}^-$	Any of these ions paired with $\text{Ag}^+, \text{Hg}_2^{2+}, \text{Pb}^{2+}$
SO_4^{2-}	SO_4^{2-} paired with $\text{Sr}^{2+}, \text{Ba}^{2+}, \text{Pb}^{2+}, \text{Ca}^{2+}$
<u>Compounds containing these ions are insoluble</u>	<u>Exceptions making compound soluble</u>
$\text{OH}^-, \text{S}^{2-}, \text{CO}_3^{2-}, \text{PO}_4^{3-}$	S^{2-} or OH^- paired with $\text{Sr}^{2+}, \text{Ba}^{2+},$ or Ca^{2+} Any of these ions with $\text{Li}^+, \text{Na}^+, \text{K}^+, \text{NH}_4^+$

Table 3.1: Solubility Chart

For Part 3:

- Summarize your results. What was the specific evidence that suggested that a chemical reaction had occurred in each case?

What you need to know for Experiment 3 Quiz

- Definition of the terms solution, miscible, immiscible, soluble, insoluble, homogeneous mixture, heterogeneous mixture, reaction, precipitate, exothermic, and endothermic and applying them to given scenarios.
- Use of the solubility chart

Sample Quiz 4

Choose the correct term(s)

1. Answer the questions below based on the following: A student mixes water and hexane and observes two phases forming.
 - a. Based on his results the two liquids are
 - i. miscible
 - ii. immiscible
 - iii. soluble
 - iv. insoluble
 - b. the two substances form a
 - i. homogeneous mixture
 - j. heterogeneous mixture
2. A student mixed two substances together in a test tube and observed that the test tube became warm. The process that occurred inside of the test tube was
 - a. Exothermic
 - b. Endothermic
3. Based on the solubility chart below, determine whether each of the following substances will be soluble or insoluble in water:
 - a. $(\text{NH}_4)_2\text{CO}_3$
 - b. AgCl

<u>Compounds containing these ions are soluble</u>	<u>Exceptions making compound insoluble</u>
$\text{Li}^+, \text{Na}^+, \text{K}^+, \text{NH}_4^+, \text{NO}_3^-, \text{C}_2\text{H}_3\text{O}_2^-$	No exceptions
$\text{Cl}^-, \text{Br}^-, \text{I}^-$	Any of these ions paired with $\text{Ag}^+, \text{Hg}_2^{2+}, \text{Pb}^{2+}$
SO_4^{2-}	SO_4^{2-} paired with $\text{Sr}^{2+}, \text{Ba}^{2+}, \text{Pb}^{2+}, \text{Ca}^{2+}$
<u>Compounds containing these ions are insoluble</u>	<u>Exceptions making compound soluble</u>
$\text{OH}^-, \text{S}^{2-}, \text{CO}_3^{2-}, \text{PO}_4^{3-}$	S^{2-} or OH^- paired with $\text{Sr}^{2+}, \text{Ba}^{2+}$, or Ca^{2+} Any of these ions with $\text{Li}^+, \text{Na}^+, \text{K}^+, \text{NH}_4^+$

Experiment 4

Separation of Components of Mixtures

PRE-LAB

Objective: Write a short statement describing the purpose of the experiment.

Background: Define the terms filtration, extraction, crystallization, and evaporation.

Procedure: Briefly write/outline what you will be doing in the experiment.

Data: Set up a data table for this experiment.

OBJECTIVE

To separate a two component heterogeneous mixture based upon the differing properties of its components.

BACKGROUND

Most of the materials we encounter in everyday life such as cement, vegetable soup, natural waters, wood, pizza, and soil are mixtures. A mixture is a blend of two or more different substances. Sometimes the blending is so uniform and complete that you can't see any differences in the composition of the mixture, even with a microscope. These are called homogeneous mixtures or solutions. They are composed of a solvent, the component found in the largest quantity, and one or more solutes, those components found in smaller quantities. Some other mixtures are obviously made of different substances because you can see variations in the properties from one region to another. These are called heterogeneous mixtures. Each separate region where the properties are uniform is called a phase. A heterogeneous mixture therefore has more than one phase present.

A mixture must be heterogeneous at the time a separation of the components is attempted. If the phases are solid and large enough one could separate them with tweezers. When panning for gold, we use differences in density to separate gold from the gravel. Iron oxide is separated from rock by using the magnetic properties of iron oxide. In all of these cases we use differences in physical properties to separate different phases from each other. In the case of homogeneous mixtures, we must change the conditions so as to create two phases so they can be separated from each other.

A large part of chemical research and industrial production is devoted to finding new ways to separate minute quantities of substances from each other. Some of these methods are described as follows:

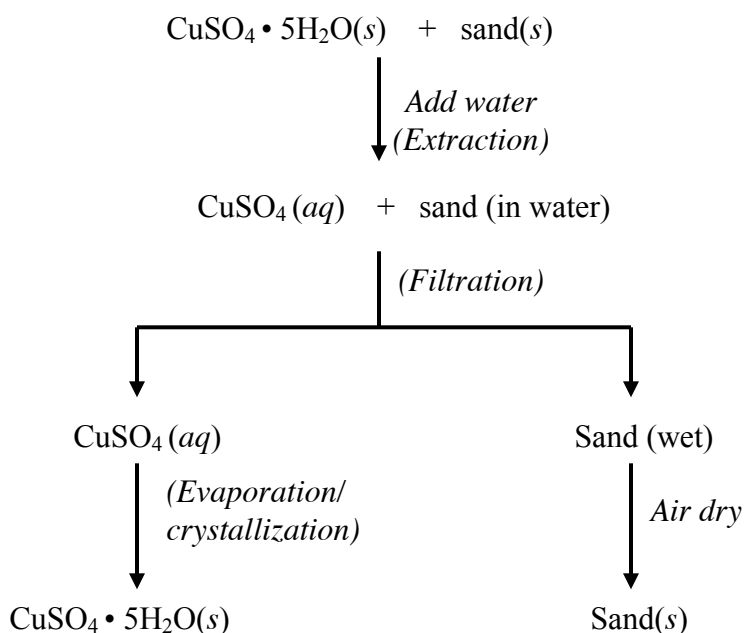
Filtration: A solid may be separated from a liquid or a gas by straining the fluid through a porous medium of some kind. If the pores are smaller than the solid phase the liquid passes through and the solid stays behind. The porous medium is usually paper.

Extraction: Soluble substances may be separated from those that are less soluble by dissolving them in a solvent. The solution can then be removed from the insoluble components by filtration or other appropriate means.

Crystallization: A solution of a solid substance may become supersaturated by evaporation of the solvent or by lowering the temperature. Under the right conditions crystals of the pure solute may form, which then can be removed by filtration from the impure solution. Some substances may be purified by crystallization from a solution, because impurities tend to be excluded from a growing crystal.

Evaporation: Is used as a means of purifying a soluble solid from a liquid by evaporating away the liquid.

In this experiment we consider the problem of separating sand from a salt. Two phases are present. (Note: there is only one state, the solid state, but there are two distinct phases). We know copper(II) sulfate is soluble in water and that sand is insoluble. These facts suggest that we might extract the salt from the sand with water and filter the sand from the mixture. Since copper(II) sulfate does not evaporate, we could recover the salt from the solution by evaporation of the solvent water and by crystallization of the salt.



We may then determine the mass of each component and calculate the mass percent composition and percent recovery:

$$\text{Percent Composition of A or B} = \frac{\text{mass of pure component A or B}}{\text{mass of mixture started with}} \times 100$$

The above calculation needs to be done once for the salt and once for the sand

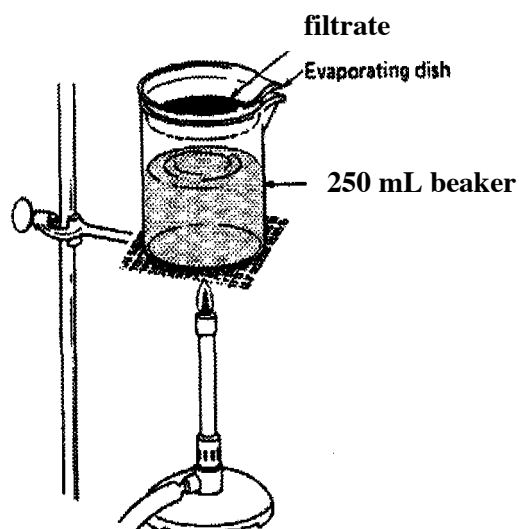
$$\text{Percent recovery} = \frac{\text{sum of mass of both components recovered}}{\text{mass of mixture started with}} \times 100$$

PROCEDURE

Wearing your protective eyewear, tare your 150 mL beaker and add a scoop full of the $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ /sand mixture to it. Record the mass in your notebook. Add 10-15 mL of deionized water to the beaker and swirl the mixture gently. Weigh and record the mass of a piece of clean filter paper. Assemble the filtering apparatus as follows: fold the piece of weighed filter paper in quarters. Open the folded paper between the folded edges so as to make a cone (your instructor will demonstrate). Place the cone-shaped filter in the funnel and wet it with deionized water from your wash bottle. Press out any air pockets between the filter and the funnel. Support your funnel with a ring attached to a support rod (rings can be found in the cabinet under the sink).

Record the mass of your clean, dry evaporating dish. Place it beneath your funnel and lower your funnel so that its stem is touching the inside of the evaporating dish. Slowly pour the contents of the beaker onto the filter. Take care that the level of liquid is always below the top rim of the filter paper. Use your wash bottle and a rubber policeman (attached to your stirring rod) to transfer **all** of the sand into the filter. Be very careful to not disturb the filter paper with your stirring rod as this may tear the filter paper. After all the liquid has drained through the filter, wash the filter with small portions of deionized water until the washings are colorless. Collect all the filtrate and washings in the evaporating dish.

Remove the filter paper from the funnel, open the filter paper containing the sand on a watch glass, write your name on a piece of tape and stick it to the watch glass. Place it in the oven or set it on the shelf assigned to your instructor. Once completely dry, record the mass of the filter paper and sample when they are dry (this may be the next period depending on what your instructor decides).



Prepare a steam bath by placing a 250 mL beaker filled about 2/3 full of tap water, on a wire screen on a 3" ring attached to a support rod. Protect the water bath by placing a large ring on top of the 3" ring. Insure that the rod and ring are tightly clamped and that your Bunsen burner flame is at the proper height. Carefully place the 250 mL beaker on the wire screen. Place the evaporating dish carefully on the beaker and heat the water to boiling. Heat the steam bath until the filtrate has completely evaporated. Do not let the steam bath boil to dryness- add more water if necessary. Once all the water has evaporated from the evaporating dish, let the set up cool completely and then carefully remove your evaporating dish. Once it has cooled to room temperature, dry the bottom of your evaporating dish containing the sample and determine the mass.

Disposal: After confirming that your calculations are accurate, the $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ solid can be disposed of in the labeled container and the sand is disposed in the trash.

IN-LAB

Data: Fill out the data table you prepared.

Calculations: Perform the calculations asked for below.

CALCULATIONS

- 1) Calculate the mass of the:
 - a) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ recovered. Remember to subtract the mass of the empty evaporating dish.
 - b) Sand recovered. Remember to subtract the mass of the clean filter paper.
- 2) Calculate the % composition, by mass, of:
 - a) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in the recovered mixture.
 - b) Sand in the recovered mixture.
- 3) Calculate the % recovery of the original mixture.

POST-LAB

Results: Create a summary results table

Conclusion: Analyze the results and errors in a written discussion. What could be the sources of error and how could these sources of error have affected your calculated percent recovery?

What you need to know for Experiment 4 Quiz

- Definition of the techniques filtration, extraction, crystallization, and evaporation.
- Calculating percent composition and percent recovery

Sample Quiz 5

1. Fill in the blank with one of the following terms: filtration, extraction, crystallization, and evaporation
A technique used to separate a soluble substance from a less soluble one by dissolving in a solvent _____
2. Answer the questions below using the following Information: A student attempts to separate 2.456 g of a sand/salt mixture just like you did in this lab. After carrying out the experiment, she recovers 1.376 g of sand and 0.658 g of salt.
 - a. What was the percent composition of sand in the mixture according to the student's data?

 - b. What was the percent recovery? _____

EXPERIMENT 5

DISTINCTIVE PROPERTIES OF CHEMICALS

PRE-LAB

Objective: Write a short statement describing the purpose of the experiment.

Background: Define the terms pure substance, mixture, element, compound, ionic compound, covalent compound, hydrates, homogeneous mixture, heterogeneous mixtures, transparent, translucent, opaque, malleable, ductile, and brittle. You may need to refer to your textbook or the internet for the definition of some of these terms.

Procedure: Briefly write /outline what you will be doing in the experiment.

Pre-lab assignment: Complete the first column of the chart on pages 60-65 by filling in either the name or chemical formula for each sample. Make sure to include the state of matter when writing a chemical formula. For all pure substances, put a check mark in the correct box to classify as metal, nonmetal, ionic, or covalent. Each sample only has one classification.

OBJECTIVE

This display of chemicals is designed to familiarize you with the identification of samples of elements, compounds, and mixtures by observing their properties and different names and formulas. You will practice making careful observations and learning how to categorize different types of matter. The chemicals have been organized into nine related groups or series. The title of each series indicates what those substances have in common. Compare and contrast the appearance of the substances within each series to see how related chemicals can have different appearances.

BACKGROUND

All matter is made up of chemicals. Matter can be classified into various groupings (see Classification of Matter flowchart on page 57). The first categorization is based upon whether the matter is a **pure substance** or a **mixture**. Pure substances are composed of only one type of chemical and can be **elements** or **compounds**. Elements are pure substances made up of atoms that are the same (they have not truly the same, as isotopes have different numbers of neutrons, but they are considered the same element). They can be further subdivided into metals or non-metals. **Compounds** are pure substances composed of two or more different elements. They can be categorized in different ways depending on the criteria used. If the criterion is the presence or absence of ions, then compounds are divided into **ionic** (those that contain ions) or **covalent** (absence of ions and made up only of non-metals). If the presence of C and H is used

as the criterion, the compounds are divided into **organic** (contains C, H, and possibly other non-metals) or **inorganic** compounds. A **mixture** is composed of two or more pure substances and it can be **homogeneous**, look the same throughout (also called a **solution**), or **heterogeneous**, look visually different. Solutions are comprised of a **solvent** and one or more **solute(s)**. The solvent is the component of the solution in the largest amount, while the solute(s) is (are) those present in smaller amounts.

The **basic unit** of a sample of matter, which relates only to a pure substance, is the smallest physical unit in which that substance exists. The basic unit of an element is an **atom** or, in a few cases, a **molecule**. The basic unit of an ionic compound is a **formula unit**, and that of a covalent compound is a **molecule**. Mixtures do not have a basic unit, as they are not pure. **Hydrates** (also known as salt hydrates) are crystalline solids that contain chemically bonded water in definite proportions. They are considered pure and are true ionic compounds. They will have the symbolism “ $\cdot x\text{H}_2\text{O}$ ” at the end of the formula, where x usually represents a small whole number. You will find more detailed definitions of these various terms and how to distinguish between them in your textbook.

PROCEDURE

Using the tables included at the end of this experiment, write the chemical formula and name of each chemical for a particular series and then complete the remaining columns. Looking carefully at the names and formulas of the chemicals will help you with the classifications. Skip a row and go onto another series. You need not go in number order. Use the following guidelines to fill in each column. Open each container carefully and smell the contents by wafting a sample with your hand to your nose. Close them tightly when done. Do not open the other containers of the other series. Use the Classification of Matter flowchart on page 57

Formula and Name of Chemical: Record the chemical symbol or formula and the complete name of the substance. Be careful to write the formulas correctly.

Physical Properties: State the physical properties of the substance. Indicate the physical state (solid, liquid, or gas) of the substance. Specify whether the substance is **transparent** (transmitting light so images can be clearly seen), **translucent** (allowing light but not detailed images to go through), or **opaque** (not transmitting any light). Do not mistake the container for the substance. If the substance is solid, report whether it is lustrous, like a metal, or not shiny (dull). Deduce from the shape and texture whether the substance can be drawn into a wire (**ductile**) or a sheet (**malleable**), or whether it is **brittle** (a substance that breaks when hammered), crystalline, or powdery. What is its color, if any? For Series 8, Organic Chemicals, **carefully** open each container and, by waving your hand over the top, see if you detect an odor. Describe it. Close the container firmly. Do some of the properties of the pure substance, such as its color, also appear in the solution? Do some substances change color in solution?

Pure Substance or Mixture: From the label on the display identify each substance as a pure substance or a mixture.

What Type of Pure Substance or Mixture: If it is a pure substance, identify it as an element or a compound.

What Type of Element: If it is an element, identify it as either a nonmetal or a metal, depending on what side of the “stairstep” in the periodic table they appear.

What Type of Compound: If it is a compound, identify it as an ionic or a covalent compound.

What type of Mixture: If it is a mixture, indicate whether it is homogeneous or heterogeneous.

Basic Unit: If it is a pure substance, indicate its basic unit. (Note: mixtures **do not** have a basic unit.)

CHEMICAL DISPLAYS

Series 1. IRON

Fe Iron (powder or wire)

Fe₂O₃ Iron(III) oxide

Fe(NO₃)₃ • 9H₂O Iron(III) nitrate
nonahydrate

Fe(NO₃)₃(aq) Iron(III) nitrate solution

Series 2. COPPER

Cu copper (powder, sheet or wire)

Cu₂O copper(I) oxide

CuO copper(II) oxide

CuSO₄ • 5H₂O copper(II) sulfate pentahydrate

CuSO₄(aq) copper(II) sulfate solution

Series 3. MERCURY

Hg mercury (liquid)

HgCl₂ mercury(II) chloride

HgI₂ mercury(II) iodide

Series 4. SULFUR

S sulfur (powder)

SO₂ sulfur dioxide

FeS iron(II) sulfide

FeSO₄ • 7H₂O iron(II) sulfate heptahydrate

Series 5. SILICON and CARBON

Si silicon (crystal)

C carbon (graphite)

SiC silicon carbide (carborundum)

SiO₂ silicon dioxide (quartz)

SiCl₄ silicon tetrachloride

Series 6. IODINE

I₂ iodine (crystal)

KI potassium iodide

KI(aq) potassium iodide solution

I₂ iodine in ethyl alcohol (C₂H₆O)

I₂ iodine in carbon tetrachloride (CCl₄)

(continued on next page)

Series 7. GASES

(Caution: do not remove the stoppers)

Cl₂ chlorine

Br₂ bromine (vapor over liquid)

NO₂ nitrogen dioxide

N₂ nitrogen

NH₃ ammonia

H₂ Hydrogen

He helium

Series 8. ORGANIC CHEMICALS

(Wear your protective eyewear. Open the containers smell carefully and then close them tightly.)

C₆H₁₄ Hexane

HC₂H₃O₂ Acetic acid

C₄H₈O₂ Ethyl acetate

C₃H₆O Acetone

C₂H₆O Ethyl alcohol (Ethanol)

C₁₀H₈O 1-naphthol

C₇H₆O Benzaldehyde

C₈H₈O₂ Methyl Benzoate

C₉H₈O Trans-cinnamaldehyde

Series 9. MIXTURES

Brass

Sand + CuSO₄ • 5H₂O

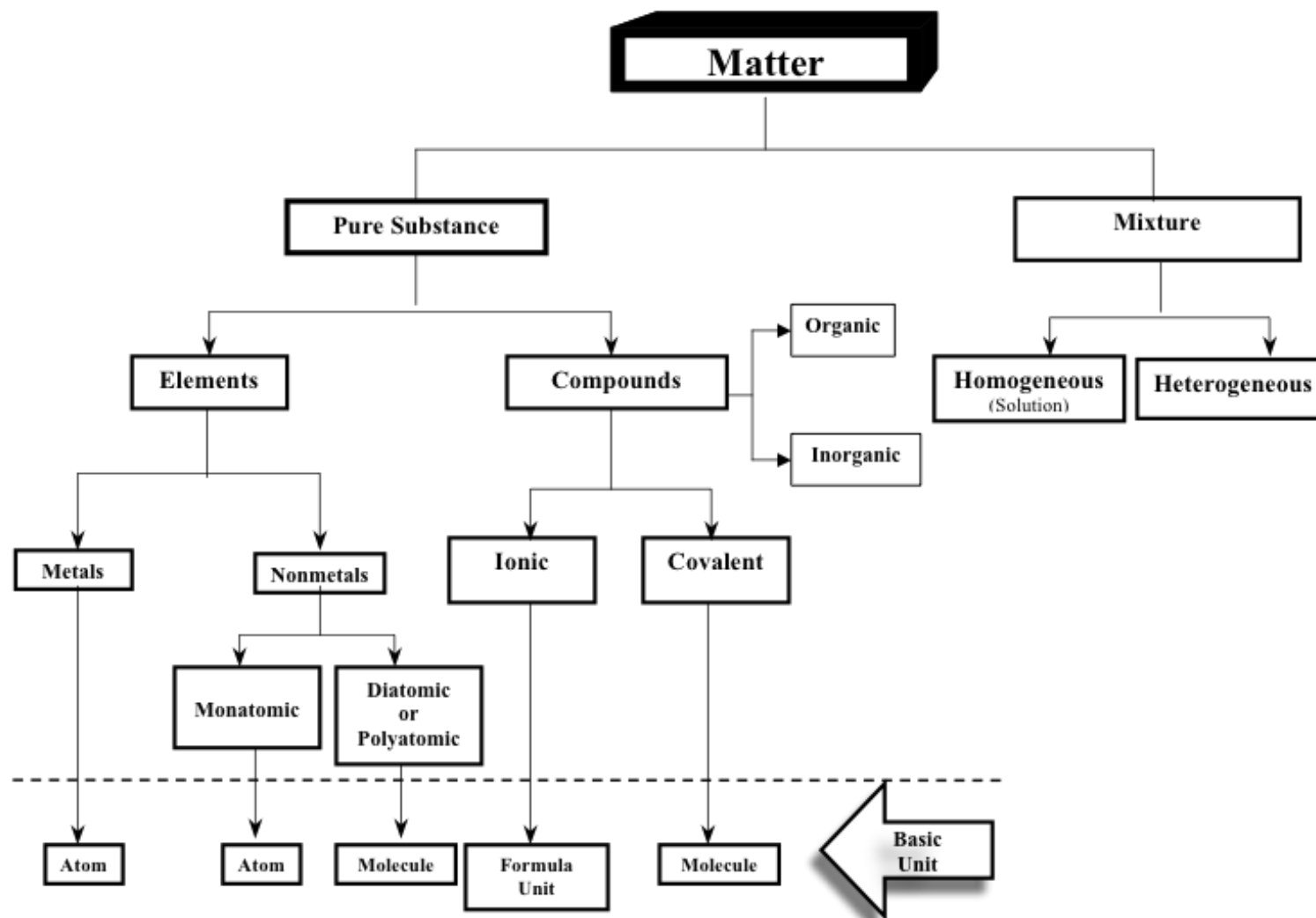
White wine

Water and BaSO₄

Oil and water

Vinegar

Classification of Matter



SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 1. IRON								
_____iron (powder, wire)								
_____iron(III) oxide								
_____iron(III) nitrate nonahydrate								
_____iron(III) nitrate solution								
Series 2. COPPER								
Cu(s) _____								
Cu ₂ O(s) _____								
CuO(s) _____								
CuSO ₄ • 5H ₂ O(s) _____								
CuSO ₄ (aq) _____								

SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 3. MERCURY								
_____mercury (liquid)								
_____mercury(II) chloride								
_____mercury(II) iodide								
Series 4. SULFUR								
S(s) _____								
FeS(s) _____								
FeSO ₄ • 7H ₂ O(s) _____								
SO ₂ (g) _____								

SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 5. SILICON AND CARBON								
Si								
C (graphite)								
SiC								
SiO ₂								
SiCl ₄								
Series 6. IODINE								
I ₂								
KI								
KI (aq)								
I ₂ in C ₂ H ₆ O								
I ₂ in CCl ₄								

SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 7. GASES								
_____ Chlorine								
_____ bromine (vapor over liquid)								
_____ nitrogen dioxide								
_____ Nitrogen								
_____ Ammonia								
_____ Hydrogen								
_____ Helium								

SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 8. ORGANIC CHEMICALS								
C ₆ H ₁₄ (l) Hexane								
HC ₂ H ₃ O ₂ (aq) _____								
C ₄ H ₈ O ₂ (l) ethyl acetate								
C ₃ H ₆ 0(l) Acetone								
C ₂ H ₆ O(l) ethyl alcohol (ethanol)								
C ₁₀ H ₈ O(s) 1-naphthol								
C ₇ H ₆ O(l) Benzaldehyde								
C ₈ H ₈ O ₂ (l) methyl benzoate								
C ₉ H ₈ O(l) trans-cinnamaldehyde								

SAMPLE OF MATTER	PURE SUBSTANCE				MIXTURE		DESCRIPTION (color; physical state; transparent/translucent/opaque; shiny/dull malleable/ductile, shape and size of particles for solids; odor for organic series only)	BASIC UNIT (for Pure substances only) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-		
	Metal	Nonmetal	Ionic	Covalent				
Series 9. MIXTURES								
_____ copper + zinc + other								
Sand + CuSO ₄ • 5H ₂ O copper(II) sulfate pentahydrate								
White wine								
H ₂ O(l) + BaSO ₄ (s) _____								
Oil + Water								
Vinegar 5% acetic acid, 95% water								

What you need to know for Experiment 5 Quiz

- Given a sample of matter, determine which of the following apply: pure substance, mixture, element, compound, ionic compound, covalent compound, hydrate, homogeneous mixture, heterogeneous mixtures
- Definition of the terms transparent, translucent, opaque, malleable, ductile, and brittle.

Sample Quiz 6

- Complete the table below just as you did in Experiment 5.

Sample of Matter	Pure Substance				Mixture		Basic Unit (if applicable) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-	
	Metal	Nonmetal	Ionic	Covalent			
CuSO ₄ •5H ₂ O							
C							
H ₂							
KI(aq)							

- Fill in the blank with one of the following terms: transparent, translucent, opaque, malleable, ductile, and brittle

A substance that can be drawn into a wire is _____

Experiment 6

Chemical Reactions and Equations

PRE-LAB

On the first page:

Objective: Write a short statement describing the purpose of the experiment.

Background: Define the following types of reactions: double replacement, single replacement, combination, decomposition, gas evolution, precipitation, acid-base, redox, and combustion. Also, define the terms precipitate, oxidation, and reduction.

On the next few pages:

Turn your lab notebook horizontally and divide it into three equally sized columns. Label the columns **Procedures/Description of Reactants; Observations of Reaction/Products; and Conclusions**. Fill in the first column by summarizing the procedure written in your lab manual. Keep it very brief. Draw horizontal lines between each reaction and do not put more than four reactions on a page (treat reactions with parts a and b as two separate reactions). When you get to lab you will fill in the rest of the information.

OBJECTIVE

To record observations before, during, and after carrying out a series of different chemical reactions. Based upon these observations and other information you have learned you will write the chemical equation for each reaction and identify its type.

BACKGROUND

In each of this set of experiments we are asking what happens when certain known substances are brought together under specified conditions. A complete answer to this question would require that we analyze the kinds and amounts of the reactants and products of each reaction and also that we measure the initial and final conditions and the energy changes observed. We do not have the time or the apparatus to do all of that, so we will simply make some qualitative observations and deduce what happened from what we know about atoms and ions.

We will also make some comparisons of reactions. By observing the relative speed of a reaction and the relative amount of energy given off as heat and light, we can judge the reactivity of certain substances compared to others under similar conditions.

Your instructor will demonstrate some of these experiments, but you will do most of them. All of the reactions in this experiment fall under one of four common reaction types. They can be recognized by the following general equations.

- a. **Double Replacement (or double displacement):** two elements (or polyatomic ions) in two different compounds exchange places with one another.

$$AB + CD \rightarrow AD + CB$$
- b. **Single Replacement (or single displacement):** one element or ion substitutes another.

$$A + BC \rightarrow AC + B$$
- c. **Combination:** Two or more reactants combine to form a new product. Also called synthesis.

$$A + B \rightarrow AB$$
- d. **Decomposition:** Two or more elements separate to form a new product

$$AB \rightarrow A + B$$

Chemical reactions can also be classified based on what products they form:

- a. **Gas evolution:** Formation of a gas in a liquid (or aqueous solution).

$$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$
- b. **Precipitation:** A reaction where two aqueous solutions mix and form a solid (precipitate).

$$2 AgNO_3(aq) + MgCl_2(aq) \rightarrow 2 AgCl(s) + Mg(NO_3)_2(aq)$$
- c. **Acid-base:** A reaction that occurs between an acid and a base forming salt and water. Also called neutralization reaction.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$
- d. **Redox:** A reaction involving the transfer of electron(s). Also called oxidation-reduction reaction.
 Oxidation: loss of electron. Element that loses electron is oxidized.
 Reduction: gain of electron. Element that gains electron is reduced.

$$K(s) + NaNO_3(aq) \rightarrow Na(s) + KNO_3(aq)$$

K is oxidized (charge goes from 0 to +1)
 Na is reduced (charge goes from +1 to 0)

 - i. **Combustion:** A reaction involving oxygen as a reactant; a type of redox reaction.

$$C_6H_{12}O_6(s) + 6 O_2(g) \rightarrow 6 CO_2(g) + 6 H_2O(g)$$

PROCEDURE

Wear your protective eyewear. Do not take more of a reagent than you need. Do not contaminate the reagent bottles. Follow the order of reactions given by your instructor.

Note: A procedure containing two parts (a and b) indicate **two separate reactions** are occurring.

Reactions 1 through 4 below will be done as demonstrations by your instructor. Do not carry them out yourself.

Reaction 1: Does iron combine with sulfur?

(Instructor Demonstration)

Mix 3.5 grams of powdered iron and 2.0 grams of powdered sulfur in a test tube. Add the sulfur followed by the iron. Mix the reactants well by rolling the tube between your hands or, alternatively, use a stirring rod. Heat the test tube and mixture in the hottest part of the burner flame until evidence of a reaction is observed (You may want to do this with the lights off). Remove the test tube from the flame and allow the reaction to go to completion. While the test tube is still hot, plunge the end in a beaker of water to crack the test tube. Observe the product. Record your observations.

Remove and save the product for a later experiment. Test the product for magnetism. Could you have separated the reactants using a magnet?

Reaction 2 a: What is the gas evolved when iron(II) sulfide reacts with an acid?

(Instructor demonstration; perform in the fume hood)

a. In a test tube, add a small piece of iron(II) sulfide (prepared in reaction 1) and “crush” it using a stirring rod. Add several drops of dilute hydrochloric acid. Watch the test tube carefully. Try to get a light “whiff” of the gas being released and notice the color of the solution over time. Record your observations.

Reaction 2 b: What happens when the gas released in a. contacts moist lead(II) nitrate?

(Instructor demonstration; perform in the fume hood)

b. Test the gas being released in reaction a. by holding a piece of filter paper moistened with lead(II) nitrate solution in the mouth of the test tube. Record your observations.

Dispose of the contents by washing them into the aqueous waste container in the hood.

Reaction 3: What happens when sodium metal is placed in water?

(Instructor Demonstration)

A pea-sized piece of sodium is placed in water. Watch and listen carefully. What gas is evolved? Place a few drops of phenolphthalein in the solution after the reaction is complete. What does the pink color signify? What color is phenolphthalein in pure deionized water? Where will this solution be disposed of?

Reaction 4: What are the products of decomposition of ammonium dichromate?

(Instructor demonstration)

Place about 0.1 gram (no more!) of ammonium dichromate in a large test tube and heat gently until a reaction commences. (You might shut the lights for this reaction and cover the top of the test tube with a glass plate to prevent the expulsion of the contents.) Watch and listen carefully. Once the reaction has started, remove the tube from the flame and allow the reaction to complete. The green product is chromium(III) oxide. Observe the condensation at the mouth of the tube. What else was formed?

Dispose of the solid waste in the container labeled "chromium(III) oxide waste" in the hood.

The student will perform all of the following reactions. Your instructor will direct you as to the order.

Reaction 5 a: Will magnesium burn in oxygen?

CARRY OUT IN THE FUME HOOD.

a. Take a small beaker with you to the hood. Hold a 2-inch strip of magnesium in the burner flame with a pair of forceps. DO NOT look directly at the flame. Let the ash fall into your beaker. Record your observations. Return to your bench and proceed to part b.

Reaction 5 b: Does the combustion product of part a. react with water?

b. At your bench, add several mL of deionized water to the beaker with the ash and mix. Test the solution by adding a drop or two of drops of phenolphthalein indicator. Record your observations. *Remember, phenolphthalein turns pink in basic solution, and is colorless in acidic or neutral solutions.*

Dispose of the contents of the beaker into the container labeled "aqueous waste" in the hood.

Reaction 6: Will iron displace copper(II) ion from copper(II) sulfate solution?

Place a small piece of iron in a test tube containing several drops of copper(II) sulfate solution. Do not disturb the test tube for at least 10 minutes. Observe what has formed. Leave the tube for another 15 minutes, shaking intermittently, and observe the color of the solution (use a white background). What can you conclude from this? Record your observations.

Dispose of the solution in the container labeled "aqueous waste" in the hood.

Reaction 7: Will copper displace silver ion from silver nitrate solution?

Place a small piece of copper (wire or sheet) in a test tube containing several drops of silver nitrate solution (tape covered bottle). Do not disturb the test tube for at least 10 minutes and observe what has formed. Leave the tube for another 15 minutes, shaking intermittently, and observe the color of the solution (use a white background). What can you conclude from this? Record your observations.

Dispose of the solution in the container labeled "aqueous waste" in the hood.

Reaction 8: Will an acid react with zinc metal?

Place a small piece of zinc in a test tube. Obtain a wooden splint and light your Bunsen burner. Add several drops of dilute hydrochloric acid to the test tube. Observe. As the gas is being released, pass a flaming splint over the mouth of the test tube. Record your observations. What possible reaction products are combustible?

Dispose of the solution in the container labeled "aqueous waste" in the hood.

Reaction 9 a: Is the decomposition of a salt hydrate reversible?

a. Place a few crystals of copper(II) sulfate pentahydrate in a test tube. Place the test tube in a clamp with its mouth sloping downwards (or use your test tube holder) and carefully heat the contents in a flame. Observe the color change and the inside surface of the tube.

Reaction 9 b: Is the decomposition of a salt hydrate reversible?

b. Cool the test tube and its contents from part a for a couple of minutes and then add back **only 2-3 drops** of water from your wash bottle directly onto the solid. Record your observations.

Wash the contents of your tube into the container labeled "aqueous waste" in the hood.

The following reactions will be carried out using a plastic well plate:

Reaction 10: Will an acid react with a metal hydroxide?

Place several drops of dilute sulfuric or nitric acid in a spot plate well and add several drops of sodium or potassium hydroxide solution. Mix and feel the bottom of the well. Record your observations. Test the resultant solution using red and blue litmus paper. What does this tell you? Where will you dispose of this?

Reaction 11: What are the reaction products when mixing lead(II) nitrate and sodium sulfate solutions?

Place a few drops of lead(II) nitrate solution in a spot plate well and add a few drops of sodium sulfate solution and mix. Record your observations.

Dispose of the solution in the container labeled "aqueous waste" in the hood.

Reaction 12: What are the reaction products when mixing lead(II) nitrate and ammonium sulfide solutions?

Place a few drops of lead(II) nitrate solution in a spot plate well and go to the fume hood where you will find the ammonium sulfide solution (sulfides have a rotten egg odor!). Add a few drops of the ammonium sulfide solution and mix. Record your observations.

Dispose of the solution in the container labeled "aqueous waste" in the hood.

Reaction 13: What are the reaction products when mixing cadmium nitrate and ammonium sulfide solutions?

Place a few drops of cadmium nitrate solution in a spot plate well, return to the hood and add 3 drops of ammonium sulfide solution. Record your observations.

Dispose of the solution in the container labeled "aqueous waste" in the hood.

IN-LAB

Fill in the observations and conclusions columns for each reaction.

In the **observations column**, answer the following questions:

1. What is the physical state of the reactants? Are they solids, and if so, are they solid pieces, chunks, granules, or powder? Are they in aqueous solution or are they pure liquids or gases? Are they colored?
2. Does the reaction occur by itself when the reactants are mixed at room temperature, or must you apply heat to start the reaction?
3. Does the reaction proceed rapidly, moderately, or slowly?
4. Is a large amount of heat and/or light given off or is there very little energy change noticeable?
5. What is the physical state of the products? Are the products solids, liquids, or gases? Are they soluble or insoluble in water? Are they colored?
6. Did any color changes occur? Did any of the solutions change color?

For reactions that have two parts, answer the set of questions twice, once for part a once for part b.

In the **conclusions column**, write down what the results of each test indicates. Use the "General Guide to Chemical Reactions" on page 80 to guide you with this section.

Example: the appearance of a flame when the glowing splint test was placed near the test tube indicates that oxygen gas was produced.

POST-LAB

- 1) Review your data to ensure that your observations and conclusions are complete and clearly written.
- 2) Complete the charts that begin on page 80.
 - Determine the correct formulas for the reactants and products for each reaction and write the molecular equation. Include all states of matter. Use the Types of Chemical Reactions supplement following this experiment and the following generalizations to direct you. After writing the correct equations, balance each of them.
 - Classify each reaction as a combination reaction (C), decomposition reaction (D), single replacement reaction (SR), or double replacement reaction (DR) by circling the correct notation.
 - Indicate if each reaction is oxidation-reduction (redox) or non-redox. If it is redox, indicate which element in the reactants gets oxidized and reduced.
 - Below each of the **single and double replacement reactions only**, write the complete ionic equation (CIE) and net ionic equation (NIE). Use your class notes and the following **Guide to Chemical Reactions Supplement** to help you.

GENERAL GUIDE TO CHEMICAL REACTIONS

1. Metallic character decreases from lower left to upper right in the periodic table.
2. Elements in the same group (vertical column) of the periodic table are chemically similar.
3. Metals rarely combine with each other. Metals and nonmetals combine to form binary (2 element) ionic compounds. In solution, the metals behave as positively charged particles (cations), and the nonmetals act as negatively charged particles (anions).
4. There are four common reaction types under which all of these reactions will fall. They can be recognized by the following general equations. More detailed information on reactions can be found in the summary starting on page 76.



5. In a replacement (or displacement) reaction, the more reactive metal displaces the less reactive one (see Single Replacement Reactions and the Activity Series in the supplement).
6. Ionic compounds contain a cation, written first, followed by an anion. The compound is neutral. A solubility table can be used to determine whether it dissolves in water.



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

7. Ions in solution may react with ions of opposite charge to form (a) an insoluble substance (precipitate) (b) a neutral substance like water, or (c) a gas like carbon dioxide (CO_2). These are generally double replacement reactions.

8. Pure compounds can change their chemical nature by decomposition, rearrangement, or polymerization (forming chains).

9. Certain metal ions have distinctive colors when in aqueous solution. Most others are colorless. Here are some colored metal ions you will encounter:

a. copper(II), Cu^{2+} : blue b. iron(III), Fe^{3+} : yellow/orange c. nickel(II), Ni^{2+} : green

10. Ions of transition metals may have more than one charge. How do you know which one forms?

a. Cu^{2+} is usually more stable than Cu^+ when forming an ion in solution
b. Fe^{3+} is usually more stable than Fe^{2+} when forming an ion in solution

11. Acid-base (pH) indicators can be used to detect the formation of an acid or base in a reaction. Two common examples are:

a. litmus: acidic solution - blue paper turns to red; basic solution - red paper turns to blue
neutral- no change of either red or blue paper
b. phenolphthalein: colorless in acidic or neutral solutions, pink in basic solutions

12. Gases are a common product of reactions and can be observed as bubbles when forming in solutions. Use the following table to help identify some common gases.

GAS	PROPERTIES	TEST
H_2 hydrogen	colorless, odorless, combustible	Burning splint test: placed near a source of hydrogen results in a “pop” due to combustion of hydrogen
O_2 oxygen	colorless, odorless, facilitates combustion	Glowing splint test: placed near a source of oxygen results in a brighter glow or appearance of a flame
NH_3 ammonia	strong noxious odor; colorless	Litmus test: detect with water moistened red paper turning blue when in contact with ammonia
N_2 nitrogen	colorless, odorless, unreactive	--
H_2O (water vapor)	colorless and odorless	detect by the presence of condensation (wetness) on cool surfaces near the reaction
H_2S hydrogen sulfide	faint yellow color and odor of rotten eggs; toxic	

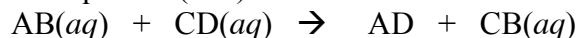
Types of Reactions

I. Double Replacement Reactions: all are non-redox

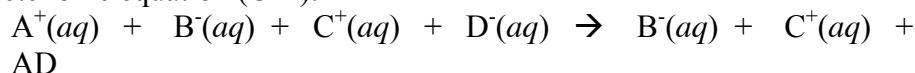
Reaction of ions in aqueous solution

General Reaction Pattern

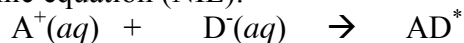
Molecular Equation (ME):



Complete ionic equation (CIE):

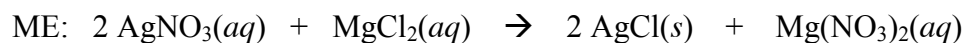


Net Ionic equation (NIE):

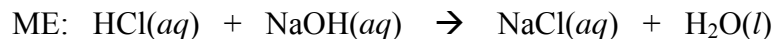
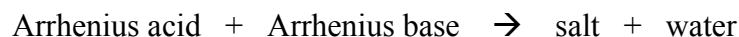


*Generally, ions must be removed from the solution in the form of a solid (precipitate), water or carbon dioxide gas. If one of these doesn't happen then there is no reaction.

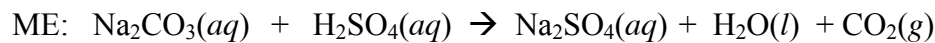
A. Precipitation Reaction: insoluble product is formed



B. Acid-Base or Neutralization Reaction: water is formed



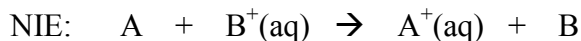
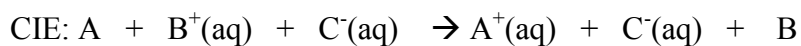
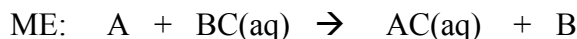
C. Gas is Formed



II. Oxidation-Reduction (Electron Transfer) Reactions

A. Single Replacement Reactions: Element reacts with ion in solution.

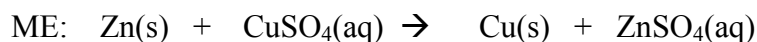
General Reaction Pattern



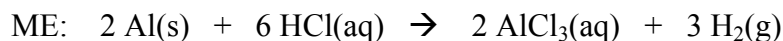
The more reactive substance displaces the less reactive one. Refer to the metal activity series on the right. The most active metal (Li) is at the top. The least active metal (Au) is at the bottom of this series.

There are several forms that a single replacement reaction may take:

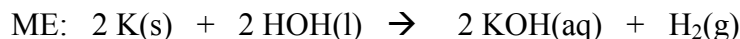
1. **metal** (more active) + **ionic compound** (containing less active metal) $\rightarrow$
metal (less active) + **ionic compound** (with more active metal)



2. **metal** (more active than H) + **acid** $\rightarrow$ **ionic compound** (with more active metal)
+ **hydrogen gas**



3. **metal** (very active) + **water** $\rightarrow$ **metal hydroxide** (a base) + **hydrogen gas**



Metal Activity Series

Li
Cs
Rb
K
Ba
Sr
Ca
Na
Mg
Be
Al
Mn
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb

H

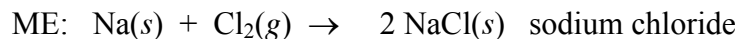
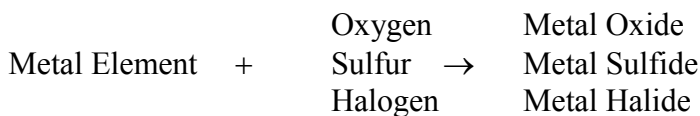
Cu
Sb
Ag
Hg
Pt
Au

B. Combination Reactions (Also called Synthesis or Addition Reactions)
Two Substances combine to form One.

General Reaction Pattern: $A + Z \rightarrow AZ$

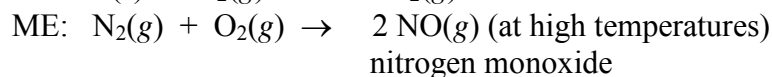
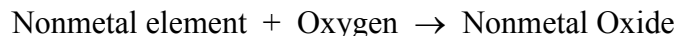
1. Metal element + Nonmetal Element $\rightarrow$ Ionic Compound
(an oxide or a salt)

Most important examples:



2. Nonmetal element + Nonmetal element $\rightarrow$ Covalent Compound

Most important example:



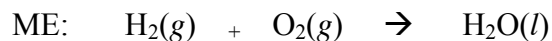
C. Decomposition Reactions

A Single Compound breaks into Two or More elements or simpler compounds.

Most decomposition reactions require the application of energy, usually in the form of heat, light, or electric current and are frequently the reverse of combination reactions.

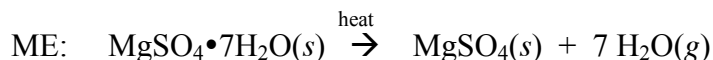
General Reaction Pattern: $AZ \rightarrow A + Z$

Decomposition reactions often involve changes in oxidation numbers and are therefore oxidation-reduction reactions.



Non-Redox Decomposition Reactions

1. Salt hydrates decompose when heated to give the salt anhydrate and water vapor.



What you need to know for Experiment 6 Quiz

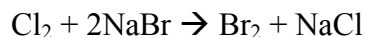
- Types of reactions (double replacement, single replacement, combination, decomposition, gas evolution, precipitation, acid-base, redox)
- Elements oxidized and reduced in a chemical reaction
- Writing molecular equation, complete ionic, and net ionic equations.

Sample Quiz 7

1. Consider the reaction : $\text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$
The reaction given above falls under what type(s) of reaction(s)? Circle all letters that apply:

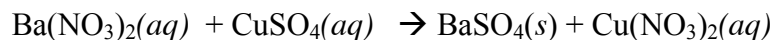
- | | | |
|-----------------------|------------------|-----------------|
| a. Gas evolution | e) combination | h)Precipitation |
| b. Single replacement | f) decomposition | i) acid-base |
| c. Double replacement | g) redox | j) combustion |

2. Consider the following redox equation:



What element is oxidized? _____ reduced? _____

3. Write the balanced complete ionic (CIE) and the balanced net ionic equations (NIE) for the following reaction



CIE: _____

NIE: _____

Experiment 6 Equation Sheet

Name _____ Section _____

Rxn	1. Balanced Chemical Equation 2. CIE (If the reaction is a single or double replacement reaction) 3. NIE (If the reaction is a single or double replacement reaction) Circle reaction type: C (combination), D (decomposition), SR (single replacement), or DR (double replacement)
1 C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
2a C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
2b C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
3 C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____

4	YOU DO NOT NEED TO WRITE THE EQUATION FOR REACTION 4	
5a	1.	
C		
D	2.	
SR		
DR	3.	
	Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____	
5b	1.	
C		
D	2.	
SR		
DR	3.	
	Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____	
6	1.	
C		
D	2.	
SR		
DR	3.	
	Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____	
7	1.	
C		
D	2.	
SR		
DR	3.	
	Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____	

8 C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
9a C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
9b C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____
10 C D SR DR	1. 2. 3. Redox: yes or no If Redox: element that is oxidized = _____ Element that is reduced = _____

Experiment 7

Determination of a Chemical Formula

PRE-LAB

Objective: Write a short statement describing the purpose of the experiment.

Background: Define the term empirical formula.

Procedure: Briefly write/outline what you will be doing in the experiment.

Pre-lab Assignment: Answer the three questions on page 87.

Data: Prepare a data table for all the information you will be collecting.

OBJECTIVE

To carry out two reactions in order to determine the correct chemical formula for a copper chloride hydrate ($\text{Cu}_x\text{Cl}_y \cdot z\text{H}_2\text{O}$).

BACKGROUND

John Dalton was an Englishman, a teacher, and an exceptional theoretical chemist. He developed and wrote the modern atomic theory at the turn of the 19th century. A fundamental component of the modern atomic theory is that the mole ratios of elements in a compound will be small whole numbers (*law of definite proportions*). The smallest whole number mole ratio is commonly referred to as the empirical formula of a compound.

One of the challenges in finding the proper chemical formula for a compound is that there may be more than one possible mole ratio for the elements in that compound. Dalton called this the *law of multiple proportions*. For example, if you were testing a compound that contained iron and sulfur, the possible chemical formula could be FeS or Fe_2S_3 . However, if you determine the mass of iron and the mass of sulfur present in a given mass of the compound, you will be able to establish the true chemical formula of the compound.

In this experiment, you will use the law of definite proportions to find the chemical formula for a hydrated compound containing copper, chlorine, and water molecules locked in the crystal structure of the solid compound. The general formula for the compound is $\text{Cu}_x\text{Cl}_y \cdot z\text{H}_2\text{O}$ where x , y and z are integers and its name is copper chloride hydrate. The letters x , y and z represent integers that will establish the proper chemical formula for this substance. First, you will gently heat a sample of the compound to drive off the water of hydration. By measuring the mass of the sample before and after heating you can determine the amount of water in the sample and identify the value of z .

Second, you will conduct a chemical reaction with the dried sample, which will produce elemental copper. By measuring the mass of copper that forms, you will have the necessary information to determine the moles of copper and chlorine in your sample, and you will be able to establish the proper chemical formula.

PROCEDURE

Wear your safety eyewear throughout this experiment.

Take a small beaker, clean dry test tube and your stirring rod to the balance. Tare your beaker on the balance and place the test tube into the beaker and weigh it. Record the mass. Remove the test tube from the beaker and transfer between 0.8 and 1.2 g of the unknown copper chloride hydrate to the test tube. Use your stirring rod to break up any large pieces of the substance by pressing the pieces against the wall of the crucible. Place the test tube back in the beaker and measure and record the mass of the test tube with compound to the nearest 0.001 g in your notebook.

To determine the mass of water:

Return to your bench, turn on your Bunsen burner and adjust the burner to get a small flame. Place the test tube in a test tube holder and move the test tube slowly back and forth over the small blue flame (but not the inner blue cone) to gently heat the sample. **Do not overheat the compound.** Note the color change, from blue-green to brownish, as the water of hydration is driven out of the crystals and condenses on the inside of the test tube. (You do not want a black solid product. This means the sample was overheated). Heat the entire length of the tube as you will need to remove any traces of water. When the sample has turned brown and you see no moisture in your tube, gently heat it for two more minutes. (Any blue-green crystals indicate the water of hydration hasn't been completely removed.)

Remove and turn off the burner. Place the test tube in your beaker and let it cool for five minutes. Return to the balance and tare the beaker. Insert the test tube and record the mass. Return to your bench and gently reheat the test tube for about 2 minutes using the clamp as you did before. Allow the sample to cool for at least five minutes. Reweigh as before and record the mass. If the mass difference from the prior weighing is **less than 0.01 g** you can assume that all of the water has been removed. This process is called heating to constant mass. If the difference is greater than 0.01 g, the sample must be heated another time. The difference between the **smallest mass** and the mass of the original sample represents the mass of water in your sample.

To determine the mass of copper:

Carefully transfer the brown solid to a clean and empty 50 mL beaker. Rinse out the test tube with two 8 mL aliquots (portions) of deionized water and pour the water into the 50 mL beaker. Gently swirl the beaker to completely dissolve the solid. Note that the color of the solution is green as the copper ions are rehydrated.

Measure out about 12 cm of aluminum wire, coil the wire, and place the wire in the beaker of solution so that it is completely immersed in the copper chloride solution. Note that the reaction produces a gas, elemental copper is forming on the surface of the aluminum wire, and the color of the solution is fading. Carefully swirl the solution periodically to expose

fresh aluminum to the solution. The reaction will take about 30 minutes to complete. (While you are waiting you can answer questions 1 and 2 under the calculations section.) When the reaction is complete, the solution will be colorless. Add 5 drops of 6 M HCl solution to dissolve any insoluble aluminum salts in the mixture, which should make the solution clear.

Use a stirring rod to scrape off as much copper as possible from the Al wire. Slide the wire up the wall of the beaker and out of the solution with the stirrer and rinse off any remaining copper with deionized water. If any of the copper refuses to wash off the aluminum wire, wash it with one or two drops of 6 M HCl solution. Put the Al wire aside.

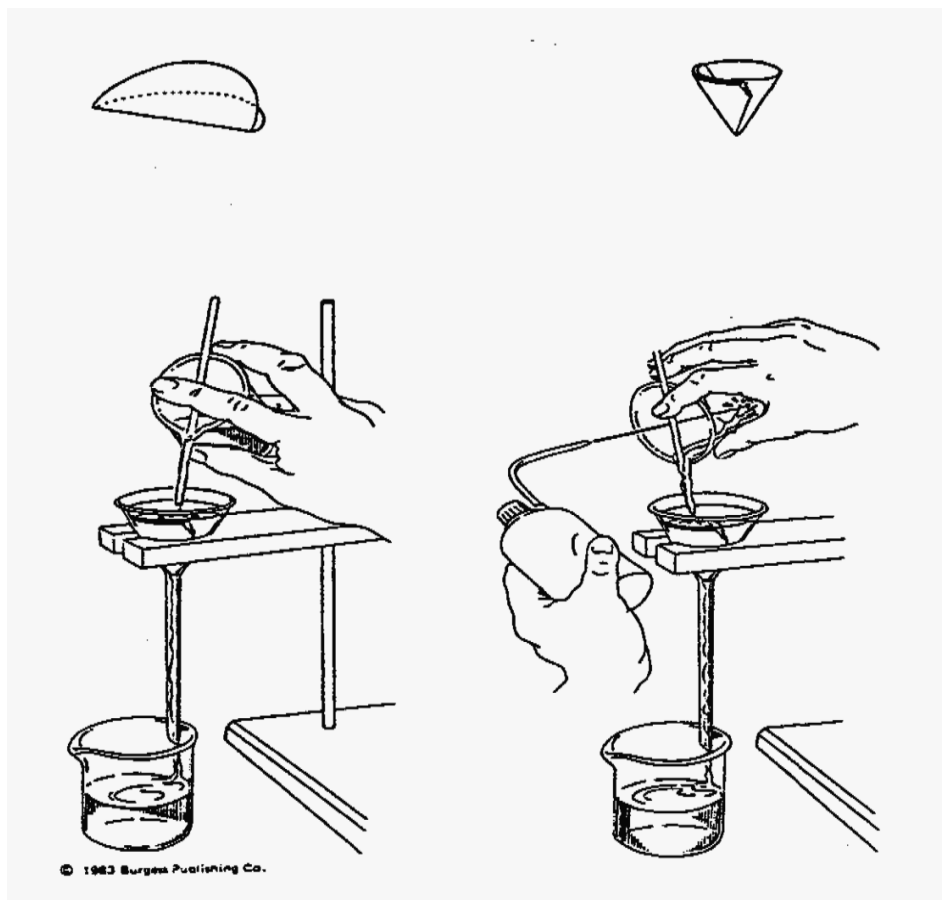


Figure 7.1
Folding Filter Paper, Use of Funnel and Wash Bottle

Set up a filtration device using a ring stand and funnel as shown in figure 7.1. Weigh and record the mass of a piece of filter paper in your lab notebook. Fold it into quarters, wet it and place it in the funnel so that it adheres to the sides. Filter the solution through the weighed filter paper, using your wash bottle to rinse the copper particles from the beaker to the filter paper in the manner shown in Figure 7.1. Using the stirring rod, carefully break up the larger pieces of copper

trying not to lose any copper. Wash the copper several times starting from the top of the filter paper. Allow the solution to drain. To insure that you have washed through any solution remaining on the solid copper on the filter, you will test the filtrate coming out of the bottom of the filter with blue litmus paper. When the blue paper no longer turns red you can be sure that the copper is clean. Unfold the filter paper with the copper, place it on a watch glass with your name taped to it.

Dry the watch glass/filter paper/copper in an oven for 10 minutes to remove any residual water. If the filter paper feels dry let it cool to room temperature and weigh it. Repeat the heating and cooling until the weight doesn't change by more than 0.01 g.

Alternatively, your instructor may suggest that you place the watch glass/filter paper/copper on the shelf assigned to your instructor to dry. At the next lab period, weigh the filter paper with dried sample of copper.

Disposal: Place any aluminum wire you remove in the container marked "solid metal waste" and place the filtrate in the aqueous waste. After you are finished with the copper put it in the solid metal waste also. Place the filtered solution into the "aqueous waste" container.

Do I Understand?

As we learned in experiment 6, chemical reactions give us clues in the form of observations that changes are taking place. What visible changes are occurring during the course of the reaction in this lab? How will we know if the reaction has gone to completion? If it hasn't gone to completion in a reasonable amount of time what might be the problem and how could you address it? Lastly, how does the fact that the blue litmus paper ultimately remains blue tell you that the copper is now "clean".

PRELAB ASSIGNMENT

- 1) In this lab, we do not measure the amount of chlorine in the sample. We need the amount of chlorine to determine the formula of the compound. How do we determine the amount of chlorine in the compound?
- 2) The mineral bonattite is known to be a copper sulfate hydrate. 2.000 grams was heated to release the water for form the anhydrate. The sample lost 0.506 g of mass. After reacting with aluminum wire, 0.595 g of copper was recovered. How many moles of water, copper and sulfate were in the compound? What is the mole ratio of these 3 components?
- 3) What is the formula of bonattite?

IN-LAB

Data: Complete the data table you prepared.

Calculations: Perform the calculations asked for on the next page.

CALCULATIONS

- 1) Write the two chemical reactions that the copper chloride hydrate ($\text{Cu}_x\text{Cl}_y \cdot z\text{H}_2\text{O}$) is undergoing in this experiment (you will not be able to balance them as you do not know x, y or z.)
- 2) Calculate the number of moles of water in your sample.
- 3) After you have the mass of your dried copper, calculate the number of moles of copper and moles of chlorine in your sample.
- 4) Determine the mole ratio of these 3 components Cu, Cl and H_2O by dividing by the smallest number. Take into account significant figures.
- 5) Predict the correct chemical formula for the compound.

POST-LAB

Results: Create a summary results table

Conclusion: Analyze the results and errors in a written discussion. The correct formula for the hydrate is $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$. How does your formula compare to the correct formula?

Now assume that during your calculations, you had rounded the ratios to the nearest whole numbers rather than finding a common multiplier and multiplying all the numbers by that number. Would your formula in this case be closer to the correct formula? Include any sources of error that may have occurred throughout the experiment, which could have led to your results deviating from the correct formula.

What you need to know for Experiment 7 Quiz

- Empirical formula
- How to find empirical formula given mass of constituent elements

Sample Quiz 8

1. Circle all the empirical formulas
 $\text{C}_2\text{H}_6\text{O}$ N_2O_4 P_2O_5
2. A 2.300 g sample of $\text{Cu}_x\text{Cl}_y \cdot z\text{H}_2\text{O}$ weighed only 1.950 g after heating. After aluminum was added, a total of 1.234 g of solid copper was collected. What is the formula of the hydrate?

Experiment 8

Qualitative Analysis of Ions

PRE-LAB

Objective: Write a short statement describing the purpose of the entire experiment.

Background: Define the terms cation, anion and positive control.

Procedure:

Part 1

Write down the procedure for part 1 of the experiment. Underneath the procedure leave some space for the in-lab calculation of the mass of salt you will need for the ion your instructor assigns you.

For the rest of the prelab, you will prepare data tables for the information you will be collecting for each part of the experiment as described below:

Part 2

Turn your notebook horizontally. Prepare a data table in your lab notebook in which to write the data you will obtain from carrying out the flame tests. In your table, include a place to record the hue (color or shade of color), intensity, and duration of the color imparted to the flame.

Part 3A

Turn your notebook horizontally. Draw a line down the center of the lab page and label the columns: procedures and observations. Briefly **outline** the **procedures** given in part 3A for each ion in the first column. You will record your observations opposite each step of the procedure in the second column in lab.

Part 3B

As above, turn your lab notebook horizontally and divide it into two columns labeled: procedures and observations. Briefly outline the procedures for each ion in the first column. You will record your observations opposite each step of the procedure in the second column in lab.

Part 4

Turn your notebook horizontally. Prepare two data tables (one for solid unknown, one for liquid (solution) unknown) in your notebook in which to record your observations and conclusions for the flame test and for the chemical tests of cations and anions for each ion test for the two unknown samples. In the conclusion section write down whether the ion you tested for is present (indicate a positive test) or absent (indicate a negative test). You will need one table for the solid unknown and one table for the solution unknown.

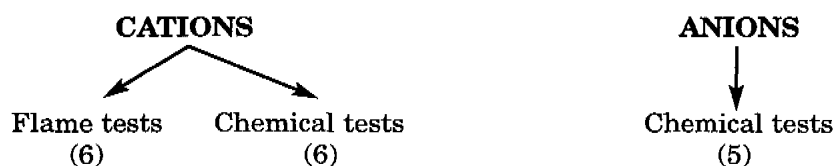
OBJECTIVE

This experiment is designed to provide you with ways to test for the presence of specific ions, both cations and anions, with the ultimate goal of using these tests to determine whether any of these ions is present in an unknown sample.

BACKGROUND

The process described above is known as "**qualitative analysis**" because you are analyzing for the qualitative, i.e. presence or absence, of specific chemicals in this case ions. This is in contrast to **quantitative analysis** which involves determining how much of certain chemicals are present. Frequently a chemist must separate the various chemicals to be analyzed before carrying out the analysis. In this experiment the ions do not need to be separated from each other because the test for each ion has been designed to be done without interference from the presence of the other ions.

In part 1 of the experiment, your class, as a whole, will prepare a set of solutions containing each of the ions that you will be testing for: these will be called the "test ion solutions". There are two different types of tests that each student will carry out depending on which ion you are testing for: **chemical or flame tests**. The chemical tests are predetermined chemical tests designed to be specific for a certain cation (NH_4^+ , Al^{3+} , Fe^{3+} , Cu^{2+} , Ni^{2+} , or Mg^{2+}) or anion (NO_3^- , Cl^- , SO_4^{2-} , PO_4^{3-} or CO_3^{2-}). Note that **cations** are positively charged ions and **anions** are negatively charged ions. The flame tests involve heating a test cation in the flame of a Bunsen burner. Each of the six cations you will test for (Li^+ , Na^+ , K^+ , Ca^{2+} , Ba^{2+} , or Cu^{2+}) will give off a color unique to that ion. In each case, the test ion will be present when carrying out the test giving an expected response--this type of test is known as a "positive control" because the results indicate that the test is positive for the ion being tested for. Once you've observed and recorded the results of the positive test for each specific cation or anion, you will perform these same tests replacing the test ion with an "unknown" sample containing unknown ions. If you observe results that are similar



to the results that were observed for that ion's positive control, then it indicates that the ion being tested for is present, i.e. a positive result. If you don't observe similar results to the positive control for that ion, the ion is assumed to be absent: i.e. a negative result.

PART 1: PREPARING TEST ION SOLUTIONS

BACKGROUND

As stated above, before you run a test to determine the presence of any chemical in an unknown, it is important to carry out the same test on a sample in which you know that chemical is present. The chemicals must be in the sample in a reasonable concentration to give an observable result when it is tested for. In this first part of the experiment, the class as a group will prepare these test ion solutions at an appropriate concentration.

It is not possible to prepare a solution containing only a cation or an anion because all ionic compounds contain both. Therefore we will prepare a solution that contains both cations and anions making sure that the presence of the ion that we are not testing for (sometimes called the counter ion) doesn't interfere with the test we are using to test for the other ion. For these tests a concentration of 10.0 mg of the test ion per mL of solution will give us enough of the test ion to carry out an accurate test. Once a complete set of test ion solutions is prepared by your class for all of the cations and anions we are testing for, they will be used by the entire class to carry out the tests.

A concentration of each test ion of 10.0 mg test ion/mL solution will be sufficient to be detected by the tests we are using. A quantity of 250.0 mL of each solution will allow for enough test ion solution for the entire class for this experiment. In order to start, you will need to calculate how much of the neutral ionic compound (or salt hydrate) you will need to weigh out to prepare 250.0 mL of 10.0 mg test ion/mL solution. Suppose that you are assigned Al^{3+} as the test ion and aluminum nitrate nonahydrate, $\text{Al}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$. Your solution map might look like this:

mL soln. $\rightarrow$ mg test ion $\rightarrow$ g test ion $\rightarrow$ mol test ion $\rightarrow$ mol compound $\rightarrow$ g compound

Once you calculate how many grams of aluminum nitrate nonahydrate needed you will dissolve it in sufficient deionized water to make 250.0 milliliters of solution using a volumetric flask. Your instructor will demonstrate.

PROCEDURE

Part 1

Your instructor will assign groups of 2-3 students to calculate and then prepare each test ion solution. After your instructor assigns your group a test ion, use the formula of the compound indicated in the table below to calculate the mass of the compound needed to prepare the solution. Show the result to your instructor for confirmation before starting. Include this calculation in your lab notebook under part 1 procedure.

Test Ion	Formula of Salt	Molar Mass of Salt (g)	Grams of Salt
Al^{3+}	$\text{Al}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$	375.15	
Fe^{3+}	$\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$	404.02	
Mg^{2+}	$\text{Mg}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$	256.4	
NH_4^+	NH_4NO_3	80.05	
Cu^{2+}	$\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$	241.5	
Ni^{2+}	$\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$	290.81	
Cl^-	NH_4Cl	53.50	
NO_3^-	NaNO_3	85.00	
CO_3^{2-}	NaHCO_3	84.01	
SO_4^{2-}	$(\text{NH}_4)_2\text{SO}_4$	132.15	
PO_4^{3-}	$\text{NH}_4\text{H}_2\text{PO}_4$	115.03	

Wearing your protective eyewear, tare a 250 mL beaker and measure the correct mass of compound needed on the balance. Return to your bench and dissolve the compound in about 150 mL of deionized water. Transfer the solution to a clean 250.0 mL volumetric flask using a funnel. Rinse out the beaker with a small amount of deionized water and add it to the flask. Bring the volume up to the mark on the neck of the flask with your wash bottle being careful not to overshoot. Using a funnel, transfer the solution you have prepared into a 250 mL dropper bottle. Take a label and prepare it as illustrated in the figure below. Write the test ion large in the center and the counter ion small below. Your instructor will direct you as to where to put the bottles.

10.0 mg/mL		Your names	
Test ion	Al^{3+}		Counter ion
Instructor's name		NO_3^-	Date

IN-LAB part 1

Show your work for the calculation of the mass of the salt you will need to make the solution for the test ion your group was assigned.

PART 2: SPECTRAL AND CHEMICAL ANALYSIS OF CATIONS

BACKGROUND

The purpose of this part is to qualitatively identify some cations by using a flame test. We know that matter will absorb and emit radiation, especially when it is hot. We can feel the radiation as heat or see it as light if the radiator is hot enough. The radiation spectrum from solid or liquid matter is more or less continuous; that is, it has a wide range of frequencies in it, as in Figure 8.1. Sunlight has a wide range of frequencies in it also, and when its components are refracted or “spread apart” by a prism we see a rainbow of colors. Matter in the atomic state is more limited. Free atoms absorb or emit radiation of particular frequencies only. Atoms excited in a flame will emit light of characteristic colors, and when the colors are spread apart by a prism or a grating, a spectrum like that in Figure 1 is obtained. The people who study such things are called spectroscopists, and they found that the patterns of “bands” emitted are a characteristic of each cation- like a fingerprint identifies a specific individual. In fact, for the simpler elements like hydrogen, the frequencies of the bands follow a fairly simple mathematical law. (Frequency is a characteristic of radiation that measures how often a segment of that radiation passes. For visible radiation, each color of light has a different frequency.) The theory that was developed to explain these and other observations is called “quantum mechanics” and, to date, this theory is the best explanation of the behavior of atoms that we have.

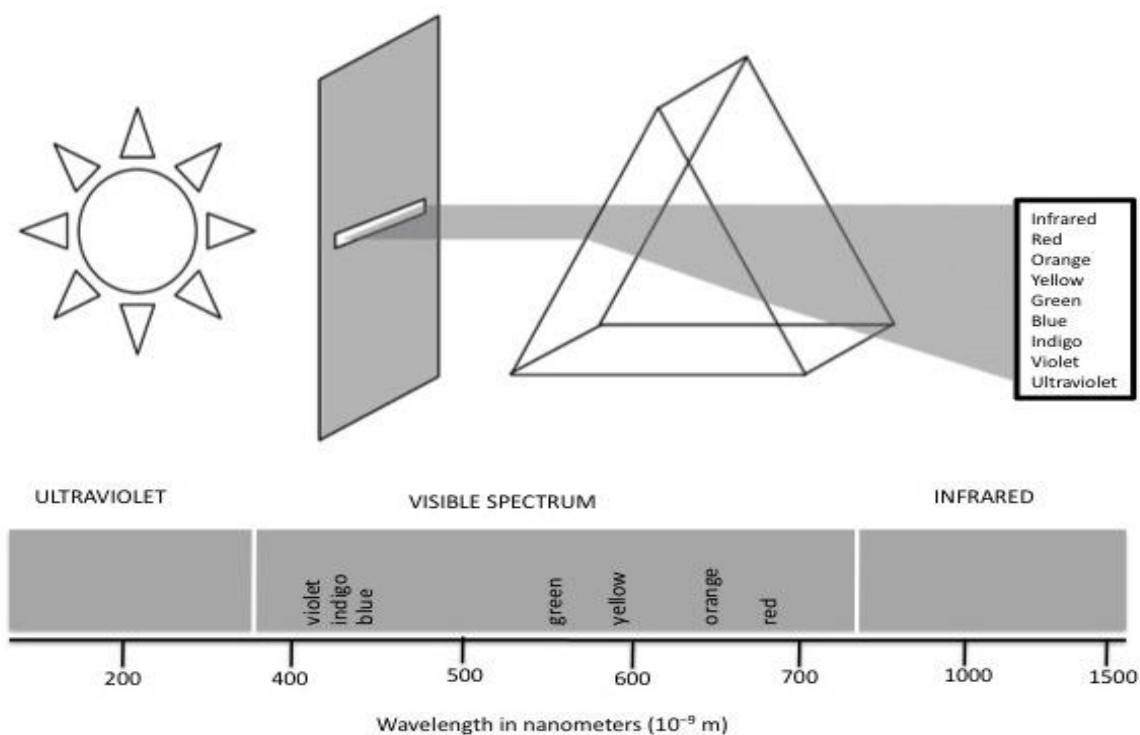
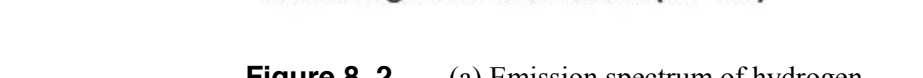


Figure 8.1. Dispersion of light by a prism: the short wavelengths are refracted (bent) more than the longer wavelengths by the transparent prism. This separates the white light into distinct bands of color.

The atom or ion can return to its lower energy by emitting a photon of light. The color of this photon depends on the difference in energy between the higher energy level and the ground state.



es of these lines depend on the total electronic str

Chemistry 60 lab
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PROCEDURE

Part 2

For this section, record your observations in the table you prepared for prelab.

Before you perform a flame test on your unknown, it is important to first observe the flame colors in situations where you know what the cation is. For this purpose you will find a set of six dropper bottles containing solutions of the following salts: LiCl, NaCl, KCl, CaCl₂, BaCl₂, and CuCl₂. Remove the cork/wire assembly attached to the bottle and place one or two drops of the solution onto the loop of the wire. Place the loop into the flame. Observe and record the hue, intensity and duration of the color imparted to the flame in a neat table in your lab notebook. Repeat each sample to ensure you know what the color looks like. Carry out the same procedure with the other five test ion solutions. Be sure to return each cork/wire assembly to the proper bottle.

IN-LAB part 2

Complete the data tables you prepared for part 2 of the experiment. Make sure to write very detailed observations.

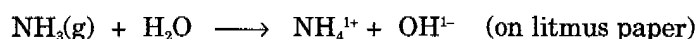
PART 3A: SPECIFIC CHEMICAL TESTS FOR CATIONS

BACKGROUND

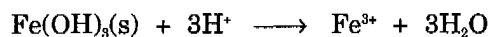
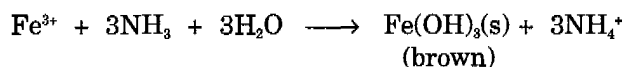
The term "analysis" means the identification of the components in a substance. When the term "qualitative" is used, it means that only visual observations are required. No information about the quantity of the component in question is obtained.

Qualitative analysis is possible because experience has shown that certain chemical reactions of ions in solution are limited to at most a few kinds of ions. Some reactions are unique to a specific ion. If we conduct an experiment in such a way that a positive reaction is possible for only one kind of ion, we can get a "yes or no" answer to the presence of the ion in question. A positive reaction in qualitative analysis must be one you can observe by sight or smell. We can observe the evolution of a gas by sight and sometimes by smell. We can observe the formation of an insoluble substance (a precipitate) and its color. We can also observe color changes in solution. Sometimes it is not possible to find a unique reaction for a single ion. In those cases we may have to rely on two or more different reactions such that a positive reaction to the first test says, "Maybe," but a positive reaction to both tests says, "Yes, for sure." Here are the rationales and net ionic equations for the cation chemical tests:

Ammonium ion (NH_4^+) will react with a strong base such as sodium hydroxide to form ammonia. Ammonia is a water-soluble gas with a distinctive odor. Warming the solution decreases the solubility of the gas, and the odor may be detected by **cautiously** sniffing the mouth of the test tube. The confirming test is to hold a piece of wet pink litmus paper **above** the mouth of the test tube. If ammonia is present, the litmus paper will quickly turn blue.



Iron(III) ion (Fe^{3+}) and aluminum ion (Al^{3+}) belong to a class of elemental ions whose hydroxides are very insoluble. Both Fe^{3+} and Al^{3+} will precipitate as the hydroxides when the solution is made slightly basic with ammonia. Iron forms a brown gelatinous precipitate.



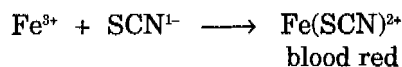
Aluminum forms a gelatinous white precipitate.



The color alone would distinguish these ions from each other but not from a number of other

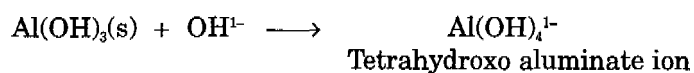
ions that also form insoluble hydroxides. Therefore, we must test the insoluble hydroxides with confirmatory tests.

Iron(III) ion will combine with thiocyanate ion (SCN) to form a soluble red complex ion.



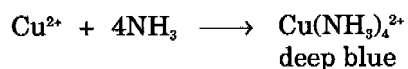
If we observe a brown precipitate when the solution is made basic with ammonia, we probably have Fe^{3+} . If we dissolve the brown precipitate with hydrochloric acid and add potassium thiocyanate to the solution, the formation of a blood red color confirms the presence of Fe^{3+} .

Aluminum belongs to a class of substances whose hydroxides are soluble in strong bases.



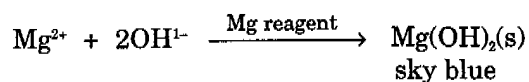
If we obtain a gelatinous white precipitate when the solution is made weakly basic with ammonia, we probably have Al^{3+} . If the white precipitate dissolves when sodium hydroxide is added, we can be quite sure that Al^{3+} is present.

Copper(II) (Cu^{2+}) and nickel(II) (Ni^{2+}) are colored ions in solution, but the exact color depends on the nature of the ions or neutral substances associated with them. Copper(II) ion is light blue in aqueous solution. If ammonia is added, the solution turns the deep blue color of tetramine copper(II) ion.



Nickel is a greenish color in solution. If we suspect that nickel is present, we can confirm the fact with a reagent that is specific for nickel. This reagent, dimethylglyoxime, forms a red gelatinous precipitate in the presence of nickel ion.

Magnesium ion (Mg^{2+}) will also form a white precipitate when the solution is made basic. Magnesium hydroxide does not usually precipitate from weakly basic solutions and the hydroxide does not dissolve when sodium hydroxide is added. If a white precipitate is observed when the solution is made strongly basic with sodium hydroxide, we can confirm that it is magnesium hydroxide by adding a dye that is specific for magnesium hydroxide. The dye gives the precipitate a sky blue color.





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

PROCEDURE

Part 3A

For this section, record your observations in the table you prepared for this section for prelab.

Use about 10-15 drops of test solution in each of the following tests, which you may do in any order. After the addition of each reagent (chemical), mix the contents in the test tube by holding it firmly at the top and flicking it with the end of the finger. If you are using a well on your well plate, mix the content by using a clean stirring rod. **You will need a hot water bath at several stages during this experiment so set one up now on the bench between you and your partner that you can share.**

Common Tips on Handling Chemicals Efficiently

- For most of this experiment **careful measurement** of chemical quantities is **not** necessary. Remembering this rule of thumb- **1 mL is about 20 drops**- allows you to quickly dispense chemicals.
- Use a clean stirring rod to **mix your solutions** before making your observations.
- **Determining if a solution is acidic or basic:** frequently you are asked to determine if a solution is acidic or basic. To sample your solution, dip your **clean** stirring rod into the mixed solution and then touch it to a piece of either red or blue litmus paper.

Test for NH_4^+ ion:

To 10-15 drops of NH_4^+ test solution (or your unknown) in a small test tube, add about 1 milliliter of dilute NaOH. Be careful to not drop NaOH on the top rim of the test tube. Immediately test the gas emitted with a piece of red litmus paper moistened with deionized water held directly **above** the mouth of the test tube (Do not touch the tube).

- **Discard the solution in the base waste container in the hood.**

Test for Al^{3+} ion:

Add 10-15 drops of Al^{3+} test solution (or your unknown) to a plastic well plate. Add dilute $\text{NH}_3(aq)$ drop wise and mix until a precipitate forms and the solution is basic to litmus. Dissolve the precipitate by adding dilute NaOH(aq) drop wise with stirring until the solution becomes clear. Observe the precipitate reform when an equal amount of dilute HCl(aq) is added.

- **Discard the solution in the aqueous waste container in the hood.**

Test for Fe^{3+} ion:

Add 10-15 drops of Fe^{3+} test solution (or your unknown) to a well on a plastic well plate. Add about 1 milliliter of dilute $\text{NH}_3(aq)$ and mix. Dissolve the brown precipitate by the drop wise addition of dilute HCl(aq). Then add 2 to 3 drops of KSCN solution; mix and observe.

- **Discard the solution in the aqueous waste container in the hood.**

Test for Cu^{2+} ion:

Add 10-15 drops of Cu^{2+} test solution (or your unknown) to a well on a plastic well plate. Add about 15 drops of dilute $\text{NH}_3(aq)$, mix and observe the color change. If you get a cloudy solution, add more $\text{NH}_3(aq)$.

- **Discard the solution in the aqueous waste container in the hood.**

Test for Ni^{2+} ion:

Add 10-15 drops of Ni^{2+} test solution (or your unknown) to a well on a plastic well plate. Add about 5 drops of dilute $\text{NH}_3(aq)$ and mix. Then add about 10 drops of dimethylglyoxime solution and mix.

- **Discard the solution in the aqueous waste container in the hood.**

Test for Mg^{2+} ion:

Add about 10 drops of the Mg^{2+} test solution (or your unknown) to a well on a plastic well plate. Place the plate on a sheet of white paper. Add no more than 2 drops of magnesium reagent dye to the solution, and mix. To the colored solution add about 5 drops of dilute NaOH(aq) and mix. The formation of a sky-blue precipitate is a positive test for magnesium ion.

- **Discard the solution in the aqueous waste container in the hood.**

Record all of your observations into your lab book.

IN-LAB part 3A

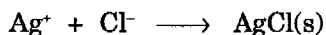
Complete the data table you prepared for part 3A of the experiment. Make sure to write your observations in detail.

PART 3B: SPECIFIC CHEMICAL TESTS FOR ANIONS

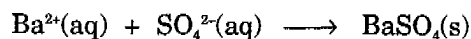
BACKGROUND

The tests for anions follow the same principles we have discussed for cations. Certain reagents are added to a solution of the test ion under controlled conditions and a specific result is obtained.

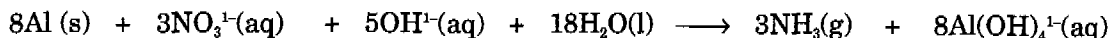
The test for **chloride (Cl⁻) ion** depends upon the formation of a white precipitate when silver ion is added. Silver ion also forms white precipitates with other anions, but not in acid solution. To prevent the interference of the other anions, we make the solution acid with nitric acid before adding the silver nitrate reagent. If a white precipitate forms when silver ion is added to the acidic test solution, chloride is present. If you now hold the solution in the sun, the silver ion will be converted into silver metal, as evidenced by the grayish tinge of the solution.



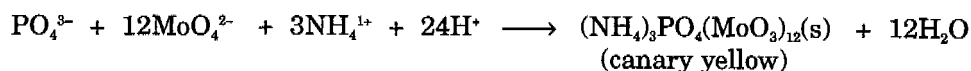
The test for **sulfate (SO₄²⁻) ion** is similar, except that we use barium ion (Ba²⁺) to precipitate the sulfate as barium sulfate (BaSO₄). Again, we must make the solution acidic first so that phosphate and carbonate do not interfere.



The test for **nitrate (NO₃⁻) ion** depends upon the formation of gaseous ammonia when aluminum metal is added to nitrate ion in basic solution.

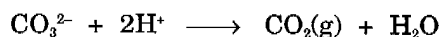
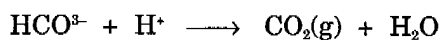


In acid solution, ammonium molybdate reacts with **phosphate (PO₄³⁻) ion** to form a bright yellow precipitate.



In the absence of phosphate, an ivory or buff-colored precipitate often forms, so the positive test requires that the precipitate be bright yellow.

Hydrogen carbonate (HCO₃⁻) ion and carbonate (CO₃²⁻) ion react with acids to form carbon dioxide and water. If an effervescence (bubbling) is observed when acid is added to a solution, bicarbonate or carbonate is probably present.



PROCEDURE

Part 3B

For this section, record your observations in the table you prepared for this section for prelab. Use the indicated amount of test solution in each of the following tests, which you may do in any order. After the addition of each reagent (chemical), mix the contents in the test tube by holding it firmly at the top and flicking it with the end of the finger. If you are using a well on your well plate, mix the content by using a clean stirring rod.

You will need a hot water bath at several stages during this experiment so set one up now. Use the bench area between you and your partner so that you can share it.

Test for the NO_3^- ion:

For part 4 (solid and solution unknowns), you may skip this test if your unknown tested positive for ammonium)

- Pour about 1 mL of the test ion solution (or your unknown) into a 150 mL beaker that was cleaned and rinsed with deionized water. Add about 5 mL of dilute sodium hydroxide solution into the beaker and mix the two by gently swirling. Make sure the resulting solution is basic by using a piece of red litmus paper. If the solution is not basic, add more sodium hydroxide solution until it is basic. Add about 5 mL of deionized water. Prepare a wet strip of red litmus paper and attach it to the bulging side of a clean watch glass. (Make sure that there is no residual base contamination from your hands on the watch glass or splattering of the solution.) Place a 1 cm long piece of aluminum wire in the solution in the beaker. Cover the beaker with the watch glass so that the strip of red litmus paper faces into the beaker. Warm the solution very **gently** on a metal ring and wire gauze over a **low** Bunsen burner flame and you will notice the wire reacting. Make sure no solution splatters on to the strip of red litmus paper or the bottom of the watch glass as this will give you a false positive result. If the litmus paper strip changes from red to blue, the original solution contains nitrate ion. **Discard the solution in the aqueous waste container in the hood.**

Test for Cl^- ion:

Add 10-15 drops of Cl^- test solution (or your unknown) to a well on a plastic well plate. Add 2 to 3 drops of dilute $\text{HNO}_3(aq)$ and mix well. Then add 2 to 3 drops of $\text{AgNO}_3(aq)$ (found in the bottle covered in silver tape). Mix and observe the results.

- **Discard the solution in the aqueous waste container in the hood.**

Test for SO_4^{2-} ion:

Add 10-15 drops of sulfate ion test solution (or your unknown) to a well on a plastic well plate. Add, while stirring, 4 to 5 drops of dilute $\text{HCl}(aq)$ or until the solution is acid to litmus. Then add 5 to 7 drops of $\text{BaCl}_2(aq)$. Mix and observe.

- **Discard the solution in the aqueous waste container in the hood.**

Test for PO_4^{3-} ion:

Add about 15 drops of the phosphate ion test solution (or your unknown) to a small test tube. Add about 1 mL of dilute nitric acid and mix well. Then add about 20 drops of $(\text{NH}_4)_2\text{MoO}_4$ solution swirl the test tube in the warm water bath for 30 seconds and observe.

- **Discard the solution in the aqueous waste container in the hood.**

Test for CO_3^{2-} (or HCO_3^-) ion:

Add 10-15 drops of carbonate ion test solution (or your unknown) to a well on a plastic well plate. Add several drops of dilute $\text{HNO}_3(aq)$. Observe the reaction upon addition of each drop.

- **Discard the solution in the aqueous waste container in the hood.**

IN-LAB part 3B

Complete the data tables you prepared for part 3B of the experiment. Be sure to write your observations in detail.

PART 4: SPECTRAL AND CHEMICAL ANALYSIS OF UNKNOWNNS

OBJECTIVE

Your mission, should you decide to accept it(!), will be to identify the presence of the test cation(s) and anion(s) in various unknown samples. **Read this section carefully as there are a lot of points riding on your results.**

PROCEDURE

Part 4

For this section, record your observations in the table you prepared for this section for prelab.

After you have completed all of the positive control tests for the test ions, **show your lab notebook to your instructor**. After initialing the data, your instructor will give you various unknowns to analyze for the presence of the test ions. You will repeat the flame test and the chemical tests you did earlier, however, instead of including the test ion solution you will replace it with an equal amount of your unknown sample in each procedure.

Solid (salt) Unknown:

You will be provided with a solid unknown that is a pure ionic compound. Record the unknown number in your lab notebook. **This compound will be made up of two ions, one cation and one anion**. Using a clean 50 mL beaker, dissolve the sample in about 30 mL of deionized water. Do not contaminate or lose this solution. Your instructor will direct you on how to proceed.

1. Flame Test:

Obtain a nichrome wire assembly and clean it as you did before using 6M HCl(aq). Repeat the cleaning process until the glowing wire imparts no color to the flame. After 5 or 6 cycles the hydrochloric acid may become dirty and need to be replaced. If the wire is excessively dirty, it may be necessary clip off the old loop and prepare another. Do not handle the wire with your hands to avoid getting salt from your skin on the wire. Do not proceed to test your unknown until the wire is absolutely clean, that is, until the glowing wire causes no color to appear in the flame. When clean, wet your wire loop with your unknown solution and insert it in the hottest part of the Bunsen burner flame. Record whether any flame color is observed. If so, describe the color and its intensity and duration. Compare your description with those you gave for the known test solutions. You may repeat a flame test on the test ion solutions if necessary. **Your unknown may or may not contain an ion that responds with a colored flame, if you don't think it does, proceed to the chemical tests.**

2. Chemical Tests:

Using appropriate amounts of your unknown solution perform all of the chemical tests for the cations and the anions as you did before. Record the results of each test on your report sheet. When finished you should have obtained 2 positive tests, one for a cation and one for an anion. Report the two ions to your instructor using a green unknown slip labeled with your name and unknown number written in ink.

Complete all tests to determine your ions before reporting to your instructor. Do not guess at which ions you have as your instructor will deduct points for incorrect ions.

Solid unknowns will not have both ammonium and nitrate ions present. Therefore, if your unknown tests positive for ammonium, you do not have nitrate, so do not test for it.

Solution Unknown:

After checking your solid unknown, your instructor will supply you with a liquid unknown solution. Record the unknown number in your lab notebook. This solution will contain four ions, **two cations and two anions**. Repeat all of the tests as you did with your solid unknown and decide what four ions you have in the solution. Again, **your unknown may or may not contain one cation that responds in the flame test but it will not contain two!** Record the results of each test in your lab notebook. When finished you should have obtained 4 positive tests, two for cations and two for anions. Report the four ions to your instructor using another green unknown slip labeled with your name and unknown number written in ink.

Complete all tests to determine your ions before reporting to your instructor. Do not guess at which ions you have as your instructor will deduct points for incorrect ions.

Solution unknowns will not have both ammonium and nitrate ions present. Therefore, if your unknown tests positive for ammonium, you do not have nitrate, so do not test for it.

IN-LAB part 4

Complete the data tables you prepared for part 4 of the experiment.

POST-LAB

Attach all of the pages of your report together in order, including your signed green unknown slips. Create a summary table of your unknown results followed by a brief discussion of the experiment.

There will be no quiz for this experiment.

Experiment 9

Mass Relationships in a Chemical Reaction

PRE-LAB

Objective: Write a short statement describing the purpose of this experiment.

Background: Define the terms theoretical yield, actual yield, and percent yield.

Procedure: Briefly write/outline what you will be doing in the experiment.

Data: Prepare a data table for all the information you will be collecting.

OBJECTIVE

Measure the amount of carbon dioxide and sodium chloride produced by the reaction of a measured amount of sodium bicarbonate with hydrochloric acid. Compare this result with the theoretical amount.

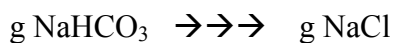
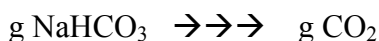
BACKGROUND

A law of chemistry states that a given chemical reaction will produce a certain definite amount of products from a given amount of reactants. If only one chemical reaction occurs, and if the reaction is complete, the amounts of products can be calculated from the amounts of reactants. This amount of product formed is called the **theoretical yield**, which is the maximum amount of product that can form.

Sodium bicarbonate and hydrochloric acid react completely to form carbon dioxide, sodium chloride, and water, according to the following equation:



If we have a known amount of pure sodium bicarbonate and it reacts with sufficient hydrochloric acid, we should be able to predict the amount of products formed. Knowing the balanced equation we can use the following road maps to calculate the theoretical yields:



We want to verify these calculations by experiment. We must measure the mass of the CO_2 from the weight loss of the reaction mixture, because CO_2 is a gas. The mass of the sodium chloride can be measured by evaporating the water and weighing the residue. This mass of CO_2 and NaCl that we will determine experimentally are called the **actual yields**. We cannot measure the water product because we are adding a lot of water with the HCl and there is no way to differentiate the product from the solvent. After obtaining the experimental data we will calculate the percent yield of each product.

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

PROCEDURE

Tare a 150 mL beaker containing a clean, dry large test tube on the balance. Remove the test tube and carefully add about 1.0 to 1.6 g of NaHCO_3 to it. Place it back in the beaker and record the new mass in your lab notebook. Go back to your bench and add a few milliliters of water from your wash bottle to wash down any salt on the sides of the test tube. This step is necessary so that there is enough water in the test tube for any produced NaCl to remain dissolved.

Fill a clean, dry 50 mL beaker with about one inch of dilute $\text{HCl}(aq)$ and place one of your clean plastic pipets in the beaker. Return to the balance with both beakers and place the 50 mL beaker (with HCl and pipet) next to the 150 mL beaker (with test tube, NaHCO_3 , and water) on the balance. Weigh and record the mass of everything together.

Return to your bench with the beakers. Remove the test tube and add a drop or two of hydrochloric acid from the beaker to it. Mix the contents by tapping the test tube and note the bubbles of CO_2 escaping. Continue adding the HCl slowly (so that the solution does not bubble over) with tapping, until no more bubbles are observed. Replace everything on the balance and reweigh. The difference between this weight and the previous one is the mass of CO_2 lost.

Weigh and record the mass of a large watch glass.

Set up a boiling water bath by filling your 250 mL beaker two thirds with water. If not already done for you, set up a snorkel device above the water bath. Place the large watch glass on top of the boiling water bath and carefully pour the contents of the test tube into the watch glass. Rinse the test tube with a few milliliters of water from your wash bottle and add the rinsings to the watch glass. Evaporate the solution in the watch glass to dryness by placing it on a beaker of boiling water. Make sure that it is dry. When all the water has evaporated, turn off the flame and let the set up cool. Dry the bottom of the watch glass and record its mass. (How will you determine the mass of the NaCl ?) While the first sample is boiling **repeat** the experiment with a second sample.

IN-LAB

Data: Complete the data table you prepared.

Calculations: Perform the calculations asked for below.

CALCULATIONS

- 1) Calculate the theoretical mass in grams of :
 - a) CO_2 that should be obtained from your first sample of NaHCO_3 .
 - b) NaCl that should be obtained from your first sample of NaHCO_3 .

- 2) Calculate the actual mass in grams of :
 - a) CO_2 you obtained after the reaction of sample one.
 - b) NaCl you obtained after the reaction of sample one.

- 3) Calculate the percent yield for:
 - a) CO_2 for sample one.
 - b) NaCl for sample one.

- 4) Repeat calculations 1 – 3 for the second sample.

POST-LAB

Results: Create a summary results table

Conclusion: Analyze the results and errors in a written discussion. Provide a written analysis of the results and explain the sources of error that could have led to an increase or a decrease in the calculated percent yield for each sample.

What you need to know for Experiment 9 Quiz

- Calculating percent yield and theoretical yield, given mass of reactant and actual yield.

Sample Quiz 9

Consider the balanced equation: $2 \text{Na} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2$

1. If 35.0 g of Na is reacted with excess H_2SO_4 , what is the theoretical mass of H_2 produced?

2. If 1.1g of H_2 is actually produced, what is the percent yield of H_2 ? _____

Experiment 10

Titrations and Solution Stoichiometry

PRE-LAB

Objective: Write a short statement describing the purpose of *each part* of the experiment.

Background: Define molarity, equivalence point, and end point.

Procedure: Briefly write/outline what you will be doing in *each part* of the experiment.

Pre-lab Questions: Answer the three questions below.

Data:

Part 1

Prepare a data table for the titration of NaOH with KHP. Include enough data columns to do at least 5 titrations.

Part 2

Create a data table for the titration of your unknown HCl with NaOH. Include enough data columns for at least 3 titrations.

Part 3

Prepare a data table for the titration of NaOH with diluted vinegar. Include enough data columns for at least 3 titrations.

PRELAB QUESTIONS

1. Calculate the molar mass of Potassium acid phthalate, $\text{KHC}_8\text{H}_4\text{O}_4$ (KHP). You will use this number to do the calculations for part 1 once you do the titration.
2. Calculate the molar mass of acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$. You will use this number to do the calculations for part 3 once you do the titration.
3. Solve the following problem:

If it takes 35.04 mL of a 0.2455 M NaOH solution to neutralize 25.00 mL of an unknown HCl solution, what is the molarity of the HCl solution?

Use the roadmap given under part 2 to calculate this problem.

PART 1: PREPARATION AND STANDARDIZATION OF NaOH(aq)

OBJECTIVE

Part 1

To prepare and determine the molarity of an NaOH solution.

BACKGROUND

Molarity (M) is the relationship between the moles of solute and the volume, in liters, of the solution. A 0.20M NaOH solution contains 0.20 mole of NaOH per 1 liter of solution. It follows that 500 milliliters of solution contains 0.10 mole of NaOH.

$$500 \text{ mL soln.} \cdot \frac{1 \text{ L soln}}{10^3 \text{ mL soln}} \cdot \frac{0.20 \text{ mols NaOH}}{1 \text{ L soln}} = 0.10 \text{ mol NaOH}$$

To make up 500 milliliters of 0.20M NaOH we must obtain 0.10 mole of NaOH and dilute it with sufficient water to make approximately 500 mL.

Sodium hydroxide, when pure, is a solid in the form of pellets or flakes, and it is difficult to weigh small quantities precisely because it absorbs CO₂ and water from the air. It is more convenient to make a large amount of a stock solution, and measure small quantities of the stock solution. The stock solution we work with is the bench top dilute NaOH(aq) which is approximately 6.0 molar, i.e. it contains 6.0 moles of NaOH per liter of solution. The volume of stock solution needed can be calculated by:

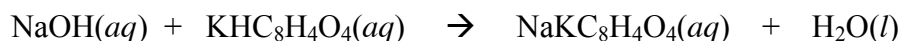
$$0.10 \text{ mol NaOH} \cdot \frac{1 \text{ L NaOH soln}}{6.0 \text{ mols NaOH}} \cdot \frac{10^3 \text{ mL NaOH soln}}{1 \text{ L NaOH soln}} = 16.67 = 20. \text{ mL NaOH soln}$$

To make 500 milliliters of 0.20M NaOH solution from 6.0M NaOH solution, we should take about 20 milliliters of stock solution and dilute it with sufficient water to make 500 milliliters and store it in a labeled plastic bottle.

For analytical purposes we need to know the concentration of our solution to at least **three significant figures**. The process of finding the accurate concentration is called **standardization**. To standardize a solution we must be able to measure something that is related to the number of moles of solute, and relate it to the volume of solution.

We want to know both the moles of solute and volume of solution with high accuracy to calculate the molarity.

Potassium acid phthalate, $\text{KHC}_8\text{H}_4\text{O}_4$ (frequently referred to as KHP), is a chemically pure salt that can be weighed accurately. It is an acid salt but only the one hydrogen written in the front of the formula is acidic (dissociates as H^+), the rest do not. The acid reacts with NaOH through a double-displacement (neutralization) reaction in a 1:1 mole ratio.



We can weigh a quantity of KHP accurately and compute the number of moles from its molar mass. Using the balanced reaction and the moles of KHP we use we can determine the moles of NaOH in the solution that we used to react with it. If we measure the volume of NaOH solution we used, we can calculate the molarity of the NaOH solution:

$$M_{\text{NaOH}} = \frac{\text{mols of NaOH}}{\text{L of NaOH}(aq)}$$

The process of adding a solution of one reagent (the titrant, in this case $\text{NaOH}(aq)$) to another until the amount added is chemically equivalent to the second reactant and no excess of either chemical is present is called **titration** (Figure 10.1). To be practical, it is necessary that the reaction be fast and complete and free of side reactions. It is also necessary that something signal when the **equivalence point** is reached. The equivalence point in the titration is when the number of moles of titrant stoichiometrically balances exactly the number of moles of the other chemical, in this case KHP. So that the analyst, you, knows when to stop titrating an indicator dye is often used in acid-base titrations. The dye changes color when the reactant to be analyzed is used up. The point in the titration at which the dye changes color is called the **end point**. Indicators must be chosen with some care so that the equivalent point and the end point are at the same place.

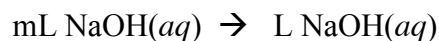
In this reaction we will use an indicator dye that we are familiar with, phenolphthalein. Recall that phenolphthalein changes from colorless to pink when a slight excess of base is present. We will add the sodium hydroxide solution slowly, while swirling, to a flask containing a measured amount of KHP dissolved in water and a few drops of phenolphthalein solution. When a pink coloration just appears and remains in the solution, the titration is stopped. The volume of sodium hydroxide solution added is recorded, and the concentration of sodium hydroxide is calculated for each trial.

Road maps to find molarity of $\text{NaOH}(aq)$:

1. Calculate moles of NaOH added:



2. Calculate liters of $\text{NaOH}(aq)$ added:



3. Divide moles of NaOH / Liters $\text{NaOH}(aq)$ to calculate the molarity of your $\text{NaOH}(aq)$ solution. Make sure to keep track of the correct number of significant figures!

THE CARE AND NURTURING OF A BURET

A buret is designed so that a relatively large change in level of the liquid is produced by a small change in the contained volume. The inside surface of the buret must be very clean so that the film of liquid on the glass never breaks when the liquid is withdrawn. Handle the buret carefully- it is fragile and expensive.

- Obtain a buret from the rack.
- Assume that the buret must be cleaned before you use it.

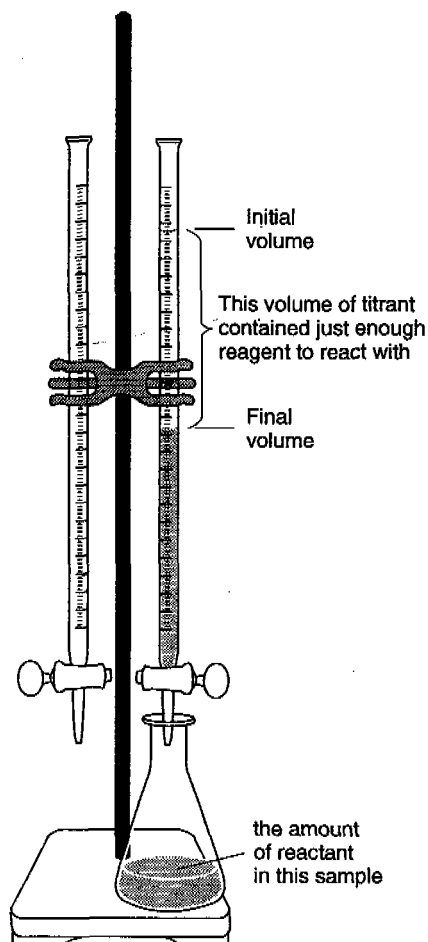


Figure 10.1

- When filling your buret, always hold it so that its top is below your eye level. Close the stopcock. Fill the buret to the top with tap water. Allow it to run through the tip. Discard this rinse water down the drain. Repeat.
- Now rinse the buret in a similar fashion twice with distilled water, and discard the rinse down the drain.
- Last, close the stopcock and add about 15 mL of your titrant (the solution to be placed to in it). Turn the buret on its side, and rotate it to rinse all of the inside surface. Discard the rinse in a waste beaker by allowing it to run through the tip. Repeat with a second 15 mL of titrant.

- Finally, fill the buret with the solution to be used using a clean funnel. Hold the buret below your eye level and fill it slowly until it is **above** the 0.00 mL mark.
- Allow the solution to run out through the stopcock into your waste beaker until the level goes **on** or **below** the 0.00 mL mark. Make sure there are no air bubbles in the tip of the buret. If there are air bubbles, remove them by opening the stopcock and allowing solution to drain into the waste beaker. Remove the funnel from the buret and place it on a clean paper towel on your benchtop.
- Allow approximately 30 seconds for drainage. Read and record the level in the buret, **estimating to the nearest 0.05 mL.**
- With the reaction vessel held below the tip of the buret, open the stopcock and allow the solution to run into the vessel. Your instructor will demonstrate the correct technique of titration.

At the end of the period:

- Discard the solution in the buret into your waste beaker. Do not return it to your stock bottle as it may be contaminated.
- If you used base in your buret, neutralize the base first by rinsing with a very dilute HCl solution. To make this dilute HCl solution, half fill a 150 mL beaker with deionized water and add 10 drops of 6 M (benchtop) HCl. Pour this solution through your buret and collect it in your waste beaker. Then rinse the buret twice with tap water and once with deionized water. The rinse water can be poured down the drain.
- Return the buret *carefully* to the rack with the stopcock open.

PROCEDURE

Part 1

Preparation of NaOH solution

While wearing your safety eyewear, measure about 20 mL of 6.0 M NaOH in a graduated cylinder. Transfer the NaOH solution to a 500 mL plastic bottle. Now add enough deionized water to reach almost to the top of the bottle (about 500 mL). Cap the bottle and mix thoroughly. The solution will change composition if it is exposed to air, so keep the bottle capped when you are not withdrawing samples. Label the bottle with your name and the contents.

Standardization of the NaOH solution

Put on your safety eyewear and clean three Erlenmeyer flasks and label them 1, 2 and 3. Measure between 1.0 and 1.5 g of KHP in each flask and record the results in your lab book (if you have less than 1 gram, add a little more and weigh again). Add 30 to 40 milliliters of deionized water to each flask and dissolve the KHP samples. **MAKE SURE ALL THE KHP IS DISSOLVED BEFORE YOU PROCEED.**

Place a clean 50-milliliters buret in the right side of a buret clamp and, after cleaning, rinse with a small quantity of the dilute NaOH solution to be standardized. Allow some to flow through the tip into your waste beaker. Pour out the rest. Fill the buret with the NaOH solution to be standardized, using a funnel. Fill it **above** the 0.00 mL mark. Open the stopcock and allow some NaOH to run through the tip to be sure that there are no bubbles remaining in the tip and that the

level in the buret is **below** the 0.00 mL mark. Read and record the initial reading of the buret to the nearest 0.05 mL (the bottom of the meniscus shows up better if a white card is held in back of the buret). You may need to remove the buret from its holder to be sure that your line of sight is parallel with the level of the meniscus. Remember that the buret calibrations are read from the top down (see figure 10.1). Check the reading with your instructor if necessary.

Add 2-3 drops of phenolphthalein solution to each of the samples to be titrated. The technique of titration is to manipulate the stopcock of the buret with your left hand (if you are right-handed) and use your right hand to continuously swirl the contents of the Erlenmeyer flask. After your instructor has demonstrated the proper way to carry out a titration (or shown you a video), titrate your first sample until a permanent faint pink color appears in the solution. Erlenmeyer flasks are constructed the way that they are so they can be swirled without slopping any solution out of the top. Be sure to rinse down the inner surfaces of the flask with deionized water from your wash bottle several times during the titration so that any titrant that splashes onto the walls is washed down into the bulk solution. Do not use a stirring rod to stir your solution (why?). The approach of the end point is signaled by a pink color that appears where the solutions are being mixed. The closer the end point, the more the pink color will persist. Finally, one drop of titrant will produce a pink color that spreads throughout the solution and persists for at least one minute. That is the end point of the titration. Read and record the final volume of the buret to the nearest 0.05 mL.

Titrate samples 2 and 3 in the same manner. Put all of the data for each flask that you titrated into a clearly labeled table in your lab report. Complete your calculations for the molarities of $\text{NaOH}(aq)$ before you clean your buret as you may need to do another titration(s). Save the NaOH solution for parts 2 and 3 of this experiment. Dispose of your titrated solutions into the “aqueous waste” container.

IN-LAB Part 1

Data: Complete the data tables you prepared for part 1

Calculations: Perform the calculations below.

CALCULATIONS

Part 1

1. Calculate the molarity of the NaOH solution by using the data from each titration. Show the results to your instructor. If the 3 calculated molarities differ from each other by more than 0.005M, you will be asked to repeat the titration until you get more precise results. Repetitively obtaining consistent results is a good indication that the result is accurate. (This is not always true however. Under what conditions might it not be?) When you do get consistent results your instructor will sign off on them.
2. Calculate an average using only those values that are precise. This average value will be the standardized molarity of your NaOH solution and it will be used for the rest of this experiment.

PART 2: DETERMINING THE ACCUARATE MOLARTIY OF AN UNKNOWN HCl SOLUTION

OBJECTIVE

Part 2

To determine the molarity of an unknown HCl solution using the molarity of the NaOH found in part 1.

BACKGROUND

Part 2

The accurate concentration of the unknown HCl solution can be determined by titrating a carefully measured volume of HCl(aq) with your standardized NaOH(aq). Write the balanced equation for this reaction in your lab book. To calculate the molarity of HCl you will use:

$$M_{\text{HCl}} = \frac{\text{mols of HCl}}{\text{L of HCl}(aq)}$$

The moles of HCl are related to the measured volume of the titrant, NaOH, used; by its molarity and by the mole ratio in the equation. The volume of HCl(aq) was determined by the titration.

Road maps:

1. Calculate moles of HCl added:
 $\text{mL NaOH}(aq) \rightarrow \text{L NaOH}(aq) \rightarrow \text{mol NaOH} \rightarrow \text{mol HCl}$
2. Calculate liters of HCl(aq) added:
 $\text{mL HCl}(aq) \rightarrow \text{L HCl}(aq)$
3. Divide moles of HCl / Liters HCl(aq) to calculate the molarity of your HCl (aq) solution. Make sure to keep track of the correct number of significant figures!



Pipetting with your mouth may make your teeth fuzzy and cheeks pucker up. Use a suction bulb.

PROCEDURE

Put on proper eyewear. Your instructor will give you about 100 mL of an unknown HCl solution. **Record the unknown number.** Obtain a 25-mL volumetric pipet to measure the volume of the HCl solution. Clean and rinse it as well. Don't forget to do a final rinse with a small amount of your unknown HCl solution. Be sure to run some through the tip of your pipet before filling.

Construct a clearly labeled data table in your lab book to accommodate the data that you will record from this part. Use a 25 mL pipet to carefully measure 25.00 mL of your HCl solution into the three flasks. Add 2 to 3 drops of phenolphthalein indicator solution to each flask.

Fill the NaOH buret with your standardized NaOH solution in the manner described previously. Titrate each of the HCl solutions to a permanent pink end point, recording the initial and final volume readings of the NaOH buret for each titration. As before, put all of the data for each flask that you titrated directly into your data table. Complete your calculations before you clean the buret and pipet as you may need to do additional titration(s). Save both the NaOH and HCl solutions. Your titrated solutions can be disposed of down the drain as they are neutral and do not contain any hazardous chemicals. When complete, discard any excess base or acid waste in the appropriate waste container.

IN-LAB part 2

Data: Complete the data tables you prepared for part 2

Calculations: Perform the calculations below.

CALCULATIONS

Part 2

Calculate three molarities for the HCl solution by using the data from each titration. As in part 1, you are looking for three precise results to ensure that the molarity you have calculated for HCl is accurate- and that is important because you will be graded on your accuracy. See your instructor if you have questions. If you feel that your results are too far apart, do another titration until you get three consistent results. You will be given enough unknown solution to do at least four titrations. Do not waste it or contaminate it- if you need more unknown you will lose credit. After confirming all of your data with your instructor, average the three most precise results. This value will be the molarity of your unknown HCl solution.

PART 3: DETERMINATION OF MASS % OF ACETIC ACID IN VINEGAR

OBJECTIVE

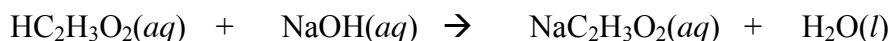
Part 3

To determine the mass percent (% composition by mass) of acetic acid in vinegar

BACKGROUND

Part 3

Acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) is the sour component in vinegar. Acetic acid is a weak organic acid and, just as was the case for potassium acid phthalate, only one of its hydrogens (written in the front of the formula) is acidic and therefore will react with base.



Assuming that the only acidic component in vinegar is acetic acid, we can titrate the acetic acid in a carefully measured mass of diluted vinegar with standardized NaOH solution to determine the mass of acetic acid present. Using the dilution factor of four (see below) we can calculate the percent mass (also known as % composition by mass) acetic acid in the undiluted vinegar. Common household vinegar is usually about 5% by mass acetic acid.

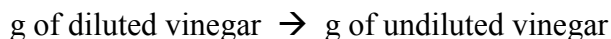
$\% \text{ composition}_{\text{acetic acid in vinegar}} = \frac{\text{mass of acetic acid (HC}_2\text{H}_3\text{O}_2\text{)}}{\text{mass of vinegar}} \times 100$

Road maps:

1. Calculate mass of acetic acid reacted:



2. Calculate mass of undiluted vinegar titrated:



3. Divide mass of acetic acid by mass of undiluted vinegar, and multiply by 100. Make sure to keep track of the correct number of significant figures!

Note: The vinegar sample you are titrating has been diluted 3 parts water to 1 part vinegar. The ratio of diluted vinegar to vinegar is therefore:

$$\frac{4 \text{ parts diluted vinegar}}{1 \text{ part vinegar}}$$

PROCEDURE

Part 3

Outfit your protective eyewear. Using a clean, dry 50 mL graduated cylinder, obtain a 20-25 mL sample of diluted vinegar. Pour it into a tared Erlenmeyer flask labeled I, place it on the balance and record the mass into your notebook. Do the same for two other flasks labeled II and III. Add 2 to 3 drops of phenolphthalein and, as before, titrate each flask with your standardized NaOH solution to a stable pink endpoint. Again, put all of the data for each flask that you titrated into a clearly labeled table in your lab notebook. Show the calculations to your instructor. After titrating your samples, they can be disposed of down the drain.

IN-LAB part 3

Data: Complete the data table you prepared for part 3

Calculations: Perform the calculations below.

CALCULATIONS

part 3

Calculate the mass percent of acetic acid in vinegar for each titration to the proper number of significant figures. Show the results to your instructor. Repeat the titration until consistent results are obtained. Average the three precise results. You will be graded on the accuracy and precision of your results.

POST-LAB

(For the entire experiment)

Results: Create a summary of results table and put the important results from all three parts of this experiment into it.

Discussion: Write a complete discussion of your results from this experiment. Include comments on the precision and accuracy of your results and any sources of error. You will be graded on the thoroughness and thoughtfulness of your analysis.

What you need to know for experiment 10 Quiz

- Calculating the molarity of an unknown solution via titration.

Sample Quiz 10

A 1.132 g sample of KHP ($\text{HKC}_8\text{H}_4\text{O}_4$) required 28.05 mL of NaOH solution to reach endpoint. Calculate the molarity of the NaOH.

APPENDIX

Laboratory Safety Rules

Your laboratory experience will be pleasant if everyone in lab observes courteous behavior and common sense. Common sense is easier to follow if you know what you are doing and why you are doing it. The following points are offered as reminders of good laboratory manners.

1. **Protect your eyes with approved protective eyewear at all times.** They must be worn at all times while in the laboratory. If you wear prescription glasses, the safety eyewear must be worn over your glasses.
2. **Dress appropriately for lab:** Suitable closed toed protective footwear is required at all times - sandals are unsafe. Tie long hair back.
3. **Let nothing touch your mouth, lips, or any other part of the face while in lab.** Food, drinks, and gum can too easily become contaminated by toxic (poisonous) substances and are not allowed in the lab. Avoid touching your face with your hands- if you must, wash them first. Do not put your nose directly over a container to smell the contents; gently waft some vapors with your hand toward your nose, instead. Wash your hands before you leave lab to avoid contaminating your food.
4. **Place all backpacks, jackets and books not being used in a cubby** on the back wall. Keep your work area free from all unnecessary clutter. Keep isles free of obstruction.
5. **Be aware of the locations of all safety equipment.** Find out where fire extinguishers, fire blankets, safety showers, and eye-wash fountains are and know when and how to use them. In case of chemical spills on skin or clothing, wash the affected area immediately with plenty of water; then tell the instructor.
6. **Follow the directions in your lab manual.** Unapproved variations in experiments are dangerous and should not be performed. You are required to carefully read the lab experiment being performed and prepare their notebooks properly before coming to lab. Pay attention to all the safety information your lab manual provides. Any questions should be addressed to the instructor before a lab begins. Use common sense and ask your instructor whenever you are unsure. If you think you have a better way of doing an experiment, discuss it with the instructor. Please do not experiment on your own yet. Never carry out any experiments in the absence of the instructor.
7. **Notify your instructor before lab of any chemicals you are allergic to.**
8. **Never leave an experiment unattended.** Always keep an eye on a lit Bunsen burner and make sure the area is safe before lighting a burner. When finished, always turn the burner off at the blue stopcock valve on the bench and not at the needle valve on the base of the burner. This valve is to regulate the size of the flame.
9. **Report all accidents to your instructor.**
10. **Practice safe techniques.** All glass tubing and rods, including stirring rods, should be free of cracks and sharp edges. Have such items replaced by your instructor. Thermometers should

not be used for stirring since the bulb end is too easily broken. Do not put your nose directly over a container to smell the contents; gently waft some vapors with your hand toward your nose, instead. When not being used, place lab stools into the knee space under your bench.

11. **Do not use any electronic devices**- cell phones, texters, ipods, etc.- while in lab. Focus and awareness is important in the laboratory.

12. **Take only as much chemical as you are instructed** from the stock bottles. **Never** return unused chemicals to a stock bottle; you might contaminate the rest of the contents of the bottle.

13. **Use the fume hood or lab snorkels** for reactions which release toxic or noxious gases- your lab manual will direct you when these are to be used.

14. **Keep your work area and the common work areas clean and dry.** You will discover that spills and breakage occur less often if you keep your glassware at the back of the bench. Look before you reach. Always return common equipment, like balances or shared reagent bottles to their original location and clean them when you are done.

15. **Read the posted instructions before using electronic balances** since they are expensive and delicate instruments. Never add chemicals to a container on a balance; remove it first, add the chemical and then return it for reading.

16. **Dispose of chemicals properly.** For your health and that of the environment, it is important that you follow the instructions given in your lab manual on where to dispose of used chemicals. Ask your instructor if you are unsure.

17. **CLEANUP:** Return all of your glassware and lock your lockers at the end of lab. Return all common equipment (rings, test tube clamps, wire gauze, burets, etc.) to where they belong. Clean your work area completely before leaving! Wipe down your bench with a wet paper towel. Wash hands with soap and water after working in the laboratory.

18. **Watch the safety video, take the safety quiz, and read the student academic integrity policy statement** before beginning any experiments. All students will be required to watch a safety video at the beginning of the semester and take a safety quiz that will be kept on file by the instructor.



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

Student Academic Integrity Policy Statement

The faculty and administration of Los Angeles Pierce College are committed to the belief that honesty and integrity are integral components of the academic process. The College expects students to be honest and ethical at all times in their pursuit of academic goals. Students who violate the code of academic conduct by which the College maintains its academic integrity will be dealt with in a manner reflecting the seriousness of these violations.

A. Violations of academic honesty and integrity occur when a student participates in any act in which he/she uses deception or fraud while performing an academic activity. Violations include, but are not limited to, the following:

1. Using study aids such as calculators, notes or electronic devices when not authorized by the instructor.
2. Cheating on examinations, assignments or experiments (allowing another student to copy one's answers or copying the answers of other students; exchanging information by any means, including verbal exchanges, sign language, hand signals, secret codes, electronic devices, passed notes, creation of a distraction for the purpose of cheating; changing answers on a previously scored test, assignment or experiment; inventing information and/or data.
3. Allowing another student to assume one's identity in order to fulfill an assignment or take a test.
4. Submitting for a grade the words, ideas, and/or written work (including laboratory notes and drawings) of another person without giving due credit to that person. This includes purchased papers or papers written by other students.
5. Falsifying or attempting to falsify attendance records and/or grade rosters.
6. Conspiring with other students to commit any of the above behaviors.

B. Disciplinary actions against a student who commits any offense violating academic honesty and integrity may include:

1. An "F" or a "0" on the examination or assignment.
2. A record of the student's violation sent to the Vice President of Student Services and placed in the student's disciplinary file.
3. Additional sanctions and/or penalties authorized by the Board of Trustees for violations of the District's Student Code of Conduct (Board Rule 91101.11) may include: restitution charges for damaged or misappropriated property; disciplinary probation from college activities or services; suspension from the College; expulsion from the College.

C. Student's Right to Appeal

Students have the right to appeal disciplinary actions through the Board of Trustees Discipline procedures or the student grievance procedures which are outlined in Administrative Regulation E-55.

D. Reporting a Violation

When an alleged incident of academic dishonesty occurs, the chemistry department **requires** that a faculty member take the following steps to report the incident:

1. Inform the student and the department chair of the nature of the alleged violation and the impending course of action.
2. Complete the Academic Dishonesty Report Form and submit it, along with any related evidence, to the Dean of Student Services. The student should also receive a copy of the form from the instructor within ten (10) working days of the incident.
3. The Dean of Student Services will forward information about the incident to the Department Chair and the appropriate Dean of Academic Affairs.
4. The Dean of Student Services will investigate the allegations and recommend any appropriate disciplinary actions, beyond actions taken by the faculty member specific to course grading.

E. Faculty Responsibilities

In order to maintain an environment free of academic dishonesty, the following recommendations are made to the faculty regarding their responsibility to uphold academic integrity:

1. Make every attempt to conduct their classroom in a manner which encourages honorable learning, to ensure student success and discourage academic dishonesty.
2. Inform students of the course requirements, grading procedures and expectations of responsible academic conduct.
3. Inform students of the College policy on Academic Integrity and the potential consequences for violations of this policy.
4. Inform students of their right to due process should they wish to contest the cheating allegation.

Student Discipline Procedures

Community college districts are required by law to adopt standards of student conduct along with applicable penalties for violation (Education Code Section 66300). The Los Angeles Community College District has complied with this requirement by adopting Board Rule 9803, Standards of Student Conduct. The purpose of this Board Rule is to provide uniform procedures to assure due process when a student is charged with a violation of these standards. All Proceedings held in accordance with these procedures shall relate specifically to an alleged violation of the established Standards of Student Conduct.

These provisions do not apply to grievance procedures, student organization councils and courts, or residence determination and other academic and legal requirements for admission and retention. Disciplinary measures may be taken by the College independently of any charges filed through civil or criminal authorities, or both.

Copies of the Student Discipline Procedures are available in the Student Services Office in the Student Services building.

Student Grievance Procedures

The Student Grievance Procedure is to provide a prompt and equitable means for resolving student(s) grievances. The grievance procedure may be initiated by a student or group of students who reasonably believe that he/she/they have been subject to unjust action or denied rights that adversely affect his/her/their status, rights, or privileges as a student. The grievance procedure is detailed in District Administrative Regulation E-55 which is available in the Student Services Office in the Student Services building to any student or applicant to the college.

The grievance procedure does not apply to the challenge process for prerequisites, corequisites, advisories and limitations on enrollment; alleged violations of sexual harassment; actions dealing with student discipline; alleged discrimination on the basis of ethnic group identification, religion, age, sex, color, sexual orientation, physical or mental disability; or an appeal for residency decision; or to eligibility, disqualification or reinstatement of financial aid.

In addition, section 76224 of the California Education Code provides:

“When grades are given for any course of instruction taught in a community college district, the grade given to each student shall be the grade determined by the instructor of the course and the determination of the student's grade by the instructor, in the absence of mistake, fraud, bad faith, or in competency, shall be final.”

The following steps should be taken to begin the grievance procedure:

Step I: Informal Resolution

All parties involved in a potential grievance should be encouraged to seek an informal remedy. The student shall make a reasonable effort to resolve the matter on an informal basis by:

1. meeting with the person with whom the student has a grievance.
2. meeting with that person's immediate supervisor (chair).
3. meeting with the College administrator of the area.
4. meeting with the College Ombudsperson to explore student rights and responsibilities, receive assistance with an informal resolution, and submit a written "Statement of Grievance."

Step II: Formal Resolution

Students unable to resolve their grievance through the informal process may file a "Formal Grievance Hearing Request Form" with the College Ombudsperson. The Ombudsperson will provide students with information about the formal grievance hearing process and their rights and responsibilities in this process.

Students pursuing a formal grievance have the right to be represented by a Student Advocate who will assist students in the formal grievance process.

Additional information and assistance with the Student Grievance Procedure may be obtained from the Student Services Office in the Student Services building.

Ombudsperson

Pursuant to the Student Grievance Procedure, the College Ombudsperson(s) has been appointed by the College President to assist the student in obtaining informal resolution of a grievance.

Each student shall make a reasonable effort to resolve a concern on an informal basis prior to requesting a grievance hearing. If an informal resolution is not obtained, the Dean of Student Services will arrange for the formation of a Grievance Hearing Committee according to District Administrative Regulation E-55. Information concerning the Ombudsperson may be obtained in the Student Services Office located in the Student Services building.

Answers to Sample Quizzes

Quiz 1

1. Graduated cylinder
2. Infinite (∞), 7
3. 6.9
4. False, false
5. Data

Quiz 2

1. 15.5 mL
2. Intensive, extensive
3. 5.684 g
4. 0.31 g/cm³
5. Higher

Quiz 3

1. Solute
2. 27.9 mL H₂O (15.0 mL alcohol)
3. a) 3.557 g b) 1.45 g/cm³

Quiz 4

1. a) ii. Immiscible b) j. heterogeneous mixture
2. a. exothermic
3. a) soluble b) insoluble

Quiz 5

1. Extraction
2. a) 56.03% b) 82.82%

Quiz 6

1.

Sample of Matter	Pure Substance				Mixture		Basic Unit (if applicable) atom, molecule, or formula unit
	Element		Compound		Homo-	Hetero-	
	Metal	Nonmetal	Ionic	Covalent			
CuSO ₄ •5H ₂ O			x				Formula unit
C		x					Atom
H ₂		x					Molecule
KI(aq)					x		

2. Ductile

Quiz 7

1. a, c, f

2. Br, Cl

3. CIE: $\text{Ba}^{2+}(aq) + 2\text{NO}_3^-(aq) + \text{Cu}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) + \text{Cu}^{2+}(aq) + 2\text{NO}_3^-(aq)$
 NIE: $\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s)$

Quiz 81. $\text{C}_2\text{H}_6\text{O}$ P_2O_5 2. $\text{CuCl} \bullet \text{H}_2\text{O}$ **Quiz 9**1. 1.53 g or 1.54 g (Depending on if 2.016g or 2.02g for H_2 Molar Mass used.)

2. 71% or 72%

Quiz 10

0.1976 M

LABORATORY SAFETY QUIZ

Name (print) _____

Course Number

Section

Instructor

I have watched the Chemistry safety video, read the Laboratory Safety Rules and the College Academic Integrity policy and understand them. I also understand that California State Law requires that I wear safety eyewear at all times while I am in the laboratory.

Signed _____ Date _____

Circle True (T) or False (F) to the following questions:

- T F (1) Experiments other than those assigned may be performed by the student, if there is free time available.
- T F (2) When in doubt about how to do something, ask your instructor.
- T F (3) A thermometer is designed so that it may be used as a stirrer.
- T F (4) Food may be eaten in the lab if class is not in progress.
- T F (5) All spilled chemicals must be cleaned up as soon as possible.
- T F (6) Books, backpacks, drinks and clothing should be placed in the cubbies at the back of the room at the beginning of lab.
- T F (7) Long hair need not be tied up if you are not using a burner.
- T F (8) If you have taken too much of a chemical, you may return the excess to the container.
- T F (9) The correct way to turn off a burner is to close the stopcock at the gas line (blue tab).
- T F (10) You may taste chemicals if it is suggested that you do so.
- T F (11) If you cut your finger on broken glass, report it to your instructor.
- T F (12) A burner may be left turned on when you leave the room if someone else will look after it.
- T F (13) Shoes must be worn at all times in the laboratory.
- T F (14) The proper way to sniff fumes is to bring the container up close to your nose.
- T F (15) You are responsible for cleaning up the tops of the tables at all times, especially at the end of each lab period.
- T F (16) Small amounts of acid or base can be poured into the sink.
- T F (17) The windows of the fume hoods should be closed when hoods are not in use.

Chem 60 Locker Item List

Quantity	Item	in	out
4	Beaker, 50 Ml		
2	Beaker, 150ml		
2	Beaker, 250ml		
3	Erlenmeyer flasks, 125ml		
1	Evaporating dish (ceramic)		
1	Funnel, narrow stem, 75mm		
1	Graduated cylinder, 10ml		
1	Graduated cylinder, 50ml		
1	Metal sample (aluminum cylinder)		
1	Plastic wash bottle with spout, cap and tube, 125 ml		
1	Stirring rod, (Plastic or Glass) with Rubber Policeman		
1	Test tube holder		
1	Test tube rack, half size		
8	Test tubes, 15 x 125mm (medium, 18 ml capacity) in rack		
2	Test tubes, 18 x 150mm (large, 27ml capacity)		
2	Transfer pipets, plastic		
1	Watch glass, 3"		
2	Watch glass, 4"		
1	Well plate, plastic		

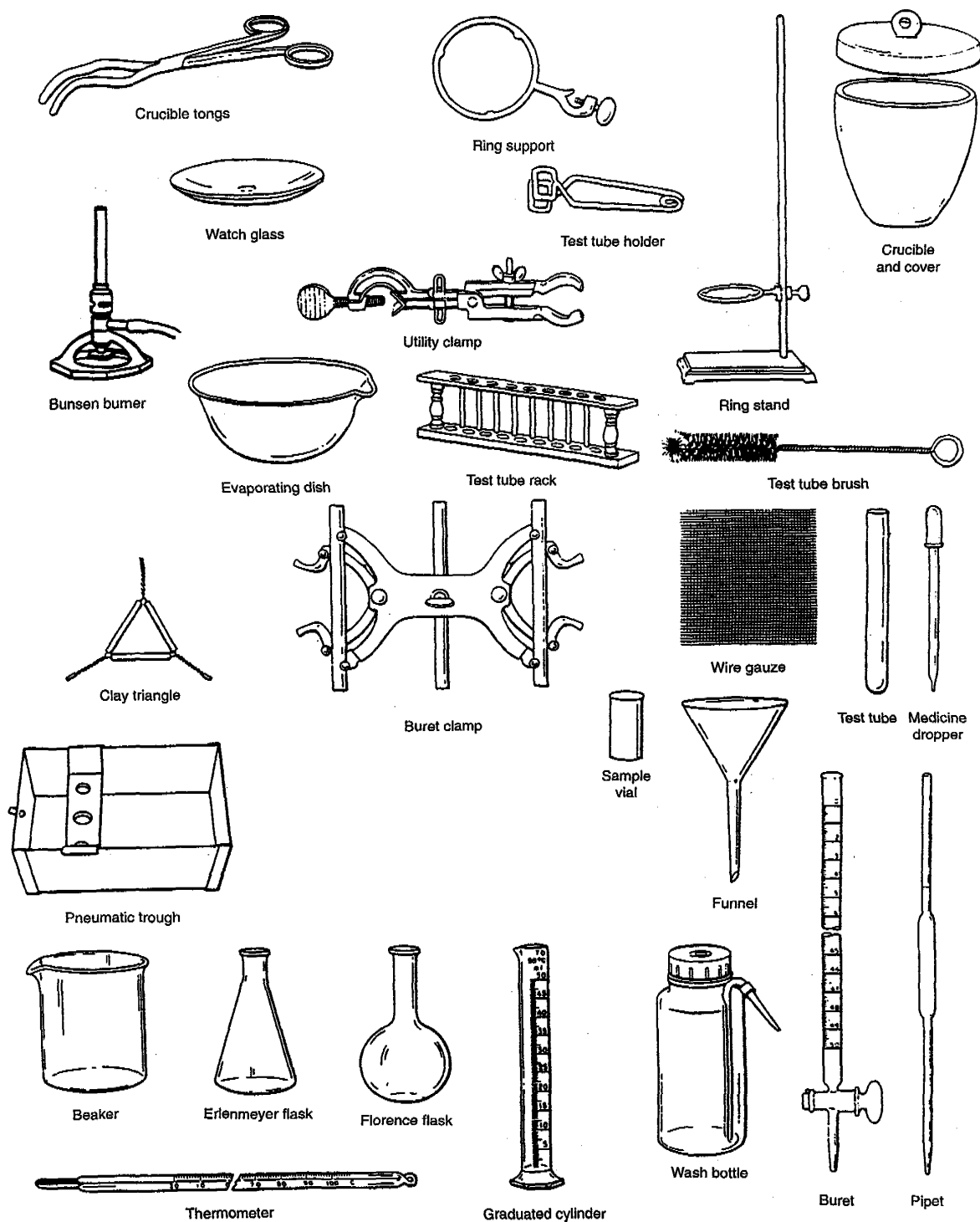
Note:

1. These are the only supplies that you should keep in your lockers. Any other equipment should be removed and placed on the cart. Pictures of the various pieces of glassware and equipment appear on the next page.
2. Do not store any chemical samples in your locker as your locker will likely be shared with another student. Place any chemical samples on the shelf in the back of the lab that is marked with your instructor's name and section number.
3. Please clean your locker materials and your lab bench area before leaving the lab.
4. Write your name, locker number and its combination below and keep this sheet in your lab manual.

is common (used by all lab sections) and will be stored in plastic containers under

Name _____ Locker # _____ Combination: _____

SOME COMMON LABORATORY EQUIPMENT



LIST OF COMMON CATIONS

POSITIVELY CHARGED IONS

1+ Charge	2+ Charge	3+ Charge	4+ Charge
✓ Ions of Group 1A Elements ✓ Hydrogen H^+ ✓ Lithium Li^+ ✓ Sodium Na^+ ✓ Potassium K^+ ✓ Rubidium Rb^+	✓ Ions of Group 2A Elements ✓ Beryllium Be^{2+} ✓ Magnesium Mg^{2+} ✓ Calcium Ca^{2+} ✓ Strontium Sr^{2+} ✓ Barium Ba^{2+}	✓ Ions of Group 3A Metals ✓ Aluminum Al^{3+}	
✓ Ammonium NH_4^+ Hydronium H_3O^+ ✓ Copper(I) Cu^+ (cuprous) Silver Ag^+	Cadmium Cd^{2+} Chromium(II) Cr^{2+} (chromous) ✓ Copper(II) Cu^{2+} (cupric) ✓ Iron(II) Fe^{2+} (ferrous) ✓ Lead(II) Pb^{2+} (plumbous) Manganese(II) Mn^{2+} (manganous) Mercury(II) Hg^{2+} (mercuric) Tin(II) Sn^{2+} (stannous) ✓ Zinc Zn^{2+}	Chromium(III) Cr^{3+} (chromic) ✓ Iron(III) Fe^{3+} (ferric) Manganese(III) Mn^{3+} (manganic)	✓ Lead (IV) Pb^{4+} (plumbic) Tin(IV) Sn^{4+} (stannic)

✓ **Learn these checked ions first.**

Names in parentheses are common names

LIST OF COMMON ANIONS

NEGATIVELY CHARGED IONS

1- Charge	2- Charge	3- Charge
√ Ions of Group 7A Elements Chloride Cl^- Bromide Br^- Fluoride F^- Iodide I^-	√ Ions of Group 7A Elements Oxide O^{2-} Sulfide S^{2-} Selenide Se^{2-}	√ Ions of Group 5A Elements Nitride N^{3-} Phosphide P^{3-} Arsenide As^{3-}
Acetate $\text{C}_2\text{H}_3\text{O}_2^-$ √ Cyanide CN^- √ Hydride H^-	Peroxide O_2^{2-}	
Dihydrogen Phosphate H_2PO_4^-	Hydrogen Phosphate HPO_4^{2-}	Phosphite PO_3^{3-} √ Phosphate PO_4^{3-}
√ Hydrogen Carbonate HCO_3^- (Bicarbonate) √ Hydrogen Sulfate HSO_4^- (Bisulfate) √ Hydrogen Sulfite HSO_3^- (Bisulfite) √ Hydroxide OH^- √ Nitrate NO_3^- √ Nitrite NO_2^-	√ Carbonate CO_3^{2-} √ Sulfate SO_4^{2-} √ Sulfite SO_3^{2-}	
* Perchlorate ClO_4^- (per-----ate: one more O than –ate) * Chlorate ClO_3^- (-----ate: one more O than –ite) * Chlorite ClO_2^- (-----ite: one less O than –ate) * Hypochlorite ClO^- (hypo-----ite: one less O than –ite)		

√ Learn these checked ions first.

* Halogens Br and I are capable of forming the same type and charge oxyanions.

COMMON MONOATOMIC IONS IN THE PERIODIC TABLE

1A	2A											3A	4A	5A	6A	7A	8A
H ⁺ hydrogen (acids)																H ⁻ hydride	He ⁰
Li ⁺ lithium	Be ²⁺ beryllium											B	C	N ³⁻ nitride	O ²⁻ Oxide O ₂ ^{2-*}	F ⁻ fluoride	Ne ⁰
Na ⁺ sodium	Mg ²⁺ magnesium	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	Al ³⁺ aluminum	Si	P ³⁻ phosphide	S ²⁻ sulfide	Cl ⁻ chlorid e	Ar ⁰
K ⁺ potassium	Ca ²⁺ calcium				Cr ²⁺ Cr ³⁺	Mn ²⁺ Mn ³⁺	Fe ²⁺ Fe ³⁺			Cu ⁺ Cu ²⁺	Zn ²⁺ zinc			As ³⁻ arsenide	Se ²⁻ selenide	Br ⁻ bromid e	Kr ⁰
Rb ⁺ rubidium	Sr ²⁺ strontium									Ag ⁺ Silver	Cd ²⁺ cadmium		Sn ²⁺ Sn ⁴⁺			I ⁻ iodide	Xe ⁰
	Ba ²⁺ barium										Hg ²⁺		Pb ²⁺ Pb ⁴⁺				

Cations: positive ions formed from metal elements:

- 1) Type I Metals: exist in a single charge form, name by using element name followed by ion; ex. H⁺ is hydrogen ion, Ca²⁺ is calcium ion
- 2) Type II Metals: exist in multiple charge forms, name by using element name followed by Roman numeral representing charge and ion;
ex. Cu⁺ is copper(I) ion and Cu²⁺ is copper(II) ion

Anions: negative ions formed from non-metals elements:

name by using element name ending in **-ide** followed by ion; ex. S²⁻ sulfide ion

* **Diatomic elemental ions** contain 2 of the same atoms in the ion; ex. O₂²⁻ is peroxide ion

Significant Figures/Order of Operation Workshop

I. How many significant figures are in each of the following numbers?

1) 5.40 ____

7) 1.2×10^3 ____

13) 75 grams ____

2) 210 ____

8) 0.00120 ____

14) 0.0005 ____

3) 801.5 ____

9) 0.0102 ____

15) $1 \text{ m} = 10^{-3} \text{ km}$ ____

4) 1,000 ____

10) 9.010×10^{-6} ____

16) 4 monkeys ____

5) 101.0100 ____

11) 2,370.0 ____

17) 4300. ____

6) 23 ____

12) 340 ____

18) 546010 ____

II. Round each number to **four** significant figures. Put the answer in scientific notation.

1) 3.682417 ____

4) 21.860051 ____

7) 375.6523 ____

2) 45.4673 ____

5) 0.000376937 ____

8) 0.098329 ____

3) 40051 ____

6) 2493.1 ____

9) 5499.9 ____

III. Round each number to one place after the decimal point:

1) 1.3511 ____

2) 32.473 ____

3) 0.0638 ____

4) 0.021 ____

IV. Perform the following operations. Include units when appropriate. Give answer in scientific notation and with correct significant figures.

1) $0.0050 \text{ m}^2 \times 0.042 \text{ m}$ ____

2) $22.0 \text{ m} + 5.23 \text{ m} + 15.5 \text{ m}$ ____

3) $0.042 \text{ kg} + 1.229 \text{ kg} + 0.502 \text{ kg}$ ____

4) $3,200 \text{ mg} + 325 \text{ mg} - 688 \text{ mg}$ _____

5) $(0.0300 \text{ mol/L}) (0.105 \text{ L})$ _____

6) $0.400 \text{ mol} / (0.500 \text{ L} + 0.025 \text{ L})$ _____

7) $2.4380 \times 10^{-3} + 8.32 \times 10^{-4}$ _____

8) $(4.32 \times 10^2)(2.1793 \times 10^{-3})$ _____

9) $(2.43 - 0.004)^2 \times 1.76 - 0.00238$ _____

V. Perform the following conversions. Show your work and give answer in scientific notation with correct units.

1) $4.390 \text{ m} \rightarrow \text{dm}$ _____

2) $0.003290 \text{ ng} \rightarrow \mu\text{g}$ _____

3) $4.5934 \text{ ML} \rightarrow \text{mL}$ _____

4) $3.398 \text{ cm}^2 \rightarrow \text{km}^2$ _____

Measurement and Calculation Workshop

1. Are measurements from all measuring devices read to the same place? Why or Why not?
2.
 - a. Are all measurements from the same measuring device read to the same place? Why or Why not?
 - b. Will all of these measurements have the same number of significant figures? Explain.
3.
 - a. If a measuring device's smallest increments (markings) are the ones place, to what place is a measurement properly read? Is that always true for this device?
 - b. How many significant figures could a properly read measurement from this device have? What would you need to know?
4. Later this semester we will use a device called a buret. It has a capacity of 50 milliliter (mL) and has smallest markings of tenths of mL. Which of these measurements represents a correct reading from this device?
 - a. 4.05 mL b. 34.53 mL c. 0.88 mL d. 25.00 mL e. all of these
5. Write these numbers in **correct** scientific notation with the indicated number () of sig. figs.

1584604 (3) =	0.000045999 (2) =
546.000 X 10 ⁵ (5) =	0.0060230 X 10 ²⁶ (4) =
6. Perform the following operations and express your answer in scientific notation with the correct number of significant figures. Include units when appropriate.

a. (156.54)(34.76)/0.0030 =	b. 1000000./15 =
c. $\frac{1}{10^{-3} \text{ mm}}$	d. $\frac{1}{10^5 \text{ } \mu\text{g}}$

e. $(13.1 \text{ m} + 5.61 \text{ m})(0.33333 \text{ m}) =$

f. $\frac{0.00034 \text{ g}}{61.1 \text{ L} - 55.97 \text{ L}}$

g. $\frac{(3.4054 \times 10^{-4} \text{ cm})(4.000 \times 10^{-3} \text{ cm})}{(6.07 \times 10^2 \text{ cm})}$

h. $(5.4 \times 10^3 \text{ kg})^2 (0.000805 \text{ kg}) =$

i. $\frac{1}{8.902 \times 10^{-5} \text{ nm}}$

j. (try this “in your head” and then check it on your calculator):

$\frac{(8.00 \times 10^4 \text{ m})(5.0 \times 10^{-2} \text{ m})(10^2 \text{ m})}{(2.0 \times 10^{-1} \text{ sec})(10^6 \text{ m})}$

7. The following masses for a nickel were determined by a student in lab using the digital balance. Do you want to use all of the trials in your calculations? Determine the correct mean (average) and average deviation. Express your final answers with correct sig. figs. Can you determine the accuracy of the data with this information? Why?

Trial #	1	2	3	4
Mass	5.305 g	5.265 g	4.285 g	5.288 g

CONVERSION FACTOR WORKSHOP

For the following problems, use familiar conversion factors or look up those you don't know. **Construct a road map for each problem first.** Show all of your work and express your answers in correct scientific notation and with correct significant figures.

1. Ether, an organic compound that used to be used as an anesthetic, has a density of 0.736 g/cm^3 at 20°C . What is the volume, in liters, of 225.0 grams of ether?
2. How many pounds (lb) does 1.0000 quart of mercury weigh? Its density is 13.55 g/ml at 25°C .
3. The floor area of a home is 1827 square feet. Find this area in square meters.
4. A French car manufacturer claims that its sedan uses only 6.0 L of gasoline to travel 100. km. How many miles per gallon of gasoline could be expected from this car? (Note: 1 mile = 1.61 km)

5. The density of balsa wood is 7.8 pounds per cubic foot. What is the weight, in kilograms, of a piece of balsa wood 4.0 inches by 6.0 inches by 20.0 inches?

6. A bottle of Lafite Rothchild wine costs \$278 and contains 24.5 fl. ounces. If a sip of wine is about 2.0 mL, how much would just one sip of this wine cost you? (1 fl. oz. = 29.6 mL.)

7. Determine the thickness, in mm., of a 7.0 cm x 3.5 cm piece of aluminum foil if it has a mass of 0.1260 g. Do you need anything else to answer this problem? If so look it up.

8. An antiseptic shampoo for dogs contains 2.35 % benzyl ammonium chloride (BAC). If the shampoo has a density of 1.10 g/mL, how many mg of BAC are in a 1.0 pint bottle? (1 qt = 2 pt exact)

9. At the U.S. Tennis Open, a tennis ball was served at the speed of 128 miles/hour. How many seconds did it take for the ball to travel 35.0 m?

PERCENTAGE WORKSHOP

Percent means parts per hundred. Percent can be expressed mathematically as:

$$\text{Percent} = \frac{\text{amount of the part}}{\text{amount of total}} \times 100$$

The “amounts” in this formula can be expressed in number of items, mass, volume, etc. but the units for both must be the same. So 35% means 35 (2 sig. figs) of the part in a total of 100 (when working with percentages, 100 is always an exact number.) This can mean 35 red balls out of 100 balls, 35 g out of 100 grams or 35 mL out of 100 mL. A percent can have any number of significant figures, depending upon how it is measured, just like any measurement (ex. 56.6% or 2.0057%). Percents can also be expressed as conversion factors. For example: 35% red balls can appear as:

$$\frac{35 \text{ red balls}}{100 \text{ balls total}} \quad \text{or inverted} \quad \frac{100 \text{ balls total}}{35 \text{ red balls}} . \text{ This means that the other 65 balls are not red,}$$

$$\text{so we can also have} \quad \frac{65 \text{ non-red balls}}{100 \text{ balls total}} \quad \text{or} \quad \frac{100 \text{ balls total}}{65 \text{ non-red balls}} .$$

$$\text{Or comparing the parts to each other,} \quad \frac{35 \text{ red balls}}{65 \text{ non-red balls}} \quad \text{or} \quad \frac{65 \text{ non-red balls}}{35 \text{ red balls}} .$$

Any of these conversion factors can be used to solve a problem using 35% red balls. Let’s say that you are told that a mixture is 27.5% copper by mass and that the rest is zinc. That means it is 72.5% zinc, by mass. Any of the following conversion factors can be used:

$$\frac{27.5 \text{ g copper}}{100 \text{ g total mixture}} \quad \text{or} \quad \frac{72.5 \text{ g zinc}}{100 \text{ g total mixture}} \quad \text{or} \quad \frac{27.5 \text{ g copper}}{72.5 \text{ g zinc}}$$

or the inverse of any of these. In fact, grams don’t have to be the unit used. It states **by mass**, so any mass unit, lb, mg, tons...., will be correct as long as you use it on both the top and bottom.

Percentage Problems: (do road maps for all problems)

- 1) a. A mixture contains 34.5 g of copper, 18.4 g of lead and 12.3 g of silver. State the percentage, by mass, of each type of metal in the mixture and create a variety of conversion factors using them.
- b. Calculate the mass, in grams, of silver in a 3.04×10^2 g sample of this mixture. (Ans. 57.5 g)

- 2) A sample of clean air is found to contain 20.45% oxygen gas by volume. How many liters of oxygen gas are present in a 65.0 μL sample of this air? (Ans. $1.33 \times 10^{-5} \text{ L}$)
- 3) A 500.0 mg pill (i.e. 500 mg/ 1 pill) contains 2.7% by mass of the active ingredient. How many g of the active ingredient would be present in 12 pills? (Ans. $1.6 \times 10^{-1} \text{ g}$)
- 4) A sunscreen preparation with a density of 2.82 g/mL contains 3.55% by mass zinc oxide (ZnO). Each tube is filled with 185.0 mL of sunscreen (i.e. 185.0 mL/tube). If you are the manufacturer and you use 5.00 lb of ZnO, how many full tubes of the sunscreen would you prepare? (Ans. 1.22×10^2 tubes)

Nomenclature: Naming Ionic Compounds

Compound	Cation symbol	Does metal have more than one charge?	Cation name	Anion symbol	Anion Name	Compound Name
NaCl	Na ⁺	No	sodium	Cl ⁻	chloride	sodium chloride
Fe(NO ₃) ₂	Fe ²⁺	Yes	iron (II)	NO ₃ ⁻	nitrate	iron (II) nitrate
AlCl ₃						
Be ₂ C						
Sn(C ₂ H ₃ O ₂) ₂						
MgO ₂						
CaHPO ₄						
K ₂ SO ₄						
NiS						
Au ₂ S						
CrPO ₄						
Ca ₃ N ₂						
CaCO ₃						

Nomenclature: Acids

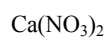
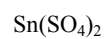
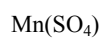
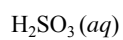
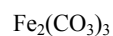
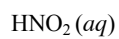
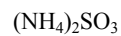
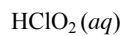
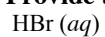
Acid	Binary or Oxyacid?	Anion	Anion name	Acid Name
$\text{HBr}(aq)$	Binary	Br^-	bromine	hydrobromic acid
$\text{HClO}_4(aq)$	Oxyacid	ClO_4^-	perchlorate	perchloric acid
$\text{HCl}(aq)$				
$\text{HNO}_2(aq)$				
$\text{H}_3\text{PO}_4(aq)$				
$\text{H}_2\text{CO}_3(aq)$				
$\text{HNO}_3(aq)$				
$\text{H}_2\text{SO}_3(aq)$				
$\text{HIO}_3(aq)$				
$\text{HI}(aq)$				
$\text{HC}_2\text{H}_3\text{O}_2(aq)$				
$\text{HF}(aq)$				

Nomenclature- Formula Writing Acids

Name	Anion Name	Anion Formula	Chemical formula of acid
perbromic acid	perbromate	BrO_4^-	$\text{HBrO}_4 (aq)$
hydrofluoric acid	fluoride	F^-	$\text{HF}(aq)$
chlorous acid			
sulfuric acid			
chloric acid			
perchloric acid			
phosphoric acid			
nitrous acid			
nitric acid			
permanganic acid			
carbonic acid			

Nomenclature: Systematic Naming

Provide the systematic name for each of the following:



Nomenclature: Writing Chemical Formulas

Provide the chemical formula for each of the following:

perchloric acid

zinc nitrate

sulfurous acid

cobalt (II) hydrogen carbonate

cadmium iodide

copper (II) sulfate

hydrofluoric acid

tetraphosphorus hexasulfide

carbonic acid

silver phosphide

tin (II) phosphate

sodium sulfite

Workshop on Chemical Calculations

Show all work, road maps(*) and put answers in correct scientific notation.

- 1) What is the mass of one mole of calcium carbonate? _____ One formula unit? _____

- 2) *What is the mass in mg of one formula unit of calcium carbonate?

- 3) *How many ions are in 4.000×10^{-4} mL of calcium carbonate if its density is 2.83 g/mL at 25°C?

- *How many kg of chromium could be extracted from 3.2×10^3 μ moles of chromium(III) oxide?

- 4) a. Estradiol contains 79.37% C, 8.88% H and the rest O. What is its empirical formula?
b. If the molar mass of estradiol is 272.37 g/mol, what is its molecular formula?

- 5) *How many atoms of copper would be produced upon reaction of exactly 4 mols of aluminum with copper(II) nitrate solution? What is assumed here?
- 6) *a. How many mg of precipitate would be formed when 3.40×10^{-3} moles of phosphoric acid react with a solution containing 2.003 g of barium hydroxide?
- b. After doing the above reaction and separating out and weighing the precipitate you recovered 0.734 g. What is the percent yield of precipitate?
- 7) In a lab experiment you separated the copper from $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ by reacting it with excess zinc. Assume that you started with 5.20×10^{-2} moles of this salt hydrate. Calculate the percent yield of copper if you obtained the following data.
- | | |
|-----------------------------------|----------|
| mass of filter paper | 0.9344 g |
| mass of filter paper + dry copper | 3.0384 g |

Stoichiometry Workshop

1. Consider the following reaction:



- a. How many moles of CO_2 are produced if 2.867 moles of O_2 react in the presence of excess of other reactants?
- b. How many moles of FeCr_2O_7 are required to produce 35 moles of CO_2 ?
- c. How many grams of Fe_2O_3 will be produced if 2.89 grams of K_2CO_3 reacts in the presence of excess of the other reactants?
- d. How many grams of CO_2 can be produced if 3.97 grams of FeCr_2O_7 and 6.42 grams of K_2CO_3 react in air? What is the limiting reactant?

- e. If the percent yield of CO_2 for the reaction in part c above is 87%, what is the actual yield of CO_2 ?
- f. What is the maximum yield and what is the experimental yield of CO_2 for part e?

Chemical Reactions Workshop

NAME _____

Solubility Info: all acids, acetate, nitrate, ammonium and group 1A compounds are soluble.

1) Write out a balanced molecular equation (ME) for these reactions. Then write a CIE and balanced NIE. State whether each is redox or not and if so, what gets oxidized and reduced.

a. acetic acid reacts with cadmium metal

b. ammonium sulfide solution reacts with barium nitrate solution to form a precipitate

c. lakes acidified by acid rain (sulfuric acid) can be neutralized by adding calcium carbonate to them

d. strontium reacts with zinc phosphate solution

2) Write the reaction for the combustion of lighter fluid (butane: C_4H_{10}) and balance it.

3) What happens to the halogen doing the displacing of another halogen in a single displacement reaction? Explain your answer and use an example reaction.

4) A solution contains one or more of the following ions: Ag^+ , Ca^{2+} , and Cu^{2+} . When sodium chloride is added to the solution, no precipitate occurs. When sodium sulfate is added to the solution, a white precipitate occurs. The precipitate is filtered off and then sodium carbonate is added to the remaining solution, producing a precipitate. Which ions were present in the original solution? Write balanced NIEs only for the formation of each of the 2 precipitates formed.

5) a. Is the decomposition of a binary compound always a redox reaction? Explain with an example.

| b. Write the reaction for the decomposition of sodium sulfate decahydrate and balance it. Is it a redox reaction? Explain.

6) What is the flaming splint test used for? What happens when it is positive? Write the balanced reaction for this positive test.

How to Create a Graph Workshop

BACKGROUND

When two properties of a substance are related to each other in a predictable way, it may be desirable to display this relationship in the form of a graph. There are various types of graphs but the most common form uses rectangular graph paper with horizontal and vertical coordinates displaying the relationship between the two variables or properties you are measuring. This is called a line graph since the data points are connected by a line or an X-Y graph because one variable is plotted on the horizontal or X-axis and the other on the vertical or Y-axis. These are sometimes referred to as the abscissa and the ordinate, respectively. The property that can be controlled, selected or set to arbitrary values is the independent variable while the property whose value depends in some way on the other one is called the dependent variable. The investigator decides on the basis of convenience which value is to be the independent variable.

GENERAL RULES FOR PREPARING A GRAPH

1. Allow plenty of space for plotting the data- try to make your data points fill most of the page of graph paper.
2. Draw the corresponding axes with a straightedge if either variable may be negative or if the graph paper does not have suitable margins. Space should be available along the margins to label the axes.
3. Determine which variable to plot on each axis. It is customary to assign the X-axis to the independent variable and the Y-axis to the dependent variable.
4. Determine the scale that is to be used on each axis. The scales should be chosen so that the data points will fill most of the page. The increments in each axis can be different from each other but should be the same within each axis.
5. Label each axis clearly. The labels should identify the variable and its numerical values with units. The reader must easily understand the labels.
6. Plot the data points and draw a smooth curve or straight line, depending on what's appropriate. The line does not have to go through every point and usually will not as experimentally determined data may contain some errors in measurement. About half of the points should fall on either side of the line.
7. Print a title above the graph. The title should clearly identify what is being illustrated by the graph.

SAMPLE GRAPH

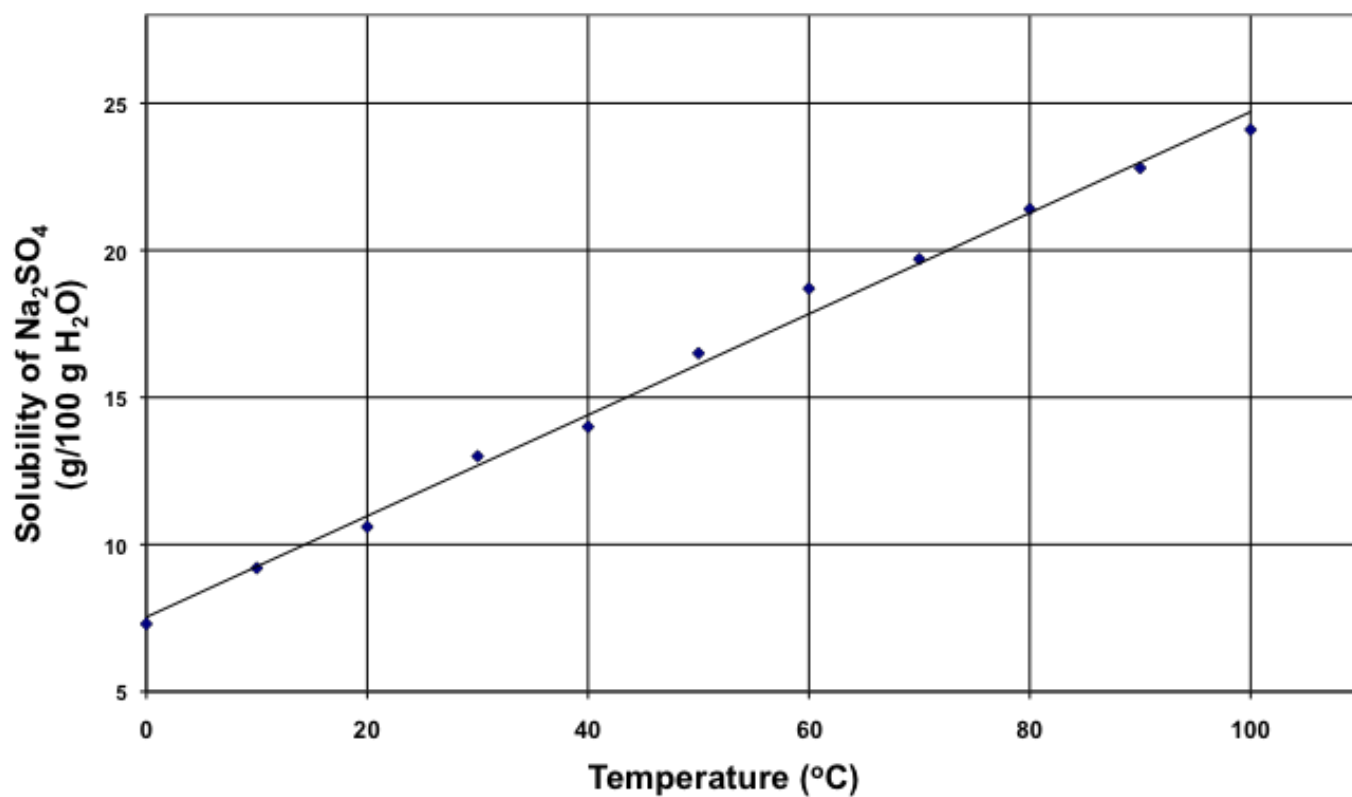
To learn how to construct a simple graph we will take the following experimentally determined data for the solubility of sodium sulfate, Na_2SO_4 , in units of grams per 100 g of water at various temperatures.

Temperature °C	0.0	10.0	20.0	30.0	40.0	50.0	60.0	70.0	80.0	90.0	100.0
Solubility g Na_2SO_4 /100g H_2O	7.3	9.2	10.6	13.0	14.0	16.5	18.7	19.7	21.4	22.8	24.1

Here temperature is chosen to be the independent variable and is plotted it on the abscissa (X-axis). Solubility is plotted on the ordinate (Y-axis). To determine scale values that will produce a graph that allows the data to fill most of the page but not go beyond its edges, we count the number of divisions on our graph paper and determine that there are 110 on the X-axis and 46 on the Y-axis. (Other graph paper may have a different number of divisions.). Dividing our temperature range of 100°C by 100 divisions gives 1.1°C per division. Plotting graph points where the divisions represent fractional units is awkward and it produces a graph that is difficult to read. Whole number ratios, particularly 1, 2, 4 and 5, or 0.1, 0.2, 0.4 and 0.5, produce graphs that are easier to plot and to read. We choose one division per degree as it gives us a graph that will extend most of the way across the page. If we start with 0°C on the third division the graph will be centered better than it would be if we started at the first line. On the Y-axis the maximum solubility of 24.1 g is divided by 46 divisions giving 0.52 grams per division. If we start numbering at zero on the Y-axis and use 1.00 gram per division, our graph will be located almost entirely in the bottom half of the page. Using 0.50 grams will give a graph that will extend above the top of the page. It is not necessary to start all graphs at zero, so our numbering on the Y-axis can start at 5.0 g. This gives a graph that fills most of the page but does not go over the boundaries.

At this point we should make a practice graph. Captions giving the units and numerical values should be entered on the margins to the left of the Y-axis and below the X-axis. For example, the caption on the left margin would indicate that the units are grams of sodium sulfate per 100 grams of water. Scale values such as 5, 10 should be clearly marked on the margins but not on every division. Too many numbers just make the graph difficult to read. Each data point is then plotted where the divisions corresponding to the temperature value and the solubility intersect. Care is essential in locating these data points or the graph will be worthless. A smooth curve is then drawn through the points. It may be extrapolated, or extended some distance beyond the first and last points giving values that are reasonably accurate. Do not draw short straight lines from point to point, but a continuous line that smoothes out minor measuring errors that occurred when the data was obtained. The data given above contains experimental errors that will result in a zigzag graph if all the data points are connected by straight lines. Any points that fall very far from the smooth line or curve should be rechecked to be sure that they were not due to an error in plotting the data. Some data will plot to give a straight line that may be drawn with a ruler. Other data will not plot as a straight line but as a smooth curve that must be hand drawn. In either case do not expect every data point to fall exactly on the line.

Solubility of Na_2SO_4 with Temperature



Give the graph a title that should be printed above the graph if possible, or located in an area of the page that does not interfere with the data points. The title should make clear to the reader what the graph represents. Make any corrections on the practice graph. As noted above, we improved the appearance of our graph by starting the values on the Y-axis at 5.0 g. rather than 0.0 g. We may also consider the possibility of a graph where the paper is vertical rather than horizontal. The data sometimes fits better in that format. Your final graph should be clear and understandable.

The best graphs are usually second or third drafts. Draw the final graph only when you are satisfied that it can't be improved. Review the "general rules for preparing a line graph" on the first page and the following suggestions both before and after the final draft of your graph. If your graph doesn't follow these rules and suggestions, make a revised copy. Graphs today are frequently done with the graphing program Excel®, but many of these same guidelines apply to constructing

FURTHER TIPS FOR GOOD GRAPHS

1. Choose scales so that all of the data points will fall on the page but will nearly fill it. A small graph on a large page is to be avoided. It is not necessary to start the scale values at zero. If none of the data is near zero the scales should begin just below the minimum data value.
2. Number the scales only at every tenth, or twentieth division so as not to crowd the axes. Make small tick marks to show that lines are indicated.
3. Make the data points stand out so that they may be clearly seen. They should not disappear when the line is drawn.
4. Make the titles on the graph clear and comprehensive. (i.e. Can the reader understand the graph without additional information?)
5. Don't clutter up the data region of the graph with notes or labels. Most labels should go in the captions.
6. If both positive and negative values appear in the data, for example Celsius temperatures, a zero or reference line may be drawn across the graph, but it should not interfere with the data points.
7. If two sets of data are plotted on the same graph, to give two lines, different symbols should be used to represent each set of data. For example use dots and triangles or different colors.
8. Proofread the final copy of the graph before submitting it. Errors in copying often sneak in. If there are any errors do not try to correct them. Make a new "final copy". It must be neat and free from errors and corrections.

GRAPHING PROBLEMS

Using the graph paper provided here or your own, prepare the following graphs. Refer to the guidelines given in the text above. Don't hesitate to make drafts. When finished with each graph show them to your instructor for feedback. Draw a straight line or best fit curve, whichever is appropriate, on each graph.

Graph 1. Calculate the Fahrenheit temperatures corresponding to the following Celsius temperatures and prepare a graph from this data.

0.0°C 10.0°C 20.0°C 40.0°C 60.0°C 80.0°C 100.0°C.

Graph 2. At a temperature of 0.0°C one mole of a gas occupies the given volumes at various pressures. Prepare a graph from this data.

Pressure,	0.25 atm.	0.50 atm.	1.00 atm.	2.00 atm.	3.00 atm.	4.00 atm.	5.00 atm.
Volume,	89.6 L.	44.8 L.	22.4 L.	11.2 L.	7.5 0 L.	5.60 L.	4.50 L.

Graph 3. Prepare a graph from the following data that gives the vapor pressure of water at various temperatures.

Temp.	0.0°C	20.0°C	40.0°C	60.0°C	80.0°C	100.0°C
Press.	4.6 torr	17.5 torr	55.3 torr	149.4 torr	355.1 torr	760.0 torr

Graph 4. The volume of a sample of air was measured at various temperatures with the following results. Pressure was constant throughout the experiment. Prepare a graph from the following data and **extend the line to zero volume** (this means that you must extend the X-axis to show negative °C temperatures.) Charles' Law predicts that this will occur at a temperature of absolute zero.

Temperature	100.0°C	70.0°C	40.0°C	10.0°C
Volume	0.650 L	0.598 L	0.545 L	0.493 L

POST ASSIGNMENT

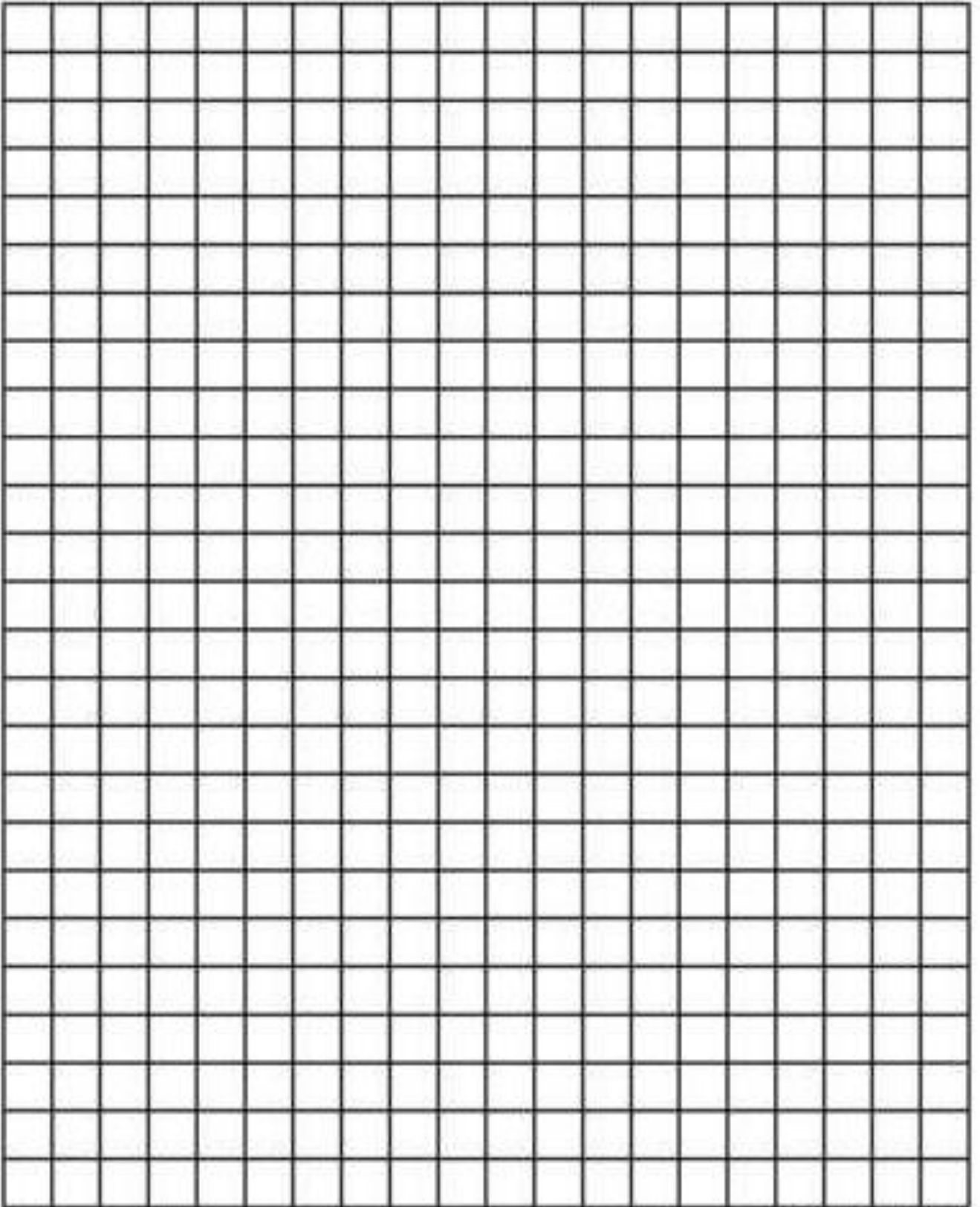
Answer the following questions by referring to the graphs that you prepared above and turn these in with your graphs. Use your straight line or best fit curve to determine the values.

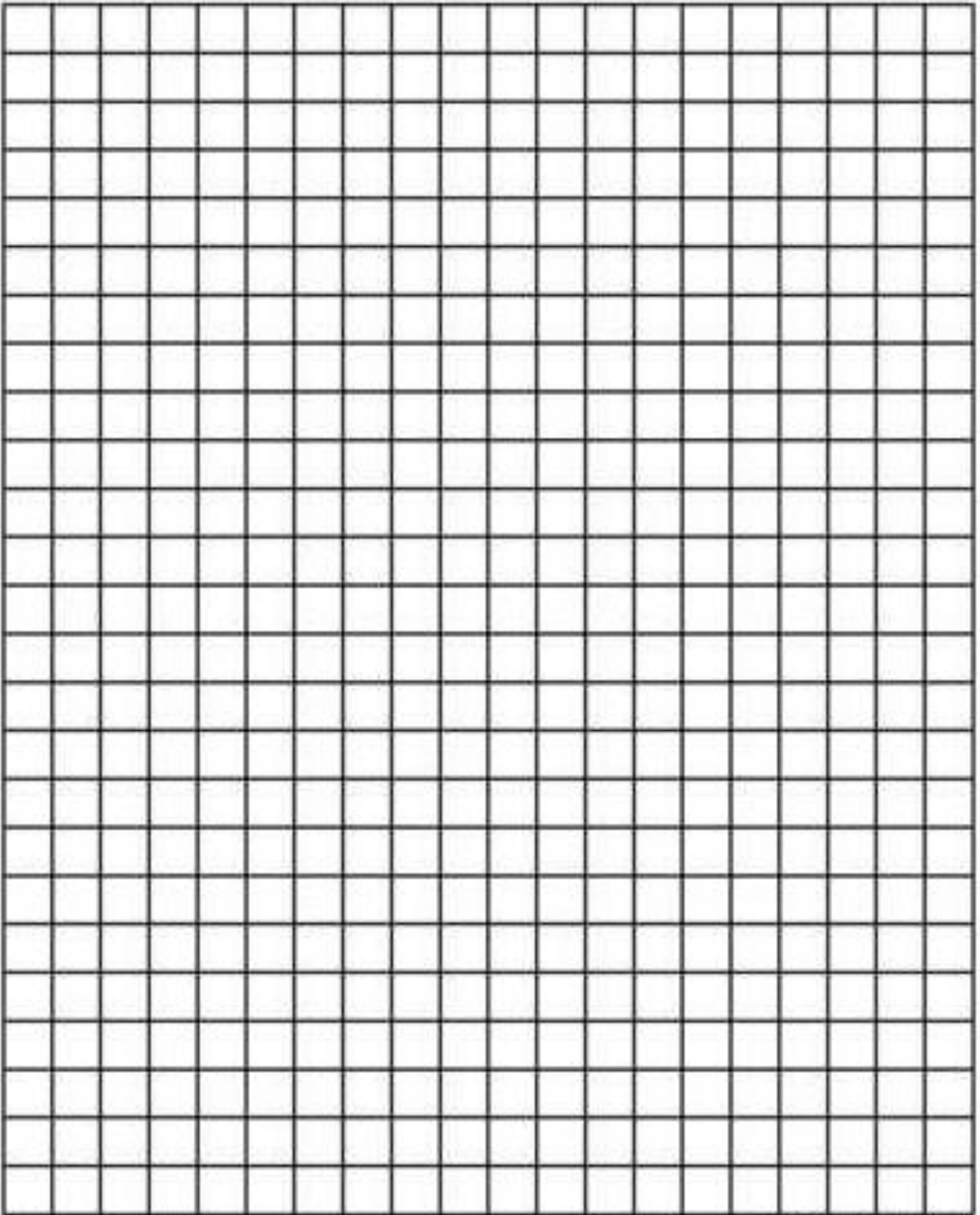
Graph 1: What is the Celsius temperature at: (a) 33.0°F; (b) 103.0°F.
What is the Fahrenheit temperature at: (c) 22.0°C; (d) 68.0°C.

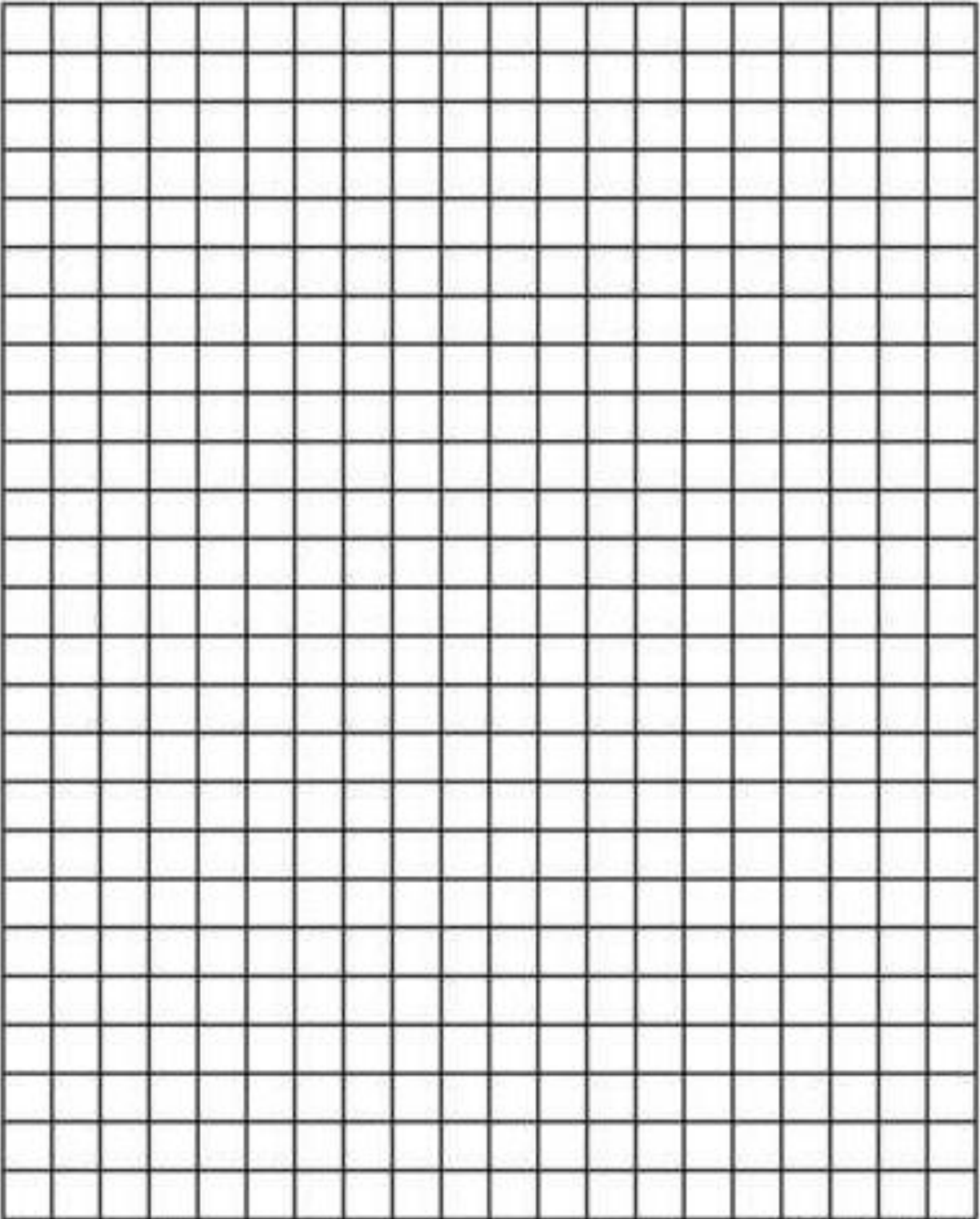
Graph 2: What is the volume of one mole of a gas at 0.0°C and: (a) 0.49 atm.; (b) 3.45 atm.
What is the pressure of one mole of a gas at 0.0°C and: (c) 15.0 L; (d) 3.5 L.

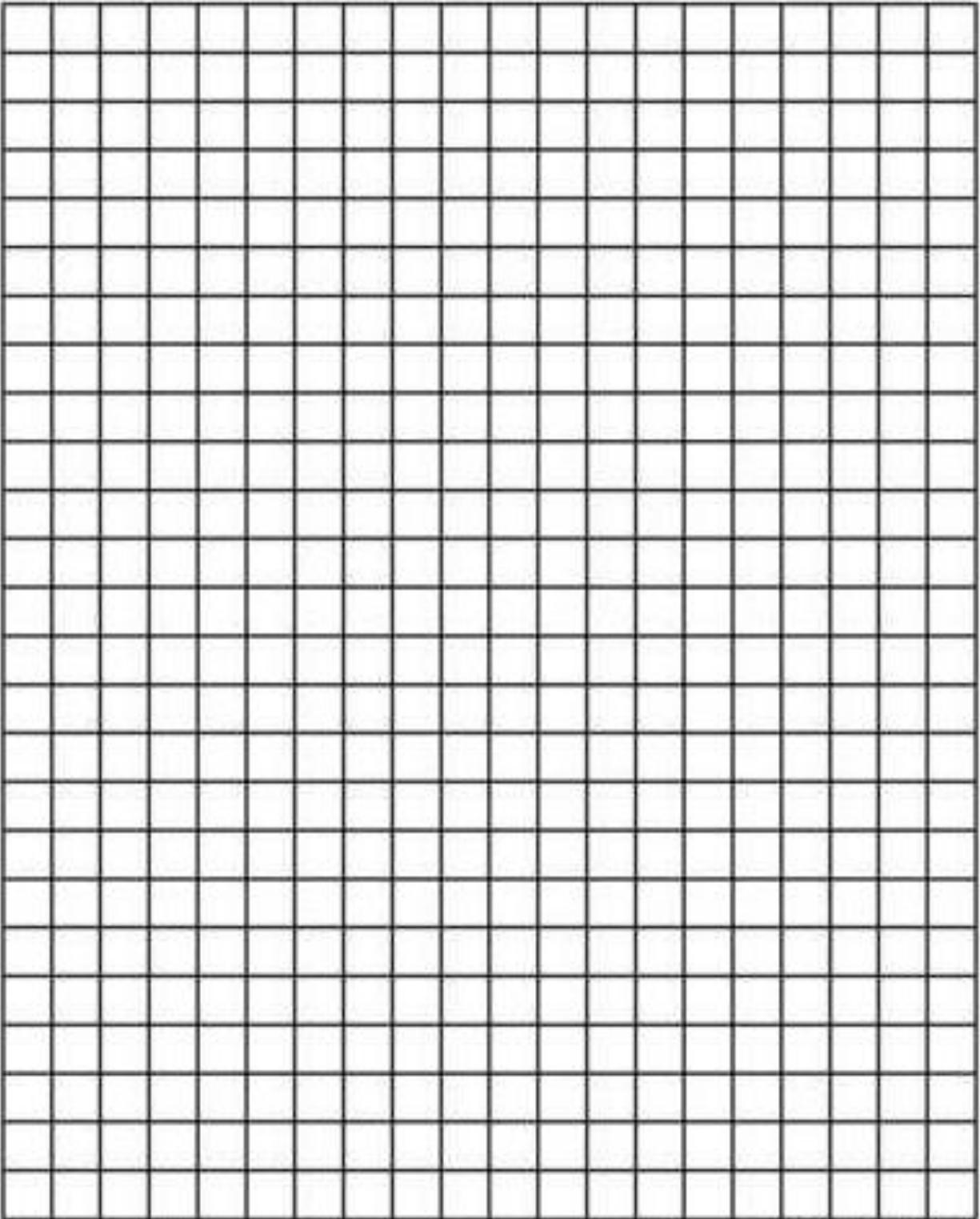
Graph 3: What is the vapor pressure of water at: (a) 22.0°C ; (b) 68.0°C .
At what temperature is the vapor pressure of water: (c) 19.0 torr, (d) 400.0 torr.

Graph 4: What would the volume of the air be at: (a) 20.0°C ; (b) 58.0°C .
What is the Celsius temperature when the volume is zero (you must extend the X-axis far out to the negative side to get this point)?



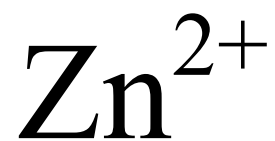






Common Ion Flash Cards:

simply cut out and learn!



manganese (II) ion

chromium (III) ion

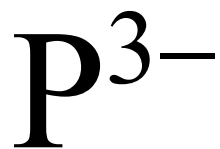
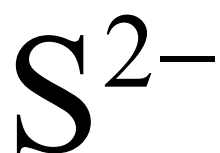
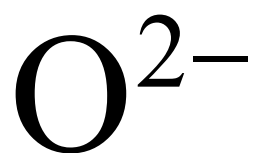
nickel (II) ion

cadmium ion

silver ion

zinc ion

chromium(II) ion



chloride ion

fluoride ion

iodide ion

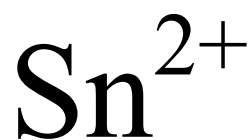
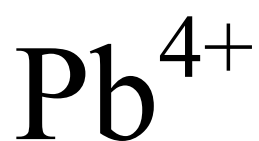
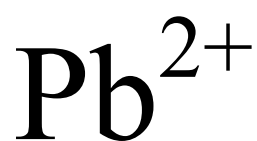
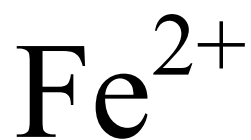
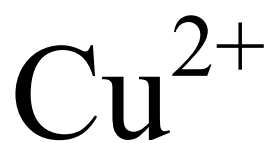
bromide ion

sulfide ion

oxide ion

phosphide ion

nitride ion



copper(II) ion

copper(I) ion

iron(III) ion

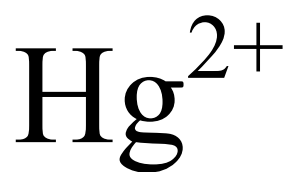
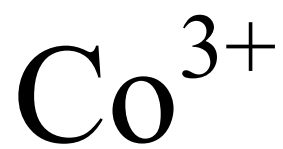
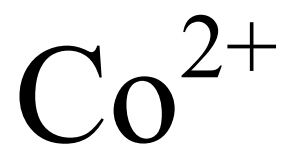
iron(II) ion

lead(IV) ion

lead(II) ion

tin(IV) ion

tin(II) ion



cobalt(III) ion

cobalt(II) ion

mercury(II) ion

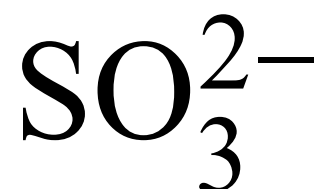
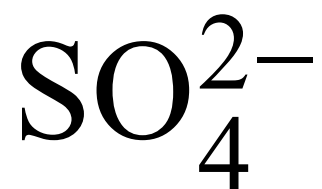
aluminum ion

hydronium ion

ammonium ion

hydroxide ion

cyanide ion



hypochlorite ion

perchlorate ion

chlorite ion

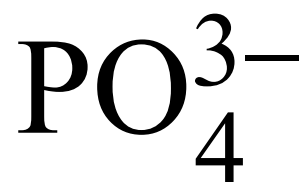
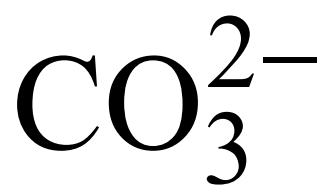
chlorate ion

nitrite ion

nitrate ion

sulfite ion

sulfate ion



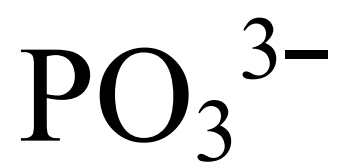
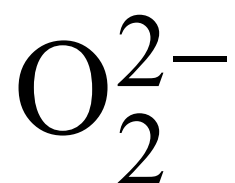
**hydrogen carbonate ion
or bicarbonate ion**

carbonate ion

phosphate ion

**Hydrogen sulfate ion
or bisulfate ion**

acetate ion



phosphite ion

peroxide ion

**dihydrogen
phosphate ion**

**Hydrogen phosphate
Ion**

**hydrogen phosphite
ion**

**hydrogen sulfite ion
or bisulfite**

