

Fresno City College

Chem 3A Lab Manual

Fall 2024

Name: _____

Instructor's Name: _____

Lab Day: _____

Lab Time: _____

Classroom: _____

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Calculator Tutorials (TI-36X Pro)

Scan the QR Code below for a YouTube Playlist with tutorial videos for the TI-36X Pro Calculator



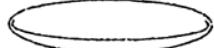
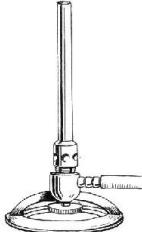
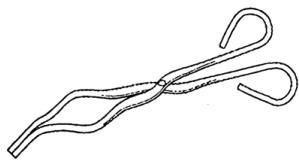
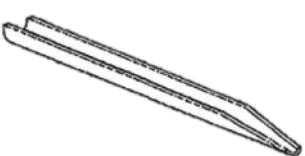
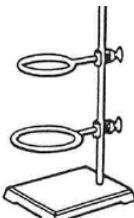
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Recording Analog Measurements

Below are guidelines for recording measurements on pieces of equipment without digital displays. Look at the numbered lines. Determine how far apart those numbers are (0.1, 1, 10, 100 etc.). If there are smaller lines, decide what value those lines represent (0.01, 0.1, 1, 2, 5, 10, etc.). Record your measurement to one decimal place beyond the value the smallest division represents.

SMALLEST LINE DIVISION	RECORD TO:	EXAMPLE	POSSIBLE EQUIPMENT
0.1	0.01	8.23 mL	Graduated Cylinder Buret Volumetric Pipet
1	0.1	8.2 mL	Graduated Cylinder Thermometer
10	1	8 mL	Beaker

Equipment List

Exp 0a – Lab Exploration

INTRODUCTION

In this experiment, you will learn where pieces of equipment are stored in the lab and get the opportunity to use some of the glassware you will be using for experiments throughout the semester. This will allow you to get comfortable with laboratory equipment, its purpose, and how to use it properly.

BACKGROUND

Buret

A buret is a specialty piece of glassware used for the purpose of delivering a volume of a liquid or solution to another vessel. Much like a graduated cylinder, the buret has gradations along the side used for measuring the volume of the liquid inside. A buret differs from a graduated cylinder in two important ways. First, there is a stopcock at the bottom of the instrument that can be opened to allow liquid to flow from the buret and closed again to stop the flow. The second important difference between a buret and a graduated cylinder is that the gradations are etched in decreasing order, instead of increasing order. Record the location of the bottom of the meniscus no matter how the graduations are ordered. Look at the example image below. The meniscus is between 46 mL and 47 mL. Begin at the smaller number (46-- mL) and count the smaller lines (0.1 mL each) until the line is close to, but not past, the bottom of the meniscus (in this example, it is 46.8- mL). Estimate between the lines the location of the bottom of the meniscus to obtain an estimated digit (46.82 mL). The estimated digit is allowed to vary slightly from person to person. The volume should be read as 46.82 mL, NOT 47.18 mL. Holding a black bar (Figure 3) slightly below the bottom of the meniscus can help to highlight it and make it easier to read.

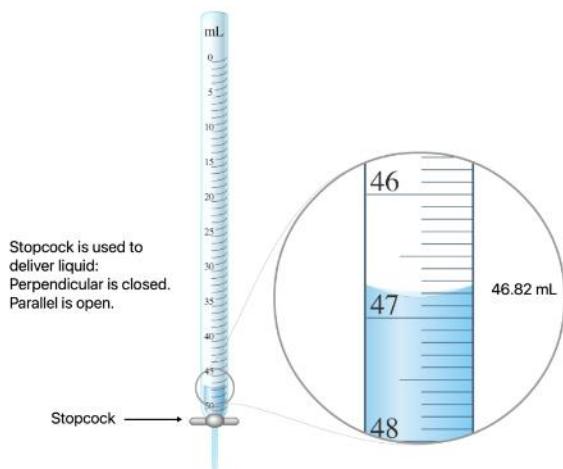


Figure 1:
Example of how to read a buret.

Volumetric Flask

Volumetric flasks are very accurate at measuring the volume of a liquid, but each volumetric flask can only measure one such volume. For example, the volumetric flask in the picture is a 10 mL volumetric flask. It can only contain 10 mL, but does so accurately to two decimal places, which must be recorded as 10.00 mL. This is the only volume that can be read from this particular flask. To use a volumetric flask, liquid is added until the bottom of the liquid's meniscus reaches the etching on the neck of the flask.



Volumetric flasks are most commonly used for preparing solutions. The solid (solute), is added to the flask first. Then the liquid (solvent) is added until the liquid level reaches the etching on the neck of the flask. It is recommended to first half-fill the flask, then swirl to dissolve your solid, before filling the rest of the way with solvent as it is much harder to mix once the flask is full.

Volumetric Pipet

A volumetric pipet is also a very accurate piece of glassware that is used to transfer a set volume from one vessel to another. Like a volumetric flask, it can only measure one volume, but does so very accurately, often to two decimal places. Liquid is drawn up through the pointed tip of the pipet, using a pipet bulb, until the bottom of the meniscus reaches the etching on the neck of the pipet. The pipette is then transferred to another vessel, and the liquid is allowed to drain out.



Your instructor will show you how to properly work the pipet bulb, but slow and controlled is key. Going too quickly can cause you to accidentally draw solution into the bulb, which damages the bulb.

SUPPLY LIST

For this experiment, follow the procedure to gather items as needed.

- Buret
- Buret Clamp
- Bench Post
- Funnel
- Beaker (50-mL or 100 mL)
- Erlenmeyer Flask
- Scoopula/spatula
- Volumetric Flask (50-mL) × 2
- Volumetric Pipet
- Pipet bulb or pump
- Wash bottle (filled with DI Water)

CLASSROOM SUPPLIES AND REACTANTS

These are shared supplies or reactants. Do not gather these items. Leave these items at their designated location in the classroom and use as needed according to the procedure.

- Food Coloring
- Sodium Chloride (Salt, NaCl)

PROCEDURE – PART 1: SCAVENGER HUNT

Use the boxes to check-off steps of the procedure as you complete each step.

1. Refer to the pictures on the “Equipment List” page of your lab manual. This equipment is typically stored either on the back wall of the classroom or in cabinets/drawers near the balances. Everything is labeled. Look around and gather the following items.
 - a. Thermometer (in a white rack, usually located on the counter near the balances)
 - b. Test tube
 - c. Beaker (50-mL or 100-mL)
 - d. Watch Glass
 - e. Graduated Cylinder (100 mL)
 - f. Erlenmeyer Flask
 - g. Volumetric Flask (50-mL)
 - h. Transfer Pipet
 - i. Volumetric Pipet (10-mL)
 - j. Pipet Bulb or Pipet Pump
 - k. Buret
 - l. Buret Clamp
 - m. Bunsen Burner
 - n. Funnel
 - o. Wash Bottle
 - p. Crucible Tongs
 - q. Scoopula
 - r. Bench Post (in wood box on back wall of classroom)
 - s. Utility clamp
 - t. Clamp holder

2. Arrange the items you found in the same arrangement they appear on the “Equipment List” page. Raise your hand to call over the instructor to have them sign off that you correctly found everything you needed.
3. Return all equipment that isn’t listed in the “Supplies List” for this experiment and gather any other items listed in that section.

PROCEDURE – PART 2: USING A BURET

1. Assemble the titration apparatus as shown in the image to the right.
2. Add approximately 15-20 mL of DI water to a 50-mL beaker. Add one drop of food coloring. Swirl the beaker gently to mix the food coloring into the water.
3. Fill the buret with colored DI water from your beaker. You can use a funnel if you want to. The water level in the buret should be below the 0 mL mark at the top of the buret. Never waste time trying to fill the buret to the 0 mL mark. Record the initial volume.
 - a. Hold the black bar (Figure 3) behind the buret, slightly below the meniscus. This will help highlight the bottom of the meniscus to make it easier to read.
 - b. Sketch the location of the meniscus on the diagram provided on the data report. Label the lines with the appropriate values.

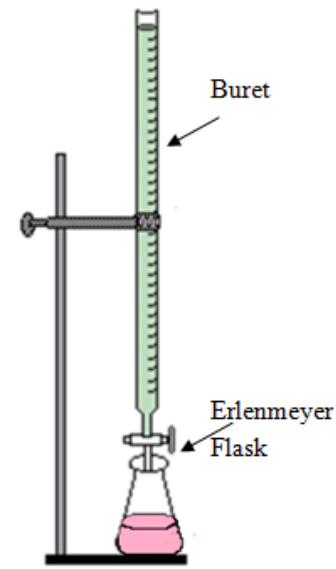


Figure 2:
Titration Apparatus

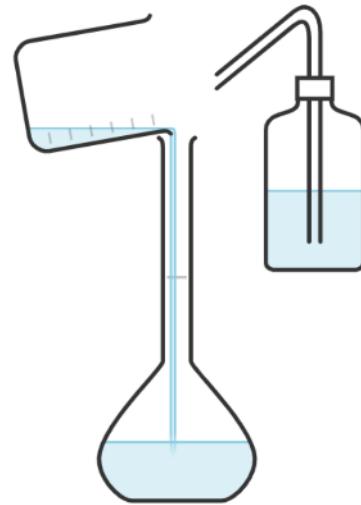


Figure 3:
Hold this black bar behind the buret, below the meniscus to make it easier to read.

4. Use a graduated cylinder to measure between 20-30 mL of DI water. Record the precise volume of DI water in the graduated cylinder on your report sheet. Sketch the location of the meniscus on the diagram provided on the data report (label the lines with the appropriate values). Transfer the measured DI water to an Erlenmeyer flask.
5. Add a drop of food coloring that is a different color than your first solution.
6. Place the Erlenmeyer flask under the tip of the buret. Open the stopcock of the buret so that it flows freely, while swirling the Erlenmeyer flask. Add approximately 5 mL of water from the buret, then close the stopcock.
7. Starting from a closed stopcock, very slowly start to open the stopcock until the water is flowing out of the buret dropwise. Add approximately 1 mL of liquid from the buret dropwise, then close the stopcock.
8. Starting from a closed stopcock, very slowly start to open until only about half of a droplet has formed on the tip of the buret. Quickly close the stopcock. Use the wash bottle to rinse the droplet into your Erlenmeyer flask.
 - a. Record the final volume on the buret and sketch the location of the meniscus on the diagram provided on the data report. Label the lines with the appropriate values.
9. Calculate the total volume of liquid delivered from the buret.
10. Raise your hand to call over the instructor. Your instructor will verify that you recorded the volume in your buret correctly, and have you demonstrate the skills in Steps 6-7.

PROCEDURE – PART 3: USING PIPETS AND VOLUMETRIC FLASKS

1. Tare the balance by hitting the “0” or tare button so that it reads 0.0000 g.
2. Place an empty 50-mL **volumetric flask** on the balance.
3. Close the doors to the balance and try not to bump into the counter. Wait for the mass on the balance to stabilize. Record the precise mass on your report sheet (record all the digits that the balance display shows).
4. Remove the volumetric flask from the balance. Place a dry 50-mL **beaker** on the balance. Tare the balance by hitting the “0” or tare button so that it reads 0.0000 g again.
5. Use a scoopula to add between 4-6 g of sodium chloride (table salt, NaCl) to the beaker. Record the precise mass on the data report sheet.
6. Add approximately 10 mL of DI water to the beaker and swirl to wet the salt. Pour the wet salt into the 50-mL volumetric flask. Not all the salt will come out.
7. While holding the beaker’s spout in the neck of the volumetric flask, spray a stream of DI water from a wash bottle into the beaker to rinse the remaining salt into the flask. Be careful not to spill or splash!
8. Rinse the beaker thoroughly into the flask to ensure all the salt is transferred, until the flask is about half to $\frac{3}{4}$ of the way full.
9. Place a stopper into the neck of the flask and swirl the volumetric flask until the sodium chloride crystals have dissolved.
10. Add one drop of food coloring.
11. Use a wash bottle to add more DI water to the volumetric flask until the bottom of the meniscus is aligned with the etching in the neck of the flask. You can use a plastic transfer pipet to carefully add water dropwise until exactly on the line.
12. Record the mass of the 50-mL volumetric flask now that it contains your prepared solution. This is your “Stock Solution.”
13. Using a 10-mL volumetric pipette, transfer 10.00 mL of your stock solution into a new 50 mL volumetric flask. Make sure you do not draw liquid up into the pipette bulb or it will damage it.



14. Use a wash bottle to add more DI water to the volumetric flask until the bottom of the meniscus is aligned with the etching in the neck of the flask. You can use a plastic transfer pipet to carefully add water dropwise until exactly on the line. This is your “Diluted Solution.”
15. Compare the color of your two solutions. Which one is darker? Which one is lighter?
16. Raise your hand to call over the instructor. They will inspect your prepared solutions, and you will demonstrate, in front of the instructor, that you understand how to use a volumetric pipet correctly.

CLEAN UP AND WASTE DISPOSAL

1. Pour all your salt solutions and water down the drain.
2. Wash your used glassware. There is soapy water near each sink.
3. Thoroughly rinse your glassware with DI water
4. You can blot the outside of the glassware dry with paper towels, but there is no need to thoroughly dry anything. You can put it back wet. Never attempt to dry the inside of narrow glassware with a paper towel!! More often than not, the paper towel gets stuck inside.
5. Return all equipment.

Exp 0a (Intro)***Lab Exploration***

DATA AND REPORT – DUE ____ / ____ / ____

SECTION _____

DATE _____

PARTNER _____

PART 1: SCAVENGER HUNT

Instructor Check: _____

DATA TABLE – PART 2: USING A BURET

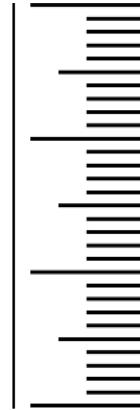
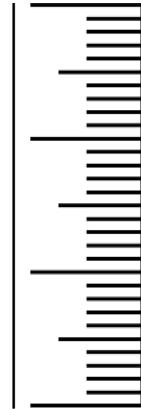
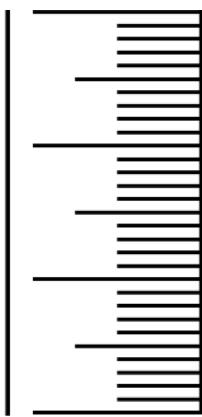
Measurements	
Graduated Cylinder	
Initial Buret Reading:	
Final Buret Reading:	
Calculations	
Volume Delivered by Buret	
Instructor Check	
Volume recorded correctly	
Adding Dropwise	
Adding Partial Drop	

SKETCH OF MENISCUS LOCATION

Graduated Cylinder

Initial Buret Reading:

Final Buret Reading:



DATA TABLE – PART 3: USING VOLUMETRIC FLASKS AND PIPETS

Measurements	
Mass of 50-mL Volumetric Flask (empty)	
Mass of sodium chloride (NaCl, salt)	
Mass of 50-mL Volumetric Flask + stock solution	
Visual comparison of stock solution and diluted solution	Observations:
Calculations	
Mass of stock solution made	
Mass of water used to make the stock solution	
Instructor Check	
Solutions made	
Pipet technique	

Exp 1a – Making Measurements

INTRODUCTION

Mass, volume, and length are fundamental measurements made in any laboratory. In the following experiment, you will be calculating the density of a metal bar based on mass, volume, and length measurements. You will also determine the density of pure aluminum shot and determine whether the metal bar is made of aluminum or some other metal.

BACKGROUND

Measuring Volume

In laboratories, a common device for measuring volumes of liquids is the graduated cylinder (see figure 1). Although it is only used for measuring liquid volumes, it is often used to indirectly measure the volume of solid objects via the “displacement method” discussed below. The volume measured is based on the lowest point of the meniscus; the curved shape created by the liquid in the graduated cylinder. Scientists always estimate one digit beyond the markings shown on the measuring device. The graduated cylinder to the right has markings for every 1 mL, so measurements should always be recorded to the nearest 0.1 mL. When the meniscus is exactly on the measurement line as shown in the picture to the right, the decimal recorded would be zero. Using this method, we would record that the volume of water in the cylinder is 43.0 mL.

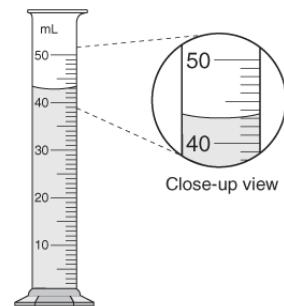


Figure 1: A 50-mL Graduated Cylinder

The volume of a geometrically regular shaped solid can be determined by first measuring its dimensions and then calculating its volume using the volume equation for that shape. For example, consider a rectangular prism (see figure 2). The volume can be determined by measuring the width, the length, and the height. The volume can then be calculated by:

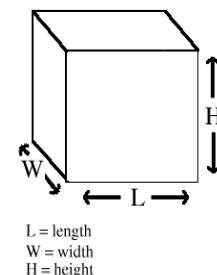


Figure 2: The Volume of a Rectangular Prism

$$V_{\text{rectangular prism}} = \text{width} \times \text{length} \times \text{height} \quad (1)$$

For a cylinder, the following volume equation is used:

$$V_{cylinder} = \pi \times radius^2 \times height \quad (2)$$

The volume of a regularly shaped and an irregularly shaped object can be determined by *displacement*. A volume of liquid is measured first without an object and then is measured again with the object. For example, if you wanted the volume of an egg, you could first put one cup of water in a measuring cup and then add the egg (fresh eggs sink). The additional measured volume is the volume of the egg. In general, the displaced volume can be calculated by:

$$Egg\ Volume = Final\ Volume - Initial\ Volume \quad (3)$$

When an object is irregularly shaped (i.e., the dimensions can't easily be measured), it is usually easiest to measure the volume by displacement.

Measuring Mass

A “balance” is used for measuring the mass of an object. The term balance is an archaic term that comes from the days when mass was measured by balancing masses (see figure 3). Modern balances measure the force exerted by an object and compare it to the force created by an object of known mass (calibration to a standard).



Figure 3: A Balance

A simple method for determining the mass of an object with an analytical balance (see figure 4) is as follows:

1. Tare the balance by hitting the “0” or tare button.
 - * If necessary, place weighing paper on the balance to protect the metal balance pan from the substance you are measuring.
 - * Make sure the balance doors are shut and you are not leaning on the counter.
 - * Push the gray bar if the digital readout does not show 0.0000 g.
2. Place the object to be massed on the balance pan.
 - * Be sure that the object is clean.
 - * Be sure to reclose the doors after adding the object.



Figure 4: An Analytical Balance

3. When the display stabilizes, record the object's mass. (Always record all decimals shown on the balance.)

If you haven't recognized it, this is the same process they use at the meat counter in the grocery store ☺.

ADDITIONAL READING

- Measurements, Problem Solving, Unit Conversions and Density

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat

SUPPLY LIST

These are items that you can gather before beginning your experiment.

- Metal bar
- Ruler
- 50 or 100-mL graduated cylinder
- Glass stirring rod
- Wash bottle with distilled water

CLASSROOM SUPPLIES AND REACTANTS

These are shared supplies or reactants. Do not gather these items. Leave these items at their designated location in the classroom and use as needed according to the procedure.

- Aluminum shot
(shot = pellets)

PROCEDURE

* Record all qualitative and quantitative data *neatly* on the “Data and Report Pages”. Remember to include both units and correct significant figures.

1. Record the unknown number or letter of your metal bar. Briefly describe its appearance.
2. Measure the mass of the metal bar using an analytical balance.
* Record all the digits visible on the balance display. (*Always record all the digits* when reading a balance.)
3. Measure the dimensions (l, w, h or diameter) of the metal bar in cm units.
* Record these values to the correct number of significant figures (remember estimate 1 place past the smallest division).
4. Measure the volume of the metal bar by displacement using water and a graduated cylinder.
 - a. Record the volume of water before and after adding the bar (remember, *always* estimate 1 place past the smallest division.)
 - b. Add the metal bar very carefully by first tilting the graduated cylinder slightly to one side, then by allowing the metal bar to slide slowly to the bottom.
* Check for air bubbles. Remove any large bubbles with the glass rod.
5. Obtain a small amount of aluminum shot (between 10 g to 30 g is fine) and measure its mass.
Record the precise mass on your report sheet.
* Make sure the sample is dry and clean
6. Briefly describe its appearance on your report sheet.
7. Measure the volume of the metal shot by displacement.
* Be sure to remove air bubbles using the glass rod.

CLEAN UP AND WASTE DISPOSAL

1. After checking your data with your instructor, dry the shot with a paper towel before putting them in the “return metal” container.
2. Dry your metal bar with a paper towel before putting it back in the original bin.
3. Empty your water bottle before putting it back.
4. Put all other equipment back where you found it.*

*There is no need to thoroughly dry the graduated cylinder prior to returning it to the shelf. Never attempt to dry the inside of a graduated cylinder with a paper towel!! More often than not, the paper towel gets stuck inside.

DATA AND REPORT – DUE ____ / ____ / ____

* Show your work and keep track of significant figures.

SECTION _____**DATE** _____**PARTNER** _____**DATA**

1. Number or letter on the bar: _____
2. Mass and appearance of the metal bar.
 - a. Mass _____
 - b. Shape _____
 - c. Color _____
3. Dimensions of the bar:
 - a. Longest side (Length) or Diameter: _____
 - b. Medium side (Height): _____
 - c. Short side (Width): _____ (not needed for cylinders)
4. Water displacement of the bar:
 - a. Initial Water Volume _____
 - b. Final Water Volume _____
5. Mass and appearance of the aluminum shot:
 - a. Mass _____
 - b. Shape _____
 - c. Color _____
6. Water displacement of the aluminum shot:
 - a. Initial Water Volume _____
 - b. Final Water Volume _____

RESULTS

* Show your work and keep track of significant figures.

7. Calculate the metal bar's volume (in cm³) using its dimensions.

8. Calculate the volume of the bar in mL using water displacement.

9. Chemists compare two values mathematically by calculating percent difference. Calculate the percent difference between the two volumes using the following equation:

$$\% \text{ Difference} = \frac{\text{Volume from dimensions} - \text{Volume by displacement}}{\text{Volume by displacement}} \times 100$$

10. How does the bar's 'volume measured by dimensions' compare to the bar's 'volume measured by displacement'? (Use complete sentences, mentioning both numerical volume measurements, and the percent difference)

11. Calculate the density of the metal bar. (Use the volume from the displacement measurement.)

12. Determine the density of the aluminum shot.

13. Compare the densities of the aluminum shot to the density of the metal bar you measured earlier. Could your bar be made of aluminum? Why or why not? If the bar is not aluminum, what could metal be?

Metal	Density (g/cm ³)
Brass	8.5
Cast Iron	6.7
Copper	9.0
Gold	19.3
Lead	11.3
Silver	10.5
Stainless Steel	7.5
Steel	8.0
Titanium	4.5
Tin	5.8

Exp 2a – Atomic Mass

INTRODUCTION

In this experiment, you will be calculating the average atomic mass of a fictional element called “Fresnocitium” which has three isotopes that are different colors: red, black, and white.

BACKGROUND

The “atomic mass” reported on periodic tables is the average atomic mass of all the different isotopes of the given element, weighted by the natural abundance. Atoms of real elements are too small for us to see with our eyeballs, or touch individually to get a sense of what it means to have a weighted average mass. Isotopes that are more abundant count towards the average atomic mass more than less abundant isotopes.

Atoms of fresnocitium are large enough to be picked up and counted with your fingers, allowing us to determine the abundance of each isotope and calculate the average atomic mass of fresnocitium with only an analytical balance. The abundance of an isotope can be calculated by:

$$\text{Abundance (\%)} = \frac{\text{number of atoms}}{\text{total number of atoms}} \times 100$$

The average atomic mass is calculated using the equation below, with the abundances in decimal form. To convert from the percent abundance into decimal form (fractional abundance), divide the percentage by 100.

$$\text{Average Atomic Mass} = (\text{mass}_1 \cdot \text{abundance}_1) + (\text{mass}_2 \cdot \text{abundance}_2) \dots$$

ADDITIONAL READING

- Atomic Mass

SUPPLY LIST

- Beaker (100 mL)
- Tray or paper plate

CLASSROOM SUPPLIES

These are shared supplies or reactants. Do not gather these items. Leave these items at their designated location in the classroom and use as needed according to the procedure.

- Beads made of various materials

PROCEDURE

1. Bring an empty 100 mL beaker over to the analytical balances.
2. Tare (zero) the balance so that the display reads 0.0000g. Place the empty beaker on the balance and then record the mass on your report sheet. You will repeat these instructions twice: Trial 1 & Trial 2. You will use this same beaker to record all the mass measurements for both trials.
3. Shake the supply bin of Fresnoccium thoroughly. Ensure the lid is snapped on and hold the lid on while shaking it.
4. Use the beaker as a scoop to obtain a sample of Fresnoccium atoms (beads). Fill the beaker with beads.
5. Return to the balances. Tare (zero) the balance, and then record the mass of the beaker containing Fresnoccium.
6. Carefully sort the different isotopes of fresnoccium (Red, Black, and White). You may obtain other beakers to hold these atoms if you like, as long as you do not use them for weighing.
7. Place all the atoms of the red isotope in your beaker.
8. Tare the balance and then record the mass of the beaker with the red isotope.
9. Remove the red atoms from the beaker. Count and record the number of red atoms. You should double-check that your counting is accurate.
10. Repeat steps 6-9 with the black atoms and then white atoms.
11. Return your atoms to the supply bin.
12. Using the same 100 mL beaker, repeat steps 3-10 to obtain a second trial of data.

CLEAN-UP

1. After your data is approved by your instructor, return all the atoms to the supply bin. With lid securely on the bin, shake the bin thoroughly.

DATA AND REPORT – DUE ____ / ____ / ____**SECTION _____**

* Show your work and keep track of significant figures.

DATE _____**PARTNER _____****DATA**

1. Mass of beaker (empty): _____

	Trial 1	Trial 2
Mass of beaker + “Nature’s Mix” of Fresnocitium		
Mass of beaker + Red Atoms		
Mass of beaker + Black Atoms		
Mass of beaker + White Atoms		
Number of Red Atoms		
Number of Black Atoms		
Number of White Atoms		

CALCULATIONS

Show your work neatly, with units, where applicable.

1. Mass of “Nature’s Mix” of Fresnocitium:

TRIAL 1**TRIAL 2**

--	--

2. Total Number of Atoms:

TRIAL 1**TRIAL 2**

--	--

3. Average Mass per Atom in Nature's Mix:

TRIAL 1**TRIAL 2**

--	--

4. Complete the table below:

	TRIAL 1		TRIAL 2	
Isotope Color	Percent Abundance of Isotope (%):	Average Mass per atom of isotope (g):	Percent Abundance of Isotope (%):	Average Mass per atom of isotope (g):
Red				
Black				
White				

5. Average Atomic Mass based on equation:

$$\text{Average Atomic Mass} = (\text{mass}_1 \cdot \text{abundance}_1) + (\text{mass}_2 \cdot \text{abundance}_2) \dots$$

TRIAL 1**TRIAL 2**

--	--

6. Compare the “average mass per atom” obtained in Question 3 to the “average atomic mass” obtained in Question 5 for Trial 1. Do these values match? Should they match? Do they match for Trial 2?

Write in complete sentences and include both numerical values obtained.

7. Are your results from Trial 1 similar to the results found in Trial 2 for Question 3? Question 5?

Write in complete sentences and include numerical values obtained.

8. If you used a much larger sample of fresnoccium, would it make a substantial difference to your calculated average atomic mass? Why or why not?

Write in complete sentences.

Exp 3a – Electrons in Atoms

INTRODUCTION

When large amounts of energy are added to elements, such as when they are heated in a flame, they emit light. The light emitted is not every color of the visible spectrum, but specific colors (wavelengths).

This happens because electrons in atoms are not just spread out evenly in the space around a nucleus, but instead occupy specific energy levels or “regions” around the nucleus. To move to a higher energy level, an electron must absorb energy and when an electron moves to a lower energy level the energy must be released. The energy absorbed can come from many sources (heat, light, electricity, etc.), but energy released is mainly given off as electromagnetic radiation (light).

BACKGROUND

A spectroscope is a device used to separate light into its distinct wavelengths (a spectrum). In this experiment, we will use a simple spectroscope to analyze the visible light emitted from electron transitions in different atoms and obtain their corresponding line spectra.

The Rydberg equation (eqn 4) shows how the wavelengths (λ) of each line in the Hydrogen spectrum are related to small whole numbers, that wavelength is quantized.

$$\frac{1}{\lambda} = \frac{1}{91.13 \text{ nm}} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad (4)$$

n_i = The “initial” shell the electron is falling from.

n_f = The “final” shell electron is falling to.

λ = The wavelength of light emitted.

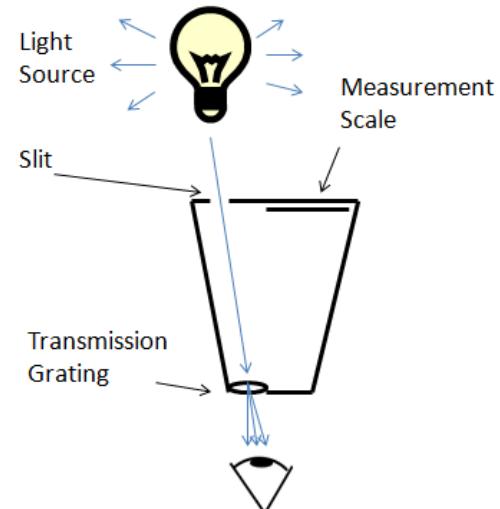


Figure 1: Spectroscope

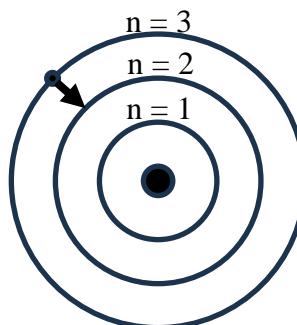


Figure 2: Bohr Model

The red line in the hydrogen spectrum corresponds to an energy change from the 3rd to 2nd level, the green line from the 4th to 2nd level, and the violet line from the 5th to 2nd level. *Energy is inversely proportional to wavelength (i.e., the energy change from the 5th to 2nd level is the largest) therefore the corresponding wavelength will be the shortest.* You will calculate the wavelengths of the red, green, and violet lines and compare them with your measured values.

ADDITIONAL READING

- The Bohr Model

SUPPLY LIST

- Spectroscopes
- Colored pencils

CLASSROOM SUPPLIES

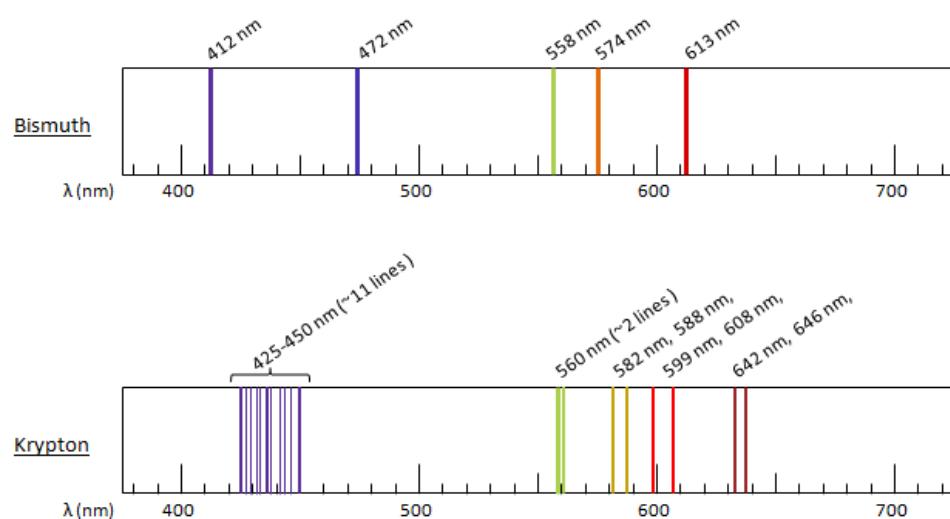
These are shared supplies or reactants. Do not gather these items. Leave these items at their designated location in the classroom and use as needed according to the procedure.

- Element lamps

PROCEDURE

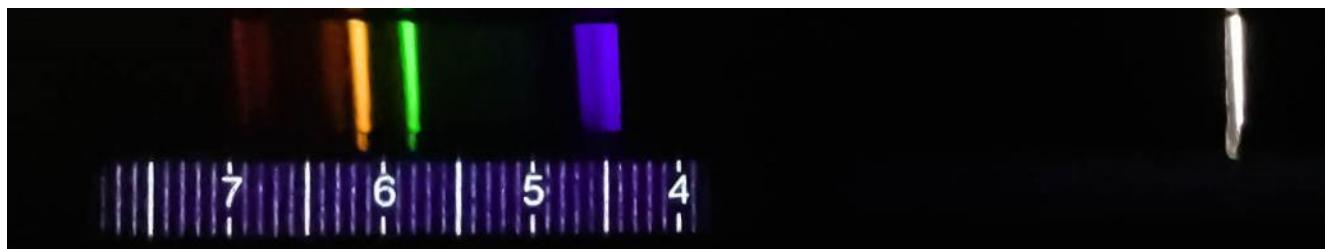
1. Aim the slit of the spectroscope at the brightest point in the lamp, then rotate your vision to see the line spectra on the measurement scale. (Once you aim the slit at the lamp, do not move the spectroscope-just move your head.)
2. Record the wavelength and color, of each strong line by drawing the spectrum on your report sheet as shown. (Use the colored pencil closest to the color of each emission line to draw the line on the chart.)

If there are many lines in the same color, instead of recording the wavelength of each, record the wavelength range of the group. If you are having a hard time reading the measurement scale, shine a light (flash light, cell phone, etc.) on that part of the spectroscope while you are looking through the viewing window.

**Figure 2: Emission Spectra for Bismuth and Krypton**

- * The scale on your spectroscope may be reversed, with the lower wavelengths on the right side of the scale. Be sure to note the actual position of the line(s) and record in the correct position on the provided charts.

* The drawings in figure 2 are what the line spectrum will look like on a high resolution spectroscope. When using the spectrosopes in this lab, the spectrum for Krypton will look something like this (figure 3).

**Figure 3: Emission Spectrum for Krypton on a Low-Resolution Spectroscope**

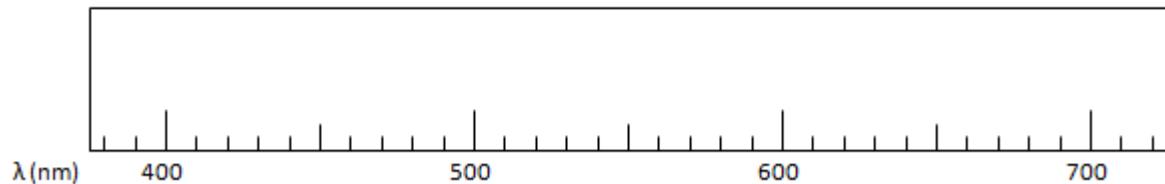
DATA AND REPORT – DUE ____ / ____ / ____**SECTION** _____

* Show your work and keep track of significant figures.

DATE _____**PARTNER** _____

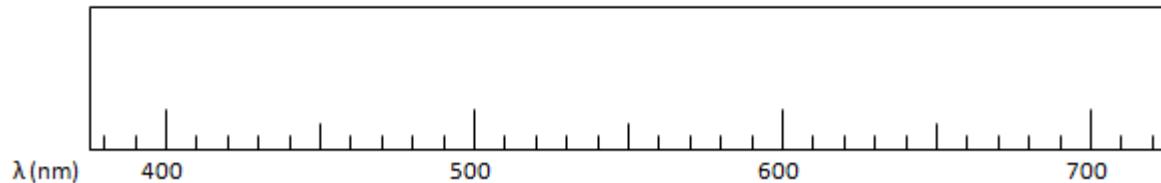
1. Helium (He)

- a. Use the spectroscope to analyze the light emitted by the helium lamp. Record your data below.



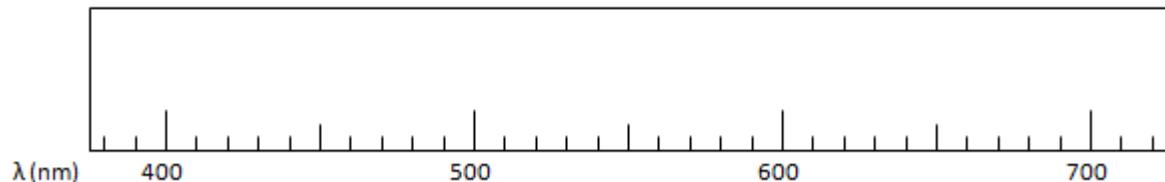
2. Nitrogen (N)

- a. Use the spectroscope to analyze the light emitted by the nitrogen lamp. Record your data below.



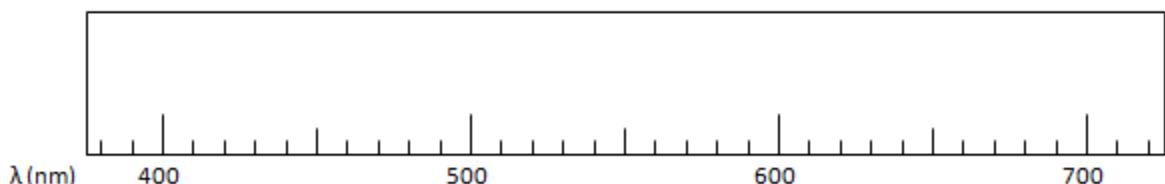
3. Neon (Ne)

- a. Use the spectroscope to analyze the light emitted by the neon lamp. Record your data below.



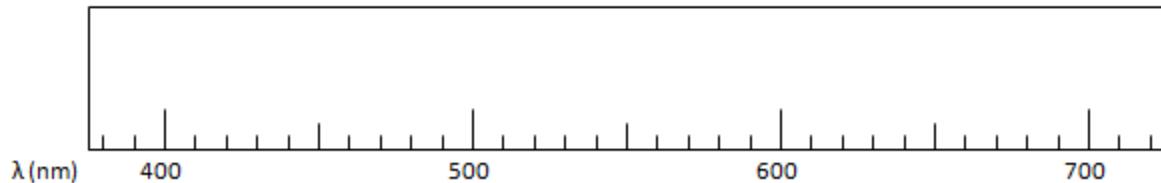
4. Mercury (Hg)

- a. Use the spectroscope to analyze the light emitted by the Mercury lamp. Record your data below.



5. Air

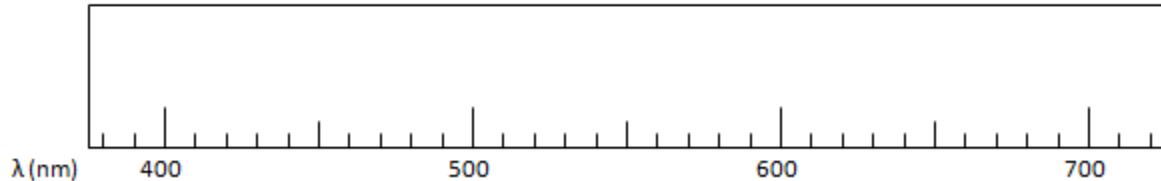
- a. Use the spectroscope to analyze the light emitted by the Air lamp. Record your values below.



- b. Of the lamps you looked at earlier, which is most likely the main component of air? Why?

6. Mystery lamp

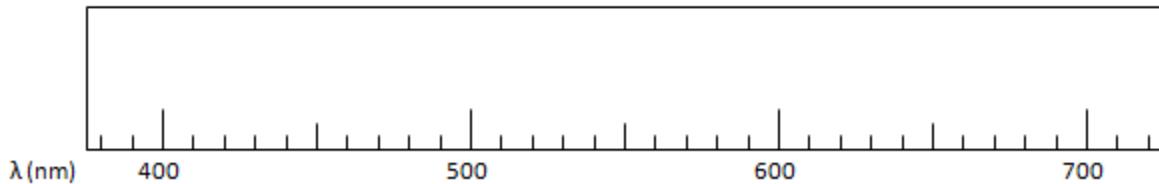
- a. Use the spectroscope to analyze the light emitted by the “Mystery” lamp. Record your values below.



- b. Of the lamps you looked at earlier, which lamp is most likely identity of the gas in the mystery lamp?

7. Hydrogen (H) – Measured wavelengths

- a. Use the spectroscope to analyze the light emitted by the hydrogen lamp. (Remember to record your data below as shown in the procedure section.)

**Calculated Wavelengths for the Hydrogen spectrum**

- a. Using the Rydberg equation calculate the wavelength of the violet line in the hydrogen spectrum.
(From $n = 5$ to $n = 2$)
- b. Using the Rydberg equation calculate the wavelength of the green line in the hydrogen spectrum.
(From $n = 4$ to $n = 2$)
- c. Using the Rydberg equation calculate the wavelength of the red line in the hydrogen spectrum.
(From $n = 3$ to $n = 2$)

- d. Use the table below to compare the values of your measured wavelengths with the values of the calculated wavelengths.

Line Color	Measured	Calculated Wavelengths	Difference
	Wavelengths from	from Rydberg Equation	(Measured – Calculated)
	Spectroscope (nm)	(nm)	
Violet			
Green			
Red			

- e. Below, the three emission lines of Hydrogen are sorted in order of increasing wavelength. Fill in the wavelengths of each line, then complete the same ranking for frequency and energy.

Increasing wavelength

Violet line (____ nm) < Green line (____ nm) < Red line (____ nm)

Increasing frequency

Increasing energy

- f. Arrange the 3 changes ($n = 3$ to $n = 2$, $n = 4$ to $n = 2$ and $n = 5$ to $n = 2$) in energy level in order of Increasing energy

- g. Does the order of energy level changes in part g match the order of energy of the emission lines in part f?

Exp 4a – Law of Constant Composition

INTRODUCTION

In this lab experiment, you will get to observe the physical properties of pure elements, and the physical properties of a compound formed from those same elements. By recording mass measurements along the way, and comparing to other students in the class, you will determine if copper and sulfur react in the same mass ratio each time.

BACKGROUND

Before modern atomic theory was developed, scientists relied upon careful mass measurements and visual observations for determining the composition of matter. In 1789, Antoine Lavoisier developed the “Law of Conservation of Mass”, which states that the mass of an element before a chemical reaction is the same as the mass of that element after the reaction. This means that mass (matter) cannot be created nor destroyed in a chemical reaction. The “Law of Constant Composition” (also known as the Law of Definite Proportion) was developed in 1797 by Joseph Proust. The law of constant composition says that all samples of a particular compound will be made up of the same elements in the same proportion or ratio by mass.

It was not until later that the English meteorologist, and chemist, John Dalton formulated the first modern description of atomic theory in 1803, which helped explain why the Law of Constant Composition held true. Dalton proposed that each chemical element is composed of atoms of a single, unique type, and though they cannot be altered or destroyed by chemical means, they can combine to form more complex structures (chemical compounds). All samples of a particular compound are made up of the same number of atoms of each element within the compound. Since Dalton reached his conclusions by experimentation and examination of the results in an empirical fashion, this marked the first truly scientific theory of the atom.

Atomic theory has been revised over the years to incorporate the existence of atomic isotopes and the interconversion of mass and energy. In addition, the discovery of subatomic particles has shown that atoms can be divided into smaller parts. However, Dalton’s importance in the development of modern atomic theory remains recognized.

In this experiment, a mixture of copper and sulfur will be heated to produce a copper and sulfur compound:



More sulfur than necessary will be used to ensure all copper reacts with sulfur. Extra sulfur will react with oxygen from the air to produce sulfur dioxide gas, leaving only the copper-sulfur compound product behind. The mass of sulfur reacted will be determined by the law of conservation of mass (and knowing the mass of copper used and copper-sulfur compound made).

ADDITIONAL READING

- Law of Conservation of Mass
- Law of Constant Composition
- Dalton's Atomic Theory

SUPPLY LIST

- | | |
|---|--|
| <ul style="list-style-type: none">▪ Copper wire▪ Test Tube▪ Bunsen Burner▪ Gas hose▪ Utility Clamp (All metal!) | <ul style="list-style-type: none">▪ Watch Glass▪ Wire Mesh▪ Flint Lighter (striker)▪ Glass stirring Rod |
|---|--|

CLASSROOM SUPPLIES AND REACTANTS

These are shared supplies or reactants.

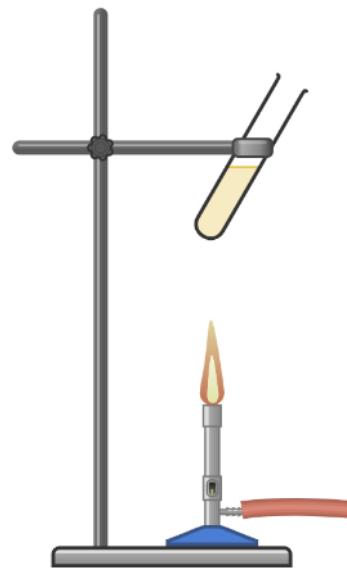
- Sulfur (solid)
- Scoopula
- Fume Hood

PROCEDURE

1. Obtain a test-tube. If the test tube is dirty, dump any solids out of the interior, rinse and dry to the best of your ability.
2. Record the mass of the empty test tube. (Remember to *always* record all digits shown on the balance.)

3. Obtain a piece of copper wire. Coil the wire around a glass stirring rod or pencil so that it looks like a spring.
4. Remove the copper wire from the stirring rod or pencil. Observe the color and tactile feel of the copper wire. Mark your observations on the report sheet.
5. Place the copper wire in the test tube.
6. Record the mass of the test tube containing the copper wire.
7. Add solid sulfur to the test tube so that the copper wire is completely covered.
8. Record the mass of the test tube containing the copper wire and sulfur.
9. Setup a Bunsen burner in the fume hood, with ring stand and utility clamp. It is important that the utility clamp is **all metal**. **Do NOT** use clamps that have cloth fibers or rubber on them. Tilt the clamp so that the mouth of test-tube is facing away from other students and the clamp is not directly in the flame. *Your instructor will demonstrate the correct setup of the clamp.*
10. Place the test tube in the utility clamp and heat the test tube until all the sulfur has reacted.
** Sulfur will melt and react with the copper. The excess sulfur will burn off as sulfur dioxide gas, which is why this should be done in the fume hood.*
11. After the test tube has finished heating, carefully remove the utility clamp from the bench post. Use the utility clamp as a handle to set it on a wire mesh square and allow it to cool back to near room temperature (about 5 minutes). Do not attempt to remove the test tube from the utility clamp until it has cooled.
** Do not set the hot test tube directly on the countertop or paper. It is extremely hot and will melt the countertop or burn paper.*
12. Once the test tube has cooled to room temperature, remove it from the utility clamp.
13. Record the mass of the test-tube containing the copper and sulfur compound.
14. Place the copper and sulfur compound on a watch-glass. It should slide out of the test tube easily.
15. Observe the color and tactile feel of the copper and sulfur compound.

Figure 1.
Bunsen Burner Setup



CLEAN-UP AND WASTE DISPOSAL

1. Place the copper and sulfur compound in the solids waste container.
2. Wash all glassware with soap and water.
3. Put back all equipment (Bunsen burners, utility clamps, etc.) back in the appropriate locations.

* Show your work and keep track of significant figures.

DATE _____

PARTNER _____

DATA

1. Description of wire before heating:

a. The wire feels: flexible rigid brittleb. The color of the wire is: dark gray light gray metallic copper____ red brown (dull) yellow

	Mass of...	Mass (g)
A	Test tube (empty)	
B	Test tube + copper	
C	Test tube + copper + sulfur	
D	Test tube + copper and sulfur compound	

2. Description of wire after heating

a. The wire feels: flexible rigid brittleb. The color of the wire is: dark gray light gray metallic copper____ red brown (dull) yellow**CALCULATIONS**

HOW	MASS OF...	MASS (g)
[B-A]	Copper wire (Initially)	
[C-B]	Sulfur (Initially)	
[D-A]	Copper and sulfur compound	
[B-A]	Copper reacted	
	Sulfur reacted	
	Excess sulfur	

1. Mass Ratio of copper / sulfur reacted based on your data (Divide): _____

MASS RATIOS OF OTHER GROUPS:

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7
Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14

2. Average mass ratio of class (excluding your data): _____

3. Calculate the percentage difference of your ratio compared to the class average:

$$\text{Percent Difference} = \frac{\text{Your ratio} - \text{Class Ratio}}{\text{Class Ratio}} \times 100$$

QUESTIONS

1. Compare the physical properties of the wire you observed before and after heating:
 - a. How are they similar?
 - b. How are they different?

2. Does the mass ratio calculated depend on the mass of copper or sulfur used in the experiment?

Explain using complete sentences.

Exp 5a – Empirical Formula

INTRODUCTION

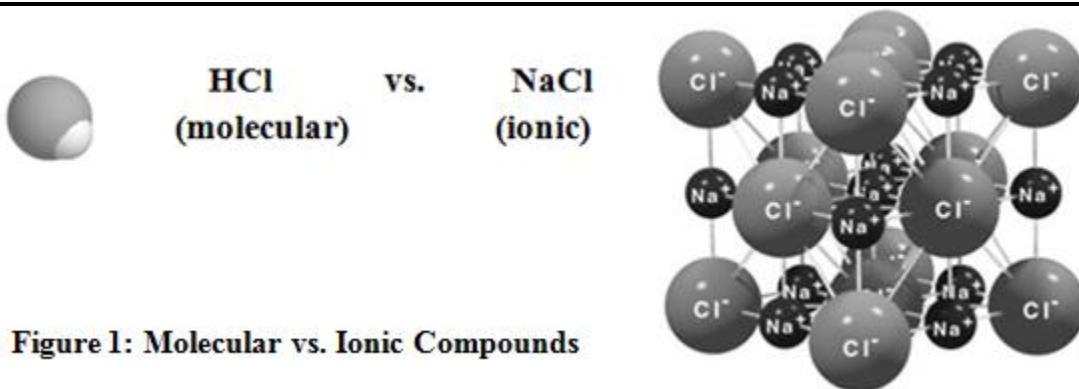
In this experiment, you will react solid magnesium with oxygen from the air to form a solid compound of magnesium and oxygen, an oxide of magnesium. Oxide compounds are a class of binary compounds of oxygen combined with another element. The goal is to experimentally determine the empirical formula (simplest whole-number ratio) of the oxide of magnesium. The unbalanced chemical equation (eqn 1) below represents the reaction of magnesium with oxygen. Note that in the formula of the oxide of magnesium x and y are used for the subscripts because they are considered unknown until they are determined using the results of this experiment.



BACKGROUND

The atom is the simplest form of an element. Compounds consist of atoms in specific ratios (water, H₂O). In fact, it is the ratio of atoms in a compound that yields its identity. This ratio of element atoms is responsible for the properties of the substance. Compounds composed of the same elements can have very different chemical and physical properties. A good example of this is hydrogen peroxide (H₂O₂) and water (H₂O). In its most concentrated form hydrogen peroxide will cause flesh to burst into flame on contact. In dilute solution, it is a disinfectant and a bleaching agent. Water does not possess these properties. Hydrogen peroxide has one more oxygen atom per molecule than water does. This small difference in the molecular formula results in large differences in chemical properties.

Compounds are generally split into two different categories: ionic and molecular. Molecular compounds are individual units with a specific number and arrangement of each type of atom, while ionic compounds are large continuous networks of atoms of the different types of elements. The number of atoms in a sample of an ionic compound varies depending on the size of crystal. An example of each type is shown below.



Since ionic compounds like NaCl (above) or Mg_xO_y have extremely large numbers of atoms that vary in quantity, the chemical formula is the smallest, simplest formula that repeats over and over. This simplest ratio is called the *empirical formula*. For molecular compounds, we write the *molecular formula*, which shows the actual number of atoms in each molecule. The molecular formula of a molecular compound is a multiple of the empirical formula of the molecular compound: the molecular formula of hydrogen peroxide H₂O₂ is twice its empirical formula HO, the molecular formula of water H₂O is once its empirical formula H₂O.

Experimental Error

As with all real-world experimentation there is a possibility that it will not work out exactly like we expect. The first step to eliminating errors is recognizing which errors are likely to occur. This experiment has a few well-known errors that may occur.

The first is incomplete combustion. It is possible that only the magnesium on the outside of the magnesium pieces reacts with oxygen. If this happens, unreacted metal will be left in the middle of the chunk of magnesium, changing the ratio of magnesium to oxygen in the final product. (Imagine if the NaCl crystal shown above had a few extra sodium atoms packed together in the middle that did not react with the chlorine gas.) To avoid this, break up the ash to expose unreacted magnesium and heat the sample over a strong flame.

A second common error that may occur is the reaction of magnesium with nitrogen instead of oxygen. The air in which the combustion of magnesium will take place is roughly 78% nitrogen and 21% oxygen. Consequently, even though the reaction with oxygen is strongly favored, it is likely that a small amount of magnesium will react with nitrogen according to the unbalanced equation (Eqn 2) shown below.



This reaction is impossible to avoid when heating in open air and will likely represent a small amount of error in your final results.

* It is very important that scientists report only the results that were observed and do not color their results with what they “hoped” to observe. (Reporting altered or “faked” data has ended more than one scientist’s career.) Even if you do not think of yourself as a scientist yet, remember that even if your results show significant errors, it is *always* better to report your actual results with an additional note about what errors you think may have occurred. It is never acceptable to try to make it seem like you observed something you did not.

ADDITIONAL READING

- Empirical and Molecular Formulas

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

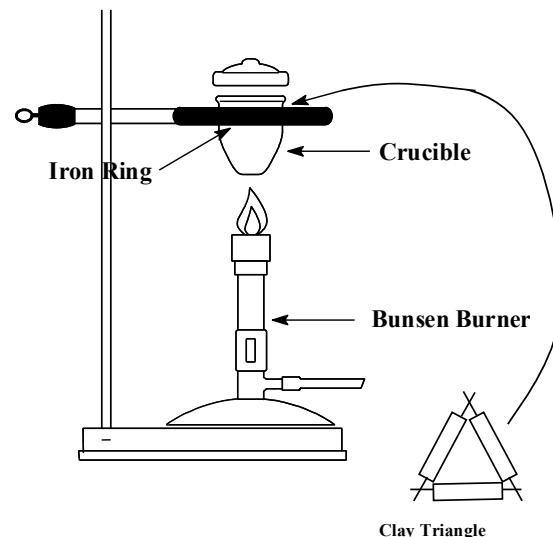
- | | | |
|----------------------------|-----------------|------------------|
| ▪ Bench Post or Ring stand | ▪ Bunsen burner | ▪ Flint Lighter |
| ▪ Iron ring | ▪ Wire mesh | ▪ Crucible Tongs |
| ▪ Clay triangle | ▪ Steel wool | |
| ▪ Crucible | ▪ Gas Hose | |

CLASSROOM SUPPLY AND REACTANTS

- Magnesium shavings
- Scoopula

PROCEDURE

1. Set up the apparatus as shown in Figure 2.
2. Obtain a porcelain crucible and lid. Be sure that the lid fits the crucible correctly. (If in doubt, ask your instructor.)
3. Dump any solids out of the interior, scrape with steel wool, and wipe with a dry paper towel.
** Don't allow the crucible to come in contact with water.*
4. Place the crucible and lid on the stand and heat over the flame until it glows slightly red-hot. Then set on the wire mesh to cool.

**Figure 2: Crucible Heating**

5. Once the crucible has cooled to room temperature (about 5 min) record the mass of the empty crucible and lid. (Remember to *always* record all digits shown on the balance.). Use the same balance to record all your mass measurements.
** After you have obtained the first mass, if you must touch the crucible at any point, you must wear gloves to avoid leaving the oils from your fingers on the crucible and adding mass.*
6. On a piece of weighing paper, obtain between 0.10 to 0.15 g of magnesium turnings. This is just an approximate amount so there is no need to write down the value on the balance. Transfer the turnings to the crucible.
7. Tare (zero) the balance. Place the crucible and lid back on the balance. Record the precise mass on your report sheet.
8. Heat the crucible with the lid slightly ajar to let air in to react with the magnesium. Use a hot flame to fully combust the magnesium for 5 minutes.
*** If the magnesium should flare do not look directly at it. ***
9. Remove the lid, continue to heat strongly for an additional 15 minutes. If the extra oxygen causes the magnesium to "flare up," place the lid back on it (ajar) temporarily.
10. After heating, set the crucible on a wire mesh and allow it to cool back to near room temperature.
It takes about 5 minutes for the crucible to cool to room temperature.
*** Do not set a hot crucible directly on the countertop or on paper. It will burn the counter or paper. Instead, when you remove it from the flame, set it on a wire mesh to allow it to cool. ***

-
11. When the heating step is complete and the crucible has reached room temperature, obtain the mass of the crucible, oxide of magnesium, and crucible lid. (This is your 1st weighing.)
* If you do not cool the crucible to room temperature, you will not be able to read a stable mass on the analytical balance. This is because the heat from the crucible causes air currents inside the balance.
 12. Reheat the crucible as before, without the lid, for 10 minutes with the lid off and cool back to room temperature (about 5 minutes).
 13. Once cooled to room temperature, re-weigh. (2nd weighing). Be sure to include the lid.
 14. If your 1st and 2nd weights are within ± 0.001 gram of each other, stop and do your calculations based on the higher of your two weights. If not, repeat step 12 until your last masses agree to within ± 0.001 gram. Completely combusted magnesium oxide should be a white powder. Because of slight contamination, you should be looking for a light gray.

CLEANUP AND WASTE DISPOSAL

15. After checking your data with your instructor, empty the oxide compound into the “solids waste” container. Scrape any excess from your crucible with steel wool before returning it to the supply bin

DATA AND REPORT – DUE ____ / ____ / ____

* Show your work and keep track of significant figures.

SECTION _____**DATE** _____**PARTNER** _____**DATA**Mass Values from Balance

1. Mass of Crucible + Lid (Empty) _____

2. Mass of Crucible + Lid + Mg _____

3. Mass of Crucible + Lid + Mg_xO_y Compound _____ (1st weighing)

_____ (2nd weighing)

RESULTSCalculated Masses

4. Mass of Mg used: _____

5. Mass of the Mg_xO_y compound: _____

(use highest of your weighings)

6. Mass of Oxygen in Mg_xO_y compound: _____

Empirical Formula

* Carry extra significant figures on values that are used in subsequent calculations.

7. Calculate the moles of Mg initially present: _____

8. Calculate the moles of O present in the compound: _____

9. Report the mole ratio without reducing. Mg _____ O _____

Example: Mg 0.006243 O 0.006151

10. Simplify the mole ratio from 9) as necessary and report the empirical formula of magnesium oxide:

Example: Mg 1.1 O 1.0

Percent Composition of Magnesium Oxide

$$\% \text{ mass} = \frac{\text{g element}}{\text{g compound}} \times 100\%$$

$$\% \text{ mass} = \frac{\text{g element}}{\text{g compound}} \times 100$$

11. Based on your experimental data (not your empirical formula), what is the percent Mg in your product?

12. Based on your experimental data (not your empirical formula), what is the percent O in your product?

POST-LAB QUESTIONS (USE VALUES FROM ONLINE REPORT)

1. Given the following analytical results: A 1.5960 g sample of the copper and sulfur compound was found to contain 0.7943 g of copper and 0.8017 g of sulfur, determine the following.
 - a. How many moles of copper and sulfur are in the sample?
 - b. What is the mass percent composition of copper and sulfur in copper and sulfur compound, based on the experimental data provided above?
 - c. What is the empirical formula for the “copper and sulfur compound”?
 - d. What is the correct name for the compound?
2. How can you tell from the chemical formula if a compound is expected to be an ionic compound or molecular compound?
3. What are structural differences between ionic compounds and molecular compounds?

Exp 6a – Molecular Modeling

INTRODUCTION

Lewis structures and ball-and-stick models are different ways to represent molecules and polyatomic ions. Lewis structures show how atoms are connected and arranged in molecules and polyatomic ions (the bonds and lone pairs) but since Lewis structures are drawn on a flat piece of paper, they do not accurately show the shape of molecules and polyatomic ions. The shape (three-dimensional representations) of molecules determines their physical and chemical properties. Ball-and-stick models are three-dimensional representations of molecules and polyatomic ions. In this experiment you will build models for molecules and draw their Lewis structures.

BACKGROUND

The model kit consists of colored balls and gray connectors. The shorter, rigid connectors (5/kit) are used to represent single bonds between two atoms. The longer, flexible connectors (4/kit) can represent double and triple bonds. Two connectors are used for a double bond (Figure 1) and three are used for a triple bond. The model kit uses different colors to represent different atoms (see table 1 below).

Table 1: Model Kit Composition

Color	Element	# holes	#/kit
white	hydrogen	1	4
black	carbon	4	3
red	oxygen	2	2
green	chlorine	1	2
orange	bromine	1	2
purple	iodine	1	2
*	nitrogen *	3	2

Many blue nitrogen balls have four holes, but one hole is plugged with a small white peg, glue, or covered with tape so that it only has three available holes. All three holes in each ball will have a connector when the model is built correctly.



Figure 1. Models showing a single bond (left) and a double bond (right)

Lewis structures use dots to represent the valence electrons, but these drawings are two dimensional and often do not accurately depict the three-dimensional shape of molecules. The molecular modeling sets help visualize the three-dimensional shape of the molecules. Valence Shell Electron Pair Repulsion

Theory allows us to assign shape names to the observed three-dimensional shapes. For example, the Lewis structure for OCl₂ drawn in Figure 2 shows the atoms connected in a straight line, which does not accurately represent the shape of the molecule. The central oxygen has four electron groups around it (Figure 2). A molecule with four electron groups around the central atom has a **tetrahedral** electron group geometry because those electron groups want to get as far apart as possible. Two of the electron groups are lone pairs. Although the lone pairs affect the three-dimensional shape, they cannot be seen. The observed molecular geometry is **bent**. The names for all the electron group geometries and molecular geometries can be found in Table 1.

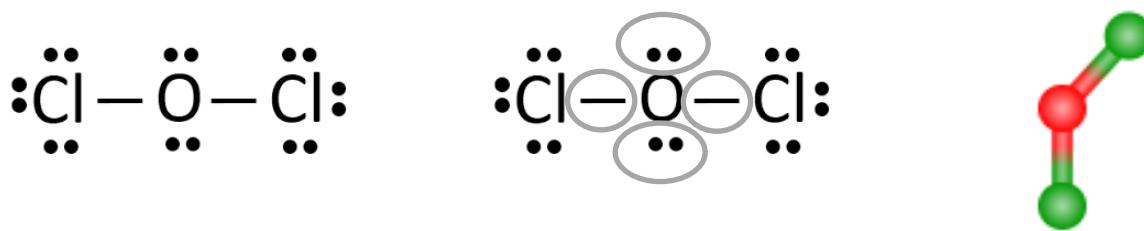


Figure 2: A) The Lewis structure for OCl₂ drawn where it appears to be linear. B) The Lewis structure for OCl₂ with electron groups circled. C) The molecular model for OCl₂ shows bent molecular geometry.

Number of Electron Groups	Electron Group Geometry	Ideal Bond Angles	0 lone pairs	1 lone pair	2 lone pairs
			Molecular Geometry		
2	Linear	180°	Linear		
3	Trigonal Planar	120°	Trigonal Planar	Bent	
4	Tetrahedral	109.5°	Tetrahedral	Trigonal Pyramidal	Bent

Table 1. Electron group geometry and molecular geometry are based on the number of electron groups.

ADDITIONAL READING

- Lewis Structures, VSEPR

SAFETY EQUIPMENT REQUIRED FOR LAB

- No safety equipment needed

SUPPLY LIST

- Molecular model kit
- Color Pencils

PROCEDURE – PART 1 – LEWIS STRUCTURE FIRST

1. Using the periodic table, add up the total number of valence electrons based on the chemical formula.
2. Sketch the Lewis Dot structure.
 - a. Each atom (except hydrogen) should have an octet (8 electrons). Hydrogen should have a duet (2 electrons).
3. Build a molecular model based on your Lewis Dot Structure. Each hole on the model will be filled when the model is correctly assembled.
4. Sketch the molecular model in the *Molecular Model* column showing the three-dimensional structure.
 - a. Show single bonds with single lines, double bonds with two lines and triple bonds with three lines.
5. Use colored pencils to color your drawing.
6. Draw the Lewis structure for each molecule. Use the structural formula and number of valence electrons as a guide.

PROCEDURE – PART 2 – MOLECULAR MODEL FIRST

7. Similar to before, only now attempt to build the model BEFORE drawing the Lewis Structure.
You will know your model is correct when each hole on the model is filled.
8. Your instructor will provide models of unknown molecules for the last part. For each molecule:
 - a. Sketch the model using the appropriate colors.
 - b. Write the chemical formula (in the first column).
 - c. Determine the number of valence electrons.
 - d. Draw the Lewis structure.

DATA AND REPORT – DUE ____ / ____ / ____

SECTION _____

* Show your work.

DATE _____

PARTNER _____

Part 1: Draw the Lewis Structure First**MOLECULAR MODELS WITH SINGLE BONDS ONLY**

#0	OCl ₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:
20	$\begin{array}{c} \ddot{\text{:O}} \\ \\ \text{:Cl}-\text{O}-\text{Cl}: \\ \\ \text{:Cl}: \end{array}$		4
			Electron Group Geometry:
			Tetrahedral
			Molecular Geometry:
			Bent

#1	H ₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:

			Electron Group Geometry:
			linear
			Molecular Geometry:
			linear

#2	HCl		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:

			Electron Group Geometry:
			linear
			Molecular Geometry:
			linear

#3	CH₂Cl₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:
			Electron Group Geometry:
			Molecular Geometry:

#4	H₂O		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:
			Electron Group Geometry:
			Molecular Geometry:

#5	NH₃		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:
			Electron Group Geometry:
			Molecular Geometry:

MOLECULAR MODELS ONE DOUBLE BOND

#6	O₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:

			Electron Group Geometry:
			linear
			Molecular Geometry:
			linear

#7	C₂H₄		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups around C:
			Electron Group Geometry around C:
			Molecular Geometry around C:

#8	C₂H₃Cl		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups around C:
			Electron Group Geometry around C:
			Molecular Geometry around C:

MOLECULAR MODELS ONE TRIPLE BOND

#9	N₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups:

			Electron Group Geometry:
			linear
			Molecular Geometry:
			linear

#10	C₂H₂		
Valence Electrons	Lewis Dot Structure	Molecular Model Sketch	Number of Electron Groups around C:
			Electron Group Geometry around C:
			Molecular Geometry around C:

Part 2: Build the Model First**MOLECULAR MODELS WITH SINGLE BONDS ONLY**

#11	H₂O₂			
	Molecular Model Sketch	Valence Electrons	Lewis Dot Structure	Number of Electron Groups around O:
				Electron Group Geometry around O:
				Molecular Geometry around O:

#12	N₂H₄			
	Molecular Model Sketch	Valence Electrons	Lewis Dot Structure	Number of Electron Groups around N:
				Electron Group Geometry around N:
				Molecular Geometry around N:

#13	NH₂OH			
	Molecular Model Sketch	Valence Electrons	Lewis Dot Structure	Number of Electron Groups around N:
				Electron Group Geometry around N:
				Molecular Geometry around N:

MOLECULAR MODELS WITH ONE DOUBLE BOND

#14	HONO			
		Valence Electrons	Lewis Dot Structure	Number of Electron Groups around N:
				Electron Group Geometry around N:
				Molecular Geometry around N:

#15	HCOOH			
		Valence Electrons	Lewis Dot Structure	Number of Electron Groups around C:
				Electron Group Geometry around C:
				Molecular Geometry around C:

MOLECULAR MODELS WITH ONE TRIPLE BOND

#16	HO₂N			
		Valence Electrons	Lewis Dot Structure	Number of Electron Groups around C:
				Electron Group Geometry around C:
				Molecular Geometry around C:

MOLECULAR MODELS WITH TWO DOUBLE BONDS

#17	CO₂			
	Molecular Model Sketch	Valence Electrons	Lewis Dot Structure	Number of Electron Groups around C:
				Electron Group Geometry around C:
				Molecular Geometry around C:

#18	C₃H₄			
	Molecular Model Sketch	Valence Electrons	Lewis Dot Structure	Number of Electron Groups around central C:
				Electron Group Geometry around central C:
				Molecular Geometry around central C:

UNKNOWN MOLECULAR MODELS

Molecule	Molecular Model	Valence Electrons	Lewis Structure
#1			
#2			
#3			
#4			
#5			

Exp 7a - Calorimetry

INTRODUCTION

Calorimetry is a technique used to measure heat given off or absorbed by a process. Since it is not possible to measure heat directly, calorimeters generally measure temperature changes, which are related to heat by the following equation:

$$q = m \cdot c \cdot \Delta T \quad (6)$$

q = heat, usually measured in J or cal

m = mass of the substance

c = specific heat capacity of the substance

ΔT = change in temperature in °C (measured as the final temperature – initial temperature)

The physical change or chemical change (reaction) is generally considered the “system”, while the solution in which the reaction takes place would be considered the “surroundings”. In the calorimeter setup for this experiment, the thermometer is placed in the solution (surroundings). Since under these conditions, significant heat is not gained or lost, but only transferred from one thing to another, the sum must equal zero, as shown in the equation below.

$$q_{system} + q_{surroundings} = 0 \quad (7)$$

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

- Calorimeter (2 Styrofoam cups, 1 cardboard lid, thermometer)
- Bench Post
- Thermometer clamp
- 100-mL graduated cylinder
- Wash Bottle filled with DI Water

CLASSROOM SUPPLY AND REACTANTS

- Citric acid (solid)
- Sodium Carbonate (solid)
- Water
- Ice
- 250-mL beaker
- Spatula
- Tongs

PART 1 – Exothermic and Endothermic Changes**BACKGROUND**

Many solids will absorb or release heat when they dissolve in liquids due to the interactions with the solvent molecules. We will dissolve the solid citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) and sodium carbonate (Na_2CO_3), separately, in water to observe this process.

Endothermic versus Exothermic

When a process releases heat, we call the process *exothermic*. (If an exothermic process happens in a beaker, it will feel warmer because it is giving off heat to your hand.)

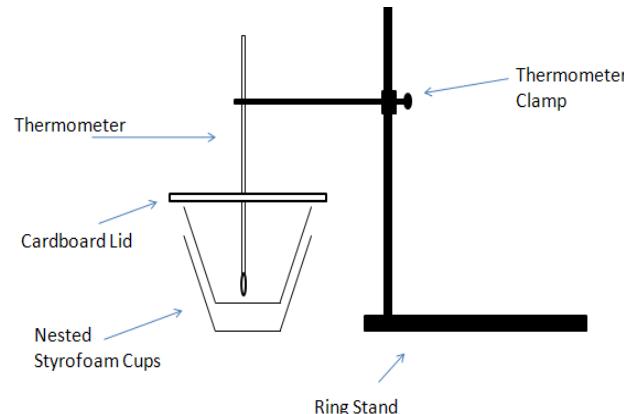
When a process absorbs heat, we call the process *endothermic*. (If an endothermic process happens in a beaker, it will feel colder because it is absorbing heat from your hand.)

PROCEDURE

1. Set up a calorimeter as shown in figure 1. Two Styrofoam cups are used to provide an extra insulating layer, which slows down the transfer of heat between the cup and the surroundings.
2. Obtain the mass of your empty calorimeter (two Styrofoam cups).
**** Do **not** include the mass of the cardboard lid in your measurement.
3. Using a graduated cylinder measure between 45.0-55.0 mL DI water. Pour into your calorimeter.
4. Record the initial temperature of the water.
5. Weigh between 4.8-5.2 g of citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) on a piece of weigh paper.

- a. Use a brush to sweep any spills inside the balance out. Wasted material can go into the aqueous inorganic waste container.
- b. Please ensure the bottle lid is screwed back on before leaving the balance area.

6. Add the citric acid to the calorimeter, then place the cardboard lid back onto the calorimeter, and *carefully* stir the mixture with the thermometer until all the solid dissolves.

**Figure 1: Calorimeter Set-**

7. Record the highest temperature (or lowest temperature if the temperature goes down) reached as the final temperature.
8. Record the mass of the calorimeter and the solution.
**** Do **not** include the mass of the cardboard lid in your measurement.
9. Empty the contents of the calorimeter into a large temporary waste beaker, then rinse and dry your calorimeter.
10. Repeat steps 3-8 using sodium carbonate (Na_2CO_3) instead of citric acid.
11. Slowly empty the contents of the calorimeter into a large temporary waste beaker. Once waste from citric acid and sodium carbonate are combined, waste can be poured down the drain.
**** Combining citric acid and sodium carbonate will produce bubbles (carbon dioxide gas).

PART 2 – Specific Heat Capacity and Enthalpy of Fusion**BACKGROUND**

When two objects at different temperatures come into contact with each other, heat from the hotter object will transfer to the cooler object until they reach the same temperature. The first law of thermodynamics (law of conservation of energy) states, that “Energy is never created or destroyed.” If we focus only on heat energy and ignore other types of energy, this could be stated as follows:

$$q_1 + q_2 + q_3 + q_4 + \dots = 0 \quad (8)$$

In other words, the heat of one object might go up and the heat of one object might go down, but the total amount of heat will stay the same. In this experiment, we will be putting ice into water and seeing if the sum of the heats is zero.

Heat from the water will transfer to the ice according to the equation below:

$$q = m \cdot c \cdot \Delta T \quad (9)$$

The ice will absorb the heat lost by the water as it melts. The heat gained by the ice will be calculated as:

$$q = n \Delta H_{fus} \quad (12)$$

For water, $\Delta H_{fus} = 6.01 \text{ kJ/mol}$

PROCEDURE

1. Using a graduated cylinder measure 45.0-55.0 mL DI water.
2. Record the mass of your calorimeter containing water.
**** Do **not** include the mass of the cardboard lid in your measurement.
3. Record the initial temperature of the water to the nearest 0.1 °C.
4. Bring your calorimeter, and cardboard lid over to the bin filled with ice.
5. Carefully transfer 3-4 squares of ice to your calorimeter. The squares of ice may be separate or frozen together, either is fine.
6. Place the lid back on the calorimeter and stir using the thermometer until all the ice melts. Record the lowest temperature reached to the nearest 0.1 °C as the final temperature.
7. Record the mass of the calorimeter containing the water and melted ice.
8. Empty the contents of the calorimeter (just water) down the sink drain.

CLEAN-UP AND WASTE DISPOSAL

9. Once you finish collecting all your data, return your calorimeter's cardboard lid to the supply bin.
10. Throw away the Styrofoam cups in the trash can.

DATA AND REPORT – DUE ____ / ____ / ____

* Show your work and keep track of significant figures.

SECTION _____**DATE** _____**PARTNER** _____

PART 1 – Exothermic and Endothermic Changes

DATA

Mass of calorimeter (empty): _____

	H ₃ C ₆ H ₅ O ₇	Na ₂ CO ₃
Initial Temperature (°C)		
Final Temperature (°C)		
Mass of calorimeter <i>and</i> solution (g)		

PART 2 – Specific Heat Capacity and Enthalpy of Fusion

DATA

	Data
Mass of calorimeter + water (g)	
Initial Temperature of water (°C)	
Final Temperature of water (°C)	
Mass of calorimeter + water + ice (g)	

CALCULATIONS – PART 1 – EXOTHERMIC AND ENDOThERMIC CHANGES

* Show your work, complete with units for $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$. I will assume you did Na_2CO_3 the same way

1. Mass of the solution

3. Heat of Solution (q_{solution})

2. Change in Temperature of
Solution

RESULTS – PART 1 – ENDOThERMIC AND EXOTHERMIC CHANGES

	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	Na_2CO_3
Mass of Solution (m)		
Specific Heat of Solution (c)	4.184 J/g°C	4.184 J/g°C
Change in Temperature of Solution(ΔT)		
Heat of solution (q_{solution})		
Heat of reaction* ($q_{\text{reaction}} + q_{\text{solution}} = 0$)		
Reaction is: Exothermic or Endothermic		

* The water in the calorimeter is considered the reaction's surroundings, therefore if heat is released by the reaction, it is absorbed by the water and vice-versa.

CALCULATIONS – PART 2 – SPECIFIC HEAT CAPACITY AND ENTHALPY OF FUSION

* Show your work, complete with units.

1. Mass of H₂O

4. Mass of Ice

2. Change in Temperature of Water

5. Moles of Ice

3. q_{water}6. q_{ice}7. q_{water} + q_{ice}**RESULTS**

	Results
Mass of H ₂ O	
Specific Heat of water (c)	4.184 J/g°C
Change in Temperature of water (ΔT) ($\Delta T = T_f - T_i$)	
q _{water} (kJ)	
Mass of Ice	
Moles of Ice (n)	
Enthalpy of Fusion (ΔH_{fus})	6.01 kJ/mol
q _{ice} (kJ)	
q _{water} + q _{ice}	

QUESTIONS

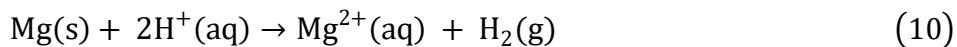
1. Is ice melting endothermic or exothermic? *Provide data or results from this experiment to support your claim.*

2. Was energy conserved (i.e. is $q_{\text{water}} + q_{\text{ice}} = 0$)? If not, provide at least one error specific to your experiment that could account for the difference.

Exp 8a – Determination of the Gas Constant R

INTRODUCTION

Magnesium reacts with acids according to the following net ionic equation.



In this experiment, magnesium will be reacted with hydrochloric acid and the volume of hydrogen gas produced will be measured. The chemical equation for this reaction says that for every 1 mol of Mg used, 1 mol of H₂ gas will form. From this volume, atmospheric pressure, and temperature, the value of "R", the ideal gas constant will be determined.

BACKGROUND

Ideal Gas Law

The ideal gas law (eqn 2) relates the variables of a gas: moles of gas (n), pressure (P), volume (V) and temperature (T). The "R" is the Ideal Gas Constant. Since R is a constant, it's value is always the same: 0.082058 L atm/(mol K). In this lab, we will be experimentally determining the Ideal Gas Constant (R) and comparing it against the accepted value.

$$PV = nRT \quad (11)$$

Dalton's Law of Partial Pressures and Vapor Pressures

Dalton's Law of Partial Pressures states that when gases are mixed, the total pressure of the mixture is equal to the sum of the individual partial pressures of the gases.

$$P_T = P_1 + P_2 + P_3 \dots \quad (12)$$

When a gas is collected “over water” water vapor will be present as part of the mixture because the water will be evaporating to some extent into the gas phase. The pressure of water vapor at different temperatures is known and is recorded on tables in your textbook and in the CRC Handbook of Chemistry and Physics®. The pressure of hydrogen alone can then be determined by applying Dalton's law.

$$P_T = P_{\text{Hydrogen}} + P_{\text{Water}} \quad (13)$$

Accuracy and Precision

Accuracy is how close your experimental value is to the accepted (“true”) value. Precision is how close the experimental values are to each other. Accuracy and precision are largely independent of each other because differences in accuracy or precision are caused by different types of experimental errors.

Systematic errors always occur in the same direction for each repetition and therefore affect accuracy. These may be due to measurement device error, such as a thermometer that always read 1 °C high, or procedural errors, such as not heating your reaction mixture long enough so that the reaction never reaches completion. The deviation from accuracy (arising from systematic errors(s)) can be quantified by calculating the percent error as shown below.

$$\% \text{ Error} = \frac{\text{experimental value} - \text{accepted value}}{\text{accepted value}} \times 100\%$$

* Other terms are often used in place of the term accepted value, such as theoretical, literature, expected or true value.

ADDITIONAL READING

- Ideal Gas Law, Mixtures of Gases, and Gas Stoichiometry

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

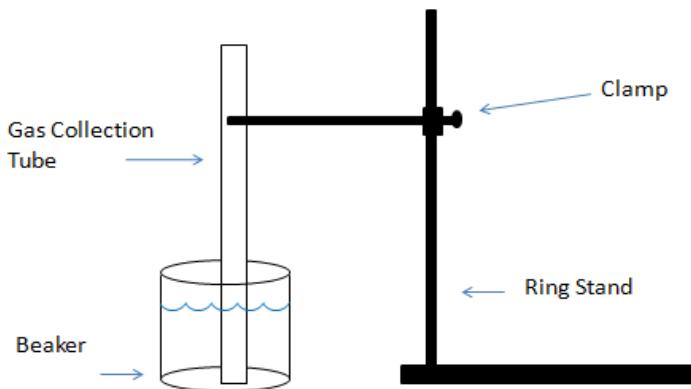
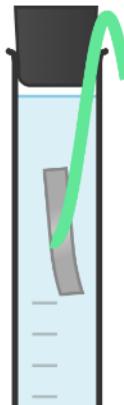
- | | | |
|---|------------------|--------------------------------|
| • 400, 600, or 800-mL beaker | • Rubber Stopper | • Utility Clamp & Clamp holder |
| • 50-mL gas collection tube
(eudiometer) | • Steel wool | • Fishing line (~15 cm) |
| | • Bench Post | |

CLASSROOM SUPPLY AND REACTANT LIST

- HCl (6 M)
- Magnesium ribbon

PROCEDURE

1. Obtain a piece of magnesium ribbon approximately 1.5 cm in length and clean it with steel wool.
* Check with your instructor if you need to do 2 trials.
2. Measure and record the precise mass of the ribbon.
3. Set up a bench stand, medium clamp, and clamp adapter as shown in figure 1.
4. Fill a large beaker about halfway with DI water.
5. Pour 8 to 10 mL of 6 M hydrochloric acid into your gas collection tube (eudiometer).
6. Fill the remainder of your tube by *slowly* pouring DI water into your tube all the way to the brim.
** Try to minimize the mixing of water with the acid. **
7. Gently bend the magnesium into a slight U-shape so that it will fit inside the tube. Be careful not to break the magnesium.
8. Tie the Mg Ribbon to one end of a piece of fishing line (~15 cm long)
9. Place the Mg Ribbon into the tube, holding the fishing line over the outer rim.
Place the rubber stopper into the tube to hold the fishing line in place.
10. While holding a finger over the hole in the rubber stopper, invert the tube into your beaker, then attach it to the bench post with the clamp as shown in figure 1.
Remove your finger from the hole once it is under water.
11. While the magnesium is reacting:
 - a. Record the temperature of the water (assumed to be the same as the temperature of the gas).
 - b. Record the pressure in the room from the classroom pressure gauge (assumed to be the same pressure as the gas in the tube) *.

**Figure 1: Gas Collection Set-up**

- * If the water in the collection tube is not the same as the water level in the beaker, it will cause the pressure in the collection tube to be different than the pressure in the room. For this experiment, we will ignore this difference, which will be a small source of error in the experimental results.
- c. Look up the partial pressure of water in the CRC Handbook®.

12. When the magnesium has finished reacting, record the volume of the gas. You will know the reaction is complete when the magnesium ribbon is no longer visible, bubbles stop forming, and the volume of gas stops changing.

CLEAN-UP AND WASTE DISPOSAL

13. After checking data with your instructor, lift the tube and remove the stopper to allow the solution to drain into the beaker.
14. With the beaker inside the sink, add sodium bicarbonate (baking soda) to neutralize the acid before dumping it down the drain.
15. Wash the beaker with soapy water. Rinse with DI water and return to the shelf. There is no need to dry the equipment thoroughly.
16. Rinse the gas collection tube with soapy water and rinse using your wash bottle (the gas collection tube does not fit in the sink well).
17. Return equipment.

CALCULATIONS

1. Use the mass of Mg to calculate the moles of hydrogen gas.
2. Use the total atmospheric pressure, and the vapor pressure of water, to calculate the pressure of hydrogen gas according to Dalton's Law of partial pressure.
3. Convert the pressure of hydrogen gas into atmospheres (atm).
4. Rearrange the equation $PV = nRT$ to solve for R. Plug in experimental values for P_{H_2} , V, n_{H_2} , and T, ensuring units match the desired units for R: $\frac{L \cdot atm}{mol \cdot K}$

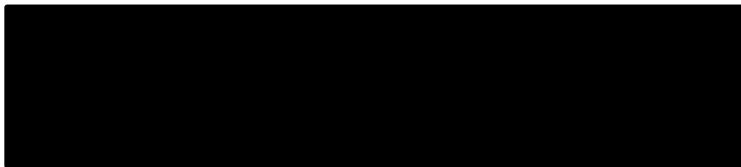


Figure 2:
Hold this black bar behind the eudiometer, below the meniscus to make it easier to read.

DATA AND REPORT – DUE ____ / ____ / ____**SECTION** _____

* Show your work and keep track of significant figures.

DATE _____**PARTNER** _____**DATA**

Mass of the magnesium ribbon _____

* Record a citation for where you found the vapor pressure of water in the space below.

Temperature of water _____

Pressure in the room _____

Vapor pressure of water at your temperature* _____

Volume of gas in collection tube _____

CALCULATIONS

Show your work, complete with units.

1. Moles of hydrogen gas based on mass of Mg used _____

2. Pressure of hydrogen gas (mmHg) _____

3. Pressure of hydrogen gas (atm) _____

4. Experimentally determined R value _____

5. % Error _____

6. If magnesium ribbon was not cleaned well, it will directly (pressure, volume, moles, temperature) which will cause the experimental R value to be (higher, lower, no change).
7. If the gas collection tube was not fully filled with water and an air bubble was trapped inside reaction started, the error would directly affect (pressure, volume, moles, temperature) which will cause the experimental R value to be (higher, lower, no change).
8. Discuss your percent error. Based on your observations, what is at least one source of error specific to your experiment? (Mention the numerical value for your percent error in your response)

Exp 9a – Solutions and Dilutions

INTRODUCTION

In this experiment, you will make two solutions. One solution will be made by dissolving a solid solute (salt, NaCl). The other solution will be made by diluting a small amount of the first solution with water to make a new solution with new concentration. You will then perform various calculations to express the concentration of salt in both solutions in different units.

BACKGROUND

Bodies of water around the world can be classified as fresh water, brackish water, saline water or briny water depending on the amount of salt dissolved in them. Saline solutions (made with table salt, NaCl) have many uses in medicine including nasal washes, eye drops, and IV therapy. Brines are solutions that vary from somewhat concentrated to saturated. Because many bacteria do not grow well in high concentrations of sodium chloride, brines are occasionally used as food preservative (pickling).

In this experiment, you will prepare two sodium chloride (NaCl) solutions of varying concentrations. You will first mix sodium chloride and water to prepare a more concentrated “stock” solution and then you will make one “working” solution by diluting the stock solution.

There are various units of concentration that can be used to describe the amount of salt in each solution:

$$\text{Molarity (M)} = \frac{\text{mol solute}}{\text{L solution}} \quad (1)$$

$$\text{Mass Percent } \left(\frac{m}{m} \% \right) = \frac{\text{mass solute}}{\text{mass solution}} \times 100 \quad (2)$$

$$\text{Molality (m)} = \frac{\text{mol solute}}{\text{kg solvent}} \quad (3)$$

$$\text{Mass Volume Percent } \left(\frac{m}{v} \% \right) = \frac{\text{mass solute in g}}{\text{mL solution}} \times 100 \quad (4)$$

Additionally, the density of each solution can be calculated by:

$$\text{Density Solution} = \frac{\text{mass solution}}{\text{volume solution}} \quad (5)$$

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat

SUPPLY LIST

- | | | |
|---------------------------|--------------------------|---------------------|
| ▪ Sodium chloride (solid) | ▪ 25-mL volumetric flask | ▪ Wash bottle |
| ▪ 10-mL volumetric pipet | ▪ Pipet bulb | ▪ One small stopper |
| ▪ 50-mL volumetric flask | ▪ Permanent marker | ▪ One large stopper |

CLASSROOM SUPPLY AND REACTANT LIST

- Sodium chloride (solid)
- Food coloring

PROCEDUREMake Solution “A”

1. Use a permanent marker (Sharpie®) to label a 50-mL volumetric flask “A.”
2. Record the precise mass of the 50-mL volumetric flask (empty) on a balance.
3. Use a piece of weigh paper to obtain between 4.5-6.0 g of NaCl.
4. Pour the NaCl into the 50-mL volumetric flask. It is okay if some spills on the countertop. Spills can be cleaned up by sweeping the salt into the trash can.
5. Record the precise mass of the 50-mL flask now that it contains salt.
6. Use a water bottle (filled with DI water) to fill the flask approximately halfway with water. Add 1 drop of food coloring. The food coloring is to make your solutions visually different.
7. With the stopper on, swirl the flask until the NaCl is mostly dissolved then fill to the line with DI-Water and mix by inverting at least 10 times.
8. If needed, carefully add more water so that the bottom of the meniscus is at the line.
9. Record the precise mass of the 50-mL flask now that it contains “Solution A”.

Make Solution “B”

1. Use a Sharpie® to label a 25-mL volumetric flask “B.”
2. Record the precise mass of the 25-mL volumetric flask (empty) on a balance.
3. Use a 10-mL volumetric pipet to transfer 10.00 mL of “Solution A” to the flask for “Solution B”.
4. Use a water bottle (filled with DI water) to fill the flask to the line with water.
5. With the stopper on, swirl the flask and mix by inverting at least 10 times.

6. Record the precise mass of the 25-mL flask now that it contains “Solution B”.

CLEAN-UP AND WASTE DISPOSAL

1. After checking your data with your instructor, all solutions can be poured down the drain.
2. Flush volumetric pipets thoroughly with DI water from wash bottle.
3. Wash volumetric flasks with soapy water. Rinse with DI water.
4. Rinse the inside of the volumetric flasks with a small amount of acetone, to help them dry faster.
5. Remove any permanent marker using acetone.
6. There may be drying racks available for the volumetric flasks. If drying racks are available, place the volumetric flasks upside down in the rack. If no rack is available, place equipment back in the original location.

CALCULATIONS

- Refer to the definitions of concentration units found in the ‘Background’ section of the experiment.
- It is recommended that you complete the calculations in the order presented in the data report, even though this is not the same order found in the Results table.
 - Rounded values should not be used in calculations. Write down the calculator output, and store values in your calculator so that you have the full number.
 - Round values for significant figures when reporting your answers in the Results table.
- Volumetric flasks and pipets are very precise. When considering significant figures in calculations, use the precise volumes below:

	Precise Volume
50-mL Volumetric Flask	50.00 mL
25-mL Volumetric Flask	25.00 mL
10-mL Volumetric Pipet	10.00 mL

DATA AND REPORT – DUE ____ / ____ / ____

* Show your work and keep track of significant figures.

SECTION _____**DATE** _____**PARTNER** _____**DATA**

	Solution A
Mass of 50-mL volumetric flask + stopper (empty)	
Mass of 50-mL volumetric flask + stopper + NaCl	
Mass of 50-mL volumetric flask + stopper + “Solution A”	

	Solution B
Mass of 25-mL volumetric flask + stopper (empty)	
Mass of 25-mL volumetric flask + stopper + “Solution B”	
Volume of “Solution A” used to make “Solution B”	

CALCULATIONS

Show your mathematical work, complete with units.

1. Mass of solute in Solution A:

2. Mass of solution for Solution A:

3. Mass of solvent for Solution A:

4. Moles of solute in Solution A:

5. Molarity (M) of Solution A:

6. Mass % (m/m%) of Solution A:

7. Molality (*m*) of Solution A:

8. Mass-Volume % of Solution A:

9. Density of Solution A:

10. Moles of solute in 10.00 mL of Solution A:

11. Mass of solute used to make Solution B:

12. Mass of solution for Solution B:

13. Mass of solvent for Solution B:

14. Molarity (M) of Solution B:

15. Mass % (m/m%) of Solution B:

16. Molality (*m*) of Solution B:

17. Mass-Volume % of Solution B:

18. Density of Solution B:

RESULTS

	Solution A	Solution B
Mass Solute		
Mass Solvent		
Mass Solution		
Moles of Solute		
Molarity (M)		
Mass Percent (m/m %)		
Mass-Volume Percent (m/v %)		
Molality (<i>m</i>)		
Density of Solution		

Exp 10a – Types of Reactions

INTRODUCTION

There are a nearly infinite number of chemical reactions possible. Because of this, it is helpful to break them down into a few different classes of reactions. In this experiment, you will carry out and observe characteristic reaction(s) for each type.

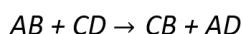
BACKGROUND

At the introductory or general chemistry level, there are two basic schemes to categorize reactions- that help us understand and predict what will happen when two chemicals are mixed. Because they describe different aspects of the reaction, they are best understood together.

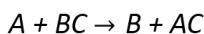
Scheme 1

What is happening with the atoms?

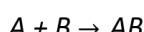
Double Displacement (“DD”) →



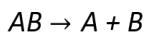
Single Displacement (“SD”)



Combination



Decomposition



Scheme 2

What overall process is taking place?

Precipitation

A “DD” reaction that forms a solid product

Acid-Base/Neutralization

A “DD” reaction between an acid and base.

Gas Evolution

A “DD” reaction that forms a gaseous product.

No Reaction

A “DD” reaction with only aqueous products.

Oxidation/Reduction (Redox)

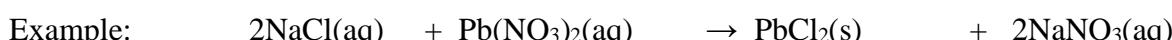
A reaction where electrons are transferred.
(Often “SD”, Combination, or Decomposition)

↳ Combustion

A Redox reaction with gaseous oxygen.

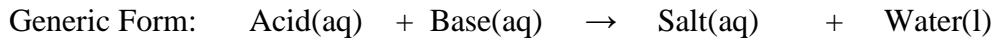
Precipitation Reactions

A precipitation reaction is a double displacement reaction where two aqueous solutions are mixed, and an ionic solid is formed in the product.

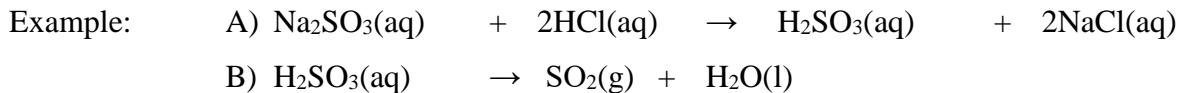
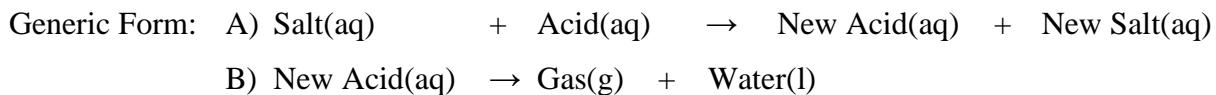


Acid-Base Reactions

An acid-base reaction (also called a neutralization reaction) is a double displacement reaction between an acid and a base where water (or another weak acid) is formed as one of the products.

Gas Evolution Reactions

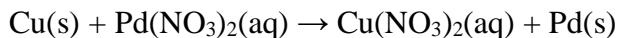
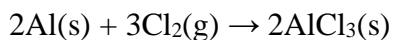
A gas evolution reaction is a double displacement reaction where a gas is formed as one of the products. Commonly the gas is formed when one of the aqueous products formed decomposes into water and a molecular gas. The common products that decompose are H₂CO₃, H₂SO₃, and NH₄OH.

No Reaction

If there is no observable change when reactants are mixed, then the reaction should be classified as “no reaction”. This can be seen in the chemical equation when a double displacement reaction forms only aqueous products.

Redox Reactions

Oxidation-Reduction (redox) reactions involve the transfer of electrons. Although identifying these reactions can involve further calculations, most can be identified simply by the presence of a neutral, uncombined element in either the reactants or products. (As in Single Displacement, Combination, and Decomposition reactions.)



* Note that if, as in the first example above, a gas is produced as a result of a redox reaction, it is best to classify the reaction as redox and not gas evolution.

Combustion Reactions

A combustion reaction is a special type of redox reaction, where a hydrocarbon (compound composed mainly of hydrogen and carbon) or a metal reacts with oxygen gas. Although these are a type of redox reaction, they are so common that they should be classified as their own type of reaction.

Generic Form: Hydrocarbon(g, l, or s) + Oxygen(g) → Carbon Dioxide(g) + Water(g)

Example: C₂H₄(g) + 3O₂(g) → 2CO₂(g) + 2H₂O(g)

Generic Form: Metal(s) + Oxygen(g) → oxide of metal(s)

Example: 2K(s) + O₂(g) → K₂O(s)

Molecular and Complete Ionic Equations

There are different formats for writing chemical equations. All the equations above are written as “molecular equations” where all products are written as intact compounds. However, in aqueous solution, strong electrolytes dissociate into separate ions. The complete ionic equations demonstrate this:

	Aqueous Strong Electrolytes	Solids, Liquids, and Gases		
<i>Molecular Equation</i>	2NaNO ₃ (aq)	PbCl ₂ (s)	H ₂ O(l)	CO ₂ (g)
	Written as separate ions:			Written as intact compounds:
<i>Complete Ionic Equation</i>	2Na ⁺ (aq) + 2NO ₃ ⁻ (aq)	PbCl ₂ (s)	H ₂ O(l)	CO ₂ (g)

ADDITIONAL READING

- Chemical Reactions, Molecular, Complete Ionic and Net Ionic Equations

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

- Test tubes
- Test Tube Rack

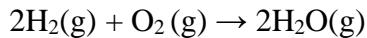
CLASSROOM SUPPLY AND REACTANT LIST

- HCl (6.0 M)
- Na₂CO₃ (1 M)
- Zinc pellets
- NaOH (3.0 M)
- KNO₃ (0.1 M)
- Red litmus paper
- BaCl₂ (0.1 M)
- CuSO₄ (0.1 M)
- Transfer pipets
- Na₃PO₄ (0.1 M)
- Na₂CO₃ (solid)
- Spatula
- CaCl₂ (0.1 M)
- KClO₃ (solid)
- Matches
- Na₂CO₃ (0.1 M)
- NH₄Cl (solid)

PROCEDURE**General Notes**

- a. All reactions are to be carried out in test tubes. All quantities may be considered approximate.
- b. It is necessary to observe the reaction mixture for at least a few minutes to determine the character of any precipitate (solid) that forms.
- c. When 1 mL of a reagent is specified, you may use approximately 20 drops from a Pasteur/transfer pipet or just pour one finger width into the test tube. There is no need to measure it precisely.
- d. When a reaction calls for 1 g of a reactant this is a pea-sized quantity. There is no need to mass it with precision.
- e. Record your own observations; photocopies or hand copies of someone else's work are not acceptable.
- f. The reagents are set up by station number around the lab room.
- g. You are free to run the reactions in any convenient order.
- h. Dump all waste into the waste containers provided. In chemistry labs, *never dump any waste down the drain* unless specifically instructed to do so.

Station 1. Transfer about 3 mL of 6.0 M HCl to a test tube. Add a few zinc pellets to the tube. Record your observations. Is the gas that evolves CO₂ or some other non-flammable gas? Or is it H₂ gas which is very flammable when it combusts with the formation of water?



* Evaluate the flammability of the gas by inverting a test tube over the one containing the reaction mixture. Capture a sample of the gas and light it with a match. If the gas is hydrogen, the combustion will yield a pop sound.

Station 2. In a test tube mix 1 mL of 0.1 M BaCl₂ with 1 mL of 0.1 M Na₂SO₄.

What happens when you mix the two reactants? What does the test tube look like after 3-5 minutes? As part of your observations look-up the solubility of sodium chloride and barium sulfate on the solubility table handout. Based on the reactants and the given product what is the formula of the insoluble product you observe in this experiment?

Station 3. In a test tube mix 1 mL of 0.1 M BaCl₂ with 1 mL of 0.1 M Na₃PO₄.

What happens when you mix the two reactants? What does the test tube look like after 3-5 minutes? As part of your observations, look-up the solubility of sodium chloride and barium phosphate on the solubility table handout. Based on the reactants and the given product what is the formula of the insoluble product you observe in this experiment?

Station 4. In a test tube mix 2 mL of 1.0 M Na₂CO₃ with 1 mL of 6.0 M HCl.

If a gas evolves, what compound is responsible for it? What happens to the solution over time? Report the overall chemical reaction that accounts for the production of any gas you observe.

Station 5. Transfer about 1 g of KClO₃ to a dry test tube. In the hood, while holding the test tube with a test tube clamp, gently warm it using a Bunsen burner. Heat until nothing appears to be happening. What does the solid look like? Do you see anything in the neck of the test tube? Record your observations associated with this station.

Station 6. In a test tube mix 1 mL of 0.1 M CaCl₂ with 1 mL of 0.1 M Na₂CO₃.

What happens when you mix the two reactants? What does the test tube look like after 3-5 minutes? As part of your observations look-up the solubility of sodium chloride and calcium carbonate in the solubility table. Based on the reactants and the given product what is the formula of the insoluble product you observe in this experiment?

Station 7. In a test tube mix 1 mL of 6.0 M HCl with 2 mL of 3.0 M NaOH.

What happens when you mix the two reactants? Is there any temperature change? What kind of reaction is this? What are the identities of the products you expect in this experiment?

Station 8. In a test tube mix 1 mL of 0.1 M KNO₃ with 1 mL of 0.1 M BaCl₂.

What happens when you mix the two reactants? What does the test tube look like after 3-5 minutes? As part of your observations look-up the solubility of potassium chloride and barium nitrate on the solubility table. Does this reaction yield a product? If not, why?

Station 9. Transfer about 3 mL of 0.1 M CuSO₄ to a test tube. Add a few zinc pellets to the tube.

What happens when you mix the two reactants? What does the test tube look like after 3-5 minutes?

As part of your observations, look-up the activity of zinc and copper on the activity series table handout.

Based on the reactants and the given products what is the identity of the solid product you observe in this experiment? What accounts for the blue color of the CuSO₄ solution?

Station 10. In a test tube mix 1 mL of 5.0 M of NH₄Cl with 1 mL of 3.0 M NaOH.

If a gas evolves, identify it. Hold a piece of moistened red litmus over the test tube without touching the liquid solution in the test tube. A color change from red to blue confirms ammonia as a product. Do you detect any odor coming from the test tube? A pungent smell is also a confirmation of ammonia (NH₃). What happens to the solution over time? Report the overall chemical reaction that accounts for the production of any gas you observe.

CLEAN-UP AND WASTE DISPOSAL

1. Station 1: Dump test tube into temporary waste beaker at the station. Use forceps to remove large zinc pellets. Rinse with DI water and place on paper towel. These zinc pellets can be used by other students.
2. After checking your data with your instructor, dump all test tubes in the aqueous inorganic waste container.
3. If large quantities of solids are lodged inside the test tube, use a water bottle filled with DI water to rinse the solids into the aqueous waste container.
4. Using a small test tube brush, wash the test tubes with soapy water. Rinse with DI water.
5. Place the clean wet test tubes upside down in the test-tube rack to drain.

DATA AND REPORT – DUE ____ / ____ / ____

SECTION _____

DATE _____

PARTNER _____

Record your observations then complete and balance the chemical reaction equation. Identify each chemical reaction equation by type. Include all the data requested in the procedure. Make your own observations; don't simply copy someone else's.

Station 1. Zn(s) and HCl(aq)

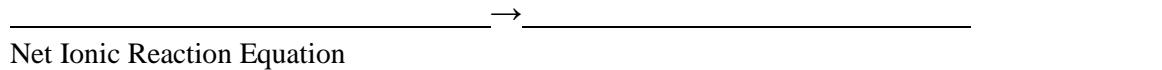
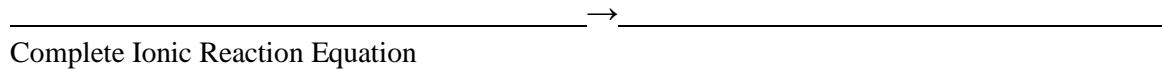
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

**Station 2.** BaCl₂(aq) and Na₂SO₄(aq)

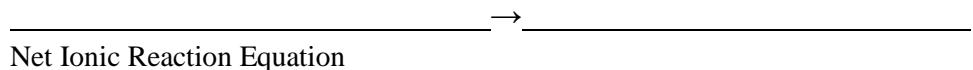
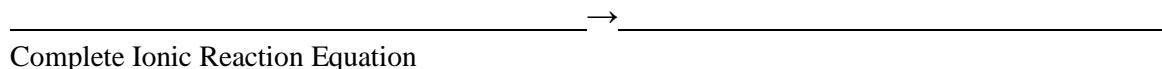
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

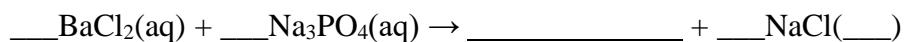


Station 3. BaCl₂(aq) and Na₃PO₄(aq)

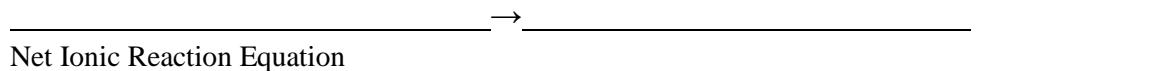
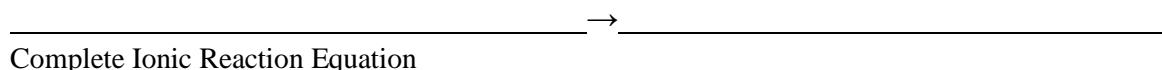
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation



Station 4. HCl(aq) and Na₂CO₃(aq)

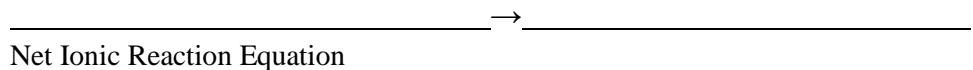
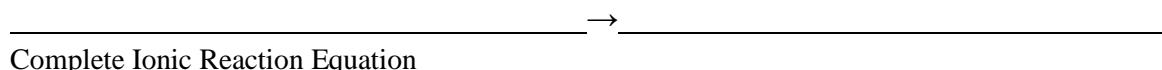
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation



Station 5. KClO₃

Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

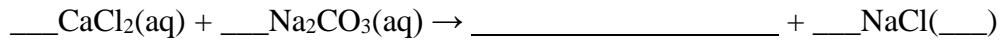
This station does not take place in aqueous solution so you will not write complete ionic or net ionic equations for this process.

Station 6. CaCl₂(aq) and Na₂CO₃(aq)

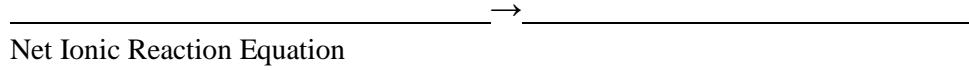
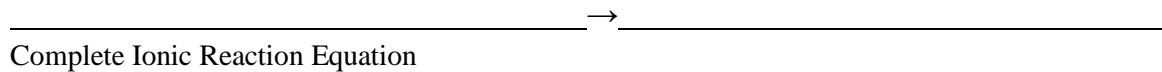
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation



Station 7. HCl(aq) and NaOH(aq)

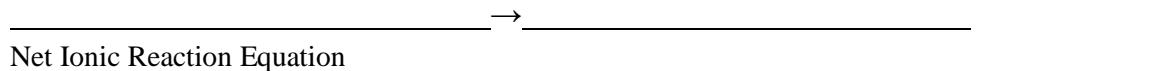
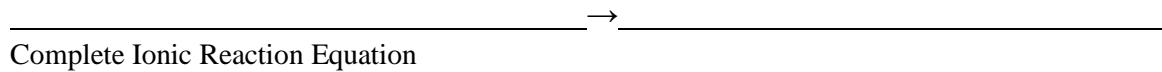
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

**Station 8.** KNO₃(aq) and BaCl₂(aq)

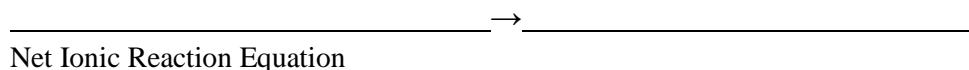
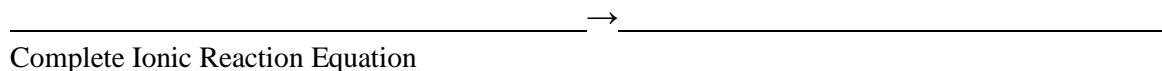
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

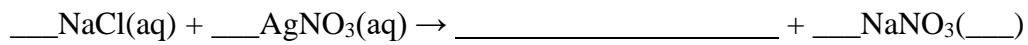


Station 9. Zn(s) and CuSO₄(aq)

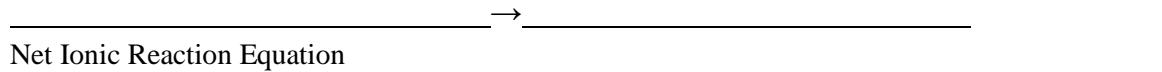
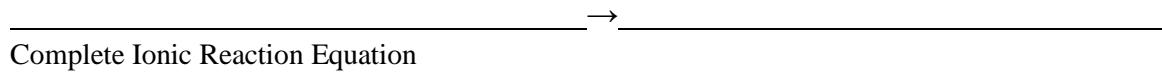
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation

**Station 10.** NH₄Cl(aq) and NaOH(aq)

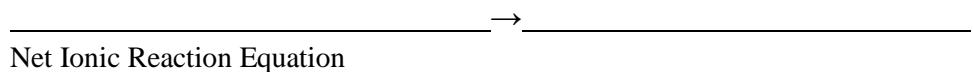
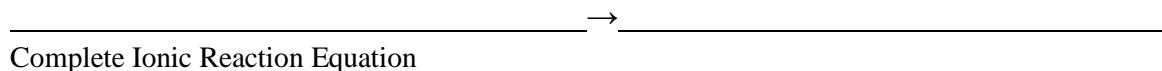
Reaction Type, Scheme 1: _____

Scheme 2: _____

Observations:



Balanced Molecular Reaction Equation



QUESTIONS

1. Identify each of the following as a salt, acid, base, or hydrocarbon.

a. Ammonium hydroxide

d. $\text{HC}_2\text{H}_3\text{O}_2$ _____

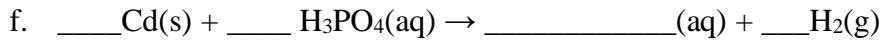
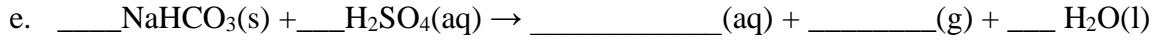
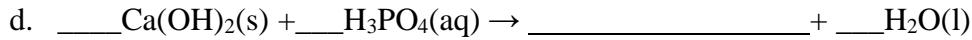
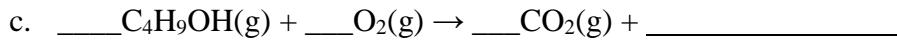
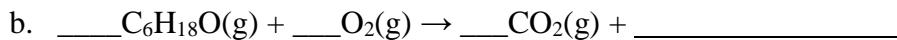
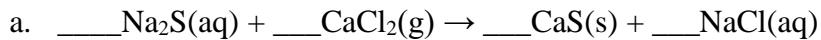
b. CaC_2O_4 _____

e. $\text{C}_2\text{H}_6\text{O}$ _____

c. $\text{Ba}(\text{ClO}_4)_2$ _____

f. $\text{H}_2\text{S}_{(\text{aq})}$ _____

2. Balance the following chemical reaction equations.



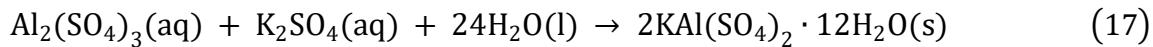
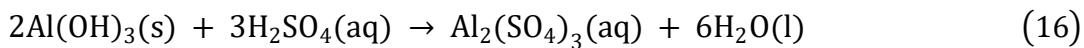
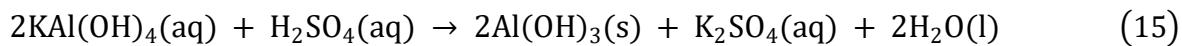
Exp 11a – Synthesis of Alum

INTRODUCTION

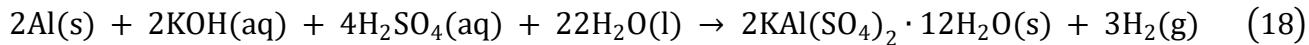
In this experiment, we will synthesize potassium aluminum sulfate dodecahydrate ($\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$) from solid aluminum. We will use stoichiometry to predict the mass of alum we can make. From the experimental results, it is possible to calculate the % yield of this reaction.

BACKGROUND

The reaction to form Alum from aluminum takes place in the following steps.



These equations can be combined to give the following total chemical equation for the process.



Some interesting facts about Alum

Potassium aluminum sulfate dodecahydrate, or “Alum” for short, has been collected and processed for centuries. It is used in industry as a water clarifier (in water purification), in the manufacture and application of dyes, in the tanning of leather, and in marble and porcelain cements. It is also the ingredient in pickles that gives them their crispness. Medicinally it is used as an astringent (a substance that draws together or constricts tissues) and a styptic (to control bleeding). Alum is used extensively in the paper industry as a sizing agent (a material that fills the spaces between the paper fibers to control its porosity and to bind other additives).

ADDITIONAL READING

- Stoichiometry and Yield Calculations

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

- | | | |
|------------------------------|-----------------------------|----------------------|
| ▪ 50-mL beaker | ▪ 100-mL graduated cylinder | ▪ Medium watch glass |
| ▪ 150, 200, or 250-mL beaker | ▪ Transfer pipets | ▪ Büchner funnel |
| ▪ 600-mL beaker | ▪ Glass-stirring rod | ▪ Side-arm flask |
| ▪ 10-mL graduated cylinder | ▪ Spatula | ▪ Filter paper |

CLASSROOM SUPPLY AND REACTANT LIST

- | | | |
|--|---------------------------|---------------|
| ▪ KOH (1.5 M) | ▪ Aluminum foil or powder | ▪ Bench stand |
| ▪ H ₂ SO ₄ (6.0 M) | ▪ Ice | ▪ Clamp |
| ▪ Alcohol solution (50 %) | ▪ Vacuum hose | |

PROCEDURE

** Warning: both potassium hydroxide and sulfuric acid are very corrosive and are very damaging to skin, eyes, clothes, etc. Be very careful when working with or around them! **

1. Record the mass of a clean dry small beaker (150, 200 or 250-mL).
2. Obtain approximately 0.4 g aluminum foil and add it to the beaker by tearing or cutting it into small pieces (about the size of a pencil eraser)
3. Record the mass of the beaker with the aluminum foil.
4. Label your beaker with your name with a permanent marker (Sharpie ®).
5. Using a graduated cylinder, add 20 mL of 1.5 M KOH to the beaker. The aluminum may take about 20 minutes to completely react with the KOH. (It has completely reacted when it is a black solution and there are no pieces of aluminum foil left.) During this time, calculate your theoretical yield of alum.

** Work in the fume hood. The hydrogen gas emitted can generate a fine mist of corrosive potassium hydroxide. **

* If the reaction slows, and does not seem to be proceeding, you may heat up the beaker briefly on a hot plate but be careful to ensure that the solution does not boil.

6. When the aluminum has completely reacted, obtain 7.5 mL of 6.0 M H₂SO₄ using a 10-mL graduated cylinder and pour into 50-mL beaker.
7. Using a transfer pipet, add the acid dropwise to the reaction mixture, while stirring with a glass-stirring rod.

* After the addition of the acid, the reaction mixture is usually greyish white.

8. Cover the beaker with a watch glass and warm the solution on a hot plate until the solid has completely dissolved. (The solution will be a transparent tan color with no visible solid chunks)
 ** Do not boil the solution! Hot concentrated sulfuric acid is very corrosive. **
9. When the solid has completely dissolved, set the beaker aside for a few minutes to cool.
10. While the beaker is cooling, either locate the ice bath set up by your instructor or make one yourself by putting ice and water into a beaker larger than the one you are using for the reaction mixture.
11. Once the beaker has cooled a little, place it in the ice bath and allow it to cool for 15 minutes.
12. Swirl the beaker when you notice the onset of crystal formation. If alum crystals are not starting to form within 5 minutes, it may be necessary to scratch the inside walls of the beaker with a glass stirring rod. This provides sites for crystallization to begin.
13. While the solution is cooling, pour 30 mL of 50 % alcohol (ethanol) into a small beaker and place it in the ice bath. The alcohol will be used to rinse your crystals out of the beaker during the filtration step. (If your instructor has the alcohol solution in ice, you can wait until you are ready to filter to obtain it)
14. Set up the vacuum filtration apparatus as shown in Figure 1, if it is not already set up in the classroom.
15. Take your beaker with alum crystals and 50 % alcohol solution to the vacuum apparatus, then:

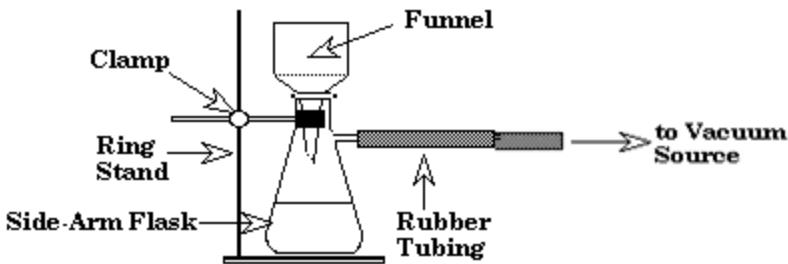


Figure 1: Vacuum Filtration Apparatus

- Turn the vacuum on so that air is being drawn through the filter paper. Leave the vacuum on through the entire process below.
 * Make sure that all the holes in the funnel are covered. Dampen the filter paper so that it stays in place.
- Swirl the beaker so that all the crystals are dislodged and dump the entire contents of the beaker quickly into the center of the filter paper within the Buchner funnel.
- Pour about 10 mL of the cooled 50 % alcohol solution into the beaker that contains the alum, then swirl the alum beaker and pour the mixture into the funnel to transfer any remaining crystals while washing the crystals already in the funnel. (Repeat until you have used all the alcohol solution.)
- Leave the vacuum on for a few minutes until the alum crystals look somewhat dry, then turn off the vacuum. You can gently poke at it with a stir rod to see how dry it is.

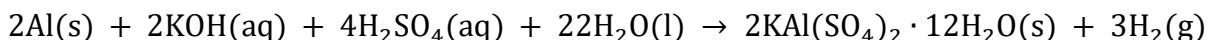
16. Obtain the mass of a clean dry watch glass, then transfer the alum crystals to the watch glass and allow them to continue to air dry.
17. Obtain a final mass for the watch glass and alum.
18. Record your experimental (actual) yield and calculate your percent yield for the synthesis of alum.
19. Discuss your percent yield. (Does this number seem reasonable? Is this what you would expect?)

CLEANUP AND WASTE DISPOSAL

1. Dispose of the alum crystals in the solids-waste container
2. Dispose of the waste from the filtration step (in the side-arm flask) in the inorganic aqueous waste container
3. Dispose of the filter paper and plastic transfer pipet in the trash can.
4. Use acetone to wipe off permanent marker on equipment.
5. Wash all glassware with soapy water. Rinse with DI water. There is no need to thoroughly dry the equipment prior to returning it to the shelves.

CALCULATIONS

1. Determine the mass of aluminum added to the reaction beaker.
2. Based on the mass of aluminum (Al), calculate the theoretical yield of alum ($\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$) that could be made:



3. Calculate the mass of alum you actually obtained in the experiment.
4. Calculate the percent yield:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

DATA AND REPORT – DUE ____ / ____ / ____

* Show your work and keep track of significant figures.

SECTION _____**DATE** _____**PARTNER** _____**REFERENCE VALUES FROM PRE-LAB QUIZ**

2. Molar Mass of aluminum _____

3. Molar Mass of Alum _____

4. Reaction ratio _____ mol Al : _____ mol $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ **DATA**

1. Mass of the reaction beaker _____

5. Mass of the beaker + aluminum _____

6. Mass of watch glass _____

7. Mass of watch glass + alum _____

RESULTS*Show your work for each of the following calculations.*

8. Mass of aluminum _____

9. Theoretical yield of Alum _____

10. Mass of Alum _____ (actual yield)

11. Percent yield _____

12. Discuss your percent yield. (Is this a reasonable value? Why or why not?)

Exp 12b – Determination of Citric Acid in Soda

INTRODUCTION

A titration is a way of carefully controlling the amounts of reactants added in a chemical reaction, so that the amount or concentration of one of the reactants may be determined from the amount and/or concentration of the other. In this experiment, the concentration of citric acid in lemon-lime soda will be determined by titration (reaction) with sodium hydroxide.

EXPERIMENTAL BACKGROUND

In a titration, one reactant is placed in a buret (titrant), while the other is placed in an Erlenmeyer flask. In this experiment, sodium hydroxide will be the titrant and the acid that it will react with will be placed in the Erlenmeyer flask.

A buret has a measurement scale that starts with the zero mark near the top and then values increase downward. When reading a buret, you are recording the location of the meniscus. Later, the amount of titrant used will be calculated by taking the difference (final – initial) of the buret readings. When reading a buret, it is important to be eye-level with the buret. When at eye-level, the lines on the buret will be parallel. It is useful to hold a white piece of paper behind the buret to make the measurement scale clear. A thick black bar on the paper helps to highlight the bottom of the meniscus. In Figure 2, notice how “39” is at the top of the photograph, and “41” is at the bottom. The bottom of the meniscus is just below the “40”. The smaller calibration lines represent 0.1 mL increments. The line below “40” represents 40.1 mL. The bottom of the meniscus is slightly less than halfway between “40.0” and “40.1” and therefore will be recorded as 40.04 mL.

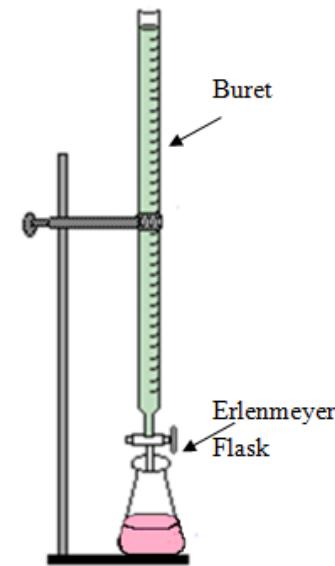


Figure 1:
Titration Apparatus

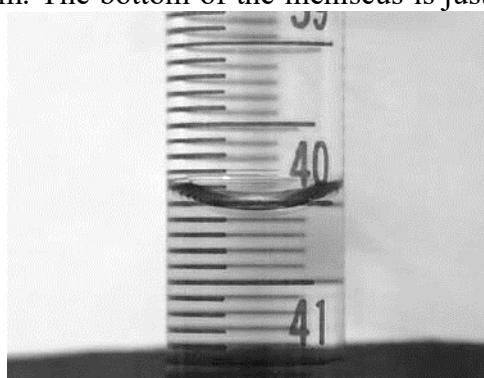


Figure 2:
Close-up of buret

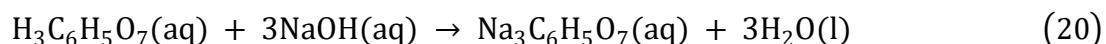
Part I- Standardization

The first part of many titration experiments is the determination of the concentration of the titrant, called *standardization*. This is especially necessary for sodium hydroxide solutions, as they will react with carbon dioxide in the air, which lowers the molarity over time.

In this experiment, the molarity of a sodium hydroxide solution will be determined by reaction with oxalic acid as shown below.

**Part II- Determination**

In the second part of the experiment, called determination, we will use the concentration of the sodium hydroxide from part I to determine the concentration of citric acid in a lemon-lime soda (6-Down Soda). The reaction of sodium hydroxide with citric acid is shown below.



If a solid is used in a titration, then water is added to the flask to dissolve the solid. Water is not a reactant. An indicator is added to the flask that will change color when the acid has completely reacted (endpoint). Proper operation of the buret will allow the experimenter to add just the right amount of sodium hydroxide to react with the acid.

The indicator used in this experiment is called phenolphthalein. Phenolphthalein is colorless in acidic conditions but turns pink in basic conditions. This means that the slightest excess of sodium hydroxide will turn the color of the solution a light pink (endpoint). (Any additional sodium hydroxide will turn the solution darker pink and represents a significant amount of experimental error.)

ADDITIONAL READING

- Solution Stoichiometry
- Acid-Base Titrations

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST**Part I and II:**

- 400 or 600-mL beaker
(temporary waste
container)
- 25-mL buret
- 125-mL Erlenmeyer flask
- 10-mL graduated cylinder
- Bench stand
- Buret clamp
- Spatula
- Wash bottle filled with DI
water

Part II:

- 100-mL beaker
- 10-mL volumetric pipet
- Pipet bulb or pump

CLASSROOM SUPPLY AND REACTANT LIST

- NaOH (~0.1 M)
- H₂C₂O₄·2H₂O (solid)
- Lemon-lime soda (6
Down)
- Phenolphthalein (1 %)

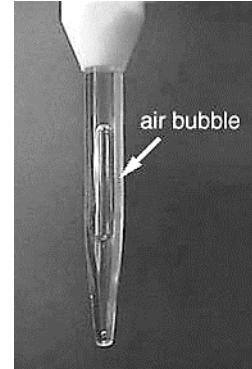


Figure 3:
**Hold this black bar behind the
buret, below the meniscus to make
it easier to read.**

PROCEDURE

Setup and Cleaning the Buret

1. Set-up the titration apparatus as shown in Figure 1.
2. Obtain a large beaker to use at your desk for temporary waste collection.
3. Rinse your buret three times using a wash bottle filled with DI-water (5 mL each) and then three times with the sodium hydroxide (NaOH) solution (5 mL each) and drain into the waste beaker before filling your buret with the NaOH solution. This ensures that only the sodium hydroxide solution is in your buret.
4. Place your buret in the clamp. Place the large beaker under the tip of the buret.
With the stopcock of the buret in the open position, add several mL of NaOH.
When the NaOH begins to flow out of the tip, close the stopcock. This should force out any air bubble in the tip and fill the tip with NaOH.
 - a. If there is an air bubble, you may remove it by fully opening the stopcock and allowing the bubble to pass through the opening. (If this doesn't work, consult your instructor for help.)



Part I- Standardization

5. Fill your buret with ~0.1 M NaOH to between the 0-3mL marks near the top of the buret, but below the 0.00 mL mark.
6. Record your initial buret reading, being sure to estimate one decimal place beyond the smallest division.
7. Bring your Erlenmeyer flask with you to the balance.
 - a. Place weigh paper on the balance and tare (zero) the balance. The mass should read 0.0000 g.
 - b. Measure out between 0.1000 and 0.1200 grams of oxalic acid dihydrate onto the weighing paper and record the precise mass on your report page.
 - c. Pour the oxalic acid into the flask, being careful not to spill.
8. Use a graduated cylinder to transfer approximately 10 mL of deionized water to your clean, well-rinsed Erlenmeyer flask, and swirl to dissolve the oxalic acid.
* If it does not completely dissolve at this point, it is not a huge problem. Just be sure that it dissolves shortly into the titration, well before the endpoint.
9. Add 2 to 3 drops of phenolphthalein indicator to the flask.

10. Place a white piece of paper under your Erlenmeyer flask to better see the color change of the endpoint.
 11. While swirling the flask, quickly add 10 mL of NaOH from the buret. After 10-mL of NaOH, add the NaOH more slowly.
 12. Watch the center of the flask as the NaOH is added. You should be able to see a pink color where the NaOH drops land. Towards the endpoint it will take more time for the pink color to dissipate back to colorless. The endpoint is reached when a persistent faint-pink color is obtained throughout the solution.
 - a. Towards the end add fractions of a drop by quickly rotating the stopcock 180° until a lasting faint pink color has been reached. Alternatively, you can get a partial drop hanging off the edge of the tip of the buret, and then touch the drop to the inside of the flask to knock it in.
 13. Read and record your final buret volume.
 14. Repeat steps 5-13 for the number of trials indicated by your instructor. *Redo titrations that are clearly overdone.* (Turn bright pink at the end).
- **Do not drain your buret once you are done with Part I. Refill with NaOH for Part II. **

Part II- Determination

15. Obtain about 40 mL of soda in a small beaker. (The soda in your lab has been opened days in advance to assure that it is de-carbonated.)
16. Use a 10-mL volumetric pipet to transfer 10.00 mL of soda to your clean, well-rinsed Erlenmeyer flask. Record the precise volume of soda used (10.00 mL).
17. Add 2 to 3 drops of phenolphthalein indicator to the flask.
18. Fill your buret with ~0.1 M NaOH to between the 0-3mL marks near the top of the buret. Record your initial buret reading being sure to estimate one decimal place beyond the smallest division.
19. Titrate the soda as outlined above in steps 10-12.
20. Read and record your final buret volume.
21. Repeat this process for the number of trials indicated by your instructor. *Redo titrations that are clearly overdone.* (Turn bright pink at the end).

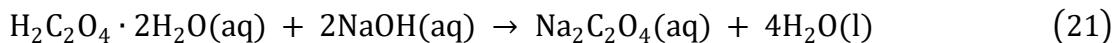
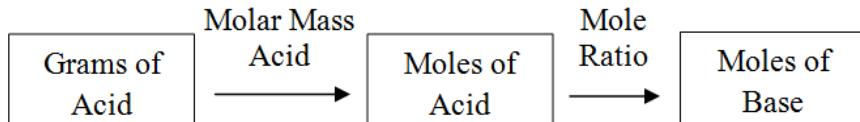
CLEANUP AND WASTE DISPOSAL

1. Pour your waste beaker into the gigantic aqueous inorganic waste beaker located in the fume hood. Your instructor will ensure that the waste is neutral before it is poured down the drain.

CALCULATIONS

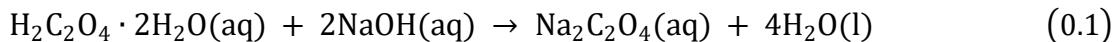
Part I: Standardization

1. Use the now known grams of oxalic acid to calculate the moles of the sodium hydroxide used.



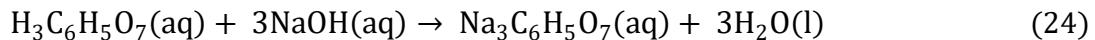
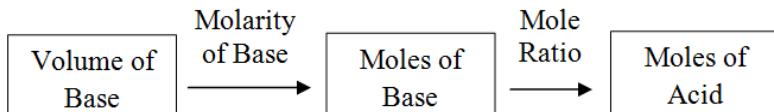
2. Calculate the net mL of NaOH added. Convert to L.
3. Calculate the molarity of sodium hydroxide from the moles of sodium hydroxide and the L of sodium hydroxide used.

$$\frac{\text{Moles Base (from above conversion)}}{\text{Liters Base (from experiment)}} = \text{Molarity of Base}$$



Part II: Determination

1. Calculate the net mL of NaOH added. Convert to L.
2. Use the net L of NaOH, and average molarity of NaOH determined from Part I to calculate the mol of citric acid reacted.



3. Calculate the molarity of citric acid from the moles of citric acid and the volume of soda.



$$\frac{\text{Moles Acid (from above conversion)}}{\text{Liters Acid (from experiment)}} = \text{Molarity of Acid}$$

•

1. Use the volume of the soda can in liters, and molarity of citric acid in 7-up ® to calculate the moles of citric acid in one can of soda. Convert moles to grams.

DATA AND REPORT – DUE _____ / _____ / _____

* Show your work and keep track of significant figures.

SECTION _____**DATE _____****PARTNER _____****REFERENCE VALUES (FROM PRE-LAB QUIZ)**

Oxalic acid dihydrate formula _____

Oxalic acid dihydrate molar mass _____

PART I- STANDARDIZATION

	Trial 1	Trial 2	Trial 3	Trial 4 (optional)
Mass of oxalic acid				
Initial Buret Reading (NaOH)				
Final Buret Reading (NaOH)				

* Show all work for trial 1 on the next page.

	Trial 1	Trial 2	Trial 3	Trial 4
Moles of sodium hydroxide				
Net Volume NaOH used (L)				
Molarity of sodium hydroxide				
Average Molarity				

Calculations Part I: Standardization:

Show your mathematical setup for ***Trial 1*** for each of the following, complete with units:

Moles of sodium hydroxide:

Net Volume NaOH used, in liters:

Molarity of sodium hydroxide:

PART II- DETERMINATION

NaOH Concentration _____ (from standardization procedure)

	Trial 1	Trial 2	Trial 3	Trial 4 (optional)
Volume of “6 Down” Soda				
Initial Buret Reading (NaOH)				
Final Buret Reading (NaOH)				

* Show all work for trial 1 on the next page.

	Trial 1	Trial 2	Trial 3	Trial 4
Volume NaOH used (L)				
Moles of Citric Acid in “6 Down” Soda				
Molarity of Citric Acid in “6 Down” Soda				
Average Molarity in “6 Down” Soda				

Additional Calculations

“6 Down” is about 10 times more concentrated in citric acid than 7 Up®, so to calculate the realistic citric acid concentration in 7 Up®, just divide by 10.

Average Molarity in 7 Up® Soda	
Grams Citric Acid in one can of 7 Up® (355 mL)	

Calculations for Part II: Determination

Show your mathematical setup for ***Trial 1*** for each of the following, complete with units:

Net Volume NaOH used, in liters:

Moles of citric acid:

Molarity of citric acid in 6-Down Soda:

Molarity of citric acid in 7-Up*:

* “6 Down” is about 10 times more concentrated in citric acid than 7 Up®, so to calculate the realistic citric acid concentration in 7 Up®, just divide by 10.

Grams Citric Acid in one can of 7 Up® (355 mL):

Exp 12a - pH of Household Items

INTRODUCTION

In this experiment, we will measure the pH of various household products using a natural indicator, anthocyanin molecules, found in the leaves of red cabbage. Anthocyanin molecules are water-soluble pigments that change color depending upon the pH of their environment. Hydrangea flowers have assorted colors, blue, light, or dark purple, pink or red, depending on the acidity of the soil. Acidic soils give colors closer to blue, and basic soils closer to pink.

Since the indicator dye of the red cabbage is water soluble, to extract it, all we need to do is place the leaves in boiling water. We will then add some of the indicator (red cabbage juice) to standards at various pH values. We will then use these colors as a reference to help us determine the pH of household products.

ADDITIONAL READING

- Acids, Bases, pH, pOH

SAFETY EQUIPMENT REQUIRED FOR LAB

- Goggles
- Lab coat
- Nitrile gloves (provided)

SUPPLY LIST

- Test-tube rack
- Small test tubes

CLASSROOM SUPPLY AND REACTANT LIST

- 100-mL beakers (to make stations)
- Plastic transfer pipets (to make stations)
- Red cabbage Juice
- Buffers (pH 2-12) or photograph of buffers
- Various household items

PROCEDURE**Part 1: Station Preparation**

This experiment will be conducted at stations set up around the room. Each lab group will be assigned one or more stations to prepare. Follow the instructions below to properly set up your assigned station(s). (Exact quantities are not important.)

1. Obtain your assigned household item from the bin.
2. Place a small amount of the item in a 100-mL beaker.
 - a. If your substance is a liquid: Place approximately 40 mL in the beaker.
 - b. If your substance is a solid: Enough to roughly cover the bottom of the beaker, then add water to approximately 40 mL and stir to begin dissolving.
 - i. For solids in the form of pills or tablets: Use a mortar & pestle to crush it into a powder.
3. Place a transfer pipet in the beaker. (The pipet will be used by all groups that come to this station.)
4. Place your beaker and household items in the room at a location that will be easy for other students to access. (The item must be placed next to the beaker so that students will know what it is.)

Part 2: Testing Household Products

5. Obtain a test tube rack and 8 test tubes.

*Note: The test tube racks have large and small holes. Tape over the bottom of the large holes so that small test tubes may be used in the large holes without falling out.

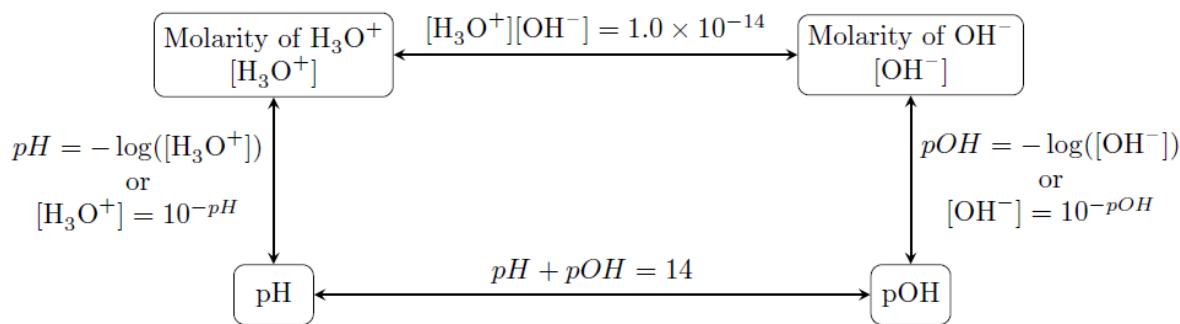
6. Using a different test tube for each station:
 - a. In your lab manual, write the location of the test tube in your rack. (so that later, you will know which product it contains).
 - b. Pipet approximately 1 mL of product into the tube. (1 mL will fill a small test tube approximately the width of a finger.)
 - c. Pipet approximately 1 mL of cabbage juice into the tube. (You may want to do this with a little force so that the product and juice are mixed upon addition.)
 - d. Record your observations of color in the data table.
7. Compare your test tubes to those of the standard solutions prepared by your instructor to estimate pH to the nearest 0.5 pH unit. (In other words, if your tube is the same color as the pH 6 standard, write 6.0, but if it is somewhere in between 6 and 7, write 6.5 for the pH.)

CLEANUP AND WASTE DISPOSAL

8. Discard your test tube solutions and the remains of the station solution into the large ‘inorganic aqueous waste’ beaker located in the rear fume hood. Your instructor will verify it is neutral before disposing of it.
9. Use acetone to remove any permanent marker from the test tubes.
10. Using a small test tube brush, wash the test tubes with soap water. Rinse with DI water. Place the clean wet test tubes upside down in the test tube drying rack.

CALCULATIONS

11. Using your pH data as a starting point, calculate the pOH, $[H_3O^+]$, and $[OH^-]$. (The following equations may be helpful.)



12. Answer the remaining Postlab questions. When calculating the pH of a certain solution, it may help to consider the following.
 - a. Is the compound dissolved in the solution an acid or base? (Will it give H_3O^+ or OH^- into solution?)
 - b. Is it strong or weak? (Will it completely dissociate? Will the molarity of H_3O^+ or OH^- be the same as the molarity of the compound?)
 - c. It is often helpful to start by finding the molarity of the compound, then the molarity of H^+ or OH^- in the solution, then use the above equations to find the pH.

DATA AND REPORT – DUE ____ / ____ / ____**SECTION _____**

* Show your work and keep track of significant figures.

DATE _____**PARTNER _____****PART 2: TESTING HOUSEHOLD PRODUCTS**

Test Tube Rack Location	Product Name	Color (after adding dye)	pH

PART 3: CALCULATIONS

Copy the product name and pH from the table on the previous page, then calculate the other information.

*Show your work and keep track of significant figures.

Product Name	[H ₃ O ⁺]	[OH ⁻]	pH	pOH

Show your work for Station 1 for each of the following:

pH to [H₃O⁺]:

pH to pOH:

pH to [OH⁻]:

QUESTIONS

1. For the products tested above,

a. Which is the most acidic?

b. Which is the most basic?

POST-LAB QUESTIONS (USE VALUES FROM ONLINE REPORT)

2. Determine the pH of each of the following solutions.

a. 1.82 g HCl in 0.500 L total solution

HCl is: A) an acid B) a base

b. 4.56 g Sr(OH)₂ in 0.500 L total solution

Sr(OH)₂ is: A) an acid B) a base

c. 1.82 g HCl and 4.56 g Sr(OH)₂ in 0.700 L total solution

1. Write a balanced chemical equation for this reaction:

2. What is the limiting reactant?

3. What mass of the excess reactant is leftover?

4. Find the pH of the solution that results from the excess reactant.

“Green Sheet” - Polyatomic Ions and Solubility Rules

Oxyanions				
borate, BO_3^{3-}	carbonate, CO_3^{2-}	nitrate, NO_3^-		
	silicate, SiO_3^{2-}	phosphate, PO_4^{3-}	sulfate, SO_4^{2-}	chlorate, ClO_3^-
		arsenate, AsO_4^{3-}	selenate, SeO_4^{2-}	bromate, BrO_3^-
			tellurate, TeO_4^{2-}	iodate, IO_3^-

Naming Rules (keep charges the same as -ate):		Examples
per-	one more oxygen (than -ate)	perchlorate, ClO_4^-
-ite	one less oxygen (than -ate)	chlorite, ClO_2^-
hypo-	one less oxygen (than -ite)	hypochlorite, ClO^-
thio-	replace one oxygen with one sulfur	thiosulfate, $\text{S}_2\text{O}_3^{2-}$

Naming Rules (the charges change):		Examples
0 H^+	Normal anion name	phosphate, PO_4^{3-}
1 H^+	Add hydrogen as prefix (charge reduced by 1)	hydrogenphosphate, HPO_4^{2-}
2 H^+	Add dihydrogen as prefix (charge reduced by 1)	dihydrogenphosphate, H_2PO_4^-

Other Common Polyatomic Ions					
Formula	Name	Formula	Name	Formula	Name
H_3O^+	hydronium	NH_4^+	ammonium	Hg_{2+}	mercury(I)
OH^-	hydroxide	CN^-	cyanide	O_{2-}	peroxide
MnO_4^-	permanganate	CrO_{42-}	chromate	$\text{Cr}_2\text{O}_{72-}$	dichromate
C_2O_{42-}	oxalate	$\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3COO^-		acetate	

Solubility in Water

Compounds containing these ions are generally soluble (aq)	... except combinations described below are insoluble (s)
Group 1 (Li^+ , Na^+ , K^+ , etc.), NH_4^+	Except Li^+ is slightly soluble with CO_3^{2-} , PO_4^{3-} , and F^- .
ClO_4^- , ClO_3^- , NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$ / CH_3COO^-	None.
Cl^- , Br^- , I^-	Except for those containing Ag^+ , Hg_2^{2+} , and Pb^{2+} .
F^-	Except for those containing Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} , and Pb^{2+} .
SO_{32-} , SO_{42-}	Except for those containing Ca^{2+} , Sr^{2+} , Ba^{2+} , Ag^+ , and Pb^{2+} .

Compounds containing these ions are generally insoluble (s)	... except combinations described below are soluble (aq)
CO_{32-} , PO_{43-}	Except those of Group 1 and NH_4^+ .
CrO_{42-} , C_2O_{42-}	Except those of Group 1 and NH_4^+ .
O_2^- , S_2^-	Except those of Group 1, NH_4^+ , Ca^{2+} , Sr^{2+} , and Ba^{2+} .
OH^-	Except those of Group 1, NH_4^+ . Except OH^- is slightly soluble with Ca^{2+} , Sr^{2+} , and Ba^{2+} .

Strong Acids and Bases

List of Strong Acids	List of Strong Bases
HCl	LiOH * $\text{Ca}(\text{OH})_2$
HClO_4	NaOH * $\text{Sr}(\text{OH})_2$
HNO_3	KOH * $\text{Ba}(\text{OH})_2$
H_2SO_4	RbOH
HBr	CsOH
HI	* Group 2 bases are slightly soluble, however they fully dissociate into component anions.

Gas Evolution Reactions

Ion Exchange Product	Decomposes?	Gase Formed
H_2S		H_2S (g)
H_2CO_3	$\rightarrow \text{H}_2\text{O}$ (l) + gas	CO_2 (g)
H_2SO_3	$\rightarrow \text{H}_2\text{O}$ (l) + gas	SO_2 (g)
NH_4OH	$\rightarrow \text{H}_2\text{O}$ (l) + gas	NH_3 (g)

“Yellow Sheet” – Periodic Table and Conversion Factors

18

Hydrogen	1	H	1.008 2.1	2	Boron	5	B	10.81 2.0	Carbon	6	C	12.011 2.5	Nitrogen	7	N	14.007 3.0	Oxygen	8	O	15.999 3.5	Fluorine	9	F	18.998 4.0	Neon	10	He	4.003 ---																				
Lithium	3	Li	6.94 1.0	4	Beryllium	4	Be	9.012 1.5	13	Boron	5	B	10.81 2.0	14	Carbon	6	C	12.011 2.5	15	Nitrogen	7	N	14.007 3.0	16	Oxygen	8	O	15.999 3.5	17	Fluorine	9	F	18.998 4.0	Neon	10	Ne	20.180 ---											
Sodium	11	Na	22.99 0.9	12	Magnesium	12	Mg	24.31 1.2	13	Aluminum	13	Al	26.98 1.5	14	Silicon	14	Si	28.09 1.8	15	Phosphorus	15	P	30.97 2.1	16	Sulfur	16	S	32.06 2.5	17	Chlorine	17	Cl	35.45 3.0	18	Argon	18	Ar	39.95 3.0	19	---	---	---	---	---				
Potassium	19	K	39.10 0.8	20	Calcium	20	Ca	40.08 1.0	21	Scandium	21	Titanium	22	Vanadium	23	Chromium	24	Manganese	25	Mn	Fe	26	Cobalt	27	Ni	28	Copper	29	Zinc	30	Gallium	31	Germanium	32	Arsenic	33	Selenium	34	Bromine	35	Krypton	36	---	---	---	---	---	
Rubidium	37	Rb	85.47 0.8	38	Sodium	38	Strontium	87.62 1.0	39	Yttrium	39	Zirconium	40	Niobium	41	Molybdenum	42	Techneium	43	Ruthenium	44	Rhodium	45	Palladium	46	Silver	47	Gold	48	Platinum	49	Indium	50	Thallium	51	Antimony	52	Te	53	Iodine	54	Xenon	54	---	---	---	---	---
Cesium	55	Cs	132.91 0.7	56	Barium	56	Lanthanum	137.33 0.9	57	Praseodymium	57	Hafnium	72	Tantalum	73	Tungsten	74	Rhenium	75	Osmium	76	Ruthenium	77	Platinum	78	Ir	79	Mercury	80	Thallium	81	Lead	82	Bismuth	83	Polonium	84	Atmosphere	85	Radon	86	Rn	222	2.4	2.6	---	---	---
Rutherfordium	87	Fr	(223) 0.7	88	Radium	88	Actinium	(226) 0.9	89	Thorium	89	Protactinium	91	Uranium	92	Neptunium	93	Plutonium	94	Americium	95	Curium	96	Berkelium	97	Californium	98	Einsteinium	99	Mendelevium	100	Curium	101	Nobelium	102	Lawrencium	103	Lu	103	---	---	---	---	---				
*lanthanide	58	Ce	140.12 1.1	59	Praseodymium	59	Neodymium	140.91 1.1	60	Praseodymium	60	Neodymium	61	Proneodymium	62	Samarium	63	Europium	64	Gadolinium	65	Terbium	66	Dysprosium	67	Hholmium	68	Erbium	69	Thulium	70	Ytterbium	71	Yttrium	71	---	---	---	---	---								
** actinide	90	Th	232.04 1.3	91	Thorium	91	Protactinium	231.04 1.5	92	Uranium	92	Neptunium	93	Plutonium	94	Americium	95	Americium	96	Curium	97	Berkelium	98	Californium	99	Einsteinium	100	Mendelevium	101	Nobelium	102	Lawrencium	103	Lu	103	---	---	---	---	---								

Common Conversions and Constants for Chemistry Courses

UNIT	SYMBOL	DEF. OR EQUIVALENT
Distance		
1 meter	m	39.37 in
1 mile	mi	1.6093 km, 5280 ft
1 yard	yd	36 in, 3 ft
1 foot	ft	12 in
1 inch	in	2.54 cm (exactly)
1 angstrom	Å	1×10^{-10} m
Volume		
1 liter	L	0.26417 gal
1 gallon	gal	4 qt, 3.785 L
1 quart	qt	2 pt, 0.946 L
1 pint	pt	2 cup
1 cup	cup	8 fl oz, 16 tbs
1	tbs	½ fl oz
1 cubic cm	cm ³ or cc	1 mL
Mass		
1 gram	g	0.002204 lbs, 0.03527
1 ton	ton	2000 lb
1 pound	lb	16 oz, 453.59 g
1 metric	Mg	1000 kg
1 kilogram	kg	1000 g, 2.2046 lb
atomic mass unit	amu	1.6605×10^{-24} g
Energy		
1 joule	J	$1 \text{ kg} \cdot \text{m}^2/\text{s}^2$, 0.23901 cal, 0.0098692 L·atm
1 calorie	cal	4.184 J (exactly)
1	kcal, Cal	1000 cal
1 electron volt	eV	1.6022×10^{-19} J
Pressure		
1	atm	760 Torr, 101325 Pa
1 Torr	Torr	1 mm Hg,
1 Pascal	Pa	1 N/m ²
1 bar	bar	1×10^5 Pa
Temperature		
Kelvin	K	°C + 273.15
Celsius	°C	(°F - 32) / 1.8
Fahrenheit	°F	1.8 · °C + 32
Misc		
1 coulomb	C	1 A·s
1 Newton	N	1 kg·m/s ²

UNIT	SYMBOL	DEF. OR EQUIVALENT
Physical Constants		
Ideal Gas Const.	R	0.082058 L·atm/(mol·K)
Avogadro's #	N _A	6.0221×10^{23}
1 mole	mol	6.0221×10^{23}
Planck's	h	6.6262×10^{-34} J·s
Speed of light in a	c	2.9979×10^8 m/s
Faraday	F	9.64846×10^4 C
Electron mass	m _e	9.109535×10^{-28} g
Electron		1.60219×10^{-19} C
Proton mass	m _p	1.672649×10^{-24} g
Neutron mass	m _n	1.674954×10^{-24} g

SI Prefixes		
tera-	T-	10^{12}
giga-	G-	10^9
mega-	M-	10^6
kilo-	k-	10^3
deci-	d-	10^{-1}
centi-	c-	10^{-2}
milli-	m-	10^{-3}
micro-	μ-	10^{-6}
nano-	n-	10^{-9}
pico-	p-	10^{-12}
femto-	f-	10^{-15}