

CHM 140: GENERAL CHEMISTRY I LAB MANUAL



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Triton College

General Chemistry I CHM 140 Lab Manual

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Glossary

Detailed Licensing

Licensing

A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.

0: LAB SAFETY ORIENTATION

PURPOSE

To introduce the fundamental principles of laboratory safety, known as **RAMP**:



Figure 0.1: The four principles of safety (Copyright; [ACS](#))

[0.1: Lab Safety Videos](#)

[0.2: Lab Safety Contract](#)

[0.3: Lab Safety Quiz](#)

[0.4: Additional Resources](#)

0.1: Lab Safety Videos

Safety Orientation Videos

Students will find below the eight Lab safety orientation videos. The first six are part of the College Lab Safety Series by the American Chemical Society. Students should watch the first seven videos before signing the safety contract and attempting the safety quiz. The eighth video is an older version of the College Lab Safety video by the American Chemical Society, which should be watched, time permitting, as it provides detailed dos and don'ts in a college lab setting.

1. ACS (American Chemical Society) College Lab Safety Series

YouTube Video

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2. Lab Safety and Techniques- Crash course safety

YouTube Video

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3. Safety Video by American Chemical Society (1991)

YouTube Video

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0.2: Lab Safety Contract

Triton College Chemistry Division

Student Safety Contract

At Triton College, safety is our number one priority. A list of rules has been developed and provided to you in this student safety contract to ensure a safe science laboratory experience. These rules must always be followed. Please be advised that you will be asked to leave the laboratory if you violate these rules. **No exceptions!!**

Download, sign, and date both contract pages before you participate in the laboratory sessions. Your signature indicates that you have read this contract and agree to abide by the rules. You will keep a copy for your records and upload or submit one to your instructor via your Blackboard or Learning Management System (LMS) as directed.

I, _____, dated _____ agree that I will comply with the following safety rules. I will:

1. Always wear safety goggles in the laboratory.
2. Always wear appropriate clothing in the laboratory. This includes long pants, shirts with sleeves, and closed shoes (i.e., no sandals). Also, I understand that long hair must be tied back, and loose or baggy clothing or jewelry must be secured.

**Note: Students violating rules one and/or two will be asked to leave the lab immediately. No second chances!! NO EXCEPTIONS!!*

3. Read laboratory instructions beforehand and ask the instructor about any part of the lab I am unsure about.
4. Use proper equipment for laboratory procedures.
5. Never perform unauthorized experiments or work alone in the laboratory.
6. Never remove chemicals or other materials/equipment from the laboratory area.
7. Read the labels on reagent bottles carefully, remove only small amounts of reagent with the proper tools, and never return unused chemicals to the bottle.
8. Dispose of broken glass and waste chemicals in the appropriate waste container.
9. Immediately inform the instructor of a chemical spill or accident (even minor ones) in the laboratory.
10. Never taste or smell any chemicals.
11. Know the location of eyewash fountains, fire extinguishers, fire blankets, safety showers, first aid kits, and exits.
12. Do not bring food or drink into the lab; refrain from eating or drinking in the laboratory. This includes chewing gum and eating candy.
13. Keep aisles clear. Books, purses, backpacks, and other personal belongings will be kept outside or under the lab benches.
14. I will responsibly conduct myself at all times in the laboratory. Horseplay, practical jokes, and pranks are dangerous and prohibited.
15. Clean my work area before leaving the lab.
16. Wash my hands before I leave the lab.

Agreement

I, _____ have read and agree to follow all the safety rules outlined in this contract. I realize I must obey these rules to ensure my safety and that of my fellow students. I will fully cooperate with my instructor and fellow students to maintain a safe lab environment. I will also closely follow all oral and written instructions provided by the instructor, the science department, and the college. I am aware that any violation of this safety contract, resulting in unsafe conditions in the laboratory or misbehavior on my part, may lead to my removal from the laboratory, a failing grade, and/or dismissal from the course.

Student Signature: _____

Date: _____

Course section and number: _____

Questions

1. Are you color blind? Yes / No
2. Do you have any allergies? Yes / No
If yes, please list specific allergies: _____
3. Do you wear contact lenses? Yes / No
4. If you are pregnant, breastfeeding, or trying to conceive, avoid working with heavy metal compounds and organic solvents. Notify your instructor and request a list of chemicals used in this course. Use this list to consult with your physician before proceeding with the course. Under your physician's supervision, you may need to purchase a respirator, which a doctor must fit and adjust to your needs.
5. Please inform your instructor of any other medical conditions or situations you feel we should be aware of to ensure your safety and the safety of others while working in the laboratory.

Please be aware that anything written in this contract will remain confidential.

0.3: Lab Safety Quiz

Multiple-Choice Questions

1. **What should you do if a chemical splashes into your eyes?**
 - A. Rub your eyes vigorously
 - B. Rinse your eyes with water for at least 15 minutes
 - C. Close your eyes and wait for the pain to subside
 - D. Apply a cold compress
2. **Which of the following is NOT appropriate lab attire?**
 - A. Closed-toe shoes
 - B. Long pants
 - C. Loose, flowing clothing
 - D. Safety goggles
3. **What is the first thing you should do if you spill a chemical on your skin?**
 - A. Wipe it off with a cloth
 - B. Rinse the affected area with water immediately
 - C. Apply a bandage
 - D. Ignore it if it doesn't hurt
4. **Where should you dispose of broken glassware?**
 - A. Regular trash bin
 - B. Recycling bin
 - C. Designated glass disposal container
 - D. Sink
5. **Which of the following statements about lab safety is TRUE?**
 - A. Eating and drinking are allowed in the lab
 - B. You should always work alone in the lab
 - C. Always read labels and safety data sheets before using chemicals
 - D. It's okay to leave a Bunsen burner unattended

True/False Questions

6. **You should always wear safety goggles when handling chemicals.**
 - True
 - False
6. **It's safe to taste chemicals in the lab to identify them.**
 - True
 - False
6. **You should know the location of safety equipment such as fire extinguishers, eyewash stations, and safety showers.**
 - True
 - False
6. **Long hair should be tied back when working in the lab.**
 - True
 - False
6. **It's acceptable to use your phone in the lab as long as you're careful.**
 - True

- False

0.4: Additional Resources

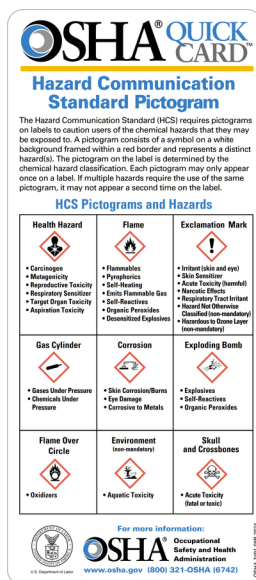


Figure 0.4.1: OSHA Hazard Communication Standard Pictogram QuickCard

1: SEPARATION OF A THREE-COMPONENT MIXTURE

PURPOSE

- To separate a three-component mixture of sand, sodium chloride (salt), and ammonium chloride by utilizing their distinct physical properties.
- To determine the mass percentage of each component in an unknown mixture of three solids (sand, sodium chloride, and ammonium chloride).

INTRODUCTION

This experiment demonstrates the separation of a mixture containing sand (SiO_2), sodium chloride (NaCl), and ammonium chloride (NH_4Cl) using various techniques based on their physical properties.

DEFINITION: MIXTURE

Mixtures are combinations of two or more substances that retain their individual chemical properties and can be separated through physical processes.

The separation strategy for this experiment involves the following principles:

1. **Sublimation:** Ammonium chloride (NH_4Cl) undergoes sublimation, transitioning directly from a solid to a gas upon heating without passing through a liquid phase. This property distinguishes it from non-sublimating substances, such as sand and sodium chloride. Ammonium chloride gas can be cooled to obtain pure solid ammonium chloride. Other substances that sublime include camphor, naphthalene, anthracene, and iodine.
2. **Solubility:** Sodium chloride (NaCl) is soluble in water, whereas sand (SiO_2) is not. When water is added to the mixture, sodium chloride dissolves, allowing it to be separated from the insoluble sand by filtration or decantation.
3. **Filtration or Decantation:** This process involves separating the sodium chloride solution from the sand by pouring the liquid away from the solid or passing it through filter paper.
4. **Evaporation:** Heating the sodium chloride solution causes the water to evaporate, leaving behind solid sodium chloride.

[1.1: Separation Of a Three-Component Mixture - Experiment](#)

[1.2: Separation Of a Three-Component Mixture - Pre-lab](#)

[1.3: Separation Of a Three-Component Mixture - Data and Report](#)

1.1: Separation Of a Three-Component Mixture - Experiment

SAFETY PRECAUTIONS

- Always **wear safety goggles** to protect your eyes.
- To avoid inhaling ammonium chloride fumes, perform the sublimation step in a **fume hood**.
- Use caution when handling hot equipment such as evaporating dishes, hot plates, and Bunsen burners. Use tongs or heat-resistant gloves, and allow the equipment to cool before handling.
- Thoroughly wash all used glassware and equipment with soap and water at the end of the experiment.
- Avoid direct contact with chemicals and wash your hands thoroughly after the experiment.
- Keep flammable materials away from the Bunsen burner.

EQUIPMENT AND CHEMICALS NEEDED

Table 1.1.1: Materials List

Chemicals & Reagents	Glassware & Supplies	Hardware & Heating Equipment
Unknown mixture (sand, NaCl, NH ₄ Cl)	Evaporating dishes (2)	Electronic balance (± 0.001 g)
Deionized water	Watch glass	Ring stand & ring clamp
Solid waste container	Glass funnel	Clay triangle
Recovered sand container	Graduated cylinder (25 mL)	Bunsen burner
	150-mL beaker	Hot plate
	Stirring rod with rubber policeman	Crucible tongs
	Cotton plug / ball	

EXPERIMENTAL PROCEDURE

Part A: Prepare the Sample

1. Obtain two evaporating dishes, then clean and dry them thoroughly.
2. Record the mass of the first clean, dry evaporating dish (labeled "first") in the data table.
3. Obtain an unknown mixture and record the unknown ID in the data table.
4. Add 2 to 3 grams of this unknown mixture to the first evaporating dish. Accurately record the combined mass of the evaporating dish and the unknown sample.
5. Obtain a glass funnel that covers the mixture inside the first evaporating dish. Insert a small piece of cotton into the stem opening. Record the mass of the glass funnel with the cotton plug.

Part B: Sublime the Ammonium Chloride

6. Set up a ring stand with a ring clamp and a clay triangle inside a **fume hood**. Place the evaporating dish containing the unknown mixture on the clay triangle.
7. Cover the mixture with the previously weighed inverted glass funnel (with cotton plug) over the evaporating dish.
8. Heat the evaporating dish with a Bunsen burner until white fumes are no longer observed. Continue heating for an additional 5 minutes, then turn off the Bunsen burner. Allow the evaporating dish and funnel to cool to room temperature for 10–15 minutes.
9. Weigh the cooled evaporating dish with the remaining sample (sand and sodium chloride) and record the mass.
10. Calculate the mass of ammonium chloride lost by sublimation by subtracting the mass of the evaporating dish and sample after heating (Step 9) from the mass before heating (Step 4).
11. Weigh the glass funnel with the sublimated ammonium chloride and cotton plug.
12. Calculate the mass of ammonium chloride collected on the funnel by subtracting the mass of the glass funnel and cotton plug before heating (Step 5) from the mass after heating (Step 11).
13. Clean the glass funnel by scraping off the ammonium chloride into the designated solid waste container as directed by your instructor.

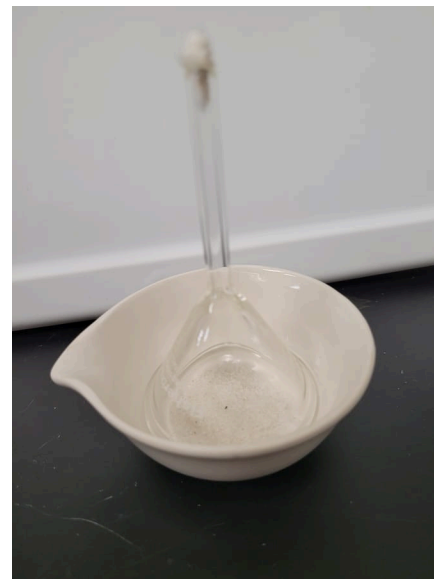


Figure 1.1.1: Apparatus setup for NH_4Cl sublimation.

Part C: Separate and Isolate Sodium Chloride

14. Add 25 mL of deionized water to the first evaporating dish containing the remaining mixture (sand and NaCl). Gently stir with a glass rod for 5 minutes to completely dissolve the sodium chloride.
15. Weigh a second clean, dry evaporating dish along with a watch glass and record the combined mass (labeled "second").
16. Carefully decant the liquid (sodium chloride solution) from the first evaporating dish into the second evaporating dish, taking care not to transfer any sand.
17. Repeat the decantation with two additional 10-mL portions of deionized water.
18. Place the second evaporating dish (containing the sodium chloride solution) on a hot plate and heat to evaporate the water.
19. To avoid splattering, when the volume reduces to approximately 5 mL (or if splattering begins), cover the evaporating dish with the previously weighed watch glass.
20. Continue heating gently until all liquid has evaporated and droplets on the watch glass are dry.
21. Allow the evaporating dish and watch glass to cool to room temperature. Weigh the cooled dish and watch glass, and record the combined mass.
22. Calculate the mass of recovered sodium chloride by subtracting the mass of the empty evaporating dish and watch glass before evaporation (Step 15) from the mass after evaporation (Step 21).

Part D: Recover the Sand

23. Dry the original (first) evaporating dish containing the wet sand on a hot plate by heating gently and stirring occasionally to prevent popping or cracking.
24. Allow the evaporating dish to cool to room temperature. Weigh the dish with the dry sand and record the mass.
25. Calculate the mass of recovered sand by subtracting the mass of the empty first evaporating dish (Step 2) from the mass of the dish with dry sand (Step 24).

CHEMICAL DISPOSAL

- Dispose of the recovered sand in the **Recovered Sand** container after it has cooled.
- Scrape the recovered ammonium chloride from the glass funnel into the designated solid waste container, then wash the funnel with soap and water.
- Rinse the recovered sodium chloride down the sink drain with plenty of running water.
- Dispose of any remaining solutions according to your instructor's guidelines.
- Clean your work area and return all equipment and glassware to their designated locations.

1.2: Separation Of a Three-Component Mixture - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Define the following separation and phase change terms:

a. **Mixture:**

b. **Decantation:**

c. **Sublimation:**

d. **Evaporation:**

e. **Freezing:**

f. **Melting:**

2. List three examples of heterogeneous or homogeneous mixtures you encounter in everyday life.

3. Explain the difference between decantation and filtration. When is decantation preferred over filtration in a laboratory procedure?

4. Describe safety precautions required when working with a Bunsen burner. Why must you allow heated glassware to cool to room temperature before measuring its mass on a balance?

5. Classify each of the following as a pure substance or a mixture. Identify whether it is an element, compound, heterogeneous mixture, or homogeneous mixture:

a. Salt (sodium chloride, NaCl) and sand mixture:

b. Sugar water solution:

c. Ammonium chloride (NH_4Cl) solid:

d. Iron (Fe) metal filings:

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1.3: Separation Of a Three-Component Mixture - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Unknown ID: _____

Record qualitative observations during heating, sublimation, decantation, and evaporation (e.g., color changes, fume formation, appearance of residues):

Table 1.3.1: Experimental Mass Measurements

Quantity	Mass (g)
Mass of the First Evaporating Dish	_____
Mass of First Evaporating Dish + Sample before Heating	_____
Mass of First Evaporating Dish + Sample after Heating	_____
Mass of Glass Funnel + Cotton Plug	_____
Mass of Glass Funnel + Cotton Plug + Sublimated NH_4Cl	_____
Mass of Second Evaporating Dish + Watch Glass	_____
Mass of Second Evaporating Dish + Watch Glass + Dried NaCl	_____
Mass of First Evaporating Dish + Dried Sand	_____

Table 1.3.2: Component Mass Calculations

Quantity	Calculation Method	Mass (g)
Mass of original sample	(Dish 1 + original sample) – (Dish 1)	_____
Mass of NH_4Cl (by loss)	(Dish 1 + sample before heating) – (Dish 1 + sample after heating)	_____
Mass of NH_4Cl (collected)	(Funnel + cotton + sublimated sample) – (Funnel + cotton)	_____

Quantity	Calculation Method	Mass (g)
Mass of NaCl recovered	(Dish 2 + watch glass + dried NaCl) – (Dish 2 + watch glass)	_____
Mass of sand recovered	(Dish 1 + dried sand) – (Dish 1)	_____
Total mass recovered	Mass of NH ₄ Cl (by loss) + Mass of NaCl + Mass of sand	_____

Show setup for all component mass calculations:

Table 1.3.3: Mass Percentage Calculations

Quantity	Calculation Method	Percent (%)
Percent Mass Recovered	(Total Mass Recovered ÷ Mass of Original Sample) × 100%	_____
Percent NH ₄ Cl	(Mass of NH ₄ Cl (by loss) ÷ Total Mass Recovered) × 100%	_____
Percent NaCl	(Mass of NaCl ÷ Total Mass Recovered) × 100%	_____
Percent Sand	(Mass of Sand ÷ Total Mass Recovered) × 100%	_____

Show setup for all mass percentage calculations:

POST-LAB QUESTIONS

1. Explain how differences in physical properties (sublimation, water solubility, density) allow for the complete separation of sand, NH₄Cl, and NaCl.

2. Identify two specific operational errors that would lead to a total percent recovery significantly below 100%.

3. Suggest an additional physical separation method (e.g., magnetism, extraction) that could be used if the mixture also contained iron filings.

4. Why is it essential to use a fume hood when heating the mixture during the sublimation step?

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2: DENSITY OF AN UNKNOWN MIXTURE

PURPOSE

- To investigate the concept of density as a fundamental property of matter and explore its application in determining the composition of an unknown ethanol–water mixture.
- To construct a standard calibration curve using mixtures of known composition and determine the percentage of ethanol in an unknown sample.

INTRODUCTION

This laboratory exercise explores the concept of density as a fundamental property of matter and investigates its relationship to the composition of mixtures. **Density**, defined as mass per unit volume ($d = \frac{m}{V}$), is an intensive property, meaning it does not depend on the amount of substance present. This experiment demonstrates how density can be used to determine the composition of an unknown mixture based on the principle that a mixture's density varies predictably with the proportion of its components.

The experiment involves creating a series of known mixtures and measuring their densities using precise volumetric glassware, including **pipettes** and **burettes**, which are necessary to obtain accurate measurements and maintain appropriate **significant figures** for precise calculations. The measured densities of the known mixtures are used to create a **calibration curve**—a graph that plots the measured property (density) against composition (percent ethanol). This calibration curve can then be used to determine the composition of an unknown mixture by **interpolation** (if the unknown's density falls within the range of known standards) or by **extrapolation** (if its density falls outside this range).

This lab will also examine the concept of **miscibility**, the property of liquids to mix in all proportions. When two liquids are miscible, the assumption is often made that their volumes are strictly additive. However, molecular interactions between ethanol and water molecules can lead to volume contraction upon mixing. You will use your experimental data to evaluate this assumption and assess its impact on density calculations.

Data will be collected, organized, graphed, and analyzed using **spreadsheet software** (such as Microsoft Excel). You will utilize spreadsheet functions to perform calculations, including the **SLOPE()** and **INTERCEPT()** functions, to define the line of best fit. Additionally, the lab introduces the concepts of **linear relationships** and **trendlines** to quantify experimental results. By the end of this experiment, you will have a solid understanding of density, calibration curves, and the application of graphical analysis to chemical data.

[2.1: Density of an Unknown Mixture - Experiment](#)

[2.2: Density of an Unknown Mixture - Pre-lab](#)

[2.3: Density of an Unknown Mixture - Data and Report](#)

2.1: Density of an Unknown Mixture - Experiment

SAFETY PRECAUTIONS

- Ethanol is highly flammable; keep all ethanol solutions away from open flames, hot plates, and spark sources.
- Wear chemical splash goggles at all times in the laboratory.

EQUIPMENT AND CHEMICALS NEEDED

Table 2.1.1: Materials Needed

Equipment	Chemicals
125-mL Erlenmeyer flask & rubber stopper	Pure ethanol (95% or absolute)
10-mL volumetric pipette & pipette bulb	Unknown ethanol-water mixture
50-mL burette, burette clamp, & ring stand	Deionized water
Electronic balance (± 0.001 g)	Nonhalogenated Organic Waste container
Two 50-mL beakers	

EXPERIMENTAL PROCEDURE

Part A: Measure Ethanol and Prepare Mixtures

- Set up a clean burette using a burette clamp and ring stand. Rinse the burette with deionized water, then fill it with deionized water to near the 0.00 mL mark.
- Record the mass of a clean, dry 125-mL Erlenmeyer flask with its rubber stopper. (*Note: A stoppered flask minimizes evaporation loss during weighing.*)
- Obtain approximately 15 mL of pure **ethanol** in a small 50-mL beaker (do **not** use an unknown mixture for this step). From the beaker, carefully pipette 10.00 mL of ethanol into the massed flask. Stopper the flask and weigh it again. Record the exact mass and volume of ethanol.
- Prepare mixtures by successive dilutions with deionized water. Record the initial reading on the burette. Deliver approximately 10.00 mL of water into the flask containing ethanol. Stopper the flask, mix thoroughly, and weigh it again. Record the exact burette reading to the nearest 0.01 mL and the new mass.
- Continuing from the previous burette reading (~ 10.00 mL), deliver additional deionized water into the flask until the burette reaches approximately the 20.00 mL mark. Stopper the flask, mix, and record the exact burette reading and combined mass.
- Repeat this addition procedure incrementally, delivering water to approximately the 30.00 mL, 40.00 mL, and 50.00 mL marks on the burette. Record the exact cumulative burette reading and mass after each addition. (*Note: Each burette reading represents the total cumulative volume of water delivered to the flask.*)
- Empty the flask into the designated waste container. Wash the flask with soap and water, rinse thoroughly with deionized water, and dry completely before proceeding to Part B.

Part B: Measure the Unknown Mixture

- Obtain approximately 25 mL of an unknown ethanol-water mixture in a clean, dry 50-mL beaker. Record the unknown ID.
- Carefully pipette 10.00 mL of the unknown mixture into the clean, dry, massed 125-mL Erlenmeyer flask with stopper.

Stopper the flask and record the combined mass.

10. Pipette an **additional** 10.00 mL of the unknown mixture into the same flask (total volume added = 20.00 mL). Stopper the flask and record the new combined mass.
11. Clean and rinse all volumetric glassware and return equipment as instructed.

CHEMICAL DISPOSAL

- Dispose of all ethanol–water mixtures and unknown solutions in the designated container labeled **Nonhalogenated Organic Waste**.

2.2: Density of an Unknown Mixture - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Explain the physical relationship between density and the mass percentage of a component in a liquid mixture. How do you expect the density of an ethanol–water mixture to change as the percentage of ethanol increases, and why?

2. Define a calibration curve and explain its utility in quantitative analysis. Differentiate between **interpolation** and **extrapolation** when reading a calibration plot, and detail how you would handle an unknown density that falls outside your standard calibration range.

3. Explain what it means for two liquids to be **miscible**. Are the total volumes of ethanol and water strictly additive when mixed? If volume contraction occurs upon mixing, explain how this non-ideal behavior affects calculated density values.

4. Discuss the advantages of utilizing spreadsheet software (e.g., Microsoft Excel) for chemical data analysis. Identify the specific spreadsheet functions used to determine the slope, y-intercept, and unknown composition from linear regression data.

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2.3: Density of an Unknown Mixture - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Known Standards

Initial Burette Reading (mL): _____

Table 2.3.1: Ethanol Reference Baseline Data

System Component	Total Mass (g)
Empty Flask & Stopper	_____
Empty Flask + Stopper + 10.00 mL Ethanol	_____

Table 2.3.2: Incremental Water Addition Data

Addition Step	Total Mass (g)	Final Burette Reading (mL)	Volume of Water Added (mL)
1 st Water Addition	_____	_____	_____
2 nd Water Addition	_____	_____	_____
3 rd Water Addition	_____	_____	_____
4 th Water Addition	_____	_____	_____
5 th Water Addition	_____	_____	_____

Part B: Unknown Mixture

Unknown ID: _____

Table 2.3.3: Unknown Mixture Mass Measurements

Volume of Unknown Added (mL)	Total Mass (Flask + Stopper + Sample) (g)
10.00	_____
20.00	_____

DATA CHECK: PHYSICAL PROPERTIES

Before leaving the laboratory, verify your measurements against physical constants:

- **Density Boundaries:** Pure ethanol density is approximately 0.789 g/mL and pure water is 1.000 g/mL. Calculated densities for all mixtures **must** fall within this range.
- **Logical Trend:** As water is added to ethanol, cumulative liquid mass and mixture density must increase monotonically.
- **Error Isolation:** A calculated density exceeding 1.00 g/mL indicates a failure to subtract the empty flask mass.

CALCULATIONS AND GRAPHICAL ANALYSIS

Setting up a Spreadsheet

1. Use the instructor-provided Excel Template or Google Sheets template.
2. Enter experimental data for known mixtures into designated columns. Ensure Cell B2 contains the mass of empty flask and stopper. Do not type units into numerical data cells.
3. Format numerical data cells to display appropriate significant figures.
4. **Total Volume Column (D):** In Cell D5, enter `=B5+C5` and copy down through Cell D10.
5. **Mass of Liquid Column (E):** In Cell E5, enter `=A5-$B$2` (using absolute reference `$B$2` for empty flask mass) and copy down through Cell E10.
6. **Density Column (F):** In Cell F5, enter `=E5/D5` and copy down through Cell F10.
7. **% Ethanol Column (G):** In Cell G5, enter `=(B5/D5)*100` and copy down through Cell G10.

DATA CHECK: SPREADSHEET LOGIC

Verify calculated spreadsheet columns before plotting:

- **% Ethanol (Column G):** Row 5 must equal 100.0%. Subsequent rows must decrease sequentially.
- **Negative or >100% Values:** Check Column C formula references for burette subtraction errors.

Creating a Calibration Graph

1. Select data in Columns F and G (Rows 5–10). Insert a **Scatter (X, Y) with markers only** plot.
2. Ensure *Density* (Column F, g/mL) is on the x-axis and *Percent Ethanol* (Column G, %) is on the y-axis. Add clear axis titles.
3. Adjust the x-axis minimum bound to 0.75 g/mL so the plot spans 0.75 to 1.00 g/mL.
4. Add a **Linear Trendline** and select **Display Equation on chart** ($y = mx + b$).

DATA CHECK: INTERPRETING THE TRENDLINE

Analyze scatter plot geometry:

- **Slope Direction (m):** The slope **must be negative**. Because ethanol is less dense than water, increasing solution density (x-axis) results in a lower mass percentage of ethanol (y-axis).
- **Linearity:** Points forming a non-linear or "V" shape indicate mass subtraction or volumetric reading errors.

Calculating Unknown % Ethanol

1. Use the `SLOPE()` and `INTERCEPT()` functions in Excel for maximum regression precision:
 - In Cell G12, enter `=SLOPE(G5:G10, F5:F10)`.
 - In Cell G13, enter `=INTERCEPT(G5:G10, F5:F10)`.
2. Calculate liquid mass, volume, and density for Trial 1 (Row 16) and Trial 2 (Row 17) of the unknown mixture.
3. Calculate % ethanol for each unknown trial using linear regression:

$$\% \text{ ethanol} = (\text{slope} \times \text{density}) + \text{intercept} \quad (2.3.1)$$

In Cell G16, enter $=G\$12 * F16 + G\$13$ and copy down to Cell G17.

4. Calculate Average % Ethanol in Cell G19 using $=AVERAGE(G16:G17)$.

DATA CHECK: REALITY TEST

- **Valid Bounds:** Average % Ethanol must fall between 0% and 100%.
- **Internal Consistency:** Compare unknown density to standard calibration data. An unknown density of 0.90 g/mL should yield a calculated % ethanol close to your 1:1 mixture standard.

POST-LAB QUESTIONS

1. Why were a volumetric pipette and burette used instead of graduated cylinders for this experiment? Explain how this selection of glassware impacts the number of significant figures in calculated density values.

2. Based on your experimental calibration curve, is the relationship between density and % ethanol strictly linear? Explain any observed non-linearity in terms of intermolecular hydrogen bonding and volume contraction.

3. Report your calculated average % ethanol for the unknown mixture. Evaluate the precision of your two unknown trials by calculating the Percent Difference between Trial 1 and Trial 2.

4. If the stock ethanol used for standard preparations was actually 95.0% (v/v) ethanol rather than pure 100% ethanol, explain how this systematic error would shift your calibration curve. How would you mathematically correct your reported unknown % ethanol?

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3: FLAME TESTS AND ATOMIC SPECTRA

PURPOSE

- To examine the relationship between elements, their emission spectra, and flame test colors.
- To use flame test colors and line spectra observations to identify unknown cation substances.

INTRODUCTION

Atomic emission spectra are unique patterns of light emitted by elements when their electrons return to lower energy levels after excitation. When energy is supplied to an atom (such as heat from a flame or electrical discharge from a high-voltage tube), its electrons absorb energy and jump to higher energy states. As they transition back to lower energy levels, they release energy in the form of electromagnetic radiation (light), producing distinct spectral lines. These emission spectra act as elemental fingerprints, allowing scientists to identify unknown substances. The flame test and spectroscopic analysis utilize these principles to determine the presence of specific elements based on their characteristic emission colors.

[3.1: Flame Tests and Atomic Spectra - Experiment](#)

[3.2: Flame Tests and Atomic Spectra - Pre-lab](#)

[3.3: Flame Tests and Atomic Spectra - Data and Report](#)

3.1: Flame Tests and Atomic Spectra - Experiment

SAFETY PRECAUTIONS

- Handle all chemicals and open flames with caution. Tie back long hair and secure loose clothing.
- Work in a well-ventilated area, preferably under a fume hood.
- Do not directly inhale fumes from the flame tests.
- Exercise extreme caution around high-voltage power supplies used for gas discharge tubes; turn off power before changing tubes.

EQUIPMENT AND CHEMICALS NEEDED

Table 3.1.1: Materials List

Equipment	Chemicals
Bunsen burner, gas tubing, & friction striker	50% Ethanol/water cation solutions (Li^+ , Na^+ , K^+ , Ca^{2+} , Sr^{2+} , Ba^{2+}) in spray bottles
Gas discharge tubes (H_2 , He, N_2 , Ne, Ar)	Three unknown cation solutions in spray bottles
High-voltage discharge tube power supply	Aqueous Heavy Metal Waste container
Handheld spectrosopes	
(Optional) Fiber optic emission spectrometer	
Utility clamp & tape	

EXPERIMENTAL PROCEDURE

Part A: Perform Flame Tests

1. Set up a Bunsen burner inside a fume hood and ignite it. Adjust the gas and air valves to produce a clear flame with a blue inner cone. Very gently mist a small amount of each aqueous ethanolic salt solution into the hottest region of the flame (the tip of the inner blue cone). *Note: Spraying too strongly will extinguish the flame; if this occurs, turn off the gas before carefully relighting the burner.* Record your observations for all known salts.
2. Test and identify **ALL three unknown salt solutions** using the same procedure. Record flame colors and determine the identity of each unknown by comparing its flame color to the known cation standards.

Part B: Analyze Emission Spectra

3. The instructor will set up several gas discharge tube stations using high-voltage power supplies.
4. **Visual / Handheld Spectroscope Observation:** Observe each discharge tube using a handheld spectroscope. Record the overall perceived color and draw or note the positions of the dominant color lines observed on the wavelength scale. Alternatively, capture a photograph through the spectroscope eyepiece.
5. **Fiber Optic Spectrometer Observation (Optional):** Align the fiber optic probe near each discharge tube to collect the light spectrum. Use a utility clamp to hold the fiber optic cable stationary during data acquisition.
6. Observe the emission spectrum of each unknown gas tube and identify the element by comparing spectral lines to the known gas standards.

CHEMICAL DISPOSAL

- Collect any excess chemical solutions or run-off in the designated **Aqueous Heavy Metal Waste** container. Do not pour heavy metal solutions down the sink drain.

3.2: Flame Tests and Atomic Spectra - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. What are three measurable properties or characteristics used to define electromagnetic radiation (light)?

2. State the wavelength range (in nanometers, nm) corresponding to the visible region of the electromagnetic spectrum.

3. Provide one example of electromagnetic radiation that falls outside the visible spectrum and cannot be detected by the human eye.

4. Distinguish between an **absorption spectrum** and an **emission spectrum**. Explain the electronic processes occurring in an atom during each process.

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3.3: Flame Tests and Atomic Spectra - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Flame Tests

Table 3.3.1: Known Cation Flame Test Observations

Cation	Observation (Flame Color)	Image / Sketch Reference
Li ⁺	_____	_____
Na ⁺	_____	_____
K ⁺	_____	_____
Ca ²⁺	_____	_____
Sr ²⁺	_____	_____
Ba ²⁺	_____	_____

Table 3.3.2: Unknown Cation Identification

Unknown ID	Observed Flame Color	Image / Sketch Reference	Identified Cation
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

Part B: Emission Spectra

Table 3.3.3: Element and Unknown Emission Line Observations

Element / Sample	Overall Color	Dominant Spectral Lines (nm) / Notes
Hydrogen (H)	_____	_____
Helium (He)	_____	_____
Nitrogen (N ₂)	_____	_____
Neon (Ne)	_____	_____
Argon (Ar)	_____	_____

Element / Sample	Overall Color	Dominant Spectral Lines (nm) / Notes
Reflected Sunlight	_____	_____
Unknown A	_____	_____
Unknown B	_____	_____

POST-LAB QUESTIONS

1. When a rubidium salt is heated in a flame, a photon of light with an energy of $3.8 \times 10^{-19} \text{ J}$ is emitted. Based on this value and Table 3.3.4, what flame color is expected for rubidium? Show your wavelength calculation ($E = \frac{hc}{\lambda}$) and explain your reasoning. ($h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}$, $c = 3.00 \times 10^8 \text{ m/s}$).

Table 3.3.4: Visible Spectrum Wavelength Ranges

~400 nm	~450 nm	~550 nm	~580 nm	~600 nm	~650 nm
Violet	Blue	Green	Yellow	Orange	Red

2. Visible light of distinct colors is emitted by excited hydrogen atoms in a discharge tube. Explain how these specific spectral line wavelengths are accounted for by the Bohr model of the atom.

3. Consider the energy level diagram of the hydrogen atom according to the Bohr model (Figure 3.3.1):

- a. Is light absorbed or emitted when an electron undergoes a transition from $n = 5$ to $n = 3$?

- b. Calculate the wavelength, **in nanometers (nm)**, of the light associated with the $n = 5 \rightarrow n = 3$ transition. Show all steps.

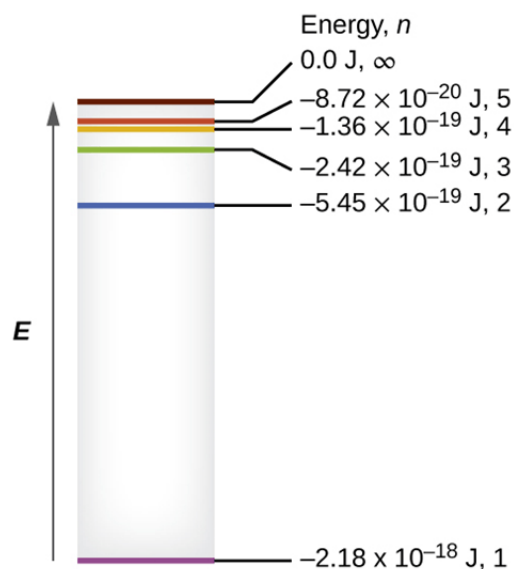


Figure 3.3.1: Hydrogen atom energy level diagram.

4. Which of the following electronic transitions in a hydrogen atom emits radiation of shorter wavelength? Explain your reasoning in terms of energy differences.

- a. Electronic transition from $n = 4$ to $n = 1$
 b. Electronic transition from $n = 4$ to $n = 3$

5. A typical microwave oven operates at a radiation wavelength of 12.2 cm.

- a. Calculate the frequency (Hz) and energy (J) of a single microwave photon.

- b. Calculate the total number of photons required to deliver 132.6 kJ of energy (the energy needed to heat 473 g of water from 26 °C to 93 °C).

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4: FOOD DYE CHROMATOGRAPHY

PURPOSE

- To use paper chromatography to separate and characterize common food dyes by measuring their R_f values.
- To analyze the synthetic dyes extracted from candy coatings and identify them by direct comparison to known standards.

INTRODUCTION

Most substances encountered in daily life are mixtures rather than pure compounds. Similarly, chemical compounds synthesized in laboratories often contain impurities such as reaction byproducts or unreacted starting materials. Consequently, a central focus in chemistry is the development of techniques to separate, isolate, and identify the components of complex mixtures.

Separation techniques exploit differences in physical properties among mixture components. For example, **filtration** separates substances in different physical states (such as insoluble solids from liquids), **centrifugation** utilizes differences in density, and **distillation** takes advantage of variations in boiling points. In this experiment, **paper chromatography** will be used to separate and analyze synthetic food dyes.

All chromatographic methods involve three essential components:

1. **Analyte:** The mixture to be separated.
2. **Mobile phase:** A liquid or gas solvent that carries the analyte through the system.
3. **Stationary phase:** A porous solid or supported liquid film that interacts with the analyte components to differential degrees.

Separation occurs because different components of the analyte mixture interact to varying degrees with both the stationary phase and the mobile phase. Components that interact more strongly with the stationary phase move more slowly, whereas components that favor the mobile phase move more rapidly. Paper chromatography is a specific type of partition/adsorption chromatography that utilizes cellulose paper as the stationary phase. The sample mixture is spotted near the bottom edge of the paper, which is then suspended upright in a covered developing tank containing a small volume of the mobile phase.

Capillary action draws the liquid mobile phase up the paper. As the solvent front advances past the sample spot, the components partition between the paper fibers and the migrating solvent. The separation is complete when the solvent front approaches the top edge of the paper.

The relative mobility of each separated spot is quantified by calculating its **retardation factor** (R_f):

$$R_f = \frac{D_{\text{dye}}}{D_{\text{solvent}}} \quad (4.1)$$

where D_{dye} is the distance traveled by the center of the dye spot from the origin line, and D_{solvent} is the total distance traveled by the mobile phase solvent front from the origin line. Under specified experimental conditions (stationary phase paper, mobile phase solvent composition, and temperature), the R_f value is a characteristic property of a compound used for identification.

4.1: Food Dye Chromatography - Experiment

4.2: Food Dye Chromatography - Pre-lab

4.3: Food Dye Chromatography - Data and Report

4.1: Food Dye Chromatography - Experiment

SAFETY PRECAUTIONS

- Synthetic food dyes can stain skin and clothing. Handle with care and clean up spills immediately.
- Ethanol is flammable. Keep extraction solvents away from hot plates, sparks, and open flames.
- Handle capillary tubes, scissors, and glass stirring rods carefully to prevent cuts or puncture injuries.

EQUIPMENT AND CHEMICALS NEEDED

Table 4.1.1: Materials List

Equipment & Supplies	Chemicals & Reagents
600-mL beaker & watch glass (or plastic wrap)	0.10% NaCl(aq) developing mobile phase
Chromatography paper (16 cm × 8 cm)	50/50 Ethanol/water extraction solvent
Porcelain spot plate & wooden toothpicks (or capillary tubes)	Standard FD&C Food Dyes (Blue 1, Blue 2, Green 3, Red 3, Red 40, Yellow 5, Yellow 6)
Test tubes (13 × 100 mm) & test tube rack	M&M's® candies (assorted colors)
30-cm metric ruler, pencil, scissors, & stapler	Deionized water

EXPERIMENTAL PROCEDURE

Part A: Prepare the Developing Tank and Paper

- Pour 25 mL of the 0.10% NaCl(aq) mobile phase into a clean 600-mL beaker. Cover the beaker tightly with a watch glass or plastic wrap. Ensure the liquid depth is less than 1.0 cm from the bottom of the beaker. Keep the chamber covered at all times so that the vapor space becomes fully saturated with solvent vapor.
- Obtain a 16 cm × 8 cm sheet of chromatography paper. Using a **pencil** and ruler, draw a light horizontal origin line 1.5 cm from the bottom long edge of the paper. Mark 8 evenly spaced tick marks along this origin line (spaced approximately 1.7–1.8 cm apart). Do not fold or staple the paper yet.

Part B: Prepare the Standard Food Dyes

- Dispense 1 drop of each standard FD&C dye solution into separate wells of a clean porcelain spot plate:

Table 4.1.2: Standard FD&C Food Dyes

Label	FD&C Dye Name	Common Chemical Name
B1	FD&C Blue No. 1	Brilliant Blue FCF
B2	FD&C Blue No. 2	Indigo Carmine
G3	FD&C Green No. 3	Fast Green FCF

Label	FD&C Dye Name	Common Chemical Name
R3	FD&C Red No. 3	Erythrosine
R40	FD&C Red No. 40	Allura Red AC
Y5	FD&C Yellow No. 5	Tartrazine
Y6	FD&C Yellow No. 6	Sunset Yellow FCF

Part C: Extract Dyes from Candy Coatings

1. Place 3 M&M's candies of a single color into a clean test tube.
2. Add approximately 3 mL of the 50/50 ethanol/water extraction solvent and swirl gently to dissolve the outer colored shell.
3. Decant the concentrated dye solution into an empty well of your spot plate as soon as the outer coating dissolves, *before the inner chocolate layer begins to melt or expose*.
4. Repeat this extraction procedure for other candy colors as assigned by your instructor.

Part D: Spot the Chromatography Paper

1. Using a clean toothpick or capillary tube for each sample, apply a tiny drop of each standard dye solution onto its designated mark on the pencil origin line. Keep spot diameters smaller than 2 mm. Allow spots to dry completely, then reapply titrations 4–5 times to build concentrated, compact sample spots. Label the dye abbreviation below each spot in pencil.
2. Spot the extracted candy dye solutions onto the remaining origin marks. Because candy extracts are more dilute, reapply sample drops 8–10 times, allowing full drying between applications to maintain small spot diameters.

Part E: Develop and Analyze the Chromatogram

1. Once all origin spots are thoroughly dry, roll the chromatography paper into a cylinder with the spots on the outside. Staple the top and bottom edges together, ensuring the paper edges **do not touch** across the seam (a 2–3 mm gap prevents edge effect distortion).
2. Carefully place the paper cylinder upright into the developing beaker. Ensure the liquid level is below the pencil origin line. Quickly cover the beaker with the watch glass.
3. Allow the mobile phase to ascend the paper until the solvent front reaches within 1–2 cm of the top edge.
4. Remove the chromatogram, lay it flat on a paper towel, and **immediately** mark the solvent front line with a pencil. Measure and record the total solvent front distance (D_{solvent}) in millimeters.
5. Circle each resolved spot with a pencil. Mark a horizontal line through the center of highest density for each spot. Measure the distance from the origin line to the center of each spot (D_{dye}) in millimeters. Calculate R_f values using:

$$R_f = \frac{D_{\text{dye}}}{D_{\text{solvent}}} \quad (4.1.1)$$

CHEMICAL DISPOSAL

- Dispose of extracted candy solutions and spent chromatography paper in the designated laboratory waste containers.

4.2: Food Dye Chromatography - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Compare and contrast the roles of the **stationary phase** and the **mobile phase** in chromatographic separations. Explain how differing chemical affinities for these two phases cause mixture components to separate.

2. Identify the specific materials serving as the stationary phase and the mobile phase in this laboratory procedure.

3. A student developed the paper chromatogram shown below for four standard dyes (Figure 4.2.1). Calculate the R_f value for each dye spot. Show your work clearly for each calculation.



Figure 4.2.1: Reference paper chromatogram.

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4.3: Food Dye Chromatography - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Characterization of Standard FD&C Food Dyes

Distance solvent front traveled (D_{solvent}): _____ mm

Table 4.3.1: Standard Dye Mobility Data

Standard Dye	Observed Color	Distance (D_{dye} , mm)	R_f Value
FD&C Blue No. 1	_____	_____	_____
FD&C Blue No. 2	_____	_____	_____
FD&C Green No. 3	_____	_____	_____
FD&C Red No. 3	_____	_____	_____
FD&C Red No. 40	_____	_____	_____
FD&C Yellow No. 5	_____	_____	_____
FD&C Yellow No. 6	_____	_____	_____

Characterization of Dyes Extracted from M&M's Candies

Distance solvent front traveled (D_{solvent}): _____ mm

Table 4.3.2: Candy Dye Identification Data

M&M Shell Color	Separated Spot Color	Distance (D_{dye} , mm)	Calculated R_f	Identified FD&C Dye
Red	_____	_____	_____	_____
Yellow	_____	_____	_____	_____
Green	_____	_____	_____	_____
Blue	_____	_____	_____	_____
Orange	_____	_____	_____	_____
Brown	_____	_____	_____	_____

POST-LAB QUESTIONS

1. Why is a lead pencil used to draw origin and solvent front lines on chromatography paper rather than an ink pen?

2. Why is it critical to keep initial sample spot diameters small (< 2 mm) when spotting the chromatography paper?

3. What is the theoretical range of possible values for the retardation factor (R_f)? Explain why an R_f value cannot be negative or greater than 1.0.

4. Discuss the physical trade-offs (advantages and disadvantages) of allowing the mobile phase solvent front to travel a longer distance up the chromatography paper.

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5: MOLECULAR GEOMETRY

PURPOSE

- To use molecular modeling kits to build three-dimensional representations of covalent molecules and polyatomic ions.
- To apply Valence Shell Electron Pair Repulsion (VSEPR) theory to determine electron domain geometries, molecular geometries, and bond angles.
- To determine central atom hybridization, count lone pairs, and predict overall molecular polarity.

INTRODUCTION

Understanding molecular geometry is essential for predicting the physical and chemical properties of molecules, including boiling points, solubilities, and chemical reactivity. In this laboratory exercise, you will utilize molecular modeling kits to construct three-dimensional representations of small molecules and polyatomic ions. By examining these physical models, you will determine electron domain and molecular geometries, identify central atom hybridization, and evaluate whether a molecule possesses an overall dipole moment (polar vs. nonpolar).

This hands-on approach reinforces concepts from Valence Shell Electron Pair Repulsion (VSEPR) theory and valence bond hybridization, offering a concrete visual framework to bridge two-dimensional Lewis structures with three-dimensional molecular architecture.

[5.1: Molecular Geometry - Experiment](#)

[5.2: Molecular Geometry - Pre-lab](#)

[5.3: Molecular Geometry - Data and Report](#)

5.1: Molecular Geometry - Experiment

SAFETY PRECAUTIONS

- Be cautious when assembling and disassembling plastic molecular models to avoid pinching fingers or damaging flexible connectors.
- Keep small model components inside storage containers when not in use.

EQUIPMENT AND CHEMICALS NEEDED

Table 5.1.1: Materials List

Equipment
Molecular modeling kits (with rigid and flexible connectors)

EXPERIMENTAL PROCEDURE

Part A: Build Molecular Models

1. Determine the total number of valence electrons and draw the complete Lewis structure for each assigned molecule or polyatomic ion.
2. Apply VSEPR theory to identify the number of electron domains around the central atom, determine the electron domain geometry, central atom hybridization, number of lone pairs, and molecular geometry (Figure 5.1.1).

Number of electron regions	Electron region geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramidal	 Bent or angular		
5	 Trigonal bipyramidal	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramidal	 Square planar	 T-shape	 Linear

Figure 5.1.1: VSEPR Electron Domain and Molecular Geometries.

3. Determine whether the molecule is polar or nonpolar based on individual bond dipoles and three-dimensional molecular symmetry.
4. Construct the 3D model using the modeling kit. Capture a photo or sketch the model showing all atoms, bonds, and electron pairs.
5. **Model Assembly Rules:** Do *not* select the central atom ball based on the color key in the kit lid. Instead, select the central ball based on the total number of electron domains surrounding the central atom. Pay close attention to hole geometry!

Single-Bonded Central Atoms:

Use Table 5.1.2 to select the appropriate central atom sphere. Use short, rigid connectors for single covalent bonds.

Table 5.1.2: Central Atom Ball Selection Guide

Electron Domains	Required Hole Geometry
2	Linear (use a 6-hole octahedral ball and utilize opposite holes)
3	Trigonal planar (use a 5-hole trigonal bipyramidal ball, leaving axial holes empty)
4	Tetrahedral (4 holes at 109.5°)
5	Trigonal bipyramidal (5 holes: 3 equatorial at 120° , 2 axial at 90°)
6	Octahedral (6 holes at 90°)

Multiple Bonds:

For double or triple bonds, select a central sphere that accommodates additional holes for multiple connectors. Use long, flexible springs for double and triple bonds, bending them to connect into the same terminal atom sphere.

Terminal Atoms:

Select terminal atom spheres of consistent color for each element type (e.g., white with 1 hole for hydrogen, red with 2 holes for oxygen, green/halogen with 1 hole for chlorine/fluorine).

5.2: Molecular Geometry - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. What is Valence Shell Electron Pair Repulsion (VSEPR) theory? Explain how it is applied to predict electron domain geometry and molecular geometry from a Lewis structure.

2. How do you determine the total number of valence electrons for a neutral molecule versus a polyatomic anion or cation? How does this count guide the placement of single bonds, octets, and lone pairs?

3. Distinguish between **electron domain geometry** and **molecular geometry**. Explain why the presence of lone pairs on a central atom alters the molecular geometry and reduces ideal bond angles.

4. Explain how to determine whether a polyatomic molecule is **polar** or **nonpolar**. Describe how electronegativity differences (bond dipoles) and molecular symmetry combine to dictate overall net dipole moment.

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5.3: Molecular Geometry - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 5.3.1: VSEPR Modeling Data – Group 1

Property / Formula	CH ₄	PH ₃	SO ₃	BCl ₃
Total Valence Electrons	_____	_____	_____	_____
Lewis Structure				
Electron Domains	_____	_____	_____	_____
Electron Domain Geometry	_____	_____	_____	_____
Central Atom Hybridization	_____	_____	_____	_____
Lone Pairs on Central Atom	_____	_____	_____	_____
Molecular Geometry	_____	_____	_____	_____
Polar or Nonpolar	_____	_____	_____	_____
3D Model Sketch / Photo				

Table 5.3.2: VSEPR Modeling Data – Group 2

Property / Formula	CO	H ₂ S	PCl ₅	SF ₄
Total Valence Electrons	_____	_____	_____	_____
Lewis Structure				
Electron Domains	_____	_____	_____	_____
Electron Domain Geometry	_____	_____	_____	_____
Central Atom Hybridization	_____	_____	_____	_____
Lone Pairs on Central Atom	_____	_____	_____	_____

Property / Formula	CO	H ₂ S	PCl ₅	SF ₄
Molecular Geometry	_____	_____	_____	_____
Polar or Nonpolar	_____	_____	_____	_____
3D Model Sketch / Photo				

Table 5.3.3: VSEPR Modeling Data – Group 3

Property / Formula	BrF ₃	XeF ₂	SF ₆	IF ₅
Total Valence Electrons	_____	_____	_____	_____
Lewis Structure				
Electron Domains	_____	_____	_____	_____
Electron Domain Geometry	_____	_____	_____	_____
Central Atom Hybridization	_____	_____	_____	_____
Lone Pairs on Central Atom	_____	_____	_____	_____
Molecular Geometry	_____	_____	_____	_____
Polar or Nonpolar	_____	_____	_____	_____
3D Model Sketch / Photo				

Table 5.3.4: VSEPR Modeling Data – Group 4

Property / Formula	XeF ₄	CO ₃ ²⁻	NO ₃ ⁻	OF ₂
Total Valence Electrons	_____	_____	_____	_____
Lewis Structure				
Electron Domains	_____	_____	_____	_____
Electron Domain Geometry	_____	_____	_____	_____
Central Atom Hybridization	_____	_____	_____	_____

Property / Formula	XeF ₄	CO ₃ ²⁻	NO ₃ ⁻	OF ₂
Lone Pairs on Central Atom	_____	_____	_____	_____
Molecular Geometry	_____	_____	_____	_____
Polar or Nonpolar	_____	_____	_____	_____
3D Model Sketch / Photo				

Table 5.3.5: VSEPR Modeling Data – Group 5

Property / Formula	SCl ₂	C ₂ H ₆	C ₂ H ₄	C ₂ H ₂
Total Valence Electrons	_____	_____	_____	_____
Lewis Structure				
Electron Domains	_____	_____	_____	_____
Electron Domain Geometry	_____	_____	_____	_____
Central Atom Hybridization	_____	_____	_____	_____
Lone Pairs on Central Atom	_____	_____	_____	_____
Molecular Geometry	_____	_____	_____	_____
Polar or Nonpolar	_____	_____	_____	_____
3D Model Sketch / Photo				

POST-LAB QUESTIONS

1. Explain how VSEPR theory predicts electron domain geometries versus molecular geometries. Identify any specific molecules modeled where electron domain geometry differed from molecular geometry and detail why lone pairs cause this distinction.

2. For molecules identified as polar, describe how individual bond dipoles and molecular geometry combine to produce a net

dipole moment. Explain why highly symmetric molecules containing polar bonds (such as SF_6 or CCl_4) are nonpolar overall.

3. Reflect on your experience constructing 3D physical models versus drawing 2D paper Lewis structures. How did building physical models improve your visual understanding of expanded octets and non-planar geometries (such as seesaw, T-shaped, or square pyramidal)?
-

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6: DETERMINATION OF EMPIRICAL FORMULA

PURPOSE

- To determine the empirical formula of a crystalline copper chloride hydrate using thermal dehydration and single-displacement reduction.

INTRODUCTION

Chemical formulas are a fundamental way to represent the composition of chemical compounds. The **empirical formula** represents the simplest whole-number ratio of atoms or ions in a compound. It is essential to distinguish this from the **molecular formula**, which specifies the actual number of atoms of each element present in a single molecule of a compound.

Many ionic compounds incorporate water molecules into their solid crystalline lattice, forming **hydrates**. Hydrates contain a specific stoichiometric ratio of water molecules relative to the anhydrous salt. This water of hydration is chemically bound within the crystal structure but can be removed by controlled heating. Once the water of hydration is completely driven off, the remaining substance is termed an **anhydrous compound**.

In this experiment, you will investigate a green-blue copper chloride hydrate. Heating the hydrate drives off the water of hydration, producing a color change to brown as anhydrous copper chloride forms. The mass difference between the original hydrate and the anhydrous compound allows you to calculate the mass of water lost.

To determine the ratio of copper to chlorine in the salt, you will perform a single-displacement reaction in aqueous solution using aluminum metal as the reducing agent. Aluminum metal reduces copper ions (Cu^{2+}) to solid elemental copper (Cu^0), while aluminum dissolves as Al^{3+} ions. The precipitated copper metal is isolated by vacuum filtration, dried, and weighed. By determining the moles of water, copper, and chlorine present in the original sample, you can establish their simplest whole-number mole ratio and write the empirical formula of the hydrate.

EXAMPLE 6.1: DETERMINING A HYDRATE EMPIRICAL FORMULA

A student heated a sample of nickel(II) sulfate hydrate and collected the following laboratory data:

- Mass of empty crucible and lid: 25.668 g
- Mass of crucible, lid, and nickel(II) sulfate hydrate: 28.225 g
- Mass of crucible, lid, and anhydrous nickel(II) sulfate: 27.389 g

Determine the empirical formula of the nickel(II) sulfate hydrate.

Step 1: Calculate the mass of the hydrated salt:

$$\text{Mass of hydrate} = 28.225 \text{ g} - 25.668 \text{ g} = 2.557 \text{ g} \quad (6.1)$$

Step 2: Calculate the mass of the anhydrous salt (NiSO_4):

$$\text{Mass of anhydrous } \text{NiSO}_4 = 27.389 \text{ g} - 25.668 \text{ g} = 1.721 \text{ g} \quad (6.2)$$

Step 3: Calculate the mass of water lost:

$$\text{Mass of } \text{H}_2\text{O lost} = 2.557 \text{ g} - 1.721 \text{ g} = 0.836 \text{ g} \quad (6.3)$$

Step 4: Calculate the moles of anhydrous NiSO_4 :

$$\text{Molar mass of } \text{NiSO}_4 = 58.69 + 32.07 + (4 \times 16.00) = 154.76 \text{ g/mol} \quad (6.4)$$

$$\text{Moles of NiSO}_4 = \frac{1.721 \text{ g}}{154.76 \text{ g/mol}} = 0.01112 \text{ mol} \quad (6.5)$$

Step 5: Calculate the moles of water (H₂O):

$$\text{Molar mass of H}_2\text{O} = (2 \times 1.008) + 16.00 = 18.016 \text{ g/mol} \quad (6.6)$$

$$\text{Moles of H}_2\text{O} = \frac{0.836 \text{ g}}{18.016 \text{ g/mol}} = 0.04640 \text{ mol} \quad (6.7)$$

Step 6: Determine the simplest mole ratio:

$$\text{Ratio for NiSO}_4 = \frac{0.01112 \text{ mol}}{0.01112 \text{ mol}} = 1.00 \quad (6.8)$$

$$\text{Ratio for H}_2\text{O} = \frac{0.04640 \text{ mol}}{0.01112 \text{ mol}} = 4.17 \approx 4 \quad (6.9)$$

The simplest whole-number mole ratio is 1:4. Therefore, the empirical formula is NiSO₄ · 4H₂O.

[6.1: Determination of Empirical Formula - Experiment](#)

[6.2: Determination of Empirical Formula - Pre-lab](#)

[6.3: Determination of Empirical Formula - Data and Report](#)

6.1: Determination of Empirical Formula - Experiment

SAFETY PRECAUTIONS

- Always wear chemical splash goggles and a lab coat in the laboratory.
- Hydrochloric acid (HCl) is corrosive and causes severe skin and eye burns. Handle with extreme care. If contact occurs, flush immediately with water and notify your instructor.
- The single-displacement reaction between aluminum wire and acid generates flammable hydrogen gas (H_2). Keep all open flames away from the reaction beaker during Part B.
- Use crucible tongs when handling heated crucibles. Allow heated equipment to cool completely on a wire gauze before weighing.

EQUIPMENT AND CHEMICALS NEEDED

Table 6.1.1: Materials List

Glassware & Hardware	Heating & Filtration	Chemicals & Reagents
50-mL beaker & watch glass	Crucible with lid & crucible tongs	Copper chloride hydrate
Glass stirring rod & dropper	Ring stand, iron ring, & clay triangle	Aluminum wire (~0.25 g)
Electronic balance (± 0.001 g)	Bunsen burner & hot plate	6 M HCl(aq) solution
	Büchner funnel, filter flask, & trap flask	95% Ethanol
	Filter paper & vacuum hose	Deionized water

EXPERIMENTAL PROCEDURE

Part A: Determine the Mass of Water by Dehydration

1. Measure and record the mass of a clean, dry crucible with its lid in the data table.
2. Add approximately 1.00 g of copper chloride hydrate to the crucible. Measure and record the combined mass of the crucible, lid, and hydrate sample.
3. Set up a Bunsen burner, ring stand, iron ring, and clay triangle as shown in Figure 6.1.1. Place the crucible containing the sample onto the clay triangle and begin heating gently with a small burner flame.
4. **CRITICAL:** Do **not overheat** the sample, or the copper chloride salt will decompose into black copper(II) oxide. Move the burner base back and forth continuously to distribute heat evenly.
5. Continue gentle heating until the green-blue crystals turn completely brown, signaling complete dehydration. Heat gently for an additional 2 minutes after the color change is complete.
6. Turn off the Bunsen burner and cover the crucible with its lid to prevent atmospheric moisture rehydration during cooling. Cool to room temperature for 15 minutes. Inspect the sample; if any green color remains, reheat gently.
7. Accurately measure and record the mass of the cooled crucible, lid, and anhydrous copper chloride.

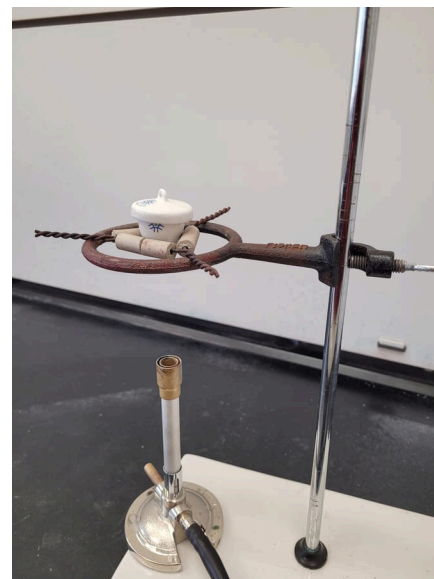


Figure 6.1.1: Thermal dehydration heating setup.

Part B: Determine the Mass of Copper by Displacement

8. Transfer the brown anhydrous copper chloride into a 50-mL beaker. Rinse the crucible with two 5-mL portions of deionized water, adding the washings to the beaker. Swirl the beaker to dissolve the salt completely, producing a clear green solution.
9. Obtain approximately 0.25 g of aluminum wire. Form the wire into a loose coil and submerge it completely in the copper solution.
10. Allow the single-displacement reaction to proceed for 25–30 minutes until the green solution turns completely colorless, indicating that all copper ions have been reduced to reddish-brown elemental copper metal.
11. Add 5 drops of 6 M $\text{HCl}(aq)$ to dissolve any insoluble aluminum hydroxide side-products and unreacted aluminum residue.
12. Use a stirring rod to scrape copper metal off the aluminum wire. Rinse the wire with deionized water into the beaker, then remove the cleaned aluminum wire.
13. Set up a vacuum filtration apparatus consisting of a Büchner funnel, rubber adapter, filter flask, safety trap flask, and vacuum hose (Figure 6.1.2).

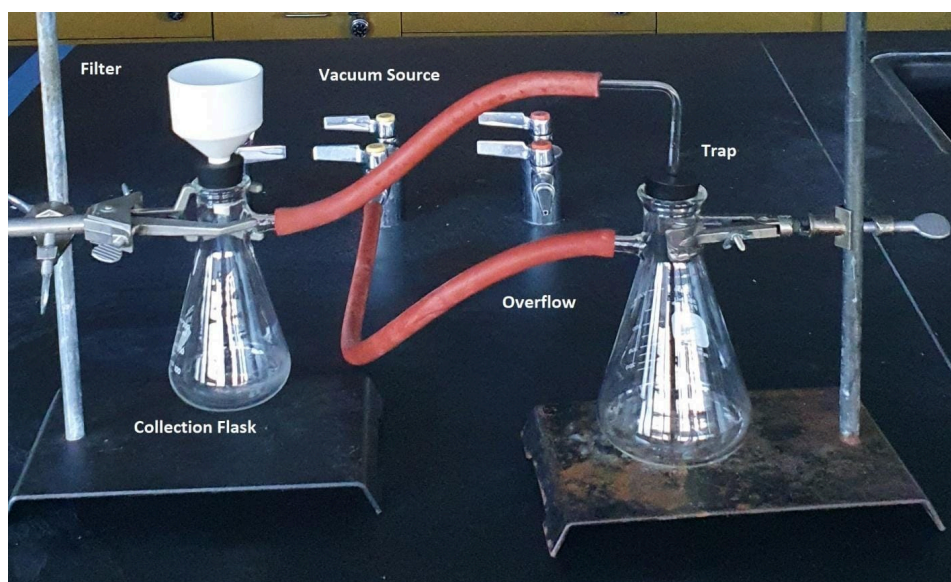


Figure 6.1.2: Vacuum filtration trap assembly.

14. Weigh a piece of dry filter paper and record its mass. Place the paper in the Büchner funnel and seat it with a few drops of deionized water.
15. Turn on the vacuum suction and quantitatively transfer the copper metal onto the filter paper. Wash the copper thoroughly with deionized water.
16. Turn off the vacuum and pour 10 mL of 95% ethanol over the copper residue to displace water. Wait 1 minute, then reapply vacuum suction for 5 minutes to air-dry the solid.
17. Carefully transfer the filter paper with recovered copper onto a watch glass and dry on a low hot plate setting.
18. Once cooled, measure and record the combined mass of the filter paper and dry copper metal.

CHEMICAL DISPOSAL

- Dispose of the liquid filtrate into the designated **Non-Halogenated Organic Waste** container.
- Place unreacted aluminum wire and dry recovered copper powder into the designated **Solid Metal Waste** container.
- Clean all glassware and return equipment to its designated station.

6.2: Determination of Empirical Formula - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Define the following chemical terms:

a. **Hydrate:**

b. **Anhydrous compound:**

c. **Empirical formula:**

2. A student heated a sample of tungsten (W) metal weighing 4.23 g in oxygen. The mass of the resulting tungsten oxide residue was 5.34 g. Calculate the empirical formula of this tungsten oxide. Show all calculations clearly.

3. Determine the empirical formula of an organic compound containing 60.0% carbon, 5.0% hydrogen, and 35.0% oxygen by mass. Show your work.

4. A compound contains 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen by mass, with an experimental molar mass of 120.0 g/mol. Determine its empirical formula and molecular formula. Show all work.
- _____

5. Why is it critical to avoid overheating the copper chloride hydrate during the thermal dehydration step? Describe the visual evidence that indicates sample overheating.
- _____

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6.3: Determination of Empirical Formula - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Mass of Water by Thermal Dehydration

Table 6.3.1: Dehydration Data

Measurement Description	Mass (g)
1. Mass of empty crucible and lid	_____
2. Mass of crucible, lid, and copper chloride hydrate	_____
3. Mass of copper chloride hydrate sample (2 – 1)	_____
4. Mass of crucible, lid, and anhydrous copper chloride	_____
5. Mass of anhydrous copper chloride (4 – 1)	_____
6. Mass of water of hydration lost (3 – 5)	_____

Show calculation for Mass of Anhydrous Copper Chloride (4 – 1):

Show calculation for Mass of Water Lost (3 – 5):

Part B: Mass of Copper by Displacement Reaction

Table 6.3.2: Copper Displacement Data

Measurement Description	Mass (g)
7. Mass of dry filter paper	_____
8. Mass of filter paper + dry recovered copper metal	_____
9. Mass of recovered copper metal (8 – 7)	_____

Show calculation for Mass of Recovered Copper Metal (8 – 7):

Part C: Mass of Chlorine in the Original Sample

$$\text{Mass of Chlorine} = \text{Mass of Hydrate (3)} - \text{Mass of Water (6)} - \text{Mass of Copper (9)} \quad (6.3.1)$$

Table 6.3.3: Chlorine Mass Determination

Measurement Description	Mass (g)
10. Mass of chlorine in sample (3 – 6 – 9)	_____

Show calculation for Mass of Chlorine (3 – 6 – 9):

Part D: Empirical Formula Determination

Table 6.3.4: Mole Calculations and Ratio Analysis

Component	Calculated Moles	Calculation Setup / Work
11. Water (H ₂ O)	_____	_____
12. Copper (Cu)	_____	_____
13. Chlorine (Cl)	_____	_____

Simplest Whole-Number Mole Ratio Calculation (Cu : Cl : H₂O):

Empirical Formula of Copper Chloride Hydrate: _____

IUPAC Chemical Name of Hydrate: _____

POST-LAB QUESTIONS

1. If a student overheated the hydrated salt, causing decomposition into black copper(II) oxide (CuO), how would this operational error affect the calculated mass of water lost and the calculated mole ratio of H_2O to Cu ? Explain clearly.

2. Explain the chemical role of aluminum wire in this experiment. Write balanced net ionic equations for the reduction of Cu^{2+} by Al^0 and the reaction of unreacted Al^0 with $\text{HCl}(aq)$.

3. If copper metal fine powder passed through the filter paper during vacuum filtration and was lost in the filtrate, how would this loss affect the calculated mass of chlorine and the resulting $\text{Cu} : \text{Cl}$ mole ratio?

4. A student obtained the following experimental data:

- Mass of hydrated salt = 5.00 g
- Mass of anhydrous salt = 3.98 g
- Mass of recovered copper metal = 1.25 g

Calculate the empirical formula of this copper chloride hydrate. Show all steps clearly.

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7: OBSERVING TWO MAJOR CLASSES OF CHEMICAL REACTIONS

PURPOSE

- To observe evidence of double displacement reactions, including precipitate formation, gas evolution, and color change.
- To observe evidence of redox reactions, such as metal displacement and the evolution of gases.
- To apply solubility rules to predict precipitate formation and confirm products based on the activity series of metals from experimental observations.
- To write balanced molecular, total ionic, and net ionic equations for double displacement and redox reactions.

INTRODUCTION

Part A: Double Displacement (Metathesis) Reactions

In double displacement reactions, elements in two compounds switch places to form new compounds.



These reactions, also known as *metathesis reactions*, involve the exchange of ions between compounds in solution, forming new compounds. Typically, a product is a solid, a gas, or a liquid such as water. When a solid forms, it's called a precipitate, as seen in equation 7.2.

The formation of solids, liquids, gases, or color changes, as well as temperature changes, can indicate the presence of chemical reactions. Exothermic reactions release heat, while endothermic reactions absorb heat, indicating a chemical reaction.

In this experiment, you'll mix various aqueous solutions and observe chemical reactions, using your observations to write balanced chemical equations.

Part B: Redox (Electron Transfer) Reactions

Redox reactions involve the transfer of electrons between chemical species. Oxidation is the loss of electrons, while reduction is the gain of electrons. These reactions are fundamental to many chemical and biological processes.

In this experiment, you will investigate redox reactions involving metals and metal ions in aqueous solutions. The relative reactivity of metals will be determined by observing whether a metal can displace the ions of another metal from the solution. This data can be used to write and balance the molecular and net ionic equation for each observed reaction and construct an activity series, which ranks metals in order of their reactivity.

[7.1: Observing Two Major Classes of Chemical Reactions - Experiment](#)

[7.2: Observing Two Major Classes of Chemical Reactions - Pre-lab](#)

[7.3: Observing Two Major Classes of Chemical Reactions - Data and Report](#)

7.1: Observing Two Major Classes of Chemical Reactions - Experiment

SAFETY PRECAUTIONS

- Always wear safety goggles and a lab coat in the laboratory.
- Hydrochloric acid and sulfuric acid are corrosive. Avoid contact with skin and eyes. In case of contact, flush the affected area immediately with plenty of water and seek medical attention if necessary.
- Sodium hydroxide is caustic. Avoid contact with skin and eyes. In case of contact, flush immediately with water.
- Metal solutions are toxic. Avoid ingestion and skin contact. Dispose of all chemical waste in designated containers.

EQUIPMENT AND CHEMICALS NEEDED

Table 7.1.1: Materials List

Equipment	Solutions	Solutions	Metal Samples
12- or 24-hole well plates & small test tubes	1.0 M Na_2CO_3	3.0 M H_2SO_4	Ca metal
Dropper bottles	6.0 M HCl	0.1 M AgNO_3	Cu metal
Test tube rack	0.1 M Na_2CO_3	0.1 M KCl	Al metal
Labels	0.1 M $\text{Cu}(\text{NO}_3)_2$	0.1 M NaNO_3	Mg metal
	0.1 M Na_3PO_4	0.1 M CuCl_2	Zn metal
	0.1 M $\text{Fe}(\text{NO}_3)_3$	0.1 M $\text{Mg}(\text{NO}_3)_2$	
	0.1 M NaOH	0.1 M $\text{Zn}(\text{NO}_3)_2$	
	1.0 M NaHCO_3		

EXPERIMENTAL PROCEDURE

Part A: Perform Double Displacement Reactions

Evidence of a chemical reaction can be observed through the formation of a precipitate, gas evolution, a color change, a pH shift, or temperature changes (exothermic or endothermic).

1. Label well plates or test tubes with the appropriate reactant combinations.
2. Using dropper bottles, add approximately 10 drops (or 1 mL in a test tube) of each solution to the corresponding well or test tube as indicated below and record your observations on the report sheet:
 - a. 1.0 M Na_2CO_3 + 6.0 M HCl
 - b. 0.1 M Na_2CO_3 + 0.1 M $\text{Cu}(\text{NO}_3)_2$
 - c. 0.1 M $\text{Cu}(\text{NO}_3)_2$ + 0.1 M Na_3PO_4
 - d. 0.1 M $\text{Fe}(\text{NO}_3)_3$ + 0.1 M Na_3PO_4
 - e. 0.1 M $\text{Fe}(\text{NO}_3)_3$ + 0.1 M NaOH

- f. $6.0\text{ M NaOH} + 6.0\text{ M HCl}$
 - g. $1.0\text{ M NaHCO}_3 + 6.0\text{ M HCl}$
 - h. $1.0\text{ M NaHCO}_3 + 3.0\text{ M H}_2\text{SO}_4$
 - i. $0.1\text{ M AgNO}_3 + 0.1\text{ M KCl}$
 - j. $0.1\text{ M KCl} + 0.1\text{ M NaNO}_3$
3. Carefully observe each reaction mixture for evidence of a chemical reaction. Record your observations in the data table. Note any precipitate formation, gas evolution, color changes, or temperature changes.
 4. For each reaction, write the balanced molecular equation, complete ionic equation, and net ionic equation. If no reaction occurred, record "NR".

Part B: Perform Redox Reactions

B1: React Metals with Acid

1. Place a small piece of each metal (Ca, Cu, Al, Mg, and Zn) into separate wells or small test tubes.
2. Add approximately 10 drops (or 1 mL) of 6.0 M HCl to each well or test tube.
3. Observe the reactions for at least 5 minutes. Record your observations in the data table, noting any gas evolution, color changes, or disappearance of the metal.
4. Write balanced molecular and net ionic equations for the observed reactions.

B2: React Metals with Metal Ions

1. Place a small piece of aluminum in three separate wells or test tubes.
2. Add approximately 10 drops (or 1 mL) of each of the following solutions: 0.1 M CuCl_2 , $0.1\text{ M Mg(NO}_3)_2$, and $0.1\text{ M Zn(NO}_3)_2$.
3. Observe reactions for at least 5 minutes. Record any color changes, precipitate formation, or changes in the appearance of the metal surface.
4. Repeat using Cu, Mg, and Zn, adding the corresponding metal ion solutions as listed in the data table.
5. Write balanced molecular and net ionic equations for the observed reactions.

CHEMICAL DISPOSAL

- Place all aqueous waste solutions into the designated **Inorganic Liquid Waste** container.
- Place all unreacted metal solids into the designated **Solid Metal Waste** container.

7.2: Observing Two Major Classes of Chemical Reactions - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Define or explain the following chemical concepts:

a. **Double displacement (metathesis) reaction:**

b. **Oxidation and reduction in terms of electron transfer:**

c. **Concept of an activity series of metals:**

d. **Difference between a single displacement reaction and a redox reaction:**

2. Using solubility rules, predict whether a precipitate will form when the following aqueous solutions are mixed. Write the **balanced** molecular, complete ionic, and net ionic equations, and identify the spectator ions. (If no reaction occurs, write "NR".)

a. **Aqueous solutions of $\text{Pb}(\text{NO}_3)_2$ and KI:**

Molecular Equation: _____

Complete Ionic: _____

Spectator Ions: _____

Net Ionic Equation: _____

b. **Aqueous solutions of NaCl and AgNO_3 :**

Molecular Equation: _____

Complete Ionic: _____

Spectator Ions: _____

Net Ionic Equation: _____

c. Aqueous solutions of CuSO_4 and NaOH :

Molecular Equation: _____

Complete Ionic: _____

Spectator Ions: _____

Net Ionic Equation: _____

3. Predict whether the following redox reaction will occur as written:



4. Identify the oxidizing agent and the reducing agent in the following reaction:



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7.3: Observing Two Major Classes of Chemical Reactions - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Double Displacement (Metathesis) Reactions

1. $1.0 \text{ M Na}_2\text{CO}_3 + 6.0 \text{ M HCl}$ (mix in a small test tube)

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

2. $0.1 \text{ M Na}_2\text{CO}_3 + 0.1 \text{ M Cu(NO}_3)_2$

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

3. $0.1 \text{ M Na}_3\text{PO}_4 + 0.1 \text{ M Cu(NO}_3)_2$

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

4. $0.1 \text{ M Na}_3\text{PO}_4 + 0.1 \text{ M Fe(NO}_3)_3$

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

5. $0.1 \text{ M NaOH} + 0.1 \text{ M Fe(NO}_3)_3$

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

6. 6.0 M NaOH + 6.0 M HCl (**mix in a small test tube**)

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

7. 1.0 M NaHCO₃ + 6.0 M HCl (**mix in a small test tube**)

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

8. 1.0 M NaHCO₃ + 3.0 M H₂SO₄ (**mix in a small test tube**)

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

9. 0.1 M KCl + 0.1 M AgNO₃

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

10. 0.1 M KCl + 0.1 M NaNO₃

Observations: _____

Molecular Equation: _____

Complete Ionic: _____

Net Ionic Equation: _____

Part B: Redox Reactions

B1: Reactions of Metals and Hydrochloric Acid

1. **Calcium + 6.0 M Hydrochloric Acid**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

2. **Copper + 6.0 M Hydrochloric Acid**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

3. **Aluminum + 6.0 M Hydrochloric Acid**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

4. **Magnesium + 6.0 M Hydrochloric Acid**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

5. **Zinc + 6.0 M Hydrochloric Acid**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

B2: Reactions of Metals and Solutions of Metal Ions

1. **Aluminum + 0.1 M Copper(II) Chloride**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

2. **Aluminum + 0.1 M Magnesium Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

3. **Aluminum + 0.1 M Zinc Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

4. **Copper + 0.1 M Magnesium Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

5. **Copper + 0.1 M Zinc Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

6. **Magnesium + 0.1 M Copper(II) Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

7. **Magnesium + 0.1 M Zinc Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

8. **Zinc + 0.1 M Copper(II) Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

9. **Zinc + 0.1 M Magnesium Nitrate**

Observations: _____

Molecular Equation: _____

Net Ionic Equation: _____

POST-LAB QUESTIONS

1. Explain why the reaction between a strong acid and a strong base is classified as a double displacement reaction.

2. Which of the double displacement reactions performed in Part A could also be classified as acid-base neutralization reactions?

3. Construct an activity series for the metals tested in this experiment (Ca, Al, Mg, Zn, Cu), including hydrogen (H_2), based on your experimental observations. Rank from most active to least active.

4. For each observed single-displacement reaction in Part B2, identify which metal element was oxidized and which metal cation was reduced.

5. Provide five distinct examples of oxidation-reduction (redox) reactions that occur in everyday life and briefly describe the chemical change occurring in each.

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8: SYNTHESIS OF ALUM FROM ALUMINUM FOIL

PURPOSE

- Synthesize alum crystals from aluminum foil through a multi-step reaction process.
- Apply stoichiometric principles to calculate theoretical and percent yields.

INTRODUCTION

Alum is the common name for potassium aluminum sulfate dodecahydrate, with the chemical formula $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$. It has various industrial and consumer applications, including water purification, dyeing mordants, and medicinal astringents.

This experiment involves the synthesis of alum from aluminum foil through a series of chemical transformations:

1. **Reaction with Potassium Hydroxide:** Elemental aluminum reacts with potassium hydroxide (KOH) to form a soluble potassium tetrahydroaluminate complex, $\text{K}[\text{Al}(\text{OH})_4]$, evolving flammable hydrogen gas.
2. **Neutralization with Sulfuric Acid:** Sulfuric acid (H_2SO_4) is added to neutralize excess base and precipitate aluminum hydroxide, which redissolves upon further acid addition to form soluble potassium and aluminum sulfates.
3. **Crystallization:** Hydrated potassium aluminum sulfate (alum) is crystallized from the aqueous solution by cooling in an ice bath.

The overall balanced equation for the synthesis of alum is:



This experiment also highlights a green chemistry application, as it utilizes post-consumer recycled aluminum foil as the starting material.

[8.1: Synthesis of Alum from Aluminum Foil - Experiment](#)

[8.2: Synthesis of Alum from Aluminum Foil - Pre-lab](#)

[8.3: Synthesis of Alum from Aluminum Foil - Data and Report](#)

8.1: Synthesis of Alum from Aluminum Foil - Experiment

SAFETY PRECAUTIONS

- Wear appropriate personal protective equipment, including safety goggles, gloves, and a lab coat.
- Potassium hydroxide and sulfuric acid are corrosive. Avoid contact with skin and eyes. Flush spills immediately with water.
- The reaction with potassium hydroxide produces hydrogen gas (H_2), which is highly flammable. Perform this step in an operating fume hood away from open flames.
- Sulfuric acid should be added slowly and carefully to the reaction mixture to control exothermic heating and avoid splattering.
- Use caution when handling hot solutions and hot plates. Use beaker tongs or heat-resistant gloves.

EQUIPMENT AND CHEMICALS NEEDED

Table 8.1.1: Materials List

Equipment & Hardware	Filtration & Heating	Chemicals & Reagents
250-mL beaker & glass stirring rod	Ice bath & spatula	1.4 M KOH solution
Hot plate & beaker tongs	Filter paper & Büchner funnel	9.0 M H_2SO_4 solution
Electronic balance (± 0.001 g)	Rubber adapter & filter flask	95% Ethanol
Scissors & ruler	Safety trap flask & vacuum hose	Aluminum foil & deionized water

EXPERIMENTAL PROCEDURE

Part A: Prepare the Aluminum Foil

1. Accurately weigh between 0.90 g and 1.20 g of clean aluminum foil on an electronic balance. Record the exact mass in the Data Table.
2. Cut the aluminum foil into small squares (approximately $5\text{ mm} \times 5\text{ mm}$) to maximize surface area, and place them into a 250-mL beaker.

Part B: React with Potassium Hydroxide

3. Place the beaker containing the aluminum pieces in an operating fume hood.
4. Carefully measure and add 50 mL of 1.4 M KOH solution to the beaker. Record all immediate observations in the report sheet.
5. Once the initial vigorous bubbling subsides, place the beaker on a hot plate set to medium heat inside the fume hood to accelerate the reaction.
6. If the volume of liquid evaporates below 25 mL, add small portions of deionized water to maintain a volume of approximately 25 mL.
7. Continue gentle heating until the evolution of hydrogen gas completely ceases (approximately 20–30 minutes). The solution may appear dark gray or black due to minor metallic impurities in the foil.
8. Remove the beaker from the hot plate and allow it to cool to room temperature.

Part C: Neutralize with Sulfuric Acid

9. Working in the fume hood, slowly and cautiously add 20 mL of 9.0 M H_2SO_4 to the cooled solution while stirring continuously with a glass rod.
10. Observe the formation of a gelatinous white $\text{Al}(\text{OH})_3$ precipitate, which will redissolve as additional sulfuric acid is added and the solution becomes acidic.
11. If the white precipitate does not fully redissolve after all acid is added, gently heat and stir the mixture on a hot plate until the solution becomes clear.

Part D: Crystallize and Filter Alum

12. Chill the clear solution in an ice bath for 15–20 minutes to induce precipitation of alum crystals.
13. Prepare a wash solution by mixing 15 mL of 95% ethanol and 15 mL of deionized water in a small beaker, then cool this mixture in the ice bath.
14. If crystals do not form after chilling, gently scratch the inner bottom glass surface of the beaker with a glass stirring rod to provide nucleation sites.
15. If crystallization still does not occur, gently boil the solution on a hot plate to reduce its volume by roughly one-third, then recool in the ice bath.
16. Assemble the vacuum filtration apparatus consisting of a Büchner funnel, rubber adapter, filter flask, safety trap flask, and heavy-walled vacuum hose (Figure 8.1.1).

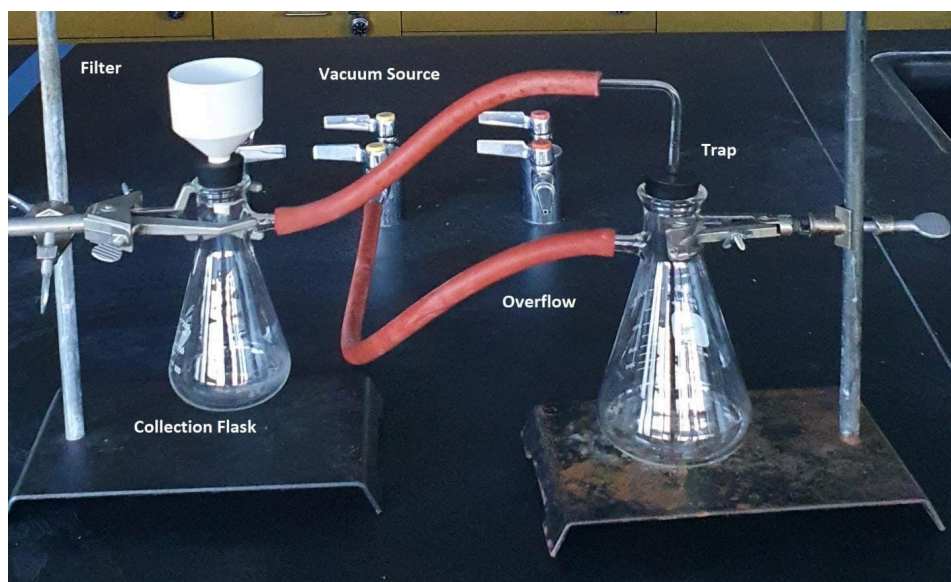


Figure 8.1.1: Vacuum filtration apparatus setup.

17. Place a disk of filter paper into the Büchner funnel, wet it with a few drops of cold ethanol/water wash, and apply vacuum suction.
18. Quantitatively transfer the alum crystals into the funnel. Rinse the beaker with small portions of the chilled ethanol/water wash to transfer all solid material.
19. Wash the filtered crystals with the remaining chilled ethanol/water solution to remove excess acid and soluble salts.
20. Draw air through the filter cake under full vacuum for 5–10 minutes to partially dry the crystals.

Part E: Collect and Weigh Alum

21. Transfer the dry alum crystals to a pre-weighed beaker or watch glass, weigh the recovered mass, and calculate the limiting reactant, theoretical yield, and percent yield.

CHEMICAL DISPOSAL

- Dispose of the liquid filtrate into the designated **Non-Halogenated Organic Waste** container.
- Place recovered alum crystals into the designated **Recovered Alum Product Jar**.
- Rinse all glassware thoroughly with deionized water and return equipment to your laboratory station.

8.2: Synthesis of Alum from Aluminum Foil - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. List several common industrial, commercial, and medical applications of alum (potassium aluminum sulfate dodecahydrate).

2. Explain how synthesizing alum from scrap aluminum foil demonstrates a "green chemistry" principle.

3. Identify at least three major chemical or physical hazards present in this experimental procedure and detail the appropriate safety measures required for each.

4. Write the overall balanced chemical equation for the synthesis of alum from elemental aluminum, potassium hydroxide, sulfuric acid, and water.

5. Define the following stoichiometric terms:

a. **Limiting reactant:**

b. **Theoretical yield:**

c. **Percent yield:**

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8.3: Synthesis of Alum from Aluminum Foil - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 8.3.1: Mass Measurements

Substance / Measurement	Mass (g)
Initial Measurement	
1. Mass of aluminum foil starting material	_____
Final Product Measurements	
2. Mass of empty collection container / beaker	_____
3. Mass of container + dry synthesized alum product	_____
4. Mass of recovered dry alum product (3 – 2)	_____

QUALITATIVE OBSERVATIONS

Observations during reaction of aluminum with KOH:

Observations during addition of H_2SO_4 neutralization:

Observations during ice-bath crystallization:

Other physical observations (crystal appearance, color, clarity):

CALCULATIONS AND STOICHIOMETRY

Determining the Limiting Reactant

Table 8.3.2: Moles of Reactants and Potential Product Yield

Reactant	Quantity Used	Reactant Moles	Potential Moles of Alum Produced
Al	_____ g	_____ mol	_____ mol
1.4 M KOH	50.0 mL	_____ mol	_____ mol
9.0 M H ₂ SO ₄	20.0 mL	_____ mol	_____ mol

Identified Limiting Reactant: _____

Calculating Theoretical and Percent Yield

The overall reaction formula for theoretical yield calculation is:



$$\text{Percent Yield} = \left(\frac{\text{Actual Mass of Alum Recovered (g)}}{\text{Theoretical Mass of Alum (g)}} \right) \times 100\% \quad (8.3.2)$$

Theoretical Yield Calculation (show complete setup with conversion factors):

Percent Yield Calculation (show complete work):

POST-LAB QUESTIONS

1. What specific experimental factors or operational errors could cause a calculated percent yield to be significantly lower than 100%?

2. What specific experimental factors could lead to a calculated percent yield greater than 100%?

3. Why must the aluminum foil be cut into small pieces before reacting with KOH? How does piece size affect reaction kinetics?

4. Explain why the synthesized alum crystals are washed with a chilled ethanol/water mixture rather than room-temperature deionized water.

5. Identify potential sources of error in this multi-step synthesis and suggest specific procedural improvements to increase yield purity and accuracy.

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9: MOLAR VOLUME OF A GAS

PURPOSE

- To determine the molar volume of CO_2 gas at room temperature.
- To use Dalton's Law of Partial Pressures and the combined gas law to determine the molar volume of CO_2 at STP.

INTRODUCTION

In this experiment, carbon dioxide gas will be generated from the thermal decomposition of sodium bicarbonate, NaHCO_3 . When solid sodium bicarbonate is heated, it decomposes into solid sodium carbonate, water vapor, and carbon dioxide gas. **You will need to write a balanced chemical equation for this reaction.**

You will heat a measured sample of sodium bicarbonate until the evolution of carbon dioxide is complete. The volume of carbon dioxide produced will be determined by measuring the volume of water displaced from a flask into a beaker. To complete the calculations in this lab, you will determine the theoretical moles of carbon dioxide evolved and use them to calculate the molar volume of carbon dioxide under laboratory room conditions.

Because the reaction produces water vapor and the gas is collected over water, water vapor will be mixed with the carbon dioxide gas in the collection flask. The experiment is designed so that the total pressure of the gas mixture (carbon dioxide and water vapor) equals atmospheric pressure, which you will measure using a laboratory barometer. The partial pressure of dry carbon dioxide in the flask can be calculated using Dalton's Law of Partial Pressures. A chart of the vapor pressure of water as a function of temperature is provided at the end of the report sheet.

Once the pressure, volume, temperature, and moles of carbon dioxide have been established under room conditions, the combined gas law can be used to convert the gas volume to standard temperature and pressure (STP, 0°C and 1.00 atm) and calculate the molar volume at STP.

The experimental apparatus is shown in Figure 9.1 in the procedure. The sample of sodium bicarbonate is placed in a hard-glass test tube connected to a water-filled flask. As carbon dioxide gas is generated, it displaces water from the flask into a beaker. The volume of water collected in the beaker is equal to the volume of gas generated at atmospheric pressure.

[9.1: Molar Volume of a Gas - Experiment](#)

[9.2: Molar Volume of a Gas - Pre-lab](#)

[9.3: Molar Volume of a Gas - Data and Report](#)

9.1: Molar Volume of a Gas - Experiment

SAFETY PRECAUTIONS

- Always wear chemical splash goggles and a lab coat in the laboratory.
- You will be using an open Bunsen burner flame. Keep all flammable materials, paper, and hair away from the heat source.
- Heated glassware remains hot for an extended period and can cause severe burns. Allow the test tube and apparatus to cool completely before touching.
- Ensure all stopper connections are tight to prevent hot gas or pressure build-up leaks.

EQUIPMENT AND CHEMICALS NEEDED

Table 9.1.1: Materials List

Equipment & Hardware	Chemicals & Supplies
Ring stand and utility clamp	Bunsen burner & striker
Large hard-glass test tube	Pinch clamp & rubber tubing
1-hole rubber stopper for test tube	400-mL beaker
Glass tubing assemblies (bent glass tubes)	250-mL graduated cylinder
500-mL Florence flask (or side-arm vacuum flask)	Thermometer or temperature probe
2-hole stopper for Florence flask	Laboratory barometer
Large rubber bulb or pipet pump	Sodium bicarbonate (NaHCO_3)

EXPERIMENTAL PROCEDURE

Part A: Assemble the Apparatus

Goal: Establish a continuous, leak-free water siphon between the flask and the beaker.

1. Assemble the gas displacement apparatus as shown in Figure 9.1.1.
2. **Fill the flask** almost completely to the top with tap water and **open the pinch clamp** on the delivery tubing.
3. **Remove the rubber stopper** from the empty test tube.
4. Using a large rubber bulb, gently apply air pressure through the tube of the test tube stopper until water fills the delivery hose completely and flows into the beaker, creating a continuous siphon.
5. **Reinsert the stopper into the test tube** while water is flowing. If the system is airtight, water flow will cease immediately. If water continues to flow, a leak is present; press all stoppers firmly. Ask your instructor if leaks persist.
6. **Close the pinch clamp** on the rubber hose while it remains completely filled with water. Do not drain the hose.

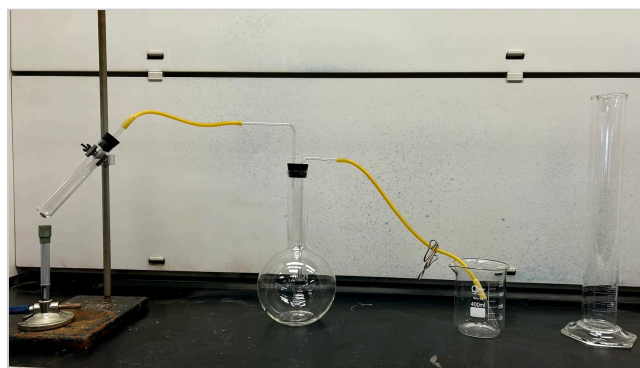


Figure 9.1.1: Molar Volume Apparatus Setup.

7. **Remove the stopper** from the test tube.

Part B: Prepare the Sample and Flask

Goal: Prepare the chemical sample and initialize the water level in the flask.

1. Ensure the large test tube is completely clean and dry inside.
2. Accurately weigh approximately 1.20 g of sodium bicarbonate on an analytical balance. Record the exact mass in the Data Table. Transfer the solid into the dry test tube. **Do not attach the test tube stopper yet.**
3. Fill the 400-mL beaker approximately half-full with water.
4. Refill the flask: **Open the pinch clamp** and raise the beaker above the level of the flask so water siphons back into the flask. Fill the flask until the water level reaches about 1 cm below the internal glass tubing.
5. **Close the pinch clamp** while keeping the hose submerged.

Part C: Equalize Initial System Pressure

Goal: Ensure initial internal system pressure matches room atmospheric pressure (P_{atm}).

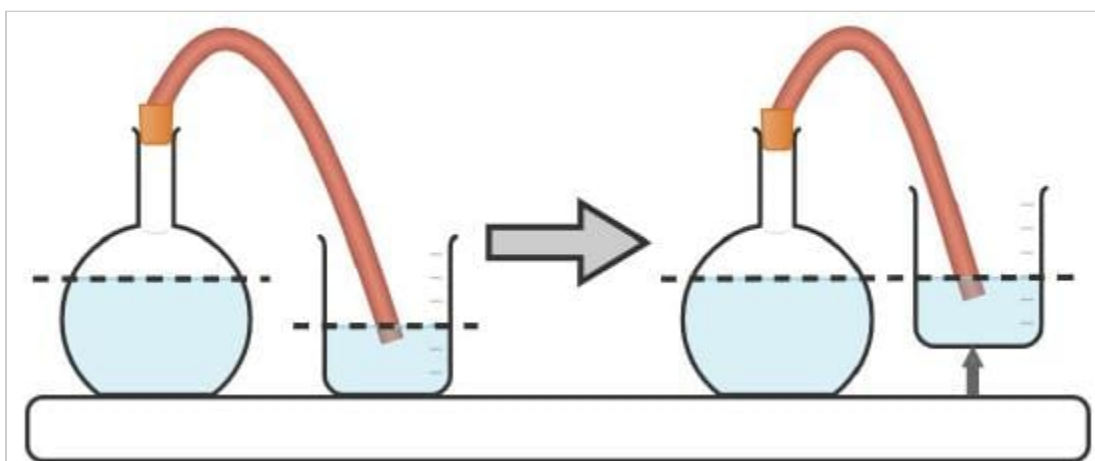


Figure 9.1.2: Pressure Equalization Procedure.

1. **Seal the System:** Firmly press the rubber stopper into the test tube containing the sample.
2. **Open the Line:** Open the pinch clamp.
3. **Leveling:** Manually raise or lower the beaker until the water level (meniscus) in the beaker is at the exact same height as the water level in the flask (Figure 9.1.2). When levels match, internal pressure equals atmospheric pressure.
4. **Lock Pressure:** Holding the beaker steady at this height, **close the pinch clamp**.

Part D: Final Setup for Gas Collection

Goal: Prepare the beaker for water collection without breaking the liquid siphon.

1. **Empty the Beaker:** Pour out the water from the beaker. Do not dry the inside; residual droplets will not affect measurements. **Crucial:** Keep the rubber hose filled with water.
2. **Position Hose:** Place the empty beaker back on the lab bench and position the end of the rubber hose deep inside the beaker.
3. **Open System:** **Open the pinch clamp.** A small amount of water may flow into the beaker or flask until equilibrium is reached; do not empty the beaker again.

Part E: Decompose Sodium Bicarbonate

Goal: Thermally decompose the sample to generate carbon dioxide gas.

1. Light the Bunsen burner. Gently heat the test tube containing the NaHCO_3 by moving the flame back and forth along the tube. After 1 minute of gentle heating, apply direct heat to the sample.
2. Observe carbon dioxide evolution as displaced water flows into the beaker. Continue strong heating until gas evolution ceases completely.
3. Remove the Bunsen burner flame and allow the entire apparatus to cool to room temperature (15–20 minutes).

Part F: Equalize Final System Pressure

Goal: Equalize final pressure to ensure collected volume reflects gas produced at atmospheric pressure.

1. **Cooling Check:** Verify the test tube has cooled completely to room temperature. Equalizing while gas is warm will introduce significant volume error upon cooling.
2. **Final Leveling:** Raise or lower the beaker until the water level in the beaker matches the water level in the flask exactly.
3. **Final Seal:** While keeping the water levels aligned, **close the pinch clamp** to trap the exact volume of displaced water at atmospheric pressure. At this point: $P_{\text{total}} = P_{\text{CO}_2} + P_{\text{H}_2\text{O}} = P_{\text{atm}}$.
4. Remove the rubber hose from the beaker.

Part G: Measure Displaced Water and Conditions

Goal: Measure displaced water volume, temperature, and atmospheric pressure.

1. Measure the volume of water collected in the beaker using a 250-mL graduated cylinder. Record this volume to the nearest 0.5 mL.
2. Measure and record the temperature of the water in the beaker. This corresponds to the gas temperature (T_{lab}) used to look up water vapor pressure.
3. Read and record the laboratory barometric pressure (P_{atm}).

Part H: Repeat for Trial 2

1. Clean the test tube, dry it thoroughly, and repeat the complete procedure (Parts A–G) with a second sample of sodium bicarbonate.

CHEMICAL DISPOSAL

- Dispose of solid sodium carbonate residue from the test tube in the designated solid chemical waste container.
- Pour excess water down the laboratory drain. Clean all glassware before returning it.

9.2: Molar Volume of a Gas - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Write the balanced chemical equation for the thermal decomposition of solid sodium bicarbonate (NaHCO_3) into solid sodium carbonate (Na_2CO_3), water vapor (H_2O), and carbon dioxide gas (CO_2). Include all state symbols.

2. Explain how Dalton's Law of Partial Pressures ($P_{\text{total}} = P_{\text{CO}_2} + P_{\text{H}_2\text{O}}$) is applied in this experiment to calculate the partial pressure of dry CO_2 gas collected over water.

3. Why is it necessary to equalize the water level in the beaker with the water level in the flask both before heating and after cooling before clamping the delivery tube?

4. A student decomposed 1.150 g of NaHCO_3 and collected 168.5 mL of CO_2 over water at 22.0°C and an atmospheric pressure of 752.0 mmHg. If the vapor pressure of water at 22.0°C is 19.8 mmHg, calculate the partial pressure of dry CO_2 in atmospheres.

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9.3: Molar Volume of a Gas - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 9.3.1: Experimental Measurements

Experimental Measurement	Trial 1	Trial 2
1. Mass of sodium bicarbonate (g)	_____	_____
2. Volume of displaced water (mL)	_____	_____
3. Temperature of water / gas (°C)	_____	_____
4. Barometric pressure (atm or mmHg)	_____	_____

Table 9.3.2: Calculations and Molar Volume Determination

Calculated Quantity	Calculation Formula / Guide	Trial 1	Trial 2
Moles of NaHCO ₃ reacted (mol)	$\text{g NaHCO}_3 \div 84.01 \text{ g/mol}$	_____	_____
Theoretical moles of CO ₂ produced (mol)	$\text{mol NaHCO}_3 \times \left(\frac{1 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} \right)$	_____	_____
Volume of CO ₂ produced at room conditions (L)	$\text{mL water} \div 1000$	_____	_____
Molar volume of CO ₂ at room conditions (L/mol)	$V_{\text{m,lab}} = \frac{\text{L CO}_2}{\text{mol CO}_2}$	_____	_____
Vapor pressure of H ₂ O at room temperature (atm)	From graph $\div 760 \text{ mmHg/atm}$	_____	_____
Partial pressure of dry CO ₂ at room conditions (atm)	$P_{\text{CO}_2} = P_{\text{atm}} - P_{\text{H}_2\text{O}}$	_____	_____
Volume of dry CO ₂ converted to STP (L)	$V_{\text{STP}} = V_{\text{lab}} \times \left(\frac{P_{\text{CO}_2}}{1.00 \text{ atm}} \right) \times \left(\frac{273.15 \text{ K}}{T_{\text{lab}}} \right)$	_____	_____
Experimental molar volume of CO ₂ at STP (L/mol)	$V_{\text{m,STP}} = \frac{V_{\text{STP}}}{\text{mol CO}_2}$	_____	_____

Average experimental molar volume of CO₂ at STP: _____ L/mol

Percent error in molar volume at STP (accepted value = 22.414 L/mol):

$$\% \text{ Error} = \left| \frac{\text{Average Experimental } V_{\text{m,STP}} - 22.414 \text{ L/mol}}{22.414 \text{ L/mol}} \right| \times 100\% = \quad \% \quad (9.3.1)$$

CALCULATION WORK SPACE

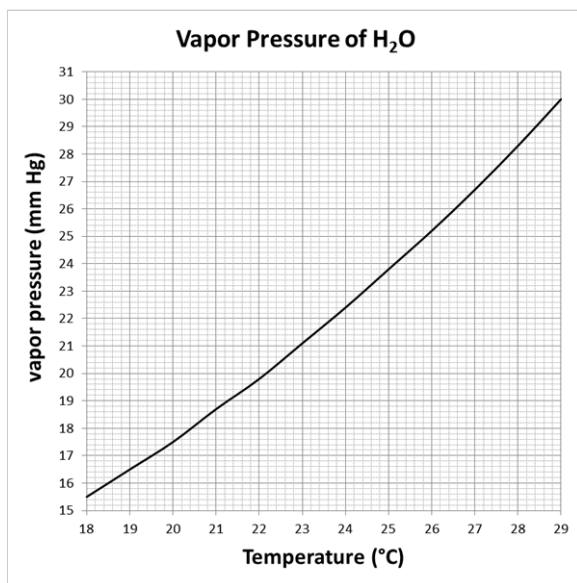


Figure 9.3.1: Vapor Pressure of Water vs. Temperature Reference Chart.

POST-LAB QUESTIONS

1. Identify two distinct potential sources of experimental error in this setup and explain specifically how each error would affect your calculated molar volume at STP (whether it would cause the calculated value to be falsely high or low).
-

2. If the experiment were performed on a day with a significantly higher room temperature, explain how this would affect the volume of water displaced, assuming the mass of NaHCO₃ remains identical. Justify your response using gas laws.
-

3. In addition to water displacement, describe two alternative laboratory methods for measuring the volume of gas generated in a chemical reaction. Outline one major advantage and one disadvantage of each method.

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10: ENTHALPY OF NEUTRALIZATION

PURPOSE

- To use coffee-cup calorimetry to determine the enthalpy of neutralization of two acids, hydrochloric acid and acetic acid, with sodium hydroxide.

INTRODUCTION

The **enthalpy of neutralization** (ΔH_{neut}) is the heat released or absorbed when an aqueous acid and base react to form one mole of water. In this experiment, the enthalpy of neutralization for two distinct acid-base reactions will be measured using a **coffee-cup calorimeter**, a constant-pressure device designed to insulate aqueous chemical reactions from ambient thermal loss.

The first system involves a **strong acid** (HCl) and a **strong base** (NaOH), both of which ionize completely in aqueous solution. The second system involves a **weak acid** ($\text{CH}_3\text{CO}_2\text{H}$) and a **strong base** (NaOH). Because weak acids exist primarily as unionized molecules in solution, a portion of the neutralization energy is consumed in breaking the covalent O – H bond to release protons. By measuring precise temperature changes during these reactions, the net heat released (q_{rxn}) can be calculated and normalized per mole of water produced.

Additionally, the heat capacity of the calorimeter apparatus ($C_{\text{calorimeter}}$) will be determined by mixing warm and cold water. Accounting for thermal energy absorbed by the polystyrene container and temperature probe ensures greater quantitative accuracy when evaluating the fundamental differences between strong and weak acid thermodynamics.

[10.1: Enthalpy of Neutralization - Experiment](#)

[10.2: Enthalpy of Neutralization - Pre-lab](#)

[10.3: Enthalpy of Neutralization - Data and Report](#)

10.1: Enthalpy of Neutralization - Experiment

SAFETY PRECAUTIONS

- Sodium hydroxide (NaOH) and hydrochloric acid (HCl) solutions are **corrosive** and cause severe skin burns or eye damage. Handle with extreme care. Flush spills immediately with running water.
- Glacial and concentrated acetic acid ($\text{CH}_3\text{CO}_2\text{H}$) are chemical irritants. Avoid skin contact and inhalation of vapors.
- Perform all measurements in a well-ventilated laboratory area. Wear safety goggles and a lab coat at all times.

EQUIPMENT AND CHEMICALS NEEDED

Table 10.1.1: Materials List

Equipment & Hardware	Chemicals & Reagents
Coffee-cup calorimeter with lid or stopper*	1.0 M Sodium hydroxide (NaOH) solution
Two stainless steel temperature probes or thermometers	1.0 M Hydrochloric acid (HCl) solution
50-mL & 100-mL beakers	1.0 M Acetic acid ($\text{CH}_3\text{CO}_2\text{H}$) solution
10-mL graduated cylinders	Deionized water
Hot plate & LabQuest interface	

*Note: Volumes used in this protocol are calibrated for small-scale Styrofoam cup calorimeters. Single-hole #6 rubber stoppers serve as insulating lids.

EXPERIMENTAL PROCEDURE

VERNIER TECH BOX: TEMPERATURE GRAPHING

1. Connect & Setup

- Connect the primary Stainless Steel Temperature Probe to **CH 1** on the LabQuest interface.
- Connect the secondary Stainless Steel Temperature Probe to **CH 2** on the LabQuest interface.
- Tap the **Meter** tab to display live digital temperature readings.

2. Configure Data Collection Mode

- Tap the **Mode** configuration box.
- Select **Time Based** collection mode.
- Set **Duration** to 300 seconds (5.0 minutes) and **Rate** to 2.0 samples/second.
- Select **OK** to save settings.

Part A: Determine the Heat Capacity of the Calorimeter

- Measure exactly 5.0 mL of room-temperature tap water into the clean calorimeter cup. Insert Temperature Probe 1 through the lid stopper into the water.
- Allow the reading to stabilize for 60 seconds and record the cold water temperature (T_{cold}) on your report sheet.

3. Measure 5.0 mL of tap water into a small beaker and heat on a hot plate until it is approximately 15–20 °C above room temperature (do not boil). Remove from the heat source and let stand for 1 minute.
4. Insert Temperature Probe 2 into the warm water, allow the reading to stabilize, and record the warm water temperature (T_{warm}).
5. Start LabQuest data acquisition to record a 10-second baseline before mixing.
6. Quickly remove the cover, pour the warm water completely into the calorimeter, and immediately re-cover the cup.
7. Swirl the calorimeter gently in a continuous circular pattern while monitoring live temperature data. Record the maximum stabilized final temperature (T_{final}).

Part B: Enthalpy of Neutralization of NaOH and HCl

1. Rinse and thoroughly dry the inner cup of the calorimeter.
2. Transfer 5.0 mL of 1.0 M NaOH into the calorimeter. Measure 5.0 mL of 1.0 M HCl into a clean, dry beaker and let stand for 3 minutes. Verify that both initial solution temperatures match within 0.5 °C.
3. Record the initial temperature (T_{initial}) and start data acquisition for a baseline.
4. Add the HCl solution to the calorimeter and replace the cover immediately.
5. Swirl gently and record the maximum peak temperature reached (T_{max}).

Part C: Enthalpy of Neutralization of NaOH and CH₃CO₂H

1. Rinse and dry the calorimeter completely.
2. Repeat the identical procedure from Part B, substituting 5.0 mL of 1.0 M CH₃CO₂H (acetic acid) for the hydrochloric acid solution.

CHEMICAL DISPOSAL

- Neutralized reaction mixtures may be poured down the drain with copious flowing water.
- Rinse all glassware and temperature probes thoroughly with deionized water before storing.

10.2: Enthalpy of Neutralization - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. State the First Law of Thermodynamics and explain the physical significance of the equation $q_{\text{system}} + q_{\text{surroundings}} = 0$ as applied to coffee-cup calorimetry.

2. A student combines 25.0 mL of 1.0 M NaOH with 25.0 mL of 1.0 M HCl in a calorimeter. Write the balanced molecular equation and identify the limiting reactant.

3. If the net heat released during the reaction in Question 2 is calculated as -1295 J , calculate the enthalpy of neutralization in kJ/mol of water formed.

4. Consider the theoretical enthalpy change for strong acid-base neutralization:

- a. Write the net ionic equation for the neutralization of $\text{NaOH}(aq)$ and $\text{HCl}(aq)$.

- b. Calculate the theoretical standard enthalpy of neutralization ($\Delta H_{\text{neut}}^{\circ}$) in kJ/mol using Table 10.2.1 standard enthalpies of formation data.

Table 10.2.1: Neutralization Thermochemical Data

Thermodynamic Quantity	$\text{H}^+(aq)$	$\text{OH}^-(aq)$	$\text{H}_2\text{O}(l)$
$\Delta H_f^{\circ} (\text{kJ/mol})$	0.0	-230.0	-285.83

Show complete $\Delta H_{\text{neut}}^{\circ}$ calculation setup:

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10.3: Enthalpy of Neutralization - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Heat Capacity of the Calorimeter

Table 10.3.1: Part A Experimental Data and Calculations

Measurement / Calculation Parameter	Result / Value	Units
1. Temperature of cold water / calorimeter before mixing (T_{cold})	_____	°C
2. Temperature of warm water before mixing (T_{warm})	_____	°C
3. Maximum final temperature of mixture (T_{final})	_____	°C
4. Temperature change of warm water ($\Delta T_{\text{warm}} = T_{\text{warm}} - T_{\text{final}}$)	_____	°C
5. Temperature change of cold water / calorimeter ($\Delta T_{\text{cold}} = T_{\text{final}} - T_{\text{cold}}$)	_____	°C
6. Heat lost by warm water ($q_{\text{warm}} = \Delta T_{\text{warm}} \times 5.0 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}$)	_____	J
7. Heat gained by cold water ($q_{\text{cold}} = \Delta T_{\text{cold}} \times 5.0 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}$)	_____	J
8. Heat gained by calorimeter cup ($q_{\text{calorimeter}} = q_{\text{warm}} - q_{\text{cold}}$)	_____	J
9. Heat capacity of calorimeter ($C_{\text{calorimeter}} = q_{\text{calorimeter}} / \Delta T_{\text{cold}}$)	_____	J/°C

Part B: Enthalpy of Neutralization of HCl and NaOH

Table 10.3.2: Part B Experimental Thermal Data

Measurement / Thermal Parameter	Result / Value	Units
1. Initial temperature of solutions before mixing (T_{initial})	_____	°C
2. Maximum temperature reached during reaction (T_{max})	_____	°C
3. Temperature change ($\Delta T = T_{\text{max}} - T_{\text{initial}}$)	_____	°C
4. Heat gained by solution ($q_{\text{solution}} = \Delta T \times 10.0 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}$)	_____	J
5. Heat gained by calorimeter ($q_{\text{calorimeter}} = \Delta T \times C_{\text{calorimeter}}$)	_____	J

Measurement / Thermal Parameter	Result / Value	Units
6. Total heat released ($q_{\text{rxn}} = q_{\text{solution}} + q_{\text{calorimeter}}$)	_____	J

Table 10.3.3: Part B Molar Enthalpy Calculations

Calculated Quantity	Calculation Setup	Result / Units
1. Balanced equation for reaction	_____	
2. Moles of HCl reacted	_____	_____ mol
3. Moles of NaOH reacted	_____	_____ mol
4. Moles of H ₂ O produced (n_{water})	_____	_____ mol
5. Enthalpy of neutralization (ΔH_{neut})	_____	_____ kJ/mol

Part C: Enthalpy of Neutralization of CH₃CO₂H and NaOH

Table 10.3.4: Part C Experimental Thermal Data

Measurement / Thermal Parameter	Result / Value	Units
1. Initial temperature of solutions before mixing (T_{initial})	_____	°C
2. Maximum temperature reached during reaction (T_{max})	_____	°C
3. Temperature change ($\Delta T = T_{\text{max}} - T_{\text{initial}}$)	_____	°C
4. Heat gained by solution ($q_{\text{solution}} = \Delta T \times 10.0 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}$)	_____	J
5. Heat gained by calorimeter ($q_{\text{calorimeter}} = \Delta T \times C_{\text{calorimeter}}$)	_____	J
6. Total heat released ($q_{\text{rxn}} = q_{\text{solution}} + q_{\text{calorimeter}}$)	_____	J

Table 10.3.5: Part C Molar Enthalpy Calculations

Calculated Quantity	Calculation Setup	Result / Units
1. Balanced equation for reaction	_____	
2. Moles of CH ₃ CO ₂ H reacted	_____	_____ mol
3. Moles of NaOH reacted	_____	_____ mol
4. Moles of H ₂ O produced (n_{water})	_____	_____ mol
5. Enthalpy of neutralization (ΔH_{neut})	_____	_____ kJ/mol

POST-LAB QUESTIONS

1. An unknown metal sample weighing 9.25 g is heated to 97.8 °C and placed into 50.0 g of water at 24.5 °C inside a calorimeter. The mixture reaches a maximum temperature of 27.4 °C. Calculate the specific heat capacity of the metal (c_{metal}) in J/g · °C.

2. A solution is prepared by dissolving 2.50 g of solid NaOH into 75.0 mL of water (both initially at 25.0 °C). The temperature rises to 33.5 °C. Is this dissolution process exothermic or endothermic? Calculate the molar enthalpy of solution ($\Delta H_{\text{solution}}$) of NaOH in kJ/mol.

3. Compare your experimental enthalpy of neutralization for HCl–NaOH (Part B) to the theoretical value calculated in Pre-Lab Question 4b. Discuss potential sources of error that account for any discrepancy.

4. Compare the experimental enthalpy of neutralization obtained for the strong acid (Part B) to that of the weak acid (Part C). Explain thermodynamically why the magnitude of ΔH_{neut} for acetic acid differs from that of hydrochloric acid.

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11: EVAPORATION AND INTERMOLECULAR ATTRACTIONS

PURPOSE

- To investigate the relationship between intermolecular forces and the rate of evaporation by measuring temperature changes.
- To compare the relative strengths of London dispersion forces and hydrogen bonding in alkanes and alcohols based on molecular structure and evaporation behavior.
- To predict and analyze the temperature change (ΔT) of different organic liquids based on their intermolecular forces and molar mass.

INTRODUCTION

Evaporation is an endothermic phase transition in which molecules at the surface of a liquid gain sufficient thermal energy to overcome attractive intermolecular forces and escape into the gas phase. Because the highest-energy molecules evaporate first, the average kinetic energy of the remaining liquid decreases, resulting in a measurable drop in temperature. The magnitude of this temperature drop is directly related to the rate of evaporation and the strength of the intermolecular forces holding the liquid molecules together.

Intermolecular forces dictate fundamental physical properties of substances, including boiling point, vapor pressure, viscosity, and surface tension. The primary intermolecular forces in pure organic liquids are London dispersion forces, dipole-dipole interactions, and hydrogen bonding. London dispersion forces exist between all molecules and increase in strength with increasing polarizability and molar mass. Dipole-dipole forces occur between polar molecules with permanent dipole moments. Hydrogen bonding—a particularly strong form of dipole interaction—occurs when hydrogen is directly covalently bonded to a highly electronegative atom (N, O, or F).

In this experiment, dual electronic temperature probes will be used to record temperature changes during the evaporation of several organic liquids, including straight-chain alkanes and aliphatic primary alcohols. Alkanes consist entirely of nonpolar carbon-carbon and carbon-hydrogen bonds and exhibit only London dispersion forces. In contrast, alcohols contain a polar hydroxyl ($-\text{OH}$) group capable of extensive hydrogen bonding in addition to dispersion forces. By comparing experimental ΔT values across these structural series, you will evaluate the relative strengths of intermolecular forces and assess the impact of molar mass and functional group structure on evaporation kinetics.

[11.1: Evaporation and Intermolecular Attractions - Experiment](#)

[11.2: Evaporation and Intermolecular Attractions - Pre-lab](#)

[11.3: Evaporation and Intermolecular Attractions - Data and Report](#)

11.1: Evaporation and Intermolecular Attractions - Experiment

SAFETY PRECAUTIONS

- Perform all experimental steps in an operating fume hood or well-ventilated laboratory area to minimize inhalation of organic vapors.
- All organic solvents used in this lab (alkanes and alcohols) are volatile and highly flammable. Keep all containers tightly closed and away from heat sources, sparks, or open flames.
- Avoid skin and eye contact. Wear safety goggles, a lab coat, and protective gloves throughout the procedure. Wash hands thoroughly before leaving the laboratory.

EQUIPMENT AND CHEMICALS NEEDED

Table 11.1.1: Materials List

Equipment & Hardware	Chemicals & Reagents
Vernier LabQuest interface	Methanol (methyl alcohol)
2 Stainless Steel Temperature Probes	Ethanol (ethyl alcohol)
6 pieces of filter paper (2.5 cm × 2.5 cm)	1-Propanol
Small rubber bands & masking tape	1-Butanol
Small test tubes & test tube rack	<i>n</i> -Pentane
Scissors & ruler	<i>n</i> -Hexane

EXPERIMENTAL PROCEDURE

Part A: Measure Ethanol and 1-Propanol

VERNIER TECH BOX: DUAL PROBE LOGGING

1. Connect Sensors

- Connect Probe 1 to **CH 1** and Probe 2 to **CH 2** on the LabQuest interface.
- Verify that the screen displays two live temperature channels.

2. Configure Data Collection

- Tap the **Mode** configuration box.
- Set **Duration** to 300 seconds (5 minutes) and **Rate** to 1 sample/second.
- Tap **OK** to save settings.

3. Data Analysis (ΔT Determination)

- After data collection stops, tap **Analyze** → **Statistics**.
- Check the boxes for both **Temperature 1** and **Temperature 2**.
- Record $T_{\max}$ and $T_{\min}$ for each probe, and calculate: $\Delta T = T_{\max} - T_{\min}$.

1. Wrap a piece of filter paper tightly around the tip of each temperature probe, ensuring it sits flush with the bottom. Secure the filter paper using a small rubber band.
2. Transfer approximately 2 mL of ethanol into one small test tube and 2 mL of 1-propanol into a second test tube (enough to fully submerge the filter paper).
3. Place Probe 1 into the ethanol tube and Probe 2 into the 1-propanol tube. Allow probes to soak in the liquids for at least 30 seconds to reach thermal equilibrium.
4. Begin data collection. Monitor live readings for 15 seconds to establish stable initial temperatures ($T_{\max}$).
5. Simultaneously remove both probes from the liquid test tubes and secure them with masking tape to the edge of the lab bench so the paper-wrapped tips extend 5 cm beyond the bench top.
6. Allow data collection to proceed for 3 minutes as evaporation occurs.
7. When both temperature curves reach a minimum plateau and begin to rise back toward room temperature, stop data collection.
8. Use the Statistics feature to determine $T_{\max}$ and $T_{\min}$ for each probe. Calculate $\Delta T = T_{\max} - T_{\min}$ and record values in the Data Table.
9. Remove and discard the used filter paper and rubber bands.

Part B: Measure 1-Butanol and *n*-Pentane

10. Based on your experimental ΔT values for ethanol and 1-propanol, predict the relative ΔT values for 1-butanol and *n*-pentane considering their molar masses and hydrogen bonding capabilities. Record your predictions in the report sheet.
11. Repeat Steps 1–9 using 1-butanol (Probe 1) and *n*-pentane (Probe 2).

Part C: Measure Methanol and *n*-Hexane

12. Based on all previous data, predict the ΔT values for methanol and *n*-hexane. Record your predictions in the report sheet.
13. Repeat Steps 1–9 using methanol (Probe 1) and *n*-hexane (Probe 2).

CHEMICAL DISPOSAL

- Dispose of used filter paper and rubber bands in the designated solid waste container.
- Dispose of residual liquid solvents into the designated **Non-Halogenated Organic Liquid Waste** container.

11.2: Evaporation and Intermolecular Attractions - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

Complete the Pre-Lab table below by analyzing the molecular structure and properties of each organic compound. Draw full Lewis structural formulas, calculate molar masses, and indicate whether each liquid is capable of intermolecular hydrogen bonding.

Table 11.2.1: Physical Properties and Structural Parameters of Test Compounds

Substance	Formula	Lewis Structural Formula	Molar Mass (g/mol)	Hydrogen Bonding? (Yes/No)
Ethanol	C ₂ H ₅ OH		_____	_____
1-Propanol	C ₃ H ₇ OH		_____	_____
1-Butanol	C ₄ H ₉ OH		_____	_____
<i>n</i> -Pentane	C ₅ H ₁₂		_____	_____
Methanol	CH ₃ OH		_____	_____
<i>n</i> -Hexane	C ₆ H ₁₄		_____	_____

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11.3: Evaporation and Intermolecular Attractions - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 11.3.1: Evaporation Temperature Data and Predictions

Substance	$T_{\max}$ (°C)	$T_{\min}$ (°C)	$\Delta T = T_{\max} - T_{\min}$ (°C)	Predicted ΔT (°C)	Brief Explanation
Ethanol	_____	_____	_____	N/A (Measured)	_____
1-Propanol	_____	_____	_____	N/A (Measured)	_____
1-Butanol	_____	_____	_____	_____	_____
<i>n</i> -Pentane	_____	_____	_____	_____	_____
Methanol	_____	_____	_____	_____	_____
<i>n</i> -Hexane	_____	_____	_____	_____	_____

POST-LAB QUESTIONS

- Although *n*-pentane (72.15 g/mol) and 1-butanol (74.12 g/mol) have nearly identical molar masses, their experimental temperature changes (ΔT) differ dramatically. Explain this difference based on the types and relative strengths of their intermolecular forces.

- Among the four alcohols tested (methanol, ethanol, 1-propanol, 1-butanol), which one exhibits the strongest overall intermolecular forces, and which exhibits the weakest? Explain this trend based on molar mass and London dispersion forces, supporting your answer with experimental ΔT data.

- Between the two nonpolar alkanes (*n*-pentane and *n*-hexane), which compound possesses stronger intermolecular forces? Justify your choice using molecular structure, polarizability, and experimental ΔT values.

4. Construct a plot of ΔT vs. molar mass for the series of four primary alcohols. Describe the observed trend and explain how it relates to molecular structure and evaporation kinetics.
-

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12: MOLAR MASS DETERMINATION USING FREEZING POINT DEPRESSION

PURPOSE

- Demonstrate the colligative property of freezing-point depression.
- Measure the change in freezing point when a known solute is dissolved in a solvent.
- Determine the molar mass of an unknown or known solute experimentally.

INTRODUCTION

When a nonvolatile solute is dissolved in a solvent, the freezing point of the resulting solution is lower than that of the pure solvent. This physical phenomenon, known as **freezing-point depression**, is a colligative property. Colligative properties depend exclusively on the total concentration of dissolved solute particles in a solution, rather than the chemical identity of the solute particles themselves.

The quantitative relationship between freezing-point depression and solution concentration is expressed by the equation:

$$\Delta T_f = K_f \times m \quad (12.1)$$

Where:

- ΔT_f is the freezing-point depression, defined as the positive difference between the freezing point of the pure solvent ($T_{f,\text{solvent}}$) and that of the solution ($T_{f,\text{solution}}$): $\Delta T_f = T_{f,\text{solvent}} - T_{f,\text{solution}}$.
- K_f is the molal freezing-point-depression constant, a property unique to each solvent. In this experiment, lauric acid serves as the solvent, with an accepted $K_f = 3.9^\circ\text{C} \cdot \text{kg/mol}$.
- m is the molality of the solution, defined as moles of solute per kilogram of solvent (mol solute/kg solvent).

This experiment is divided into two parts:

- Determining the Freezing Point of Pure Lauric Acid ($\text{CH}_3(\text{CH}_2)_{10}\text{COOH}$):** The freezing point of the pure solvent is measured experimentally by constructing and analyzing its cooling curve.
- Determining the Freezing-Point Depression and Molar Mass of Benzoic Acid ($\text{C}_6\text{H}_5\text{COOH}$):** A precisely weighed mass of benzoic acid is dissolved in a known mass of lauric acid. By comparing the freezing points of the pure solvent and the solution, ΔT_f is calculated. Using ΔT_f , K_f , and the mass of the solvent, the experimental molar mass of the solute is determined.

[12.1: Molar Mass Determination Using Freezing Point Depression - Experiment](#)

[12.2: Molar Mass Determination Using Freezing Point Depression - Pre-lab](#)

[12.3: Molar Mass Determination Using Freezing Point Depression - Data and Report](#)

12.1: Molar Mass Determination Using Freezing Point Depression - Experiment

SAFETY PRECAUTIONS

- Always wear chemical splash goggles and a lab coat in the laboratory.
- Exercise extreme caution when handling hot equipment, hot plates, and test tubes containing molten organic compounds.
- Avoid direct skin contact with hot lauric acid or benzoic acid, which can cause thermal burns or mild skin irritation.
- Handle all glass test tubes and temperature probes with care to prevent breakage.

EQUIPMENT AND CHEMICALS NEEDED

Table 12.1.1: Materials List

Hardware & Equipment	Hardware & Equipment	Chemicals & Solvents
Vernier LabQuest interface	Ring stand & utility clamp	Lauric acid (solvent)
Stainless Steel Temperature Probe	Large test tubes (25 mm × 150 mm)	Benzoic acid (solute)
400-mL beaker (water bath)	Paper towels & wire stirrer	Tap water (cooling bath)
Hot plate	Electronic balance (± 0.001 g)	

EXPERIMENTAL PROCEDURE

Part A: Set Up and Configure Data Collection

VERNIER TECH BOX: COOLING CURVE SETUP

1. Configure Collection Parameters

- Connect the Stainless Steel Temperature Probe to **CH 1** on the LabQuest interface.
- Tap **Mode** on the screen.
- Set **Duration** to **1200 s** (20 minutes) and **Rate** to 1 sample/second.
- Tap **OK**.

2. Adjust Graph Display Axes

- Tap the Y-axis (Temperature) scale values.
- Set **Top** to 100 °C and **Bottom** to 0 °C.
- This scaling ensures the full phase-change curve is visible without manual scrolling.

Part B: Determine the Freezing Point of Pure Lauric Acid

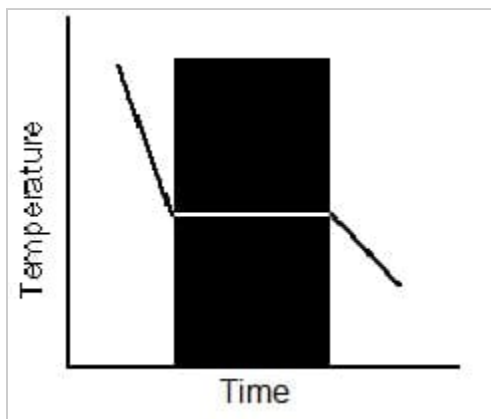


Figure 12.1.1: Cooling curve for pure lauric acid featuring a constant-temperature freezing plateau.

1. Fill a 400-mL beaker with approximately 300 mL of room-temperature tap water (20–25 °C) and place it on a ring stand setup.
2. Obtain a test tube containing melted pure lauric acid from the instructor's hot water bath and clamp it securely using a utility clamp.
3. Insert the temperature probe into the liquid lauric acid and allow it to equilibrate for 45 seconds. Keep the test tube suspended above the water bath.
4. Start data collection on the LabQuest interface.
5. Lower the test tube into the water bath so the water level is higher than the lauric acid level in the tube. If the initial lauric acid temperature drops below 50 °C before immersion, return it to the hot bath to remelt.
6. Gently move the temperature probe up and down in a continuous motion to stir the liquid as it cools. Continue collecting data for 20 minutes.
7. After data collection completes, raise the test tube out of the water bath and remove the beaker of water (reserve the water for Part C).
8. Return the test tube to the hot water bath to melt the solid lauric acid. Do not force the probe out while solid. Once melted, carefully wipe the probe with a paper towel.
9. Analyze the plateau region on the cooling curve to locate the precise freezing point.

VERNIER TECH BOX: FINDING THE FREEZING POINT

Pure Solvent (Plateau Method)

- Identify the flat, horizontal plateau where temperature remains constant during phase change.
- Drag to highlight **only** the flat plateau region on the graph.
- Tap **Analyze** → **Statistics**.
- Record the **Mean** temperature as the freezing point ($T_{f,\text{solvent}}$).

10. Record the mean plateau temperature as the freezing point of pure lauric acid on your report sheet. Save the LabQuest data file.

Part C: Freezing Point of a Benzoic-Acid/Lauric-Acid Solution

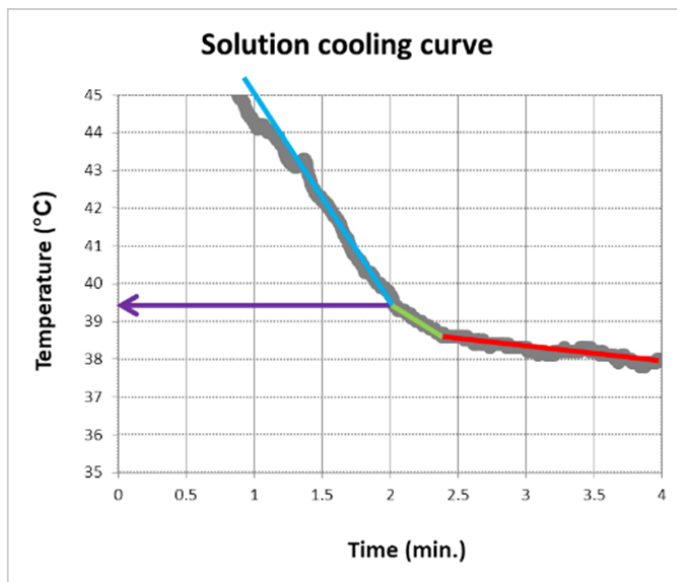


Figure 12.1.2: Solution cooling curve showing the break point marking the onset of freezing.

11. Obtain a test tube containing a melted mixture of approximately 1.0 g benzoic acid in 8.0 g lauric acid.
12. Record the **exact masses** of lauric acid and benzoic acid from the container label or instructor's data log.
13. Repeat Steps 2–8 from Part B to record the cooling curve for the solution.
14. Identify the freezing point from the solution cooling curve.

VERNIER TECH BOX: FINDING THE FREEZING POINT

Solution (Break Point Method)

- Solutions do *not* exhibit a flat plateau; the cooling curve continues sloping downward as concentration increases during freezing.
- Locate the **Break Point**: the sharp change in slope where rapid liquid cooling transitions into a shallower freezing line.
- Tap that exact break point on the curve to read the freezing temperature ($T_{f,\text{solution}}$).

15. Record the break-point temperature as the freezing point of the solution ($T_{f,\text{solution}}$).

CHEMICAL DISPOSAL

- Return all test tubes containing lauric acid and lauric-acid/benzoic-acid mixtures to the designated collection rack in the hot water bath. Do not pour organic chemicals down the drain.
- Dispose of water bath water into the laboratory sink. Clean all equipment before storing.

12.2: Molar Mass Determination Using Freezing Point Depression - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Define colligative properties and list three specific examples of colligative properties other than freezing-point depression.

2. Explain why the addition of a nonvolatile solute causes the freezing point of a solution to be lower than that of the pure solvent. Describe this relationship in terms of chemical potential and particle concentration.

3. What is the molal freezing-point-depression constant (K_f)? What are its SI units, and what physical factor determines its numerical value?

4. Describe how a cooling curve is constructed. How do the cooling curves of a pure solvent and a solution differ, and why does a solution lack a flat horizontal plateau?

5. A student dissolves 0.850 g of an unknown nonvolatile nonelectrolyte in 10.00 g of lauric acid ($K_f = 3.9^\circ\text{C} \cdot \text{kg/mol}$). The freezing point of pure lauric acid is 43.80°C and the freezing point of the solution is 39.15°C . Show the complete step-by-step calculation to determine the molar mass of the unknown solute.

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12.3: Molar Mass Determination Using Freezing Point Depression - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Freezing Point of Pure Lauric Acid

Freezing point of pure lauric acid ($T_{f,\text{solvent}}$): _____ °C

Part B: Freezing Point of Solution and Mass Data

Table 12.3.1: Experimental Measurements

Experimental Measurement	Recorded Value & Units
1. Mass of lauric acid solvent (g)	_____ g
2. Mass of lauric acid solvent (kg)	_____ kg
3. Mass of benzoic acid solute (g)	_____ g
4. Freezing point of benzoic-acid/lauric-acid solution ($T_{f,\text{solution}}$)	_____ °C

CALCULATIONS

Show complete step-by-step calculation setups with units for full credit.

1. Calculate Freezing-Point Depression (ΔT_f):

$$\Delta T_f = T_{f,\text{solvent}} - T_{f,\text{solution}} \quad (12.3.1)$$

2. Calculate Molality (m) of the Solution:

$$m = \frac{\Delta T_f}{K_f} \quad (K_f = 3.9^\circ\text{C} \cdot \text{kg/mol}) \quad (12.3.2)$$

3. Calculate Moles of Benzoic Acid Solute:

$$\text{Moles of solute} = m \times \text{mass of lauric acid (kg)} \quad (12.3.3)$$

4. Calculate Experimental Molar Mass of Benzoic Acid:

$$\text{Experimental Molar Mass} = \frac{\text{Mass of benzoic acid (g)}}{\text{Moles of benzoic acid (mol)}} \quad (12.3.4)$$

5. Calculate Accepted Molar Mass of Benzoic Acid ($\text{C}_6\text{H}_5\text{COOH}$):

6. Calculate Percent Error:

$$\% \text{ Error} = \left| \frac{\text{Experimental Molar Mass} - \text{Accepted Molar Mass}}{\text{Accepted Molar Mass}} \right| \times 100\% \quad (12.3.5)$$

Table 12.3.2: Summary of Results

Calculated Quantity	Final Result
Freezing-point depression (ΔT_f)	_____ °C
Molality of solution (m)	_____ mol/kg
Moles of benzoic acid solute	_____ mol
Experimental molar mass of benzoic acid	_____ g/mol
Accepted molar mass of benzoic acid	_____ g/mol
Percent error	_____ %

POST-LAB QUESTIONS

1. Identify two specific experimental sources of error in this procedure and explain how each would affect your calculated molar

mass (whether it would cause the calculated value to be falsely high or low).

2. Explain why the plateau region of the pure lauric acid cooling curve represents its freezing point, whereas the solution cooling curve lacks a flat horizontal plateau.
- _____

3. How would selecting a different solvent with a larger molal freezing-point-depression constant (K_f) affect the precision of the experimental molar mass determination?
- _____

4. Explain on a molecular level how solute-solvent intermolecular interactions lower the freezing point of a solution compared to the pure solvent.
- _____

5. Describe a practical commercial or industrial application of freezing-point depression (e.g., de-icing roads or automotive anti-freeze) and explain the chemical principles involved.
- _____

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13: APPENDIX

13.1: Vernier LabQuest 3 Instructions

13.2: Periodic Table of the Elements

13.3: Units and Conversions

13.4: Activity Series

13.1: Vernier LabQuest 3 Instructions

This appendix contains step-by-step instructions for the common tasks used in General Chemistry II. Refer to these sections whenever a lab procedure directs you to "Calibrate" or "Setup" a sensor.

A. Temperature Probe

Measuring Temperature

The Stainless Steel Temperature Probe is auto-identified by the LabQuest.

1. Plug the probe into **CH 1**.
 2. The screen will display the live temperature in °C.
 3. **Note:** You typically do *not* need to press "Start/Collect" to read a temperature. Just use the live meter readout.
 4. **Calibration:** These probes are factory-calibrated and rarely need adjustment. If the reading seems wrong (e.g., room temperature is 80°C), check that the probe is not set to °F or K in the units menu.
-

B. pH Sensor

Two-Point pH Calibration

pH probes drift over time. For accurate K_a or titration data, you must calibrate before use.

1. Connect the pH Sensor to **CH 1**.
 2. Tap the red **pH** box on the meter screen → Select **Calibrate**.
 3. Tap **Calibrate Now**.
 4. **First Point (Acidic):**
 - Rinse the probe with distilled water and place it in the **pH 4.00** buffer solution.
 - Wait for the voltage reading to stabilize.
 - Enter **4.00** in the "Value 1" field.
 - Tap **Keep**.
 5. **Second Point (Neutral):**
 - Rinse the probe and place it in the **pH 7.00** buffer solution.
 - Wait for the voltage to stabilize.
 - Enter **7.00** in the "Value 2" field.
 - Tap **Keep**.
 6. Tap **OK**. The probe is now calibrated.
-

C. Drop Counter (Automated Titration)

Calibrating Drops/mL

The "Drop Counter" counts drops, but you need to know Volume (mL). Because every plastic reagent reservoir tip is slightly different, you must tell the device exactly how large your drops are.

1. Connect the Drop Counter to **DIG 1**.

2. Tap the **Mode** box → Select **Drop Counting**.
3. Tap the red **Drop Count** box → Select **Calibrate** → **Calibrate Now**.
4. **The Procedure:**
 - Place a clean **10-mL graduated cylinder** directly under the drop tip.
 - Open the bottom valve fully to establish a steady, slow drop rate (about 1-2 drops per second).
 - Allow the drops to fall until the cylinder is nearly full (between 9 and 10 mL).
 - **Close the valve completely.**
5. **The Calculation:**
 - Read the exact volume in the graduated cylinder (e.g., 9.45 mL).
 - Enter this value into the LabQuest.
 - Tap **Keep**.
6. The device now calculates a precise "Drops/mL" ratio for your specific setup.

D. Manual Data Entry

"Events with Entry" Mode

Use this mode when you want to type in one variable (like Concentration or Volume) while the sensor records the other (like Absorbance or pH).

1. Tap the **Mode** box (default is usually "Time Based").
2. Select **Events with Entry** from the dropdown menu.
3. **Configure the X-Axis:**
 - **Name:** Enter the variable you will type (e.g., "Concentration").
 - **Units:** Enter the units (e.g., "mol/L" or "M").
4. Tap **OK**.
5. **How to Collect Data:**
 - Tap **Start/Collect**.
 - When the sensor reading stabilizes, tap **Keep**.
 - A keyboard will appear. Type in the value for your X-axis (e.g., "0.05") and tap **OK**.
 - Repeat for all samples. Tap **Stop** when finished.

E. Spectrometer (SpectroVis Plus)

Calibration & Setup

Step 1: Calibration (Blanking)

1. Allow the lamp to warm up for 2-3 minutes.
2. Fill a cuvette 3/4 full with your blank solvent (usually water).
3. Insert the cuvette so the clear sides face the arrow/light path.
4. Tap the red **Absorbance** box → **Calibrate**.
5. Tap **Finish Calibration** → **OK**. (The reading should be 0.000).

Step 2: Finding $\lambda_{\max}$ (Peak Wavelength)

1. Insert a cuvette containing your most concentrated colored solution.
2. Tap **Start/Collect**. A spectrum graph will appear.
3. Tap **Stop**.
4. Tap the highest point on the curve. The wavelength (nm) and absorbance will appear on the right. Note this wavelength.

Step 3: Setting Up for Beer's Law

1. Tap the **Meter** tab (gauge icon).
2. Tap the **Mode** box.
3. Select **Events with Entry** (See Appendix A.5).
4. The device will now lock onto the $\lambda_{\max}$ you found in Step 2.

F. Linear Fit (Regression)

Finding Slope and Intercept

To find the best-fit line for your data (e.g., Beer's Law or Kinetics):

1. Tap the **Graph** tab.
2. Tap **Analyze** → **Curve Fit**.
3. Select the checkbox for your data series (usually "Absorbance").
4. From the "Fit Equation" dropdown, select **Linear**.
5. The screen will display the equation form $y = mx + b$:
 - **m (Slope)**: The rate of change or molar absorptivity.
 - **b (Intercept)**: The y-value at zero.
 - **r (Correlation)**: A value close to 1.000 indicates a good fit.
6. Tap **OK** to display the line on the graph.

G. Time-Based Settings (Kinetics)

Changing Duration & Rate

For Kinetics or Temperature experiments, you often need to adjust how long the sensor records.

1. Tap the **Meter** tab.
2. Tap the **Mode** box (it usually says "Time Based").
3. **Duration**: Change this to match your reaction time (e.g., change "180 s" to "600 s").
4. **Rate/Interval**:
 - **Rate**: Samples per second (e.g., 2 samples/s).
 - **Interval**: Time between samples (e.g., 0.5 s/sample).
5. Tap **OK**.

H. Managing Data Runs

Storing & Hiding Runs

If you want to do multiple trials (e.g., Trial 1 and Trial 2) without losing your first graph:

1. **Store a Run**: Tap the **File Cabinet Icon** (next to the "Run 1" label). The screen will clear and prepare for "Run 2." Your old data is safe.
2. **Viewing Multiple Runs**:
 - Tap the **Graph** tab.
 - Tap the **Run Number** button (e.g., "Run 2").
 - Select **All Runs** to see them overlayed.
3. **Strike Through Data (Outliers)**:

- Tap the **Table** tab.
- Tap and drag to select the "bad" data cells.
- Tap **Table** (menu) → **Strike Through Data**.
- The graph will ignore these points, but the data remains visible with a line through it.

I. Graph Analysis Tools

Statistics & Tangents

1. Statistics (Mean/Max/Min)

1. On the **Graph** tab, drag across the region of interest.
2. Tap **Analyze** → **Statistics**.
3. Check the box for your data column.
4. A box will appear showing the Mean (Average), Min, and Max for that selection.

2. Tangent (Finding Slope at a Point)

Use this for Titrations to find the steepest part of the curve (Equivalence Point).

1. Tap **Analyze** → **Tangent**.
2. Tap anywhere on the graph curve.
3. A crosshair will appear. The number displayed is the **Slope** at that exact instant.
4. For titrations, the Equivalence Point is where the slope is **highest** (maximum value).

J. Exporting Data

Emailing Files for Excel

To analyze your data in Excel or Google Sheets, you must email the text file to yourself.

1. Ensure the LabQuest is connected to WiFi (Connections icon).
2. Tap **File** → **Email**.
3. **Option A: Text File (Recommended for Excel)**
 - Tap the **Text** icon.
 - Enter your email address and subject.
 - Tap **Send**. (You can open the .txt file in Excel).
4. **Option B: Screenshot (For Lab Reports)**
 - Tap the **Screenshot** icon (Camera).
 - Enter your email and send.

13.2: Periodic Table of the Elements

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13.3: Units and Conversions

The [International System of Units](#) (SI) contains seven BASE UNITS that each represent a different kind of physical quantity.

- SI unit: **meter (m)**
- The meter is defined as "the length of the path traveled by light in vacuum during a time interval of $1/299,792,458$ of a second."
- SI unit: cubic meter (m^3)
- The cubic meter is defined as the volume of a cube with edges one meter in length. A *liter* is exactly $1/1000^{\text{th}}$ of a cubic meter.
- SI unit: kilogram (kg)
- The kilogram is defined as being equal to the mass of the *International Prototype of the Kilogram* (manufactured in 1799).
- SI unit: pascal (Pa)
- The pascal is a derived from other units

$$1 \text{ Pa} = 1 \frac{\text{kg}}{\text{m} \cdot \text{s}^2} \quad (13.3.1)$$

- SI unit: kelvin (K)
- The kelvin is defined as the fraction $1/273.16$ of the thermodynamic temperature of the triple point of water (exactly 0.01°C).
- SI unit: joule (J)
- The joule is derived from other units

$$1 \text{ J} = 1 \text{ kg} \frac{\text{m}^2}{\text{s}^2} \quad (13.3.2)$$

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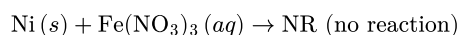
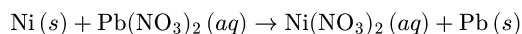
13.4: Activity Series

Single-replacement reactions only occur when the element that is doing the replacing is more reactive than the element that is being replaced. Therefore, it is useful to have a list of elements in order of their relative reactivities. The **activity series** is a list of elements in decreasing order of their reactivity. Since metals replace other metals, while nonmetals replace other nonmetals, they each have a separate activity series. The table below is an activity series of most common metals and of the halogens.

Table 13.4.1: Activity Series of Metals in Aqueous Solutions

Most Active (Easily Oxidized—Readily Lose Electrons)		
lithium	Li	These metals displace hydrogen from water $\text{Ca}(s) + 2 \text{H}_2\text{O}(l) \rightarrow \text{Ca}(\text{OH})_2(aq) + \text{H}_2(g)$ These elements are very reactive and react readily to form compounds.
potassium	K	
barium	Ba	
calcium	Ca	
sodium	Na	
magnesium	Mg	These metals displace hydrogen from acids $\text{Zn}(s) + \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$
aluminum	Al	
zinc	Zn	
chromium	Cr	
iron	Fe	
cadmium	Cd	
nickel	Ni	
tin	Sn	
lead	Pb	
hydrogen	H	
copper	Cu	These metals do not displace hydrogen from acids or water. These elements are more stable and form compounds less readily than do those higher in the table.
silver	Ag	
mercury	Hg	
platinum	Pt	
gold	Au	
Least Active		

For a single-replacement reaction, a given element is capable of replacing an element that is below it in the activity series. This can be used to predict if a reaction will occur. Suppose that small pieces of the metal nickel were placed into two separate aqueous solutions: one of iron (III) nitrate and one of lead (II) nitrate. Looking at the activity series, we see that nickel is below iron, but above lead. Therefore, the nickel metal will be capable of replacing the lead in a reaction, but will not be capable of replacing iron.



In the descriptions that accompany the activity series of metals, a given metal is also capable of undergoing the reactions

described below that section. For example, lithium will react with cold water, replacing hydrogen. It will also react with steam and with acids, since that requires a lower degree of reactivity.

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- [16.5: The Activity Series- Predicting Spontaneous Redox Reactions](#) by Henry Agnew, Marisa Alviar-Agnew is licensed [CK-12](#). Original source: <https://www.ck12.org/c/chemistry/>.

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