SCHOOL-BASED ASSESSMENT

LABORATORY MANUAL

CHEMISTRY

(4th May 2026)

MALAYSIAN INDEPENDENT CHINESE SECONDARY SCHOOLS (MICSS)

Preface

This manual is compiled based on the "Senior Middle Level Chemistry Curriculum Standards" set by the Chemistry subject under Unified Curriculum Committee of Malaysian Independent Chinese Secondary School (MICSS) Working Committee, with reference to both national and international secondary chemistry curricula. Within this framework, experiments are positioned as the core of high school science learning, serving as a bridge between theory and practice. Designed as a tool for school-based assessment by Examination Department, this manual aims to establish a scientific and systematic evaluation system for experimental teaching.

Its key features are as follows:

（1） Guided by the principle of “using assessment to enhance teaching and learning,” the manual provides a structured evaluation framework covering the entire experimental process – from operation procedures and data collection to analysis and reasoning. It emphasizes not only the accuracy of data but also the cultivation of scientific thinking and inquiry skills.

（2） The experiments selected align with curriculum standards while reflecting the distinctive characteristics of independent Chinese secondary schools, balancing foundational skills development with opportunities for extended inquiry. Each experiment is designed to holistically assess students’ scientific literacy.

We extend our sincere gratitude to the experts, scholars, and experienced teachers who reviewed this manual and provided valuable feedback on its content and design.

We believe that true science education lies at the intersection of hands-on practice and deep intellectual engagement. May this manual guide students in exploring the scientific world, fostering a rigorous and inquisitive mind, and finding joy in discovery and growth.

We welcome any feedback or corrections regarding shortcomings in this manual.

Dong Zong Examination Department
January 2026

Laboratory Safety Guidelines

Introduction

Safety in the laboratory is our highest priority. These guidelines are designed to protect you, your colleagues, and the environment. No experiment is so important that it must be performed at the expense of safety. You are responsible not only for your own safety but also for the safety of those around you. Read, understand, and follow these rules at all times.

Section 1: Personal Safety and Preparation

Personal Protective Equipment (PPE):

- **Safety Glasses/Goggles**: Must be worn at all times in the lab, even if you are not performing an experiment.
- **Lab Coat**: A properly fastened lab coat must be worn to protect your skin and clothing.
- **Gloves**: Wear appropriate, chemically resistant gloves for all procedures involving hazardous chemicals. Inspect them for holes before use. Remove them before touching common surfaces (e.g., door handles, keyboards, phones).
- **Closed-Toe Shoes**: Shoes must completely cover the foot. No sandals, flip-flops, or open-toe shoes are permitted.
- **Appropriate Clothing**: Wear clothing that covers and protects your skin. Avoid loose sleeves, dangling jewellery, and scarves. Tie back long hair.

Personal Hygiene:

- **No Eating or Drinking**: Never consume food, beverages, or chew gum in the laboratory.
- **No Applying Cosmetics**: Do not apply lip balm, makeup, or wear contact lenses.
- **Wash Your Hands**: Wash your hands thoroughly with soap and water after handling chemicals, before leaving the lab, and after removing gloves.

Section 2: Before You Begin an Experiment

Know the Hazards:

- Read the entire experimental procedure and any associated Safety Data Sheets (SDS) **before** starting.

- Identify the location and proper use of all safety equipment: eyewash station, safety shower, fire extinguisher, fire blanket, and first-aid kit.
- Plan your work. Understand the steps and anticipate potential hazards.

Chemical Handling:

- **Never Taste or Smell Chemicals**: To detect odours, use your hand to gently waft the vapours toward your nose.
- **Label Everything**: All containers must be clearly labelled with the contents and hazard warnings.
- **Use the Fume Hood**: Any procedure that produces volatile, toxic, or flammable vapours must be conducted in a properly functioning fume hood.

Section 3: During the Experiment

General Conduct:

- **Work Attentively**: Conduct yourself in a responsible and professional manner at all times. No running, horseplay, or practical jokes.
- **Keep a Clean Workspace**: Clutter leads to accidents. Clean up spills immediately and dispose of waste properly.
- **Minimize Distractions**: Avoid using personal mobile phones for non-work purposes.

Specific Procedures:

- **Heating Substances**: Never heat a closed container. Point the mouths of test tubes being heated away from yourself and others. Use boiling chips or broken porcelain pieces to prevent bumping.
- **Glassware**: Check for chips or cracks before use. Do not use damaged glassware. Learn the proper procedure for inserting and removing glass tubing from stoppers.
- **Waste Disposal**: Dispose of all chemical and biological waste in the appropriately labelled containers. Never pour chemicals down the sink unless explicitly instructed to do so.

Section 4: In Case of Emergency

Know the Emergency Procedures:

- **Spills**: Alert others in the area. For minor, non-hazardous spills, clean them up immediately using the appropriate spill kit. For major or hazardous spills, evacuate the area and notify the instructor/lab manager immediately.
- **Fire**: Alert everyone in the lab. For a small, contained fire (e.g., in a beaker), smother it with a watch glass or use a fire extinguisher if you are trained. For a larger fire, **EVACUATE** immediately and activate the fire alarm.

Chemical Splash:

- **On Skin/Clothing**: Immediately flush the affected area with copious amounts of water in the safety shower for at least 15 minutes. Remove contaminated clothing while under the shower.
- **In Eyes**: Immediately use the eyewash station. Hold eyelids open and flush with water for at least 15 minutes.
- **Injury**: Report all injuries, no matter how minor, to the instructor or lab manager immediately.

Section 5: After the Experiment

Clean-Up:

- Clean all glassware and equipment and return it to its proper storage location.
- Wipe down your bench space with disinfectant or soapy water.
- **Waste Disposal**: Ensure all chemical and biological waste has been disposed of according to the provided instructions.
- **Personal Hygiene**: Wash your hands thoroughly with soap and water before leaving the laboratory.

A Final Reminder:

If you are ever unsure about the safety of a procedure, **STOP** and **ASK** your instructor or lab supervisor. Do not proceed until you are certain it is safe to do so.

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Confirming the Copper (II) Oxide Chemical Formula Using Stoichiometric Analysis and Gravimetric Method

Problem Statement

What is the chemical formula of copper (II) oxide as determined through stoichiometric analysis and gravimetric method?

Objectives

1. To determine the chemical formula of copper (II) oxide through experimental observation.
2. To determine the chemical formula of copper (II) oxide through experimental data.
3. To enhance the observation and data analysis capability.

Theory and Principle

Heating is required to provide the activation energy needed for copper to react with oxygen in forming copper (II) oxide. The mass of the copper powder is weighted before and the reaction. The increase in mass corresponds to the oxygen absorbed during the formation of copper (II) oxide. The Cu:O molar ratio is determined by comparing the moles of copper consumed to the moles of oxygen gained using their respective atomic masses.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Copper powder	-	2.0 g
Oxygen gas	Pure	-
Porcelain boat plate	-	1
Hard glass tube	-	1
Electronic balance	-	1
Bunsen burner	-	1
Retort stand	-	1
Spatula	-	1

Precautionary Steps

1. Stop heating once the copper powder undergoes a complete change from reddish brown to black, indicating the formation of copper (II) oxide.
2. Allow the porcelain boat plate to cool to room temperature before measuring it mass to prevent thermal drift error.
3. Cool the porcelain boat plate to room temperature in a desiccator (if available in laboratory) to prevent moisture absorption.
4. Repeat the heating-cooling-weighing cycle until the mass of the porcelain boat plate and its contents [copper (II) oxide] remains constant to confirm complete oxidation.

Procedure

1. Weigh the mass of an empty and dry porcelain boat plate using electronic balance. Record the mass as m_1.
2. Weigh the mass of porcelain boat plate containing 2.0 g of fine copper powder using electronic balance. Record the mass as m_2.
3. Place the porcelain boat plate containing fine copper powder in a hard glass tube.
4. Heat the hard glass tube strongly while passing through a stream of oxygen gas over the sample.

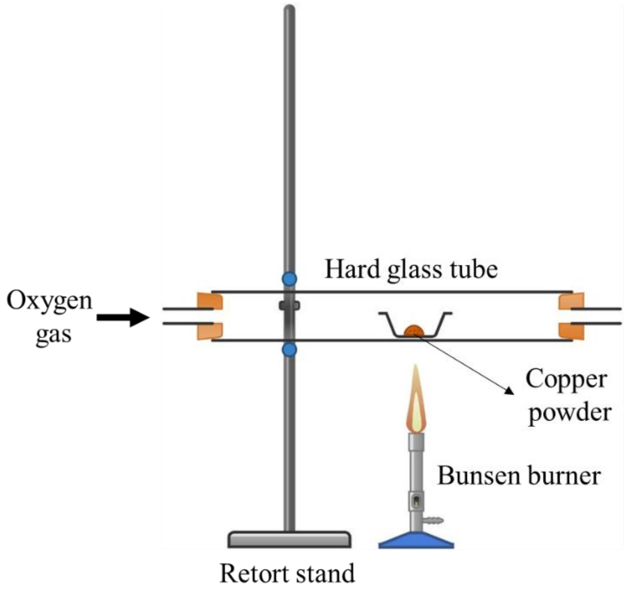


Figure 1.1: Experimental setup for heating copper powder in a hard glass tube under a continuous flow of oxygen gas.

5. Stop heating after 20 minutes. Allow the hard glass tube and porcelain boat plate to cool to room temperature in a desiccator to prevent moisture absorption.
6. Weigh the mass of cooled porcelain boat plate and its copper (II) oxide using electronic balance.

7. Repeat the heating, cooling and weighting process until the mass of cooled porcelain boat plate and its copper (II) oxide remains constant. Record the mass as m_3.
8. Use the recorded data to determine the copper (II) oxide chemical formula.

Results

Recorded data	Mass, m (g)
The mass of porcelain boat plate, m_1	
The mass of porcelain boat plate containing 2.0 g of fine copper powder, m_2	
The mass of cooled porcelain boat plate and its copper (II) oxide, m_3	
The mass of fine copper powder	
The mass of copper (II) oxide	
The mass of oxygen gas used	

Cu:O mole ratio in copper (II) oxide:

(Relative molecular mass: Cu = 63.5, O = 16)

Chemical formula of copper (II) oxide:

Discussion

1. Why might the experimental result deviate from the theoretical ratio? Explain your answer.

2. How to confirm the copper and oxygen has completely reacted to form copper (II) oxide?

3. If the final mass decreased after heating, how would this affect the empirical formula calculation?

4. Why is pure oxygen gas flowed over the sample instead of relying on ambient air? How would using air affect the result?

5. Can we use the similar method to determine the chemical formula of magnesium oxide? Explain your answer.

6. An electronic balance can measure mass up to four decimal places (e.g., 0.1234 g). When recording the mass of a single chemical sample for a laboratory experiment, how many decimal places should typically be recorded? Justify your answer.

7. When measuring the mass of a sample three times, the readings are slightly different each time. How do you determine that the mass has stabilized and can be considered constant？

Confirming the Zinc Sulfate Chemical Formula Using Stoichiometric Analysis and Gravimetric Method

Problem Statement

What is the chemical formula of zinc sulfate is as determined through stoichiometric analysis and gravimetric method?

Objectives

1. To determine the chemical formula of zinc sulfate through experimental observation.
2. To determine the chemical formula of zinc sulfate through experimental data.
3. To enhance the observation and data analysis capability.

Theory and Principle

The chemical formula of a compound can be determined through a displacement reaction between a metal a salt solution. In this experiment, a fixed mass of zinc powder reacts with excess copper (II) sulfate solution displacing copper metal. By weighing the masses of zinc consumed and copper formed, the Zn:Cu molar ratio can be calculated. This ratio combined with known SO_4^{2-} ion charge, allows derivation of zinc sulfate chemical formula.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Zinc powder	Analytical grade	1.5 – 2.0 g
Copper (II) sulfate solution	l.0 mol L^{-1}	80.0 mL
Distilled water	-	130.0 mL
Alcohol	-	20.0 mL
Hydrochloric acid	1.0 mol L^{-1}	20.0 mL
Electronic balance	-	1
Tong	-	1
Evaporating dish	10 cm in diameter	1

Glass rod	-	1
Measuring cylinder	25 mL	1
Desiccator	-	1
Tripod stand	-	1
Bunsen burner	-	1
Wire gauze with asbestos	-	1
Spatula	-	1

Precautionary Steps

1. Use decantation method to decant solution with the following precautions for its operation:
   - Stir the solution for 2 – 3 minutes to ensure homogeneity, then let it stand undisturbed for 10 – 15 minutes. This allows complete sedimentation of insoluble particles.
   - Carefully decant the supernatant liquid into an evaporating dish by guiding it with a glass rod. Ensure the sediment remains undisturbed at the bottom of the original container.
2. Gently dry the copper powder using a Bunsen burner or dry the copper powder in a drying oven (if available in laboratory) at 60 – 70 °C for 20 minutes to remove moisture while minimizing oxidation.

Procedure

1. Weigh the mass of an empty and dry evaporating dish using electronic balance. Record the mass as m_1.
2. Weigh the mass of evaporating dish containing 1.5 – 2.0 g of zinc powder using electronic balance. Record the mass as m_2.
3. Add 80.0 mL of saturated copper (II) sulfate solution to evaporating dish, then stir the mixture using a glass rod to ensure complete reaction.
4. Pour out the upper layer (clear liquid) from the evaporating dish. Use distilled water to wash the evaporating dish inner wall 2 – 3 times, then decant the solution out. Repeat the aforementioned step using 5.0 mL of alcohol.
5. Dry the evaporating dish and its copper powder gently using a Bunsen burner on the wire gauze with asbestos and tripod stand.
6. Use a tong to transfer the evaporating dish to desiccator to cool to room temperature.

7. Repeat the heating, cooling and weighting process until the mass of cooled evaporating dish and its copper powder remains constant.
8. Wash the copper powder using 20.0 mL of 1.0 mol L^{-1} of hydrochloric acid. Decant the liquid and wash it again with distilled water.
9. Oven dry the copper powder for 5 minutes at 80 – 100 ^{o}C or air dry it completely. Record the mass as m_3.

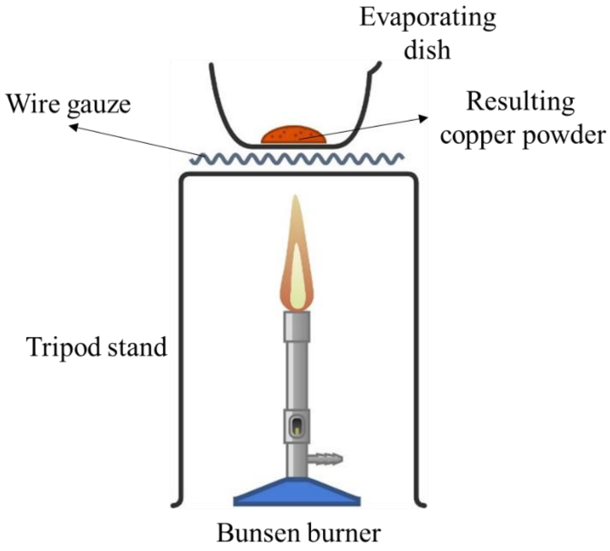


Figure 2.1: Experimental setup for drying resulting copper powder.

Results

Recorded data	Mass, *m* (g)
The mass of evaporating dish, m_1	
The mass of evaporating dish containing zinc powder, m_2	
The mass of cooled evaporating dish and its copper powder, m_3	
The mass of zinc powder	
The mass of copper powder	

Zn:Cu mole ratio:

(Relative molecular mass: Zn = 65, Cu = 63.5)

Chemical formula of zinc sulfate:

Discussion

1. What is the impact on the mass of copper powder obtained if the evaporating dish inner wall is not wash with distilled water in step 4? Explain your answer.

2. State two advantages using distilled water followed by alcohol to wash the obtained copper powder.

3. Why evaporating dish needed to be heated on the wire gauze with asbestos?

4. How to test the purity of copper powder obtained? Explain your answer.

5. If copper (II) sulfate solution was contaminated with sodium chloride, how would this affect the calculated formula? Explain your answer.

6. During the reaction between zinc powder and copper (II) sulfate solution, describe your observations and explain the changes.

7. Give an example of a displacement reaction that is applicable in daily life.

Electrical Conductivity of Ionic and Covalent Compounds

Problem Statement

What is the difference between the electrical conductivity of sodium thiosulfate pentahydrate powder and paraffin wax?

Objectives

1. To determine the electrical conductivity of sodium thiosulfate pentahydrate powder and paraffin wax using electrolysis.
2. To determine the electrical conductivity of sodium thiosulfate pentahydrate powder and paraffin wax in different states.

Theory and Principle

Ionic compounds conduct electricity only when molten or dissolved (due to mobile ions), while most covalent compounds are non-conductive except for special cases like graphite (with delocalized electrons) or polar molecules that ionize in water. Solid ionic compounds and typical molecular covalent compounds don't conduct because their charges/electrons can't move freely. This explains why ionic compounds work well as electrolytes, whereas covalent materials generally act as insulators. A simple conductivity test can demonstrate these properties - ionic solutions complete a circuit, while pure covalent substances won't.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium thiosulfate pentahydrate powder	-	5.0 g
Paraffin wax	-	5.0 g
Batteries	1.5 V	2
Ammeter	-	1
Connecting wires	-	2
Rheostat	-	1
Crucible	-	2

Carbon electrode	-	4
Pipeclay triangle	-	1
Tripod stand	-	1
Bunsen burner	-	1
Spatula	-	1

Procedure

1. Connect the batteries to the carbon electrode by connecting wires.
2. Put a few spatula of sodium thiosulfate pentahydrate powder into the crucible.
3. Turn on the switch and observe the ammeter.
4. Melt the sodium thiosulfate pentahydrate powder using hot water bath and test the electrical conductivity under molten state. Record the observation.
5. Repeat the steps above using paraffin wax to replace sodium thiosulfate pentahydrate powder.

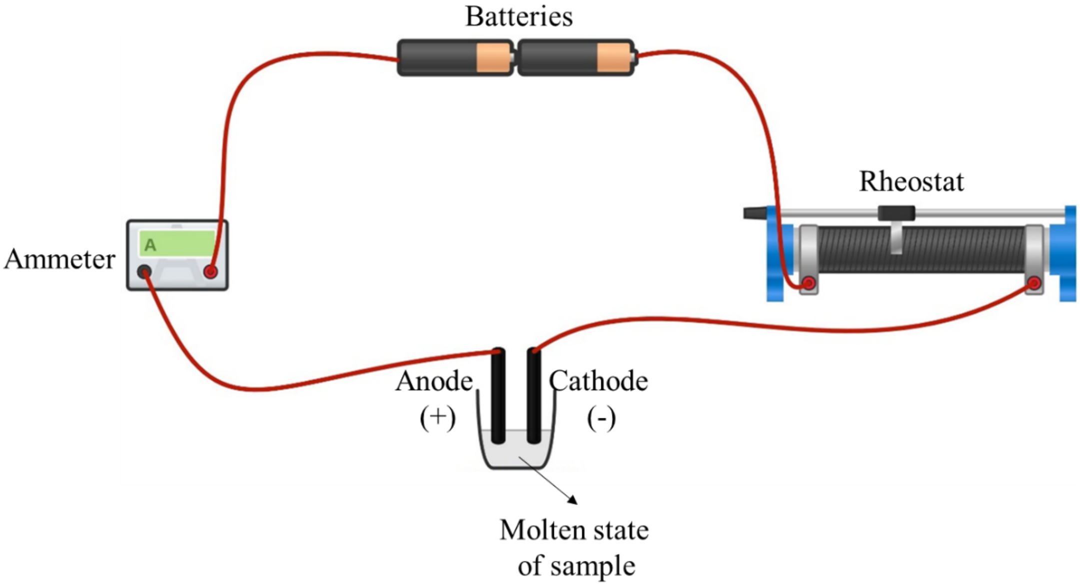


Figure 3.1: Experimental setup for testing electrical conductivity of ionic and covalent compound at molten state.

Results

Chemical compounds	Electrical conductivity	
	Powder	Molten
Sodium thiosulfate pentahydrate		
Paraffin wax		

Discussion

1. Based on your result, how does the bonding type influence a compound's ability to conduct electricity?

2. How did the conductivity change when Sodium thiosulfate pentahydrate was tested in powder and molten state? Explain your answer.

3. In another experiment, hydrochloric acid was used for replacing paraffin wax and found that hydrochloric acid was able to conduct electricity after dissolved in water. Explain you answer.

Solubility of Ionic and Covalent Compounds in Water and Organic Solvent

Problem Statement

What is the difference between the solubility of sodium chloride and hexane in water and organic solvent?

Objectives

1. To determine the solubility of sodium chloride and hexane solution in water and organic solvent.

Theory and Principle

The solubility of compounds depends on their bonding type and the solvent's polarity. Ionic compounds dissolve well in polar solvents like water due to strong ion-dipole interactions, where water molecules stabilize the separated ions. However, they are insoluble in nonpolar organic solvents because these solvents cannot overcome the ionic lattice energy. On the other hands, non-polar covalent compounds are insoluble in water but dissolve well in non-polar organic solvents due to similar London dispersion forces. The principle "like dissolves like" summarizes this behavior: polar solvents dissolve polar or ionic compounds, while non-polar solvents dissolve non-polar compounds. This explains why ionic compounds conduct electricity in water (free ions) but not in organic solvents, whereas covalent compounds vary based on polarity and ionization ability [Experiment 3].

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium chloride	-	4.0 g
Hexane	-	4.0 mL
Water	-	10.0 mL
Cyclohexane	-	10.0 mL
Test tube	15 mm × 150 mm	4
Measuring cylinder	10 mL	2

Precautionary Steps

1. The highly flammable solvent like cyclohexane and hexane, the essential precaution is the strict elimination of all ignition sources, including open flames, hot surfaces, and static electricity – by working in a fume hood to prevent fire or explosion.

Procedure

1. Put 2.0 g of sodium chloride into a test tube containing 5.0 mL of water and shake it vigorously.
2. Observe and record the solubility of sodium chloride in water.
3. Put 2.0 g of sodium chloride into a test tube containing 5.0 mL of cyclohexane and shake it vigorously.
4. Observe and record the solubility of sodium chloride in cyclohexane.
5. Repeat the aforementioned steps using 2.0 mL of hexane to replace sodium chloride.

Results

Chemical compounds	Solubility	
	Water	Cyclohexane
Sodium chloride		
Hexane		

Discussion

1. Based on your observations, which types of compounds were soluble in water? Explain you answer using polarity and bonding concepts.

2. Why water is known as “universal solvent” for ionic compounds except for covalent compounds?

3. In another experiment, ethanol was used to dissolved both sodium chloride and hexane. Predict your observation and explain your answer.

4. What is the reliable method to separate and recycle the mixture of cyclohexane and hexane? Justify your answer.

Melting and Boiling Points of Ionic and Covalent Compounds

Problem Statement

What is the difference between the melting point and boiling point of sodium chloride and paraffin wax?

Objectives

1. To determine the melting point and boiling point of sodium chloride and paraffin wax.

Theory and Principle

Ionic compounds have high melting and boiling points due to strong electrostatic forces in their giant ionic lattices, requiring significant energy to break. Their melting and boiling points depend on ion charge, size and lattice energy. In contrast, covalent compounds vary: simple molecules have melting and boiling points because of weak intermolecular forces (van der Waals, H bonds), while giant covalent networks exhibit extremely high melting and boiling points from their rigid covalent bonding.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium chloride	-	5.0 g
Paraffin wax	-	5.0 g
Crucible	-	2
Pipeclay triangle	-	1
Tripod stand	-	1
Bunsen burner	-	1
Thermometer	100 ^{o}C	1

Precautionary Steps

1. It is recommended to use an alcohol-filled thermometer instead of a mercury thermometer for safety and environmental considerations.

Procedure

1. Put 5.0 g of sodium chloride into a crucible and heat.
2. Observe and record changes of sodium chloride.
3. Repeat the aforementioned steps using 5.0 g of paraffin wax to replace sodium chloride.

Results

Chemical compounds	Observations
Sodium chloride	
Paraffin wax	

Discussion

1. How could you experimentally confirm whether a solid compound is ionic or covalent based on melting and boiling point tests?

2. Magnesium oxide was used in another experiment and found that it has higher melting point than sodium chloride. Explain the phenomenon.

3. Water and hydrogen sulfide are both covalent compounds, why water has a much higher boiling point (100 °C) than hydrogen sulfide (-60 °C)? Explain your answer.

Qualitative Analysis of Selected Metal Cations Using Chemical Test

Problem Statement

How can specific reagents be used to differentiate between Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+} and Ag^{+}, and their observable changes during the qualitative analysis?

Objectives

1. To perform qualitative analysis to identify Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+} and Ag^{+}.
2. To apply chemistry knowledge and experimental skills in identifying unknown chemical compounds.

Theory and Principle

Qualitative analysis of common cations involves systematically separating and identifying ions based on their selective precipitation reactions with group-specific reagents, like hydrochloric acid, hydrogen sulfide and ammonium carbonate. Cations are classified into five groups: Group I forms insoluble chlorides (Ag^{+}, Pb^{2+}), Group II yields acid-insoluble sulfides (Cu^{2+}, Cd^{2+}), Group III cations produce base-insoluble sulfides, but their hydroxides have varying solubility. For example, Fe^{3+} hydroxide is insoluble in base, while Zn^{2+} hydroxide is amphoteric and dissolves in excess base, Group IV precipitates as carbonates (Ca^{2+}, Ba^{2+}), while Group V remains soluble (Na^{+}, K^{+}). Identification relies on solubility principles (Ksp), pH-dependent reactions and characteristic observations, like precipitate colors (white silver chloride, black copper sulfide) or complex formation (blue $[Cu(NH_3)_4]^{2+}$). This method combines classical wet chemistry techniques with fundamental concepts of equilibrium and coordination chemistry for practical cation analysis.

Materials and Apparatus

A. Fe^{2+} and Fe^{3+}

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Iron (II) sulfate solution	0.2 mol L^{-1}	8.0 mL
Hydrochloric acid	0.2 mol L^{-1}	10.0 mL
Hydrogen peroxide	60%	10.0 mL

Potassium hexacyanoferrate (II) solution	0.1 mol L^{-1}	5.0 mL
Sodium hydroxide solution	0.2 mol L^{-1}	10.0 mL
Acidified potassium thiocyanate	0.1 mol L^{-1}	10.0 mL
Potassium hexacyanoferrate (III) solution	0.1 mol L^{-1}	5.0 mL
Iron (III) chloride solution	0.1 mol L^{-1}	5.0 mL
Test tube	12 mm × 120 mm	3
White drip plate	-	1
Dropper	-	7
Test tube holder	-	1

B. Zn^{2+} and Cu^{2+}

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Zinc sulfate solution	0.2 mol L^{-1}	3.0 mL
Copper (II) sulfate solution	0.2 mol L^{-1}	4.0 mL
Concentrated sodium hydroxide solution	2.0 mol L^{-1}	20.0 mL
Concentrated ammonia solution	6.0 mol L^{-1}	10.0 mL
Distilled water	-	-
Test tube	12 mm × 120 mm	6
Bunsen burner	-	1
Dropper	-	5
Filter paper	-	-
Test tube rack	-	1

C. Ag^{+}

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Silver chloride solution	0.1 mol L^{-1}	5.0 mL
Hydrochloric acid	0.1 mol L^{-1}	10.0 mL

Nitric acid	0.1 mol L^{-1}	5.0 mL
Sodium chloride solution	0.1 mol L^{-1}	5.0 mL
Ammonia solution	0.1 mol L^{-1}	10.0 mL
Sodium hydroxide solution	0.1 mol L^{-1}	10.0 mL
Test tube	12 mm × 120 mm	5
Dropper	-	5
Test tube rack	-	1

Precautionary Steps

1. Fresh preparation of iron (II) salt solutions is required before each analytical step to minimize atmospheric oxidation of Fe^{2+} ions to Fe^{3+}.
2. For optimal visual assessment, hold the test tube in front of a plain white paper to enhance color contrast during observations.
3. If laboratory not available, the white drip plate can be replaced with a white tile.
4. Heat the test tube gently and evenly to equally heat the resulting contents. Also, focus the heat just below the surface of the liquid to avoid bumping when heating liquids.
5. When handling concentrated alkalis, prioritize safety to avoid severe burns or injuries. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Work in a well-ventilated area (e.g., a fume hood) to reduce inhalation risks.
6. The acidified reagents are used in the experiment to ensure complete reaction of all carbonate and sulfite ions, preventing the formation of interfering precipitates.
7. Add the reagents dropwise to avoid missing initial observations as the experiment require monitoring of two sequential reaction stages.
8. Carefully handle the chemical wastes generated in the experiment.

Procedures

A. Fe^{2+} and Fe^{3+}

Test for Fe^{2+}

1. Add 3.0 mL of freshly prepared iron (II) sulfate solution into a test tube.
2. Insert the dropper below the iron (II) sulfate solution surface and slowly inject sodium hydroxide into it. Observe and record the changes in the test tube.
3. Re-observe and record the changes in the test tube after a few minutes.
4. Again, add 3.0 mL of freshly prepared iron (II) sulfate solution into a new test tube, then add 2.0 drops of acidified potassium thiocyanate solution. Observe the changes in the test tube.
5. Add 1.0 mL of hydrogen peroxide into the mixture. Observe and record the changes in the test tube.
6. Add 2 drops of freshly prepared iron (II) sulfate solution on the white drip plate, follow by 1 drop of hydrochloric acid and 1 drop of potassium hexacyanoferrate (III) solution. Observe the changes on the white drip plate.

Test for Fe^{3+}

1. Add 3.0 mL of iron (III) chloride solution into a test tube. Observe and record the colour of solution.
2. Then, add sodium hydroxide solution dropwise to the test tube. Observe and record the changes in the test tube.
3. Add 2 drops of iron (III) chloride solution on the white drip plate, follow by 2 drop of acidified potassium thiocyanate. Observe the changes on the white drip plate.
4. Wash the white drip plate. Next, add 2 drops of iron (III) chloride solution, follow by 1 drop of hydrochloric acid and 1 drop of Potassium hexacyanoferrate (II) solution. Observe the changes on the white drip plate.

B. Zn^{2+} and Cu^{2+}

Test for Zn^{2+}

1. Add 3.0 mL of zinc sulfate solution into a test tube.
2. Add concentrated sodium hydroxide solution dropwise to the test tube until a reaction is observed. Split the resulting precipitate equally into 2 test tubes and label as A and B, respectively.
3. Add excess concentrated sodium hydroxide solution to test tube A. Observe and record the changes in the test tube.

4. Add concentrated ammonia solution to test tube B. Observe and record the changes in the test tube.

Test for Cu^{2+}

1. Add 2.0 mL of copper (II) sulfate solution into a test tube. Observe and record the colour of solution.
2. Add concentrated sodium hydroxide solution dropwise to the test tube. Observe and record the changes in the test tube.
3. Filter and wash the resulting precipitate using distilled water. Heat the washed resulting precipitate in a new test tube. Observe and record the changes in the test tube.
4. Add 2.0 mL of copper (II) sulfate solution into a test tube, follow by 2.0 mL of concentrated ammonia solution. Observe and record the changes in the test tube.
5. Then, add concentrated ammonia solution dropwise to the test tube continuously. Observe and record the changes in the test tube.

C. Ag^{+}

1. Add 1.0 mL of silver nitrate solution to 2 different test tubes, respectively, and label as A and B.
2. Add few drops of hydrochloric acid and small volume of nitric acid to test tube A. Observe and record the changes in the test tube.
3. Add few drops of sodium chloride solution and small volume of nitric acid to test tube B. Observe and record the changes in the test tube.
4. Add 1.0 mL of silver nitrate solution, follow by few drops of hydrochloric acid to a new test tube and label as C.
5. After precipitation, add ammonia solution to the test tube C with agitation. Observe and record the changes in the test tube.
6. Add 1.0 mL of silver nitrate solution, follow by sodium hydroxide solution dropwise to a new test tube and label as test tube D. Observe the changes in the test tube.
7. Allow the test tube D stand for 3 minutes. Observe and record the changes in the test tube.
8. Add 1.0 mL of silver nitrate solution, follow by sodium hydroxide solution dropwise to a new test tube and label as test tube E.
9. Once precipitate forms, add ammonia solution dropwise to the test tube E with agitation. Observe and record the changes in the test tube.

Results

A. Fe^{2+} and Fe^{3+}

Test for Fe^{2+}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Slowly add sodium hydroxide solution to iron (II) sulfate solution			
2	Slowly add sodium hydroxide solution to iron (II) sulfate solution after a few minutes			
3	Add acidified potassium thiocyanate solution to iron (II) sulfate solution			
4	Add hydrogen peroxide solution after acidified potassium thiocyanate solution is added with iron (II) sulfate solution			
5	Add acidified potassium hexacyanoferrate (III) solution to iron (II) sulfate solution			

Test for Fe^{3+}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Colour of iron (III) chloride solution			
2	Add sodium hydroxide solution to iron (III) chloride solution			
3	Add acidified potassium thiocyanate solution to iron (III) chloride solution			
4	Add hydrochloric acid and potassium hexacyanoferrate (II) solution to iron (III) chloride solution			

B. Zn^{2+} and Cu^{2+}

Test for Zn^{2+}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Add concentrated sodium hydroxide solution dropwise to zinc sulfate solution			
2	Add concentrated sodium hydroxide solution in test tube A			
3	Add concentrated ammonia solution in test tube B			

Test for Cu^{2+}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Colour of copper (II) sulfate solution			
2	Add concentrated sodium hydroxide solution to copper (II) sulfate solution			
3	Heat the resulting precipitate			
4	Add concentrated ammonia solution dropwise to copper (II) sulfate solution			
5	Add excess concentrated ammonia solution to copper (II) sulfate solution			

C. Ag^{+}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Test tube A Silver nitrate solution + hydrochloric acid + nitric acid			
2	Test tube B Silver nitrate solution + sodium chloride solution + nitric acid			

3	Test tube C Silver nitrate solution + hydrochloric acid + ammonia solution, then agitate			
4	Test tube D Silver nitrate solution + sodium hydroxide solution, then allow to stand			
5	Test tube E Silver nitrate solution + sodium hydroxide solution, then add ammonia solution dropwise with agitation once precipitation occurs			

Discussion

1. The metal cations Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+} and Ag^{+} can be identified using sodium hydroxide solution as a reagent. State the observations for each metal cations.

2. If your test for Fe^{3+} with acidified potassium thiocyanate showed a faint red color instead of deep red, what could be the cause?

3. Evaluate the reliability of using sodium hydroxide solution and ammonia solution to distinguish between Zn^{2+} and Cu^{2+}. Which test is more definitive and why?

4. Why did Ag^+ form a white precipitate with hydrochloric acid but dissolve in ammonia solution, while Cu^{2+} formed a blue complex with ammonia solution?

5. Design a flowchart for a laboratory to separate and identify these metal cations in a mixed sample.

Qualitative Analysis of Selected Anions Using Chemical Test

Problem Statement

How can specific reagents be used to differentiate between S^{2-}, SO_4^{2-}, SO_3^{2-}, CO_3^{2-} and NO_3^-, and their observable changes during the qualitative analysis?

Objectives

1. To perform qualitative analysis to identify S^{2-}, SO_4^{2-}, SO_3^{2-}, CO_3^{2-} and NO_3^-.
2. To apply chemistry knowledge and experimental skills in identifying unknown chemical compounds.

Theory and Principle

The qualitative analysis of anions (S^{2-}, SO_4^{2-}, SO_3^{2-}, CO_3^{2-} and NO_3^-) relies on selective chemical reactions that produce observable changes such as precipitate formation, gas evolution or color shifts. The S^{2-} forms a black precipitate with lead (II) nitrate and releases hydrogen sulfide gas with acid. The SO_4^{2-} yields a white precipitate with barium chloride in acidic medium. The SO_3^{2-} produces sulfur dioxide gas when treated with acid, which turns acidified potassium dichromate paper green. The CO_3^{2-} reacts with acid to release carbon dioxide gas, turning limewater milky. The NO_3^- is identified by the brown ring test, where it forms a nitrosyl complex with iron (II) sulfate and concentrated sulfuric acid. These distinct reactions allow for systematic differentiation among the anions.

Materials and Apparatus

A. S^{2-}, SO_4^{2-} and SO_3^{2-}

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Iron (II) sulfide	Solid	-
Dilute sodium sulfide solution	-	3.0 mL
Dilute lead (II) nitrate solution	-	Small amount
Dilute barium chloride solution	-	10.0 mL
Dilute sodium sulfite solution	-	3.0 mL
Dilute sodium sulfate solution	-	3.0 mL
Concentrated hydrochloric acid	2.0 mol L^{-1}	10.0 mL

Concentrated sulfuric acid	2.0 mol L^{-1}	10.0 mL
Test tube	12 mm × 120 mm	9
Dropper	-	6
Blue litmus paper	-	4
Test tube rack	-	1

B. CO_3^{2-} and NO_3^-

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Calcium carbonate	Solid	3 – 4 pieces
Potassium nitrate	Powder	Small amount
Saturated iron (II) sulfate solution	-	5.0 mL
Dilute potassium nitrate solution	-	Small amount
Saturated clear limewater	-	5.0 mL
Concentrated hydrochloric acid	2.0 mol L^{-1}	5.0 mL
Concentrated sulfuric acid	-	Small amount
Concentrated nitric acid	-	Small amount
Copper plate	-	2 pieces
Test tube	12 mm × 120 mm	10
Rubber stopper with delivery tube	-	1
Bunsen burner	-	1
Test tube holder	-	1
Test tube rack	-	1

Precautionary Steps

1. Conduct the experiment in a fume hood or fume cabinet due to the reactions will produce hydrogen sulfide gas and nitrogen dioxide gas.
2. Heat the test tube gently and evenly to equally heat the resulting contents. Also, focus the heat just below the surface of the liquid to avoid bumping when heating liquids.
3. Concentrated acids must always be handled in a fume hood to protect against toxic fumes, exothermic reactions and chemical burns. Always wear appropriate PPE – gloves, eye protection (goggles or a

face shield) and a lab coat. Acids might release hazardous vapors or gas. These fumes can cause severe respiratory irritation, lung damage or even chemical pneumonia if inhaled.

4. Dispose the remaining chemical solution to waste bottles after finishing the experiment.
5. Wash the unreacted or incomplete reacted iron (II) sulfide and copper plate, and reuse in next experiment.

Procedures

A. S^{2-}, SO_4^{2-} and SO_3^{2-}

Test for S^{2-}

1. Place a small piece of iron (II) sulfide in a test tube and add 2.0 mL of concentrated sulfuric acid. Observe and record the changes in the test tube.
2. Also, put the moist blue litmus paper near to the mouth of test tube. Observe and record the changes on the litmus paper.
3. Add 3.0 mL of dilute sodium sulfide solution, follow by 1 – 2 drops of dilute lead (II) nitrate solution to the new test tube. Observe and record the changes in the test tube.

Test for SO_4^{2-}

1. Add 3.0 mL of concentrated sulfuric acid to the third test tube containing a small volume of dilute barium chloride solution. Observe and record the changes in the test tube.
2. Then, add a small volume of concentrated nitric acid into the mixture. Re-observe and record the changes in the test tube.

Test for SO_4^{2-} and SO_3^{2-}

1. Add 3.0 mL of dilute sodium sulfate solution and dilute sodium sulfite solution, respectively, to separate test tubes.
2. Then, add a small volume of dilute barium chloride solution to each test tube. Observe and record the changes in the test tubes.
3. Next, add a small volume of concentrated hydrochloric acid to each test tube and gently agitate. Re-observe and record the changes in the test tubes.
4. Meanwhile, put the moist blue litmus paper near to each mouth of test tube. Observe and record the changes on the litmus paper.

B. CO_3^{2-} and NO_3^{-}

Test for CO_3^{2-}

1. Place a few pieces of calcium carbonate to a clean test tube. In a separate test tube, prepare 3.0 – 4.0 mL of saturated clear limewater.
2. Then, add 2.0 – 3.0 mL of concentrated hydrochloric acid to the test tube containing calcium carbonate. Immediately seal the test tube with a rubber stopper fitted with a delivery tube and connect the other end to a second test tube containing saturated clear limewater.
3. Observe and record the changes in both test tubes.

Test for NO_3^{-}

1. Add a small amount of potassium nitrate powder to a new test tube, follow by a small volume of concentrated sulfuric acid and copper plate.
2. Heat the mixture gently. Observe and record the changes in test tube.
3. In a new test tube, add 2.0 mL of freshly prepared saturated iron (II) sulfate solution. Then introduce 1.0 mL of dilute potassium nitrate solution, while gently agitating the mixture to ensure homogeneous mixing.
4. Tilt the test tube at 45° and slowly add 1.0 mL of concentrated sulfuric acid along the wall of the test tube. Do not shake and slowly put the test tube upright.
5. After a moment, observe and record the changes in both test tube.

Results

A. S^{2-}, SO_4^{2-} and SO_3^{2-}

No.	Procedures	Observations	Chemical equations	Conclusions
1	Add concentrated sulfuric acid to iron (II) sulfide			
2	Add few drops of dilute lead (II) nitrate solution to sodium sulfide solution			
3	Add a small volume of dilute barium chloride			

	solution to concentrated sulfuric acid then add concentrated nitric acid			
4	Add a small volume of dilute barium chloride solution to dilute sodium sulfate solution then add hydrochloric acid			
5	Add a small volume of dilute barium chloride solution to dilute sodium sulfite solution then add hydrochloric acid			

B. CO_3^{2-} and NO_3^-

No.	Procedures	Observations	Chemical equations	Conclusions
1	Add concentrated hydrochloric acid to calcium carbonate			
2	Add concentrated sulfuric acid and copper plate to a small amount of potassium nitrate powder, and heat			
3	Add dilute potassium nitrate solution to freshly prepared saturated iron (II) sulfate solution, follow by concentrated sulfuric acid			

Discussion

1. How did the addition of concentrated hydrochloric acid or concentrated nitric acid help to differentiate between SO_4^{2-} and SO_3^{2-} in your tests? Support your answer with observed results.

2. If an unknown sample gave positive results for both S^{2-} and SO_3^{2-}, how would you resolve this ambiguity?

3. When testing for CO_3^{2-} with acid, why was it important to use limewater as confirmatory test rather than relying solely on effervescence?

4. The brown ring formed but disappeared immediately in some students' test. What does it indicate about the sample or technique?

5. Design a flowchart for an laboratory to separate and identify these anions in a mixed sample.

Identification of an Unknown Chemical Compound Using Qualitative Analysis

Problem Statement

How can the identity of an unknown chemical compound be determined through systematic qualitative analysis of its physical and chemical properties?

Objectives

1. To determine the identity of an unknown chemical compound by analyzing its physical and chemical properties.
2. To conduct an experiment for the systematic identification of ionic compounds.

Theory and Principle

Qualitative analysis identifies ions through systematic chemical tests based on their unique reactivity patterns. The process relies on selective precipitation, characteristic color changes, gas evolution and solubility differences to distinguish between ions. For cations, NH_4^+ is detected through ammonia gas liberation when heated with sodium hydroxide solution, while Al^{3+} forms an amphoteric hydroxide soluble in excess sodium hydroxide solution and Mg^{2+} produces a non-amphoteric hydroxide. The SO_4^{2-} is identified by forming an acid insoluble white precipitate with Ba^{2+} and Cl^- yields a silver precipitate soluble in ammonia. These tests exploit fundamental chemical principles, including solubility product constants, amphoterism, and complex ion formation. The methodology follows a logical sequence to prevent interference between tests, beginning with NH_4^+ detection before proceeding to precipitation reactions. Each confirmatory test is designed to produce distinctive visual results that provide unambiguous identification of the target ion while accounting for potential cross-reactivity through controlled experimental conditions.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Unknown solution (labeled as 'X')	-	10.0 mL
Barium nitrate solution	0.5 mol L^{-1}	10.0 mL
Nitric acid	0.5 mol L^{-1}	10.0 mL
Silver nitrate solution	0.5 mol L^{-1}	5.0 mL
Concentrated sodium hydroxide solution	2.0 mol L^{-1}	10.0 mL
Test tube	15 mm × 150 mm	3
Dropper	-	4
Red litmus paper	-	1
Measuring cylinder	5 mL	1
Filter paper	-	1
Filter funnel	-	1
Test tube rack	-	1

Precautionary Steps

1. Perform all steps involving barium nitrate solution in a fume hood, as it is toxic and poses inhalation risks. Wear gloves and safety goggles throughout the experiment.
2. When handling concentrated alkalis, prioritize safety to avoid severe burns or injuries. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Work in a well-ventilated area (e.g., a fume hood) to reduce inhalation risks.

Procedures

1. Add 3.0 mL of Solution 'X' to the test tube, follow by few drops of dilute nitric acid. Observe and record the changes in the test tube.
2. Then, add some barium nitrate solution to the same test tube. Observe and record the changes in the test tube.
3. If white precipitate is formed, add excess barium nitrate solution to the same test tube and filter it.

4. Add silver nitrate solution and dilute nitric acid to the filtrate. Observe and record the changes of the mixture.
5. If no precipitate formed, add silver nitrate solution and dilute nitric acid to the same text tube. Observe and record the changes in the test tube.
6. Add 1.0 mL of Solution 'X' to the new test tube, follow by excess concentrated sodium hydroxide and heat. Put the moist red litmus paper near to the mouth of test tube. Observe and record the changes on the litmus paper.
7. Add 3.0 mL of Solution 'X' to another test tube, follow by concentrated sodium hydroxide dropwise until excess. Observe and record the changes in the test tube.

Results

No.	Procedures	Observations	Conclusions
1	Add few drops of dilute nitric acid to the Solution 'X'		
2	Add barium nitrate solution to the same test tube		
3	Add silver nitrate solution and dilute nitric acid to the filtrate		
4	Heat the mixture of Solution 'X' with excess concentrated sodium hydroxide solution		
5	Add concentrated sodium hydroxide solution dropwise to the Solution 'X'		
6	Add excess concentrated sodium hydroxide solution dropwise to the Solution 'X'		

Discussion

1. Design a flowchart to separate and identify these five ions in a mixture.

2. Why was it critical to test for NH_4^+ first before analyzing other cations (Al^{3+} and Mg^{2+})? How would ammonia gas interfere with later tests?

3. Explain the purpose of adding excess concentrated sodium hydroxide solution when testing for Al^{3+} and Mg^{2+}. How did their reactions differ?

4. When testing for SO_4^{2-}, why was dilute nitric acid added to the barium nitrate solution test? What would happen if you omitted the acid?

5. Your unknown gave:
    a. No gas with sodium hydroxide solution (ruling out NH_4^+)
    b. White precipitate with sodium hydroxide solution (insoluble in excess)
    c. White precipitate with dilute nitric acid added to the barium nitrate solution

    What is the likely identity? Justify step-by-step.

Characterization of Carbon Compounds Using Sodium Carbonate and Sodium Bicarbonate

Problem Statement

How can the properties of carbon compounds being characterize using sodium carbonate and sodium bicarbonate?

Objectives

1. To determine sodium carbonate and sodium bicarbonate properties.
2. To conduct an experiment by applying basic experimental skills.

Theory and Principle

The characterization of carbon compounds using sodium carbonate and sodium bicarbonate is based on their reaction with acidic organic compounds. Carboxylic acids, being strong acids (p*K*a ~ 4 – 5), react with both sodium carbonate and sodium bicarbonate, producing effervescence due to carbon dioxide gas release. Phenols, which are weaker acids (p*K*a ~ 10), react only with sodium carbonate, forming phenoxide without carbon dioxide gas evolution, while alcohols, being neutral, do not react with either reagent. This test helps distinguish between carboxylic acids, phenols and alcohols based on their acidic strength. Sodium bicarbonate is more selective and reacts only with stronger acids (like carboxylic acids), whereas sodium carbonate reacts with both carboxylic acids and phenols. The presence or absence of effervescence (carbon dioxide gas) serves as a key observation. Limitations include weak phenols not reacting clearly and some substituted phenols (e.g., picric acid) behaving like carboxylic acids. This method is useful in qualitative organic analysis for preliminary functional group identification.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium carbonate	-	8.0 g
Sodium bicarbonate	-	8.0 g
Concentrated hydrochloric acid	2.0 mol L^{-1}	10.0 mL

Phenolphthalein indicator	-	5.0 mL
Saturated clear limewater	-	20.0 mL
Test tube	18 mm × 180 mm	8
Delivery tube	-	1
Electronic balance	-	1
Dropper	-	1
Bunsen burner	-	1
Retort stand with clamp	-	1
Balloon	-	2
Rubber band	-	1

Precautionary Steps

1. Concentrated acid must always be handled in a fume hood to protect against toxic fumes, exothermic reactions and chemical burns. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Acid might release hazardous vapors or gas. These fumes can cause severe respiratory irritation, lung damage or even chemical pneumonia if inhaled.
2. Ensure the concentrated hydrochloric acid used in the experiment is in excess.
3. Ensure the same molarity of sodium carbonate and sodium bicarbonate used in experiment before put inside the balloon.

Procedures

A. Test for Solubility of Sodium Carbonate and Sodium Bicarbonate

1. Add approximately 1.0 g of sodium carbonate and sodium bicarbonate, respectively, into two separate test tube.
2. Then, add few drops of water to both test tubes and agitate. Touch the bottom of both test tubes.
3. Continue add 10.0 mL of water, follow by 1 – 2 drops of phenolphthalein indicator to both test tubes. Observe and record the changes in the test tubes.

B. Test for Thermostability of Sodium Carbonate and Sodium Bicarbonate

1. Place 5.0 g of sodium carbonate to the new test tube, then connect this test tube to another test tube containing saturated clear limewater using a delivery tube.
2. Heat the test tube containing sodium carbonate using a Bunsen burner. Observe and record the changes in the test tube containing saturated clear limewater.
3. Repeat step 1 and 2 using sodium bicarbonate. Observe and record the changes in the test tube containing saturated clear limewater.

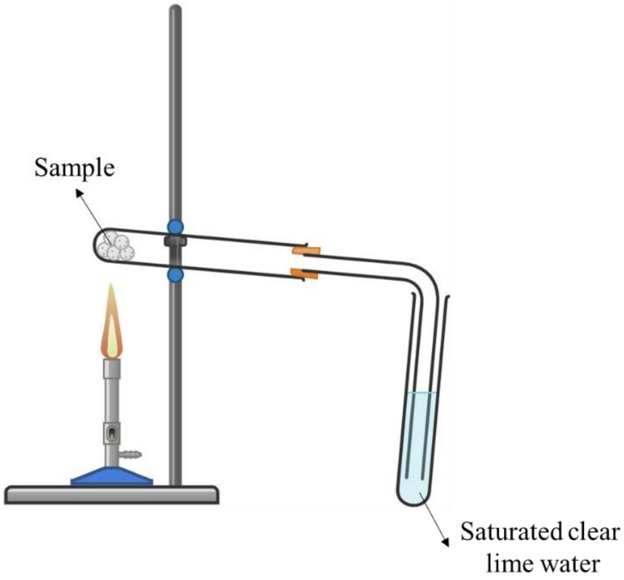


Figure 9.1: Experimental setup for testing the thermal stability of sodium carbonate and sodium bicarbonate.

C. Test for Amount of Gas Produces with Acid

1. Add 5.0 mL of concentrated hydrochloric acid to two separate test tubes, respectively.
2. Inside the balloon, place 1.06 g and 0.84 g of sodium carbonate and sodium bicarbonate, respectively.
3. Stretch the balloon over the mouth of test tube and seal it tightly.
4. Meanwhile, pour the sodium carbonate and sodium bicarbonate to the concentrated hydrochloric acid. Compare both reactions reactivity, and observe and record the changes on the balloon.

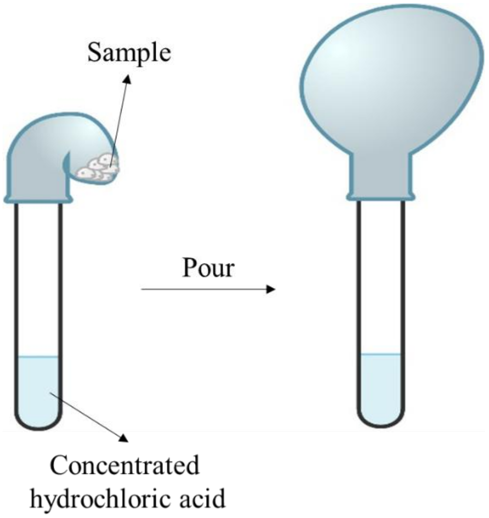


Figure 9.2: Experimental setup for the reaction of concentrated hydrochloric acid with sodium carbonate and sodium bicarbonate.

Results

Chemical compounds	Procedures	Observations	Conclusions
Sodium carbonate	Reaction with water		
	Saturated clear limewater clarity		
	Reactivity and balloon expansion		
Sodium bicarbonate	Reaction with water		
	Saturated clear limewater clarity		
	Reactivity and balloon expansion		

Discussion

1. Why different observations occur when sodium carbonate and sodium bicarbonate dissolves in water?

2. Why did concentrated hydrochloric acid produce effervescence with both sodium carbonate and sodium bicarbonate?

3. How does the p*K*a of a compound determine whether it reacts with sodium bicarbonate or only sodium carbonate?

4. Which carbonate salts used in the experiment is a more selective reagent in identifying acids? Explain your answer.

5. How could you using sodium carbonate and sodium bicarbonate to distinguish between aliphatic and aromatic carboxylic acids? *(Related to Organic Chemistry – Carboxylic acids)*

Preparation and Characterization of Sulfur Dioxide Gas

Problem Statement

How can sulfur dioxide gas be prepared in the laboratory and characterized using qualitative analysis methods?

Objectives

1. To prepare the sulfur dioxide gas in the laboratory.
2. To characterize the prepared sulfur dioxide gas using qualitative analysis.
3. To determine the physical and chemical properties of the prepared sulfur dioxide gas.

Theory and Principle

To prepare sulfur dioxide gas, we start with sodium sulfite and add a dilute acid like hydrochloric acid or sulfuric acid. The reaction happens quickly, producing sulfur dioxide gas along with salt and water. We can collect the gas by downward displacement of air since it's heavier than air. The sulfur dioxide gas has several identifying features. It has a strong, choking smell like burnt matches. When we bubble it through water, it makes the solution acidic, turning blue litmus paper red. Two important chemical tests confirm its presence: it turns purple potassium manganate (VII) solution colourless and it also decolourises the brown colour of bromine solution or iodine solution.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium sulfite	Powder	20.0 g
Concentrated sulfuric acid	75%	50.0 mL
Fuchsin solution	Magenta or pink colour	5.0 mL
Acidified potassium manganate (VII) solution	0.01 mol L^{-1}	5.0 mL
Litmus solution	-	Small amount
Dilute iron (III) chloride solution	-	5.0 mL
Barium chloride solution	-	5.0 mL

Concentrated sodium hydroxide solution	3.0 mol L^{-1}	Small amount
Round-bottomed flask	250 mL	1
Separating funnel	50 mL	1
Beaker	50 mL	1
Bunsen burner	-	1
Absorbent cotton	-	Small amount
Retort stand with clamp	-	2
Gas bottle	200 mL	1
Test tube	18 mm × 180 mm	7
Glass piece	7 cm × 7 cm	2
Glass delivery tube	Right angle	2
Benchclamp	-	1
Rubber stopper with holes	-	1
Dropper	-	1

Precautionary Steps

1. To control rate of reaction, concentrated sulfuric acid should be added dropwise to the sodium sulfite powder.
2. Concentrated acid must always be handled in a fume hood to protect against toxic fumes, exothermic reactions and chemical burns. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Acid might release hazardous vapors or gas. These fumes can cause severe respiratory irritation, lung damage or even chemical pneumonia if inhaled.
3. When handling concentrated alkalis, prioritize safety to avoid severe burns or injuries. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Work in a well-ventilated area (e.g., a fume hood) to reduce inhalation risks.
4. Must conduct the experiment in a fume hood and under well-ventilated area as hydrogen sulfide gas has an irritating odor and is toxic. Avoid direct contact with the gas and wear appropriate personal protective equipment, including gloves and safety goggles.
5. The mouth of test tube must seal with sodium hydroxide solution-soaked absorbent cotton to prevent gas leakage during the experiment.

Procedures

Preparation of Sulfur Dioxide Gas

1. Place approximate 20.0 g of sodium sulfite powder to the round-bottomed flask.
2. Add approximate 50.0 mL of 75% concentrated sulfuric acid to the separating funnel.
3. Hold the round-bottomed flask using a retort stand with clamp, then seal it with a rubber stopper with holes.
4. Hold the separating funnel with a benchclamp, fix with the retort stand and connect with round-bottomed flask through the rubber stopper.
5. Fill the test tube with 5.0 mL of Fuchsin solution and hold by a new retort stand with clamp.
6. Seal the mouth of test tube with concentrated sodium hydroxide solution-soaked absorbent cotton and connect to the round-bottomed flask using a glass delivery tube through the rubber stopper.
7. Allow the 75% concentrated sulfuric acid add dropwise to the sodium sulfite powder and heat the round-bottomed flask. Observe and record the changes of Fuchsin solution.
8. Collect the sulfur dioxide gas using a gas bottle.
9. Add 20.0 mL of distilled water to a beaker and deliver the collected sulfur dioxide gas to it in order to prepare sulfur dioxide solution.

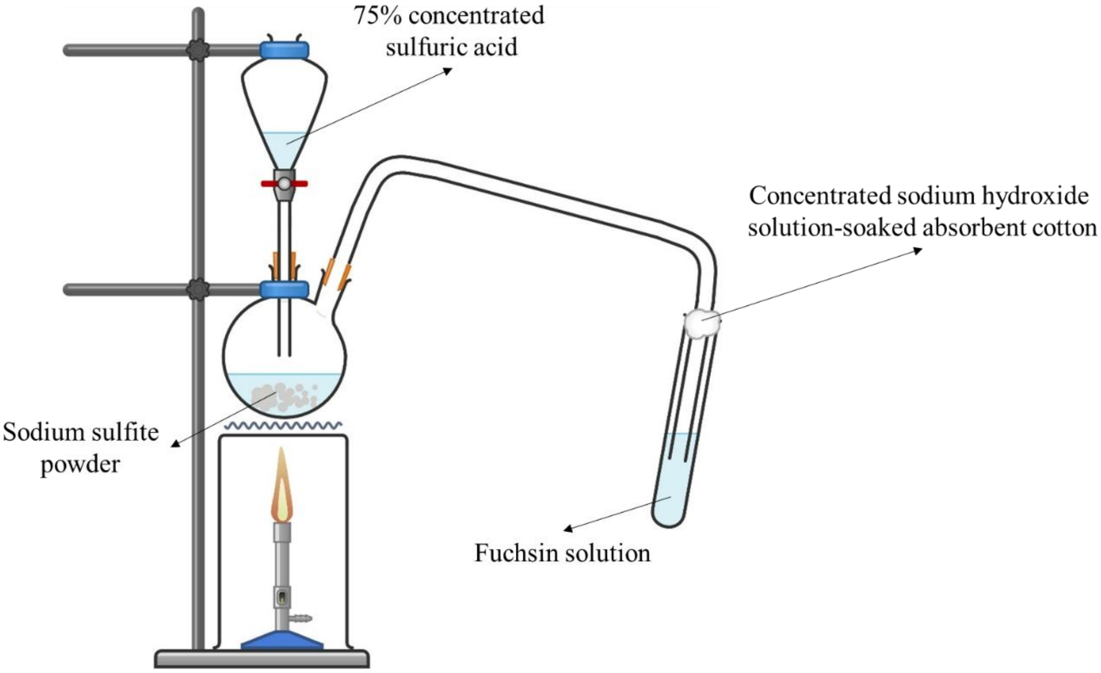


Figure 10.1: Experimental setup for sulfur dioxide preparation.

Test for Sulfur Dioxide Gas

1. Add a few drops of sulfur dioxide solution to the test tube containing litmus solution using a dropper. Observe and record the changes in the test tube.
2. Prolong the exposure of sulfur dioxide solution or add few more drops of sulfur dioxide solution to the test tube containing litmus solution using a dropper. Re-observe and record the changes in the test tube.
3. Add 3.0 mL of sulfur dioxide solution to a new test tubes, follow by few drops of acidified potassium manganate (VII) solution. Observe and record the changes in the test tube.
4. Repeat 3 using dilute iron (III) chloride solution and barium chloride solution, respectively, replacing acidified potassium manganate (VII) solution. Observe and record the changes in the test tubes.

Results

No.	Procedures	Observations	Conclusions
1	Sodium sulfite powder reacts with concentrated sulfuric acid		
2	Pass into the Fuchsin solution		
3	Add few drops of sulfur dioxide solution to litmus solution		
4	Add acidified potassium manganate (VII) solution in dropwise		
5	Add dilute iron (III) chloride solution in dropwise		
6	Add barium chloride solution in dropwise		

Discussion

1. What is the role of sulfur dioxide solution as an oxidizing and reducing agent in the **Results** (3) – (6) reactions? Provide balanced chemical equations.

2. Can the reagents below use to selectively eliminate sulfur dioxide gas in the mixture with carbon dioxide gas? Explain your answer.
   a) Sodium hydroxide solution b) Sodium carbonate solution c) Barium chloride solution

   d) Acidified potassium manganate (VII) solution

3. A student accidentally put the petal in the gas bottle containing sulfur dioxide gas. Predict and explain your observation.

4. How does the sulfur dioxide gas contribute to acid rain? Write out your answer with relevant equations.

5. When barium chloride is added dropwise to sulfur dioxide solution with the presence of hydrogen peroxide, formation of white precipitate is observed. The white precipitate does not dissolve in dilute hydrochloric acid. Explain the observation and suggest the product name.

Determination of the Heat of Neutralization

Problem Statement

How does the temperature change during the neutralization of hydrochloric acid and sodium hydroxide in a calorimeter?

Objectives

1. To enhance the knowledge in heat of neutralization.
2. To determine the heat of neutralization between strong acid and strong base.
3. To analyze the deviation and causes in heat of neutralization experiment.

Theory and Principle

The heat of neutralization measures the energy change (-57.1 kJ mol^{-1} for strong acid-base reactions) when H^+ and $OH^–$ ions combine to form water. Using calorimetry, we track temperature changes to calculate the released energy through $Q = mc\Delta t$, then determine molar heat by accounting for moles of water produced. Strong electrolytes yield consistent results due to complete dissociation, while experimental errors like heat loss may cause deviations. This fundamental thermochemical concept has practical applications in industrial processes and provides a foundation for studying more complex reactions. The experiment demonstrates core principles of energy transfer in chemical systems while emphasizing careful technique for accurate measurements.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Hydrochloric acid	0.5 mol L^{-1}	150.0 mL
Sodium hydroxide solution	0.55 mol L^{-1}	150.0 mL
Beaker	500 mL	1
	100 mL	1
Crumpled paper	-	-
Cardboard	Hole in center	1
Thermometer	10 °C – 100 °C	1

Measuring cylinder	50 mL	2
Glass rod	-	1

Precautionary Steps

1. It is recommended to use an alcohol-filled thermometer instead of a mercury thermometer for safety and environmental considerations.
2. Crumpled papers in the experiment can be replaced by cotton wools or Styrofoam.

Procedures

1. Place crumpled paper into the 500 mL beaker, then carefully nest the 100 mL beaker inside it, ensuring the rims of both beakers are aligned at the same level.
2. Fill the space between the beakers using crumpled paper and cover the 500 mL beaker with cardboard to achieve the purpose of heat preservation, insulation and minimize heat loss during experiment.
3. Pour 50.0 mL of 0.5 mol L^{-1} hydrochloric acid to 100 mL beaker. Measure and record the initial temperature of the hydrochloric acid. Wash the thermometer after taking the temperature.
4. Measure 50.0 mL of 0.55 mol L^{-1} sodium hydroxide solution using a measuring cylinder. Measure and record its initial temperature. Wash the thermometer after taking the temperature.
5. Pour the 50.0 mL of 0.55 mol L^{-1} sodium hydroxide solution to the hydrochloric acid and gently stir the mixture using a glass rod. Cover the rims of both beakers using the cardboard immediately.
6. Insert the thermometer through the hole in center on the cardboard. Observe and record the changes on the thermometer. The highest temperature indicates the termination temperature of the reaction.
7. Triplicate the aforementioned steps above and record the data of the experiment.

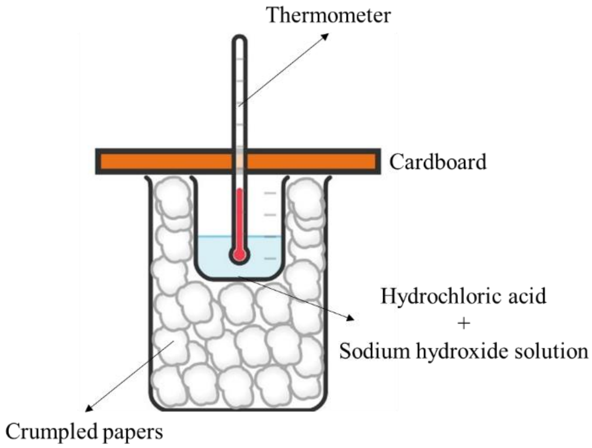


Figure 11.1: Experimental setup for determining the heat of neutralization of hydrochloric acid and sodium hydroxide solution.

Results

Trials	Initial temperature (°C)			Final temperature (°C)	Difference (°C)
	Hydrochloric acid	Sodium hydroxide solution	Average, t_1	t_2	$t_2 - t_1$
1					
2					
3					

Heat of neutralization in this experiment =

Discussion

1. Why is it important to use equal volumes of acid and base solutions in the experiment?

2. How would the temperature change differ if using 1.0 mol L^{-1} of hydrochloric acid instead of 0.5 mol L^{-1} in same volume?

3. Why the heat of neutralization between weak acids and weak bases are lower than of strong acids and strong bases?

4. Slightly excessive of base concentration was used in the experiment rather than acid. Explain your answer.

5. If you repeated the experiment with 100 mL of 0.5 mol L^{-1} hydrochloric acid and 100 mL of 0.55 mol L^{-1} sodium hydroxide solution, how would the temperature change differ? Explain your answer.

Factors Affecting Rate of Reaction

Problem Statement

How do changes in the concentration of hydrochloric acid, the temperature of the reaction, and the particle size (surface area) of calcium carbonate affect the rate of reaction between calcium carbonate and hydrochloric acid?

Objectives

1. To investigate the effect of hydrochloric acid concentration on the rate of reaction.
2. To investigate the effect of reaction temperature on the rate of reaction.
3. To investigate the effect of calcium carbonate particle size (surface area) on the rate of reaction.
4. To practice measuring reaction rates using the gas collection method.

Theory and Principle

The reaction between calcium carbonate and hydrochloric acid is a classic example of a reaction that produces carbon dioxide. The rate of reaction is a measure of how quickly reactants are consumed or products are formed. In this experiment, the rate is measured by monitoring the volume of carbon dioxide gas produced over time. According to Collision Theory, for a reaction to occur, particles must collide with sufficient energy (greater than the activation energy) and/or correct orientation. The factors being investigated affect the reaction rate by altering the frequency and/or energy of successful collisions:

- Concentration: Increasing the concentration of hydrochloric acid increases the number of H^+ ions per unit volume. This leads to a higher frequency of collisions with the marble surface, thus increasing the reaction rate.
- Temperature: Increasing the temperature provides the particles with more kinetic energy. A greater proportion of collisions will then possess energy equal to or greater than the activation energy, significantly increasing the rate.
- Particle size (surface area): Breaking a solid into smaller pieces increases its total surface area. This provides more contact points for collisions with H^+ ions to occur, leading to a higher reaction rate.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Calcium carbonate	Granules	14.0 g
	Large	2.0 g
	Powder	2.0 g
Hydrochloric acid	0.5 mol L^{-1}	50.0 mL
	0.75 mol L^{-1}	150.0 mL
	1.0 mol L^{-1}	50.0 mL
Conical flask	250 mL	3
Gas syringe with stand and clamps	100 mL (gas syringe)	1
Delivery tube with bung	-	1
Measuring cylinder	50 mL and 100 mL	1 each
Stopwatch	-	1
Electronic balance	-	1
Water baths	30 ^oC, 40 ^oC and 50 ^oC	100 mL each
Tongs	-	1
Spatula	-	1

Precautionary Steps

1. Concentrated acid must always be handled in a fume hood to protect against toxic fumes, exothermic reactions and chemical burns. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Acid might release hazardous vapors or gas. These fumes can cause severe respiratory irritation, lung damage or even chemical pneumonia if inhaled.
2. Point the mouth of the conical flask away from yourself and others.

Procedures

General Setup

1. Connect the conical flask using delivery tube to a gas syringe.
2. Ensure the bung fits tightly into the conical flask and the gas syringe is at zero.
3. Measure a fixed mass of the required calcium carbonate size using the electronic balance.

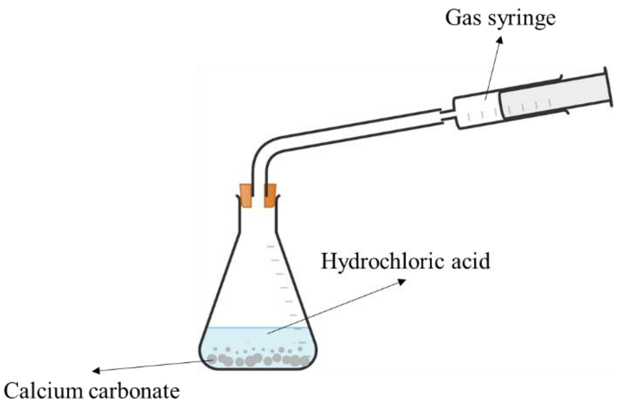


Figure 12.1: General experimental setup for investigating the effect of concentration, temperature and particle size (surface area) on the reaction rate of calcium carbonate with hydrochloric acid.

A. Investigating the Effect of Acid Concentration

1. Using a measuring cylinder, pour 50.0 mL of 0.5 mol L^{-1} hydrochloric acid into the conical flask.
2. Add the 2.0 g of granules of calcium carbonate to the acid, quickly insert the bung, and start the stopwatch.
3. Record the volume of carbon dioxide gas in the syringe at 30 seconds intervals for 5 minutes.
4. Repeat the experiment using 50.0 mL of 0.75 mol L^{-1} hydrochloric acid and then 50.0 mL of 1.0 mol L^{-1} hydrochloric acid, keeping the mass and size of calcium carbonate and the temperature constant.

B. Investigating the Effect of Temperature

1. Place 50.0 mL of 0.75 mol L^{-1} hydrochloric acid into the conical flask and place it in a water bath for a few minutes to reach 30 °C.
2. Add the 2.0 g of granules of calcium carbonate to the warm acid, quickly insert the bung, and start the stopwatch.
3. Record the volume of carbon dioxide gas in the syringe at 30 seconds intervals for 5 minutes.
4. Repeat the experiment with the same acid and calcium carbonate at 40 °C and 50 °C, respectively.

C. Investigating the Effect of Particle Size (Surface Area)

1. Pour 50.0 mL of 0.75 mol L^{-1} hydrochloric acid into the conical flask at room temperature.
2. Add 2.0 g of large size calcium carbonate, quickly insert the bung, and start the stopwatch.
3. Record the volume of carbon dioxide gas in the syringe at 30 seconds intervals for 5 minutes.

4. Repeat the experiment using 2.0 g of granules and powder calcium carbonate, keeping the acid concentration and volume constant.

Results

A. Investigating the Effect of Acid Concentration

Acid concentration ($mol\ L^{-1}$) / Time (seconds)	Volume of carbon dioxide (mL)									
	30	60	90	120	150	180	210	240	270	300
0.5										
0.75										
1.0										

B. Investigating the Effect of Temperature

Temperature (°C) / Time (seconds)	Volume of carbon dioxide (mL)									
	30	60	90	120	150	180	210	240	270	300
30										
40										
50										

C. Investigating the Effect of Particle Size (Surface Area)

Particle size / Time (seconds)	Volume of carbon dioxide (mL)									
	30	60	90	120	150	180	210	240	270	300
Powder										
Granules										
Large size										

Discussion

1. Plot the results obtained in each experiments and answer the following questions.
    a) What does the slope (gradient) of the tangent to the curve at any point represent?
    b) Why is the initial slope considered the most reliable measure of the initial reaction rate?
    c) Why do both graph of high acid concentration and low acid concentration eventually level off at the same final volume?

2. What were the main sources of error or uncertainty in this experiment? Classify them into random errors or systematic errors.

3. Based on your experimental data for the reaction between calcium carbonate and hydrochloric acid, compare the effects of acid concentration and particle size on the reaction rate. Which factor demonstrated a more significant influence? Justify your answer.

4. Suggest two specific improvements to the experimental procedure or apparatus that would increase the reliability or accuracy of the results. For each, explain how it would lead to better data.

Determination of the Percentage of Acetic Acid in Vinegar

Problem Statement

How the percentage of acetic acid can be determined in vinegar using acid-base titration method?

Objectives

1. To determine the acetic acid percentage in vinegar using acid-base titration method.
2. To develop experimental and data processing skills.

Theory and Principle

The percentage of acetic acid in vinegar is determined through acid-base titration, where a standardized sodium hydroxide solution neutralizes the acetic acid in a 1:1 molar ratio ($CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$), with phenolphthalein indicator signalling the endpoint by turning pink, allowing calculation of the acetic acid concentration based on the volume and molarity of sodium hydroxide solution used.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Vinegar	~ 0.5 mol L^{-1}	100.0 mL
Standardized sodium hydroxide solution	0.5 mol L^{-1}	100.0 mL
Phenolphthalein indicator	-	Small amount
Pipette	25 mL	1
Burette	50 mL	1
Conical flask	250 mL	3
Retort stand with clamp	-	1

Procedures

1. Use 25 mL of pipette transfer 25.0 mL of vinegar to 250 mL of conical flask. Then, add few drops of phenolphthalein indicator into it.
2. Use standardized sodium hydroxide solution to rinse the burette 2 – 3 times. Add the standardized sodium hydroxide solution to the burette and record the initial reading.
3. Titrate the standardized sodium hydroxide solution to the vinegar in dropwise while shaking the conical flask.
4. When the mixture in the conical flask turns pink and remains stable for half a minute without decolorization, it indicates the endpoint of the titration.
5. Record the final reading on the burette and calculate the volume of standardized sodium hydroxide solution used.
6. Triplicate the aforementioned steps.

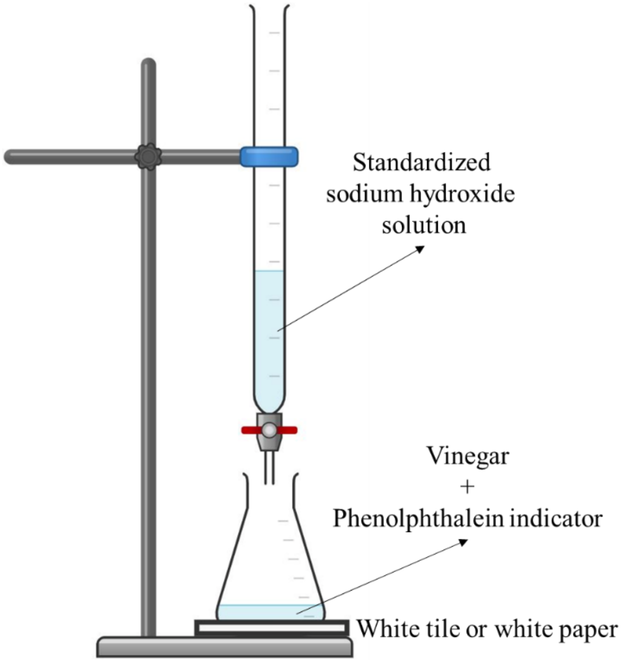


Figure 14.1: Experimental setup for the titration of vinegar with standard sodium hydroxide solution to determine its acetic acid content.

Data Analysis

a. Mass of vinegar = Volume of vinegar × Density of vinegar
where, the density of vinegar can be measured by specific gravity

b. Mass of acetic acid in vinegar = Amount of acetic acid × Molar mass of acetic acid
where, the amount of acetic acid and the amount of standardized sodium hydroxide solution used in neutralization is same

c. Percentage of acetic acid in vinegar = $\frac{\text{Mass of acetic acid in vinegar}}{\text{Mass of vinegar}} \times 100\%$

d. Calculate the percentage of acetic acid in vinegar according to the triplicated experiment results.

Results

Concentration of standardized sodium hydroxide solution = mol L^{-1}

Solutions	Descriptions	Replications		
		1	2	3
Standardized sodium hydroxide solution	Initial reading of burette / mL			
	Final reading of burette / mL			
	Volume used / mL			
	Average of volume used / mL			
Vinegar	Density / g mL^{-1}			
	Volume / mL			
Percentage of acetic acid in vinegar				

Discussion

1. Why using phenolphthalein instead of methyl orange as indicator when titrating sodium hydroxide solution to vinegar?

2. How does the endpoint differ from the equivalence point in this titration?

3. Why the mass of sodium hydroxide solution can be used in calculating the mass of acetic acid?

4. How do dilution errors impact the final percentage of acetic acid and what strategies can be used to prevent them?

5. How would the percentage change if impure vinegar (with other acids) was used?

6. Wine (contains acetic acid) is commonly used in cooking to tenderize meat and add flavour. Besides these two purposes, what is another objective of using wine in this process?

Electrolysis and Electroplating

Problem Statement

How can the principles of electrolysis be applied to electroplate a metal object with another metal?

Objectives

1. To enhance the knowledge in electrolysis and electroplating.
2. To understand simple operation of cyanide-free zinc plating.
3. To nourish experimental skills and observations.

Theory and Principle

Electrolysis and electroplating are electrochemical processes that utilize electric current to drive chemical reactions. Electrolysis involves the decomposition of an electrolyte – either in molten form or dissolved in water – into its constituent elements by passing an electric current through it. This process occurs in an electrolytic cell, where oxidation takes place at the anode (positive electrode) and reduction occurs at the cathode (negative electrode). Faraday's laws govern electrolysis, stating that the amount of substance deposited or liberated is directly proportional to the quantity of electricity passed, and the masses of different substances deposited are proportional to their equivalent weights. A common example is the electrolysis of water, which produces hydrogen gas at the cathode and oxygen gas at the anode.

Electroplating is a practical application of electrolysis used to deposit a thin layer of metal onto the surface of another material. In this process, the object to be plated serves as the cathode, while the plating metal acts as the anode, and both are immersed in an electrolyte solution containing ions of the plating metal. When current flows, metal ions from the solution are reduced and deposited onto the cathode, forming a uniform coating. The quality of electroplating depends on factors such as current density, plating time, electrolyte composition, and surface preparation. This technique is widely used to enhance appearance (e.g., gold or silver plating), improve corrosion resistance (e.g., chrome plating on steel), or increase conductivity in electronic components.

Both processes rely on redox reactions and the movement of ions under an applied voltage. While electrolysis breaks down compounds into simpler substances, electroplating uses the same principles to build up metal layers on surfaces. These processes have significant industrial applications, from metal purification and refining to manufacturing and decorative finishes. Understanding their underlying principles is essential for optimizing efficiency and achieving desired results in practical applications.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Brine	-	Appropriate amount
Concentrated sodium hydroxide solution	2 mol L^{-1}	50.0 mL
Concentrated hydrochloric acid	1 mol L^{-1}	50.0 mL
Phenolphthalein indicator	-	Small amount
Starch-potassium iodide test solution	-	Appropriate amount
Zinc chloride	Solid	3.5 g
Potassium chloride	Solid	12.5 g
Boric acid	Solid	1.5 g
Dilute nitric acid	0.01 mol L^{-1}	50.0 mL
Distilled water	-	Appropriate amount
Zinc plate	-	1
Iron plate	-	1
U-tube with branches	-	1
Beaker	150 mL	2
Single hole rubber stopper	-	2
Graphite electrode	-	1
Iron rod	-	1
Iron nail	-	1
Test tube	12 mm × 120 mm	2
Electronic balance	-	1
Thermometer	-	1
Low voltage direct current power supply	6 – 12 V	1

Sandpaper	-	1
Glass rod	-	1
Connecting wires	-	A number of
Retort stand with clamp	-	1
Forceps	-	1
Spatula	-	1

Precautionary Steps

1. The U-tube should contain just enough brine to submerge the electrodes, avoiding excess solution.
2. If a U-shaped tube is not available in the laboratory, an electrolytic tank can be used as a substitute.
3. It is recommended to use an alcohol-filled thermometer instead of a mercury thermometer for safety and environmental considerations.
4. Only refined brine should be used in electrolysis.

Procedures

A. Electrolysis of Brine

1. Support the U-tube with branches using retort stand with clamp. Fill the U-tube with brine and add small amount of phenolphthalein indicator into it.
2. Each mouth of U-tube cover by single hole rubber stopper. Insert graphite electrode and iron rod through the single hole rubber stopper as anode and cathode, respectively.
3. Connect both electrodes to the low voltage direct current power supply (6 – 12 V) using connecting wires. Observe and record the changes in the U-tube.
4. After 2 – 3 minutes of electrolysis, collect the gas produced in cathode using a test tube and ignite it. On the other hands, use the starch-potassium iodide test solution to test the gas produced in anode. Observe and record the phenomena.

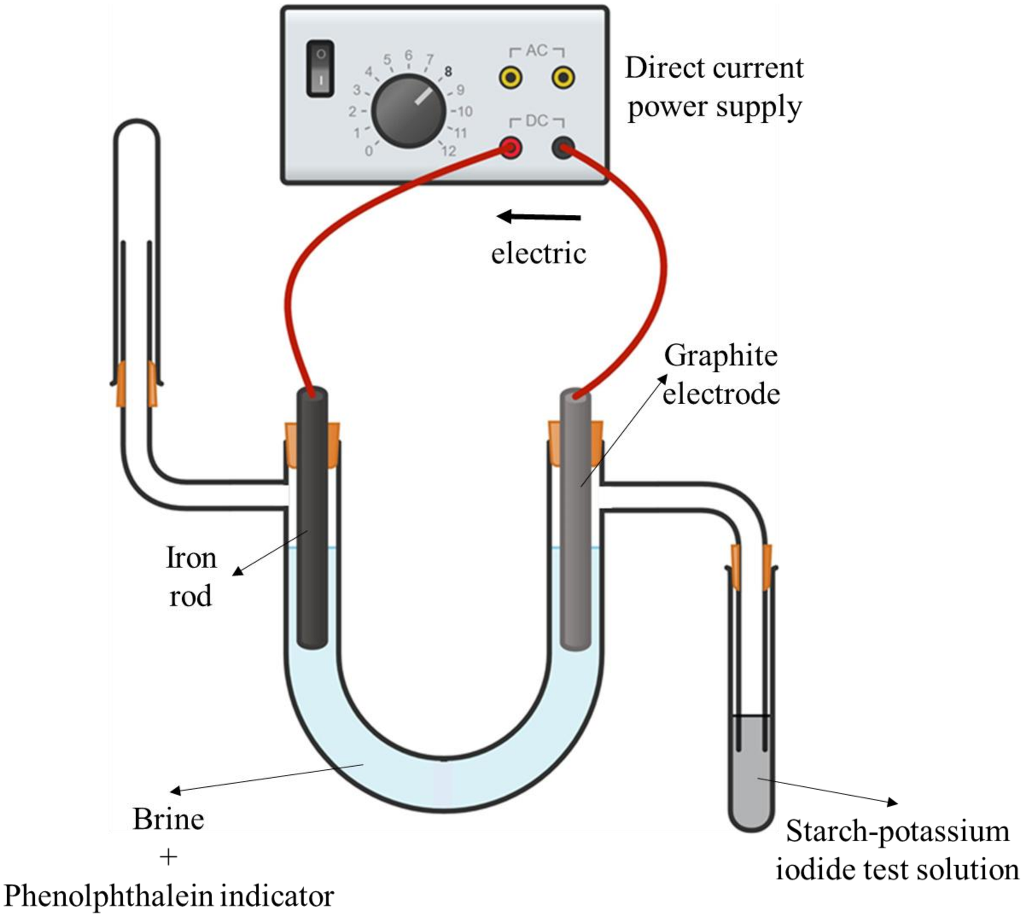


Figure 14.1: Experimental setup for the electrolysis of brine using a graphite anode and an iron cathode.

B. Zinc Electroplating of An Iron Nail

1. Add 50.0 mL of distilled water to the beaker, follow by 3.5 g of zinc chloride, 12.5 g potassium chloride and 1.5 g boric acid. Stir the mixture until dissolve and adjust the pH of the solution to 5 – 6.
2. Polish the iron nail until it surface become smooth using fine sandpaper. Use water to clean the iron nail.
3. Submerge the iron nail to the 80^{o}C of 2 mol L^{-1} sodium hydroxide solution for 5 minutes to remove oil stains. Use water to clean the submerged iron nail.
4. Then, submerge the iron nail to 45^{o}C of concentrated hydrochloric acid for 2 minutes. Clean the submerged iron nail with water.
5. Continue to submerge the iron nail to the 1:100 (volume ratio) of dilute nitric acid for 3 – 5 seconds. Clean the submerged iron nail with water.

6. Add electrolyte (Step 1) to a beaker. Use zinc plate as anode and iron nail as cathode, connect them to a 6 V direct current power supply.
7. Allow the electroplating process occur for 20 – 25 minutes. Observe and record the changes in the beaker.

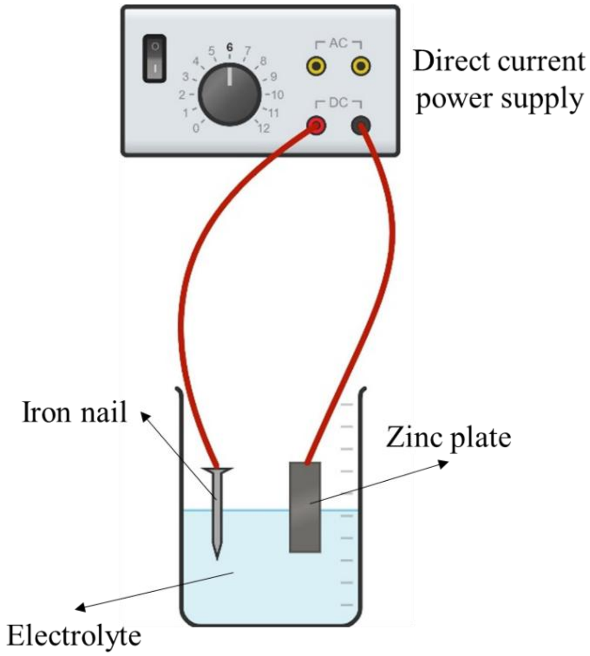


Figure 14.1: Experimental setup for the electroplating of iron nail using zinc plate.

Results

A. Electrolysis of Brine

No.	Procedures	Observations	Conclusions
1	Changes in the U-tube after connecting to the low voltage direct current power supply (6 – 12 V)		
2	The gas collected in cathode and ignite it		
3	The gas produced in anode and tested with the starch-potassium iodide test solution to test		

B. Zinc Electroplating of An Iron Nail

No.	Procedures	Observations	Conclusions
1	Changes on zinc plate		
2	Changes on iron nail		

Discussion

1. What gases are produced at the cathode and anode during the electrolysis of brine? Identify each gas and explain its formation.

2. Why brine is needed during the electrolysis? Explain your answer.

3. Why concentrated sodium hydroxide solution is used to wash iron nail before electroplating started? Explain your answer.

4. How did electrode cleanliness (e.g., sanding the nail) impact the electroplating results? Explain your answer.

5. Copper and gold could be tested for electroplating, and how might results differ in terms of their appearance and conductivity?

Chemical Properties of Aldehydes and Ketones

Problem Statement

Which chemical tests can reliably differentiate aldehydes from ketones, and what is the underlying chemical principle behind each test?

Objectives

1. To deepen the understanding of the chemical properties of aldehydes.
2. To deepen the understanding of the chemical properties of ketones.

Theory and Principle

Aldehydes and ketones are carbonyl compounds characterized by the presence of a polar C=O group, where the carbon atom carries a partial positive charge (δ^+) and the oxygen a partial negative charge (δ^-). This polarity makes the carbonyl carbon highly susceptible to nucleophilic attack, forming the basis for most reactions of these compounds. The key difference between aldehydes (R-CHO) and ketones (R-CO-R') lies in their structure: aldehydes have at least one hydrogen atom attached to the carbonyl carbon, while ketones have two alkyl or aryl groups. This structural distinction makes aldehydes generally more reactive than ketones due to reduced steric hindrance and greater electrophilicity of the carbonyl carbon.

Nucleophilic addition reactions are fundamental to both aldehydes and ketones, where nucleophiles such as Grignard reagents, cyanide ions, or alcohols attack the electrophilic carbonyl carbon. Aldehydes react more readily in these additions because their carbonyl carbon is less sterically hindered and more electron-deficient. Oxidation reactions further highlight their differences: aldehydes can be easily oxidized to carboxylic acids using mild oxidizing agents like Tollens' reagent or Fehling's solution, whereas ketones resist oxidation unless subjected to vigorous conditions that cleave carbon-carbon bonds. Reduction reactions, on the other hand, convert both aldehydes and ketones to alcohols, with aldehydes yielding primary alcohols and ketones forming secondary alcohols.

Aldol condensation, another critical reaction, occurs in compounds with α-hydrogens, leading to the formation of β-hydroxy carbonyl compounds. Additionally, aldehydes and ketones exhibit keto-enol tautomerism, where they interconvert between keto (C=O) and enol (C=C-OH) forms, with the enolate ions serving as key intermediates in further reactions like halogenation.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Copper (II) sulfate solution	2%	Small amount
Propanone	Analytical grade	5.0 mL
Ethanal	40%	Small amount
Silver nitrate solution	4%	5.0 mL
Sodium hydroxide solution	10%	10.0 mL
Bromine water	-	10.0 mL
Ammonia water	2%	Small amount
Schiff’s reagent	-	2.0 mL
Acidified potassium manganate (VII) solution	0.01 mol L^{-1}	10.0 mL
Beaker	250 mL	2
Bunsen burner	-	1
Test tube	15 mm × 150 mm	8
Dropper	-	2
Test tube holder	-	1

Precautionary Steps

1. Thoroughly clean all glassware (including inner wall) before performing the silver mirror test.
2. Use fresh silver ammonia water during the silver mirror test.
3. During the silver mirror test, avoid shaking the glassware while in the hot water bath and do not add excess ammonia water to prevent disrupting silver deposition.
4. When preparing Fehling’s solution, avoid adding excess copper (II) sulfate solution to prevent the formation of black precipitate (copper oxide) instead of the desired red precipitate (copper (I) oxide) upon heating during Fehling’s test.

Procedures

A. Silver mirror test

1. Add 1.0 mL of 4% silver nitrate solution to a clean test tube, follow by dropwise addition of ammonia solution with gentle shaking until the initial brown precipitate completely dissolves (Tollens' reagent).
2. Add 3 drops of 40% ethanal to the prepared Tollens' reagent, shake gently, and place in a 70 – 95°C water bath for few minutes. Observe and record the changes in the test tube.

B. Fehling's solution test

1. Add 2.0 mL of sodium hydroxide solution to a test tube, follow by 4 – 5 drops of copper (II) sulfate solution. Gently shake the mixture until the formation of blue suspension (Fehling's solution).
2. Then, add 0.5 mL of 40% ethanal to the prepared Fehling's solution, shake gently, and heat the mixture until boil. Observe and record the changes in the test tube.

C. Bromine water test

1. Add 2.0 – 3.0 mL of bromine water to a test tube, follow by 1 – 2 drops of ethanal and allow the mixture to stand for few minutes. Observe and record the changes in the test tube.

D. Acidified potassium manganate (VII) solution test

1. Add 2.0 – 3.0 mL of acidified potassium manganate (VII) solution to a test tube, follow by 1 – 2 drops of ethanal and allow the mixture to stand for few minutes. Observe and record the changes in the test tube.

E. Schiff's reagent test

1. Add 2.0 mL of Schiff's reagent to a test tube, follow by few drops of ethanal. Observe and record the changes in the test tube.

Repeat all the aforementioned tests using propanone without changing the amount of solutions or reagents used and the specific reactions' conditions.

Results

No.	Tests	Observations	
		Ethanal	Propanone
1	Silver mirror test		
2	Fehling's solution test		
3	Bromine water test		
4	Acidified potassium manganate (VII) solution test		
5	Schiff's reagent test		

Discussion

1. Write out the chemical equations for all reactions (except Schiff's reagent test) that produced positive test results in the experiment.

2. Why different observations were recorded between ethanal and propanone in the experiment? Explain your answer.

3. If methanoic acid, fructose and sucrose are used in the experiment, predict their observations and explain your answer.

4. Suggest why some ketones might give false positives in certain carbonyl tests if the reagent concentrations are not carefully controlled.

5. Compare the oxidation products of aldehydes and ketones. Why can aldehydes be oxidized to carboxylic acids, while ketones require stronger conditions?

Preparation of Ethyl Ethanoate

Problem Statement

How can ethyl ethanoate be prepared in the laboratory from ethanol and acetic acid?

Objectives

1. To understand the principle of esterification.
2. To master the laboratory method in preparing ethyl ethanoate.

Theory and Principle

The preparation of ethyl ethanoate is based on the principle of acid-catalyzed esterification, specifically the Fischer esterification reaction, where ethanoic acid and ethanol react reversibly in the presence of concentrated sulfuric acid. The sulfuric acid serves a dual role: it acts as a catalyst by protonating the carbonyl group of the acetic acid, enhancing its electrophilicity for nucleophilic attack by ethanol, and it functions as a dehydrating agent to absorb the water produced, thereby shifting the equilibrium toward ester formation in accordance with Le Chatelier’s principle. The mixture is heated under reflux to accelerate the reaction while preventing the loss of volatile components. After the reaction, the crude product is purified by washing with a sodium carbonate solution to neutralize any unreacted acid.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Ethanoic acid (acetic acid)	Glacial	20.0 mL
Ethanol	Glacial or 95%	25.0 mL
Concentrated sulfuric acid	98%	5.0 mL
Sodium carbonate solution	Saturated	10.0 mL
Boiling chips or ceramic fragments	-	A few
Distillation flask	250 mL	1
Tripod stand with wire gauze	-	1
Bunsen burner	-	1
Measuring cylinder	25 mL	1

Conical flask	100 mL	2
Separating funnel	100 mL	1
Thermometer	100 ^{o}C	1
Condenser	-	1

Precautionary Steps

1. When preparing ethyl acetate, the reaction mixture should not be heated to an excessively high temperature, as this can promote side reactions that form impurities.
2. To prevent bumping and ensure even boiling, add boiling chips or ceramic fragments to the flask before heating begins. Once boiling commences, immediately reduce the heat to maintain a gentle reflux.
3. Gently swirl the mixture of solution as it is exothermic and prevent splashing.
4. Concentrated acid must always be handled in a fume hood to protect against toxic fumes, exothermic reactions and chemical burns. Always wear appropriate PPE – gloves, eye protection (goggles or a face shield) and a lab coat. Acid might release hazardous vapors or gas. These fumes can cause severe respiratory irritation, lung damage or even chemical pneumonia if inhaled.
5. It is recommended to use an alcohol-filled thermometer instead of a mercury thermometer for safety and environmental considerations.

Procedures

1. Assemble a reflux apparatus (round-bottom flask + condenser in a vertical position).
2. Add 25.0 mL of ethanol to the flask. While gently swirling, slowly add 5.0 mL of concentrated sulfuric acid. Allow to cool.
3. Add 20.0 mL of glacial ethanoic acid and a few boiling chips. Swirl to mix.
4. Heat the mixture under reflux for 20 – 30 minutes. This allows the reaction to proceed at an elevated temperature without losing the volatile reactants or product.
5. Set up a distillation apparatus. Transfer the reaction mixture from the reflux flask to the distillation flask (you can use the same flask if you reassemble the apparatus).
6. Distill the mixture and collect the distillate that boils between 75 – 80 °C (the boiling point range of ethyl ethanoate). Collect this distillate in a clean, dry conical flask or a beaker.

7. Transfer the crude distillate (which contains ethyl ethanoate, water, and unreacted ethanol/acid) to a separating funnel.
8. Add 10.0 mL of saturated sodium carbonate solution to the separating funnel. Vigorous shaking with frequent venting is required to release carbon dioxide gas produced by neutralizing unreacted acid.
9. Allow the layers to separate fully. The lower aqueous layer will be the sodium carbonate solution. Drain and discard this lower layer.
10. Transfer the organic layer (ethyl ethanoate, upper layer) to a dry measuring cylinder and record its volume.
11. Calculate the percentage yield using the density provided (ethyl ethanoate = 0.9 g cm^{-3}) and the theoretical yield based on your limiting reagent.

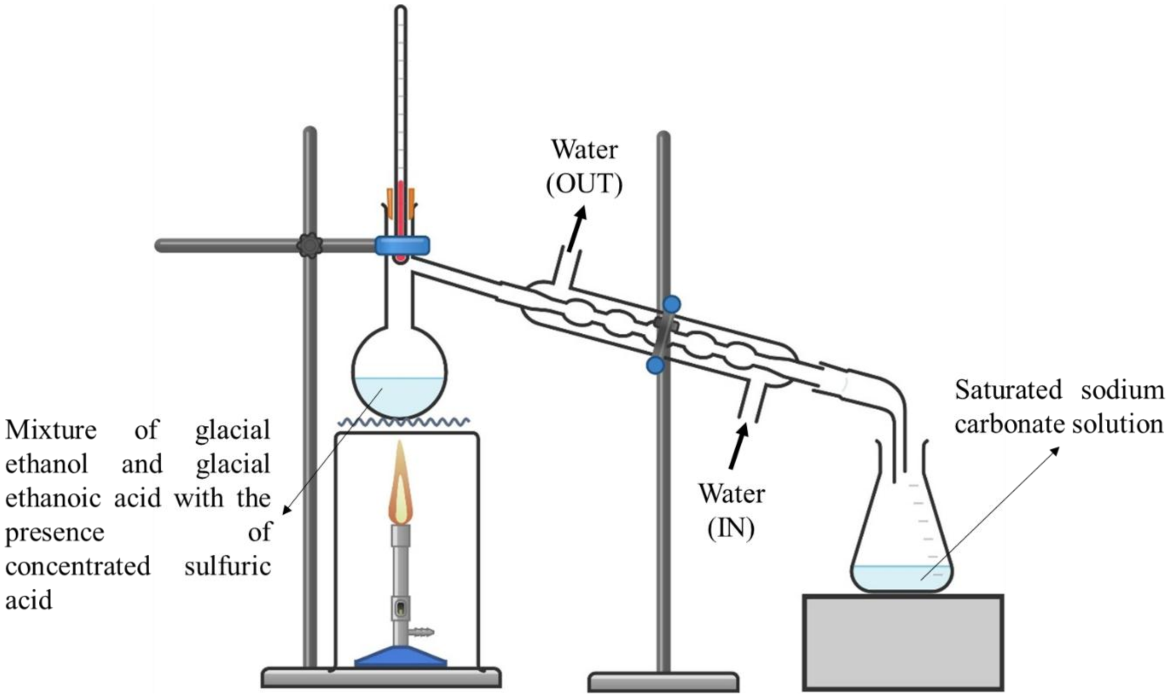


Figure 16.1: Experimental setup for ethyl ethanoate preparation.

Results

Percentage yield of ethyl ethanoate = %

Discussion

1. The reaction mechanism is described as acid-catalyzed. Using the formation of ethyl ethanoate as an example, write the balanced chemical reaction.

2. The collected ethyl ethanoate was channeled to the saturated sodium carbonate solution. Explain the purpose of this step, write out the balanced chemical equation that occurs and the observations expected.

3. The experimental yield of ethyl ethanoate is often less than the theoretical yield. Discuss two possible reasons of loss that could account for this and suggest improvements to the procedure to minimize each one.

4. Describe one chemical test (other than smell) that could be used to confirm the identity of the product as an ester. State their reagents used and the positive results you would observe.

5. In the advance experiment, the ester is washed with sodium carbonate solution and then with water. Explain why the washing step with water is necessary after the sodium carbonate wash. What impurity is being removed?

Preparation of Nylon-66

Problem Statement

How can nylon-66 be prepared in the laboratory from 1, 6-hexanediamine and adipoyl chloride?

Objectives

1. To understand the principle of polymerization through the nylon-66 preparation.

Theory and Principle

The preparation of nylon-6,6 is fundamentally governed by the principles of step-growth condensation polymerization, which requires an exact 1:1 stoichiometric balance between its two monomers (1, 6-hexanediamine and adipic acid) to achieve high molecular weight. This precise balance is ensured by first forming an intermediate, known as "nylon salt" (hexamethylenediammonium adipate), which crystallizes from a solution containing equimolar amounts of the diamine and diacid, guaranteeing equal reactivity at both ends of each molecule. In industry, the polymerization itself proceeds through melt polycondensation, where the dried nylon salt is heated to high temperatures (270 – 280 °C) under inert atmosphere and pressure. As the reaction advances, temperature increases and pressure is gradually reduced to facilitate the removal of water (the condensation by-product) driving the equilibrium toward polymer formation.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities
Sodium hydroxide	Pellet	3.20 g
1, 6-hexanediamine	-	2.32 g
Distilled water	-	100.0 mL
Deionized water	-	Suitable amount
Adipoyl chloride	-	3.66 g
Cyclohexane or Heptane	-	100.0 mL
Dilute hydrochloric acid	3%	Suitable amount
Beaker	50 mL	1
	500 mL	2

Measuring cylinder	10 mL	1
	250 mL	1
Forceps	-	1
Electronic balance	-	1
Glass rod	-	1

Precautionary Steps

1. The 1, 6-hexanediamine is corrosive. Its vapours are irritating to the eyes and upper respiratory tract. In case of skin contact, thoroughly rinse with copious amounts of running water. If swallowed, immediately rinse the mouth and drink milk or egg white. Exposure to high heat, open flames, or contact with oxidising agents may cause combustion.
2. The highly flammable solvent like cyclohexane or heptane, the essential precaution is the strict elimination of all ignition sources, including open flames, hot surfaces, and static electricity – by working in a fume hood to prevent fire or explosion.

Procedures

1. Add 100.0 mL of distilled water to dissolve 2.32 g of 1, 6-hexanediamine and 3.20 g of sodium hydroxide in a beaker. Label the beaker as A.
2. Add 100.0 mL of cyclohexane or heptane to dissolve 3.66 g of adipoyl chloride in a beaker. Label the beaker as B.
3. Slowly pour the mixture from Beaker B down the side of the glass rod into Beaker A and a translucent film can be seen forming at the interface.
4. Using a glass rod, carefully hook the polymer film at the interface and slowly, steadily draw it upward, winding the continuous strand of nylon onto the rod until no more material forms or one of the reactant layers is depleted**.**
5. Use 3% dilute hydrochloric acid to wash the polymer film, follow by deionized water to neutralize it.
6. Dry the polymer film.
7. Measure the mass obtained and calculate the percentage yield.

Results

Mass of nylon-66 = g

Percentage yield of nylon-66 = %

Discussion

1. Name the type of polymerization in the above experiment and explain your answer.

2. The synthesized nylon is often washed with dilute hydrochloric acid and then water. Explain the purpose of this acid wash?

3. Why dilute hydrochloric acid was used in the above experiment. Explain your answer.

4. Nylon was one of the first synthetic polymers to be commercialized. Based on your knowledge, give two fundamental chemical principles make such a material so useful and versatile.

5. In industrial, a small amount of acetic acid is often added to the reaction mixture. Explain its function briefly.

Experiment 1 Confirming the Copper (II) Oxide Chemical Formula Using Stoichiometric Analysis and Gravimetric Method

Hypothesis

If the copper powder is fully oxidized in air, the resulting compound's mass increase will correspond to a 1:1 Cu:O molar ratio and confirming the formula of copper (II) oxide.

Variables

Manipulated variable: The mass of oxygen

Responding variable: The chemical formula of copper (II) oxide

Controlled variable: The mass of copper powder

Discussion

1. Why might the experimental result deviate from the theoretical ratio? Explain your answer.

Reasons	Explanations
Incomplete oxidation	Causing incomplete oxidation of copper to copper (II) oxide
Insufficient heating	
Insufficient oxygen gas flow	
Copper loss due to spattering	
Moisture absorption	Might causing formation of copper hydroxide as intermediate
Overheating	Causing decomposition of copper (II) oxide into oxygen and copper metal

2. How to confirm the copper and oxygen has completely reacted to form copper (II) oxide?

Method	Observation
Add dilute hydrochloric acid	If no solid residue remains after the reaction, indicates complete conversion to copper (II) oxide

3. If the final mass decreased after heating, how would this affect the empirical formula calculation?

Result	Reason	Impact
Final mass decreased after heating	Indicates less amount of oxygen absorbed during reaction	High Cu-to-O ratio (> 1:1)

4. Why is pure oxygen gas flowed over the sample instead of relying on ambient air? How would using air affect the result?

Reasons	Impact
Ambient air has lower oxygen concentration than pure oxygen gas	Slower or incomplete reaction under ambient air
Copper is an inactive metal	

5. Can we use the similar method to determine the chemical formula of magnesium oxide? Explain your answer.

Answer	Explanation
Cannot	Magnesium is a reactive metal reacts vigorously under heating.

6. An electronic balance can measure mass up to four decimal places (e.g., 0.1234 g). When recording the mass of a single chemical sample for a laboratory experiment, how many decimal places should typically be recorded? Justify your answer.

Decimal places	Justifications
4	To preserve the complete raw data from the measurement. Any rounding or data processing should be done in subsequent calculations, not at the point of recording.
	All the digits displayed, including the last one, are considered significant figures. If you round the value to three decimal places, you would incorrectly reduce it to 3 significant figures, thereby losing the precision of the electronic balance was designed to provide.

7. When measuring the mass of a sample three times, the readings are slightly different each time. How do you determine that the mass has stabilized and can be considered constant?

Criteria	Explanations
Precision criterion	The mass can be considered constant if the variation between the three readings falls within the balance's stated precision or tolerance.
Stability assessment	The values show small, random fluctuations around a central value. This indicates the measurement has stabilized.

Conclusion

This experiment confirmed the empirical formula of copper (II) oxide with a 1:1 CuO ratio, matching theoretical ratios. Minor deviations from the ideal ratio were likely due to incomplete oxidation or moisture absorption. This shows the effectiveness of stoichiometric analysis and gravimetric method for determining chemical compositions.

Experiment 2 Confirming the Zinc Sulfate Chemical Formula Using Stoichiometric Analysis and Gravimetric Method

Hypothesis

If the copper element is fully displaced by zinc element in saturated copper (II) sulfate solution, the resulting copper powder mass increase will correspond to a 1:1 Zn:Cu molar ratio and confirming the formula of zinc sulfate.

Variables

Manipulated variable: The mass of zinc powder

Responding variable: The chemical formula of zinc sulfate

Controlled variable: The volume of saturated copper (II) oxide

Discussion

1. What is the impact on the mass of copper powder obtained if the evaporating dish inner wall is not wash with distilled water in step 4? Explain your answer.

Impact	Explanation
Mass of the final product will find larger than expected	Surface of the obtained copper powder may contaminate by the zinc sulfate crystal after drying

2. State two advantages using distilled water followed by alcohol to wash the obtained copper powder.

Washing agents	Advantages
Distilled water	Remove water soluble impurities (e.g., residual salts or acids) without introducing new impurities
Alcohol	Fasten the drying process under gentle heating with minimum oxidation

3. Why evaporating dish needed to be heated on the wire gauze with asbestos?

Reason
Heat equally during the drying process of copper powder

4. How to test the purity of copper powder obtained? Explain your answer.

Methods	Explanations
Visual inspection	Pure copper powder appears as uniform particles in reddish-brown
	Decolorization to black or green spots indicates the presence of CuO or Cu_2O, respectively; Presence of heterogenous particles suggesting non-copper contaminations
Acid test	Pure copper powder does not react with non-oxidizing acid (e.g., hydrochloric acid or sulfuric acid)
	Bubbling (H_2 gas) suggests reactive metals (e.g., Zn) as impurities
Oxide detection	Pure copper powder remains unchanged with acetic acid
	A blue-green solution forms if CuO or Cu_2O is present due to the formation of copper acetate
Heating test	Pure copper powder may darken slightly due to surface oxidation, but retains metallic properties
	Mass gain significantly suggest oxidation of impurities (e.g., Zn oxidize to ZnO); Release of colored gases (e.g., SO_2 from copper (II) sulfate solution)

5. If copper (II) sulfate solution was contaminated with sodium chloride, how would this affect the calculated formula? Explain your answer.

Impact	Explanations
No effect	Sodium chloride is a spectator ion in the reaction
	Does not directly participate in the displacement reaction (not reacting with zinc or copper)

6. During the reaction between zinc powder and copper (II) sulfate solution, describe your observations and explain the changes.

Observations	Explanations
The blue colour of the solution gradually fades and eventually becomes colourless	The blue colour is due to $[Cu(H_2O)_6]^{2+}$ in solution. As these ions are reduced to copper metal, their concentration decreases, causing the colour to fade.

A dark red-brown precipitate form	Cu^{2+} gain electrons from Zn atoms and are reduced to elemental Cu.
The zinc powder appears to dissolve or diminish in quantity	Zn atoms lose electrons to form Zn^{2+}, which enter the solution. This oxidation causes the solid zinc to gradually dissolve.
The reaction mixture may feel warm to the touch	The reaction is exothermic, releasing heat energy as Zn displaces Cu^{2+}.

7. Give an example of a displacement reaction that is applicable in daily life.

Examples
Gold mining and recycling, galvanized iron nails and roofing, cooking in iron cookware, lead pipe replacement with copper, alkaline batteries, traditional photographic development **OR** iron removal from groundwater.

Conclusion

This experiment confirmed the empirical formula of zinc sulfate with a 1:1 $ZnSO_4$ ratio, matching theoretical ratios. Minor deviations from the ideal ratio were likely due to slight oxidation when heating copper powder obtained. This shows the effectiveness of stoichiometric analysis and gravimetric method for determining chemical compositions.

Experiment 3 Electrical Conductivity of Ionic and Covalent Compounds

Hypothesis

Sodium thiosulfate pentahydrate can conduct electricity in its molten state but not in powder form, while paraffin wax cannot conduct electricity in its solid form and molten state.

Variables

Manipulated variable: Type of chemical compound used

Responding variable: Electrical conductivity of chemical compound used

Controlled variable: Electrode used in electrolysis

Discussion

1. Based on your result, how does the bonding type influence a compound's ability to conduct electricity?

Type of bonding	Conductivity
Ionic compounds or ionic bond	Can conduct electricity when molten or dissolved because their ions become mobile charge carriers, but remain non-conductive as solids where ions are fixed in place.
Covalent compounds or covalent bond	Cannot conduct in any state because their electrons are tightly bound in molecular bonds with no free ions or electrons.

2. How did the conductivity change when sodium thiosulfate pentahydrate was tested in powder and molten state? Explain your answer.

State	Conductivity
Powder form or solid	Sodium thiosulfate pentahydrate is in ionic lattice structure whereby the Na^+ and $S_2O_3^{2-}$ are bonded strongly by ionic bond. Consequently, the ions cannot move freely and only can vibrate in place. Hence, sodium thiosulfate pentahydrate in powder form cannot conduct electricity.
Molten	Ionic structure of sodium thiosulfate pentahydrate has been broken down releasing free moving Na^+ and $S_2O_3^{2-}$. The mobile ions can move towards oppositely charged electrodes. Therefore, sodium thiosulfate pentahydrate in molten state can conduct electricity.

3. In another experiment, hydrochloric acid was used for replacing paraffin wax and found that hydrochloric acid was able to conduct electricity after dissolved in water. Explain you answer.

Explanation
Due to hydrochloric acid undergoes ionization producing free moving ions that can act as charge carriers.

Conclusion

Ionic compounds can conduct electricity in molten state but not in solid state (powder form), while covalent compounds do not conduct electricity in all states.

Experiment 4 Solubility of Ionic and Covalent Compounds in Water and Organic Solvent

Hypothesis

Sodium chloride is soluble in water but insoluble in cyclohexane. On the other hands, hexane is insoluble in water but soluble in cyclohexane.

Variables

Manipulated variable: Type of chemical compound used

Responding variable: Solubility of chemical compound used in water and cyclohexane

Controlled variable: Mass of sodium chloride used **OR** Volume of hexane, water and cyclohexane used

Discussion

1. Based on your observations, which types of compounds were soluble in water? Explain you answer using polarity and bonding concepts.

Observation	Explanations	
	Polarity concept	Bonding concept
Ionic compound is soluble in water	The positively charged or δ^+ hydrogen atoms in water molecules attract anions, while the negatively charged or δ^- oxygen atoms attract cations.	Strong ion-dipole interactions overcome the ionic compound's lattice energy, allowing the solid to dissolve.

2. Why water is known as “universal solvent” for ionic compounds except for covalent compounds?

Explanations	
Ionic compounds	Covalent compounds
Polar molecules effectively dissolve ionic compounds by forming strong ion-dipole interactions, whereby water's positively charged or δ^+ hydrogens attract anions, while its negatively charged or δ^- oxygen attracts cations, overcoming the ionic lattice energy.	Non-polar covalent compounds lack of charge separation and cannot form strong interactions with water molecules.

3. In another experiment, ethanol was used to dissolve both sodium chloride and hexane. Predict your observation and explain your answer.

Prediction	Explanations
Sodium chloride dissolved poorly in ethanol	Polarity of ethanol is insufficient to overcome sodium chloride strong ionic lattice energy.
Hexane miscible with ethanol	Weak London dispersion forces between its non-polar hydrocarbon chain and hexane.

4. What is the reliable method to separate and recycle the mixture of cyclohexane and hexane? Justify your answer.

Method	Justifications
Fractional distillation	The difference in boiling points between cyclohexane (≈ 80.7 °C) and hexane (≈ 68.7 °C). The 12 °C difference is adequate for separation using a fractionating column, which provides multiple vaporization-condensation cycles.

Conclusion

Ionic compounds are soluble in water, while covalent compounds are soluble in organic solvents.

Experiment 5 Melting and Boiling Points of Ionic and Covalent Compounds

Hypothesis

Sodium chloride has higher melting point and boiling point than paraffin wax.

Variables

Manipulated variable: Type of chemical compound used

Responding variable: Melting and boiling points of chemical compound

Controlled variable: Mass of sodium chloride and paraffin wax used

Discussion

1. How could you experimentally confirm whether a solid compound is ionic or covalent based on melting and boiling point tests?

Chemical compounds	Temperature required	Explanations
Ionic compound	Higher temperature	Due to strong electrostatic forces in their lattice structure
Covalent compound	Lower temperature	Due to weaker intermolecular forces

2. Magnesium oxide was used in another experiment and found that it has higher melting point than sodium chloride. Explain the phenomenon.

Explanations	
Charge density	Ionic size
Magnesium oxide contains Mg^{2+} and O^{2-}, sodium chloride contains Na^+ and Cl^-	Magnesium oxide has smaller ionic size than sodium chloride
Magnesium oxide has greater charge than sodium chloride leads to stronger electrostatic attraction, increasing the energy required to break down its ionic lattice	Smaller ions allow closer packing, enhancing the Coulombic forces and increasing lattice energy

3. Water and hydrogen sulfide are both covalent compounds, why water has a much higher boiling point (100 °C) than hydrogen sulfide (-60 °C)? Explain your answer.

Explanations	
Attraction force	Molecular size and shape
Water forms hydrogen bonds due to: • High electronegativity of oxygen which strongly attracts shred electrons • Polar O-H bonds creating large dipole moment • Lone pairs on oxygen that interact with hydrogen atoms of neighboring molecules	Although hydrogen sulfide is larger and heavier than water, the hydrogen bonding in water overwhelmed the effect
Weak dipole-dipole forces in hydrogen sulfide due to: • Sulfur is less electronegativity than oxygen • S-H bonds are less polar • Primarily relies on weaker dipole-dipole forces ad London dispersion forces (van der Waals) • No hydrogen bonding (larger sulfur size and low electronegativity)	

Conclusion

Ionic compounds have higher melting and boiling points, while covalent compounds have lower melting and boiling points.

Experiment 6 Qualitative Analysis of Selected Metal Cations Using Chemical Test

Hypothesis

If characteristic chemical tests are applied to aqueous solutions containing Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+}, and Ag^{+}, then each metal cation will produce distinct visual changes due to their unique chemical properties, allowing for accurate identification.

Variables

Manipulated variable: Type of reagents used

Responding variable: Observed reactions

Controlled variable: Concentration of metal cations **OR** Reagent purity used

Discussion

1. The metal cations Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+}, and Ag^{+} can be identified using sodium hydroxide solution as a reagent. State the observations for each metal cations.

Metal cations	Observations
Fe^{2+}	Light green solution initially forms a white precipitate, which rapidly transitions to a greyish-green intermediate and finally settles as a reddish-brown precipitate
Fe^{3+}	Yellow solution forms reddish brown precipitate
Zn^{2+}	Colourless solution forms white precipitate but it dissolves in excess sodium hydroxide solution
Cu^{2+}	Blue solution forms blue precipitate
Ag^{+}	Colourless solution forms white precipitate

2. If your test for Fe^{3+} with acidified potassium thiocyanate showed a faint red color instead of deep red, what could be the cause?

Causes	Explanations
Low Fe^{3+} concentration	The sample may contain trace amounts of Fe^{3+} below the test's detection limit
Insufficient of acidification	The blood-red $[Fe(SCN)]^{2+}$ complex forms best in acidic conditions
Diluted or expired reagent	Old or overly diluted potassium thiocyanate may produce weak color
Redox interference	Reducing agents may convert $Fe^{3+} \rightarrow Fe^{2+}$, thus reducing color intensity
Competing ions	PO_4^{3-} or F^- can mask Fe^{3+} by forming colorless complexes

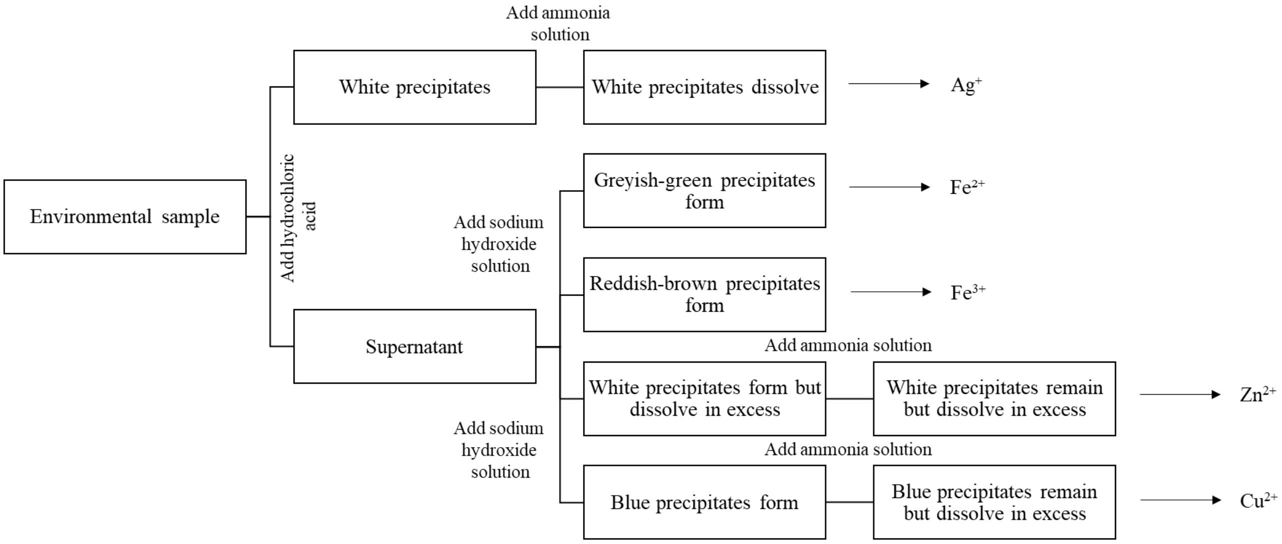
3. Evaluate the reliability of using sodium hydroxide solution and ammonia solution to distinguish between Zn^{2+} and Cu^{2+}. Which test is more definitive and why?

Reliability and Definitive	
Sodium hydroxide	Ammonia solution
Lower reliability and definitive • Require adding it in excess in dropwise • Other amphoteric ions, such as Al^{3+} and Pb^{2+}, may mimic Zn^{2+}	Higher reliability and definitive • Distinct colour changes provides unambiguous identification • Less interference from other ions compared to sodium hydroxide solution

4. Why did Ag^+ form a white precipitate with hydrochloric acid but dissolve in ammonia solution, while Cu^{2+} formed a blue complex with ammonia solution?

Metal cations	Explanations	
	Hydrochloric acid	Ammonia solution
Ag^+	• Forms white precipitate (silver chloride) due to low solubility product • Silver chloride lattice energy exceeds its hydration energy, making it insoluble	• Silver chloride dissolves to form colourless complex $\{[Ag(NH_3)_2]^+\}$ • Ammonia solution ligands stabilize Ag^+ through strong dative bonding (high stability constant)
Cu^{2+}	• Cu^{2+} do not reacts with hydrochloric acid due to copper is less reactive than hydrogen • Thus, cannot copper displace hydrogen	• Forms blue precipitate [copper (II) hydroxide] initially, then dissolves in excess ammonia solution to form deep blue complex $\{[Cu(NH_3)_4]^{2+}\}$ • Cu^{2+} d^9electronic configuration favours tetraammine coordination, absorbing light in the orange region, thus it appears in blue

5. Design a flowchart for an laboratory to separate and identify these metal cations in a mixed sample.



Conclusion

This experiment successfully demonstrated the identification of Fe^{2+}, Fe^{3+}, Zn^{2+}, Cu^{2+} and Ag^{+} through distinct chemical reactions, leveraging principles of precipitation, complexation and redox chemistry.

Experiment 7 Qualitative Analysis of Selected Anions Using Chemical Test

Hypothesis

If characteristic chemical tests are applied to solid or aqueous solutions containing S^{2-}, SO_4^{2-}, SO_3^{2-}, CO_3^{2-} and NO_3^-, then each anion will produce distinct visual changes due to their unique chemical properties, allowing for accurate identification.

Variables

Manipulated variable: Type of reagents used

Responding variable: Observed reactions

Controlled variable: Concentration of anions **OR** Reagent purity used

Discussion

1. How did the addition of concentrated hydrochloric acid or concentrated nitric acid help to differentiate between SO_4^{2-} and SO_3^{2-} in your tests? Support your answer with observed results.

Anions	Observations	Explanations
SO_4^{2-}	No reaction	Hydrochloric acid or nitric acid cannot protonate sulfate further
SO_3^{2-}	Bubbling and pungent gas release	Hydrochloric acid or nitric acid protonate sulfite, releasing sulfur dioxide

2. If an unknown sample gave positive results for both S^{2-} and SO_3^{2-}, how would you resolve this ambiguity?

Anions	Methods	Observations
SO_3^{2-}	Deliver sulfur dioxide to acidified potassium dichromate (VI) solution	Orange solution tuns green
	Deliver sulfur dioxide to acidified potassium manganate (VII) solution	Purple solution decolorizes
S^{2-}	Deliver hydrogen sulfide to lead (II) acetate solution	Forms black solid (lead sulfide)

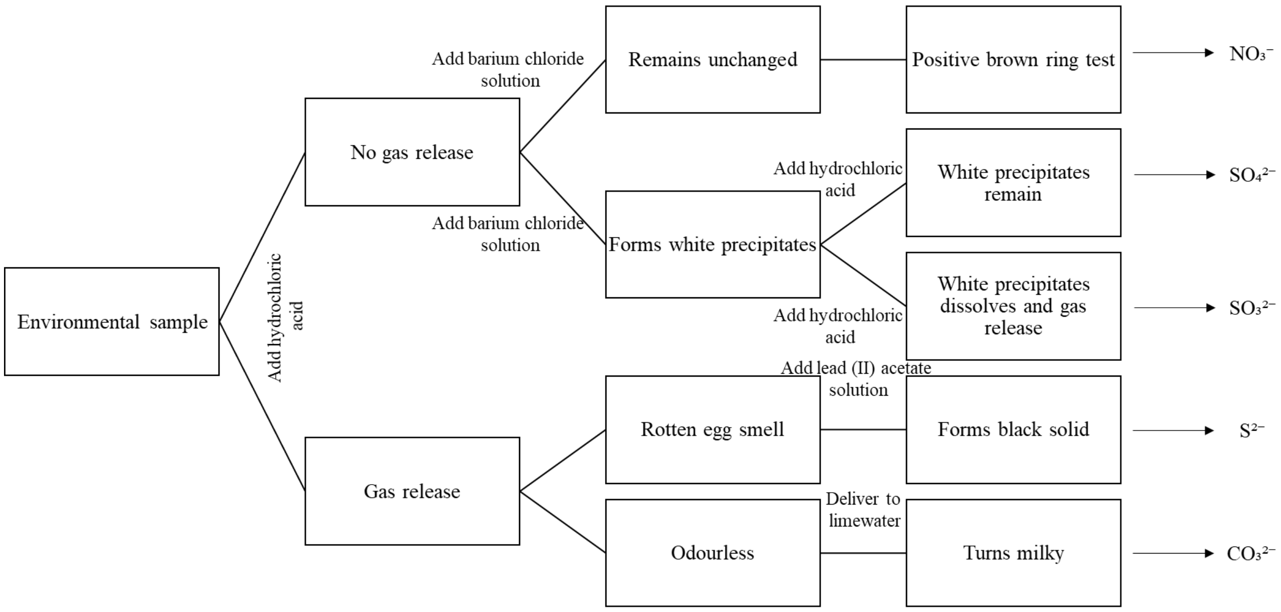
3. When testing for CO_3^{2-} with acid, why was it important to use limewater as confirmatory test rather than relying solely on effervescence?

Reasons	Explanations
False positives	Other anions can also produce bubbles when treated with hydrochloric acid
No gas identification	Bubbles could be carbon dioxide, hydrogen sulfide or even air trapped in the sample
Unique of carbon dioxide	Only carbon dioxide forms insoluble white precipitate with limewater

4. The brown ring formed but disappeared immediately in some students' test. What does it indicate about the sample or technique?

Causes	Explanations
Improper layering technique	Adding concentrated sulfuric acid too quickly or shaking the test tube, causing the brown ring disperses before stabilizing
NO_2^- interference	NO_2^- is a stronger reductant reacts faster than NO_3^-, producing fleeting nitrogen oxide
Degraded reagents	Old iron (II) sulfate solution has insufficient Fe^{2+} to reduce nitrate ions into nitric oxide, which will further react with iron (II) sulfate solution to form brown ring
	Dilute sulfuric acid has insufficient acidity to form brown ring
Low NO_3^- concentration	Trace NO_3^- produces a faint ring that dissipates

5. Design a flowchart for an laboratory to separate and identify these anions in a mixed sample.



Conclusion

This experiment successfully demonstrated the identification of S^{2-}, SO_4^{2-}, SO_3^{2-}, CO_3^{2-} and NO_3^- through distinct chemical reactions, leveraging principles of precipitation, complexation and redox chemistry.

Experiment 8 Identification of an Unknown Chemical Compound Using Qualitative Analysis

Hypothesis

If an unknown solution (NH_4^+, Al^{3+}, Mg^{2+}, SO_4^{2-} and Cl^-) is subjected to systematic qualitative analysis – including tests for solubility, flame color, precipitation reactions, and gas evolution – then its identity can be determined by matching observed chemical and physical properties to known reference data, because each ion or functional group exhibits unique and predictable reactivity patterns.

Variables

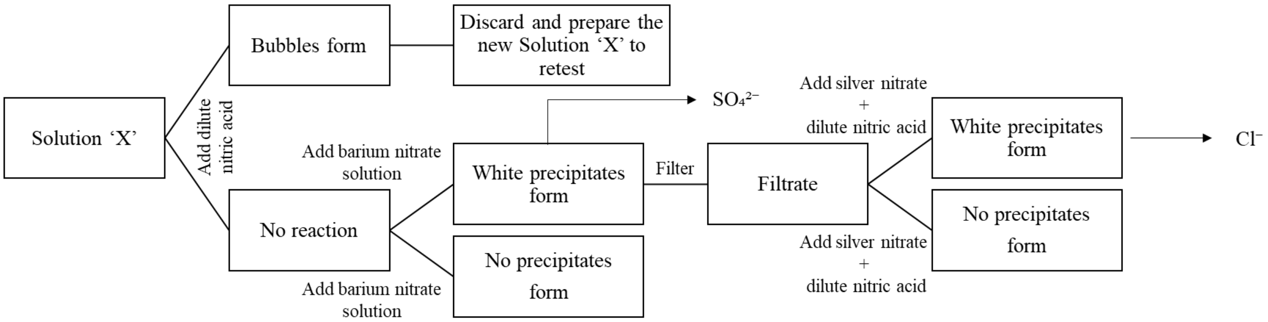
Manipulated variable: Type of reagents used

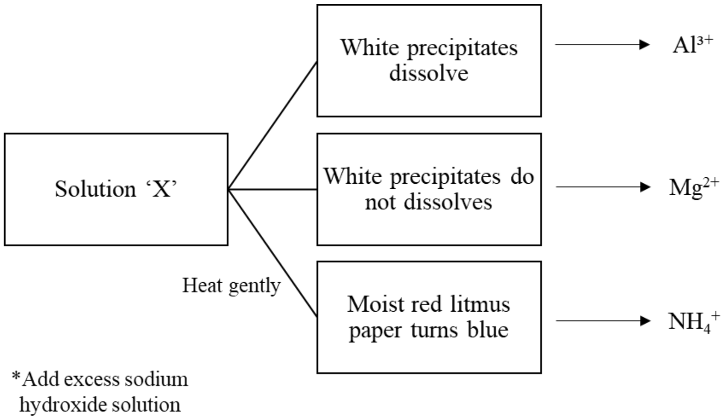
Responding variable: Observed reactions

Controlled variable: Unknown Solution 'X'

Discussion

1. Design a flowchart to separate and identify these five ions in a mixture.





2. Why was it critical to test for NH_4^+ first before analyzing other cations (Al^{3+} and Mg^{2+})? How would ammonia gas interfere with later tests?

Reasons	Explanations
Release of ammonia gas	Ammonia gas could interfere with later tests for Al^{3+} and Mg^{2+} by forming complex ions masking their presence, leading to false negative result

3. Explain the purpose of adding excess sodium hydroxide solution when testing for Al^{3+} and Mg^{2+}. How did their reactions differ?

Reasons	Explanations
Confirms amphoterism	• Al^{3+} forms aluminium hydroxide which dissolves in excess sodium hydroxide solution due to its amphoteric nature • Mg^{2+} forms magnesium hydroxide which is not amphoteric and remains insoluble
Eliminate false positives	Excess sodium hydroxide solution ensures complete precipitation and subsequent dissolution of Al^{3+}, avoiding partial reactions

4. When testing for SO_4^{2-}, why was dilute nitric acid added to the barium nitrate solution test? What would happen if you omitted the acid?

With dilute nitric acid

Reasons	Explanations
Eliminates interfering ions	• CO_3^{2-} and SO_3^{2-} can form white precipitates with Ba^{2+} • Dilute nitric acid acidifies the medium converting CO_3^{2-} and SO_3^{2-} to carbon dioxide and sulfur dioxide, respectively • Only SO_4^{2-} forms barium sulfate which is insoluble in acids
Confirms specificity	Only barium sulfate present under acidic condition, confirming SO_4^{2-} unambiguously

Omit dilute nitric acid

Reasons	Explanations
False positives	The formation of white barium carbonate and barium sulfite precipitates could form, mimicking barium sulfate, leading to incorrect SO_4^{2-} identification
Inconclusive results	Cannot distinguish between CO_3^{2-}, SO_3^{2-} and SO_4^{2-} based on the white precipitate solely

5. Your unknown solution gave:
   a. No gas with sodium hydroxide solution (ruling out NH_4^+)
   b. White precipitate with sodium hydroxide solution (insoluble in excess)
   c. White precipitate with dilute nitric acid added to the barium nitrate solution

   What is the likely identity? Justify step-by-step.

Compound identity	Explanations
Magnesium sulfate	No gas release when heating unknown solution with sodium hydroxide indicating absence of NH_4^+
	Formation of white precipitate in the addition of sodium hydroxide solution indicating the presence of Al^{3+} and Mg^{2+}
	However, the white precipitate dissolves in the excess of sodium hydroxide solution indicating the presence of Mg^{2+} due to magnesium hydroxide non-amphoteric properties
	SO_4^{2-} is acid stable salt which would not dissolves in dilute nitric acid

Conclusion

This experiment identified an unknown compound through qualitative analysis by testing its reactions with specific reagents. The systematic approach, based on solubility rules and amphoteric behavior, confirmed the compound identity through distinct observations like precipitate formation and gas evolution.

Experiment 9 Characterization of Carbon Compounds Using Sodium Carbonate and Sodium Bicarbonate

Hypothesis

If an organic compound is a sodium carbonate or sodium bicarbonate, then its solubility in water, thermal stability and reaction with dilute hydrochloric acid can be used to characterize it.

Variables

Manipulated variable: Type of carbonate and bicarbonate salt used

Responding variable: Observed reactions (water solubility, thermal stability and reaction with dilute hydrochloric acid)

Controlled variable: Mass of carbonate and bicarbonate salt used (water solubility and thermal stability), **OR** Molarity of carbonate and bicarbonate salt used **OR** Concentration of dilute hydrochloric acid used (reaction with dilute hydrochloric acid)

Discussion

1. Why different observations occur when sodium carbonate and sodium bicarbonate dissolves in water?

Observation	Explanation
Sodium carbonate more soluble than sodium bicarbonate	CO_3^{2-} in sodium carbonate are more strongly hydrated (surrounded by water molecules) than HCO_3^- ions in sodium bicarbonate

2. Why did concentrated hydrochloric acid produce effervescence with both sodium carbonate and sodium bicarbonate?

Carbonate and bicarbonate salts	Explanations
Sodium carbonate	CO_3^{2-} react with H^+ from concentrated hydrochloric acid, forming carbonic acid, which instantly decomposes into carbon dioxide gas and water $Na_2CO_{3(s)} + 2HCl \rightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Sodium bicarbonate	HCO_3^- directly react with H^+, releasing carbon dioxide gas faster than sodium carbonate because it requires only 1 proton $NaHCO_{3(s)}+HCl \rightarrow NaCl_{(aq)}+H_2O_{(l)}+CO_{2(g)}$

3. How does the p*K*a of a compound determine whether it reacts with sodium bicarbonate or only sodium carbonate?

Acidity	Explanations
p*K*a ~ 4 – 5 (e.g., carboxylic acids)	React with both sodium carbonate and sodium bicarbonate, producing carbon dioxide gas because they are strong enough to protonate CO_3^{2-} and HCO_3^-
p*K*a ~ 10 (e.g., phenols)	React only sodium carbonate, forming phenoxide without carbon dioxide gas release, since they are too weak to protonate HCO_3^- but can react with the stronger base, CO_3^{2-}
p*K*a ~ 16 and above (e.g., alcohols)	Show no reaction with either reagent because their low acidity to protonate CO_3^{2-} or HCO_3^-

4. Which carbonate or bicarbonate salts used in the experiment is a more selective reagent in identifying acids? Explain your answer.

Selective reagent	Explanation
Sodium bicarbonate	Reacts only with strong acids, such as carboxylic acids, which can protonate its weak base, HCO_3^- to release carbon dioxide gas (produce effervescence)

5. How could you using sodium carbonate and sodium bicarbonate to distinguish between aliphatic and aromatic carboxylic acids? *(Related to Organic Chemistry – Carboxylic acids)*

Tests	Aliphatic carboxylic acids	Aromatic carboxylic acids
React with sodium carbonate	Rapid effervescence	Rapid effervescence
React with sodium bicarbonate		Slow or no effervescence

Conclusion

This experiment shown that sodium carbonate dissolved easily and resisted to heat, but sodium bicarbonate dissolved less and decomposed when heated. Both produced carbon dioxide gas with dilute hydrochloric acid.

Experiment 10 Preparation and Characterization of Sulfur Dioxide Gas

Hypothesis

If sulfur dioxide gas is prepared by the reaction of sodium sulfite with sulfuric acid, then the gas can be effectively characterized by its pungent odor, acidic properties and colour change of reagents.

Variables

Manipulated variable: Mass of sodium sulfite used

Responding variable: Odor smell and colour change of chemical reagents

Controlled variable: Concentrations of the materials used

Discussion

1. What is the role of sulfur dioxide solution as an oxidizing and reducing agent in the **Results** (3) – (6) reactions? Provide balanced chemical equations.

Oxidizing agent	Observations	Chemical equations
With saturated hydrogen sulfite solution	Forms yellow or milky white precipitate	$SO_2 + 2H_2S \rightarrow 3S + 2H_2O$

Reducing agent	Observations	Chemical equations
With acidified potassium permanganate solution	Purple acidified potassium manganate (VII) solution decolorizes completely	$5SO_2 + 2KMnO_4 + 2H_2O \rightarrow K_2SO_4 + 2MnSO_4 + 2H_2SO_4$
With dilute iron (III) chloride solution	Yellowish-brown solution becomes pale green	$SO_2 + 2FeCl_3 + 2H_2O \rightarrow 2FeCl_2 + H_2SO_4 + 2HCl$
With barium chloride solution, follow by hydrogen peroxide and dilute hydrochloric acid	Forms white precipitate in dilute hydrochloric acid	H_2SO_3 (from sulfur dioxide solution) $+ H_2O_2 \rightarrow H_2SO_4 + H_2O$ $BaCl_2 + H_2SO_4 \rightarrow BaSO_4$ (white precipitate) $+ 2HCl$

2. Can the reagents below use to selectively eliminate sulfur dioxide gas in the mixture with carbon dioxide gas? Explain your answer.
   a) Sodium hydroxide solution b) Sodium carbonate solution c) Barium chloride solution

   d) Acidified potassium manganate (VII) solution

Elimination of sulfur dioxide gas	Reagents	Explanations
Can	Acidified potassium manganate (VII) solution	Only sulfur dioxide gas reacts and decolonization of the reagent occurs indicate the sulfur dioxide gas absorption
Cannot	Sodium hydroxide solution	Can react with both sulfur dioxide gas and carbon dioxide gas
	Sodium carbonate solution	
	Barium chloride solution	

3. A student accidentally put the petal in the gas bottle containing sulfur dioxide gas. Predict and explain your observation.

Observation	Explanations
Petal's colour fades or turn white over time	• Sulfur dioxide gas acts as reducing agent and bleaching agent • Reacts with coloured organic pigments in the petal, breaking their conjugated double bonds • The bleaching occurs through reduction
Petal becomes limp and shriveled	• Sulfur dioxide dissolves in moisture on the petal, forming sulfurous acid • The formed acid damages cell membrane, causing water loss and wilting
Petal may turn brown or translucent	Denatures proteins and destroys cellular structure, leading to decay

4. How does the sulfur dioxide gas contribute to acid rain? Write out your answer with relevant equations.

Sulfur dioxide contributes to acid rain	Explanations	Chemical equations
Emissions of sulfur dioxide gas into the atmosphere	Burning of fossil fuels, industrial processes and vehicles	$S + O_2 \rightarrow SO_2$
Oxidation of sulfur dioxide gas to sulfur trioxide gas	In the atmosphere, sulfur dioxide gas reacts with oxygen and hydroxyl radicals	$2SO_2 + O_2 \rightarrow 2SO_3$ $SO_2 + OH\bullet \rightarrow HOSO_2\bullet$ $HOSO_2\bullet + O_2 \rightarrow HO_2\bullet + SO_3$
Formation of sulfuric acid	Sulfur trioxide dissolves in water vapour (cloud or rain) to form sulfuric acid	$SO_3 + H_2O \rightarrow H_2SO_4$
	Sulfur dioxide directly in water to form sulfurous acid, which further undergo oxidation	$SO_2 + H_2O \rightleftharpoons H_2SO_4$ $H_2SO_3 + \frac{1}{2}O_2 \rightarrow H_2SO_4$
Acid rain falls to earth	The sulfuric acid dissociates in rainwater, lowering its pH	$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

5. When barium chloride is added dropwise to sulfur dioxide solution with the presence of hydrogen peroxide, formation of white precipitate is observed. The white precipitate does not dissolve in dilute hydrochloric acid. Explain the observation and suggest the product name.

Reaction	Explanation	Product name
Reacts with water	Sulfur dioxide dissolves in water to form sulfurous acid, which partially dissociates into HSO_3^- (bisulfite ion) and SO_3^{2-} (sulfite ion)	Barium sulfate
Presence of hydrogen peroxide	Hydrogen peroxide act as an oxidant in oxidizing SO_3^{2-} to SO_4^{2-} (sulfate ion)	
Reacts with barium chloride solution	SO_4^{2-} reacts with Ba^{2+} and forms barium sulfate	
Reacts with dilute hydrochloric acid	Barium sulfate does not dissolve in dilute hydrochloric acid due to its low solubility of product and SO_4^{2-} is a weak base which does not react with H^+ from dilute hydrochloric acid	

Conclusion

This experiment involved the preparation of sulfur dioxide gas by reacting sodium sulfite with concentrated sulfuric acid, followed by systematic characterization tests to confirm its identity and properties.

Experiment 11 Determination of the Heat of Neutralization

Hypothesis

If a strong acid and a strong base are neutralized in aqueous solution, then the experimentally determined heat of neutralization will approximate the theoretical value of -57.1 kJ mol^{-1}.

Variables

Manipulated variable: Type of acid and base used

Responding variable: Heat of neutralization between acid and base

Controlled variable: Volume of acid and base

Discussion

1. Why is it important to use equal volumes of acid and base solutions in the experiment?

Reasons	Explanations
Ensures comparable heat capacity	Equal volumes (assuming similar densities) mean both solutions contribute equally to the heat capacity (*mc*) of the system.
Maintains stoichiometric balance	Equal volumes of concentration solutions ensure moles of H^+ = moles of OH^-, preventing excess unreacted ions that could alter heat measurements.
Minimizes experimental error	Unequal volumes could lead to incomplete mixing or uneven heat distribution, increasing uncertainty in Δt.

2. How would the temperature change differ if using 1.0 mol L^{-1} of hydrochloric acid instead of 0.5 mol L^{-1} in same volume?

Changes	Explanations
Heat of neutralization remains unchanged	Because it is an intensive property for strong acid and strong base reaction.
Greater Δt	Twice as many H^+ ions react with OH^-, releasing double the heat energy (since heat if neutralization is per mole of H_2O formed).

3. Why the heat of neutralization between weak acids and weak bases are lower than of strong acids and strong bases?

Reasons	Explanations
Ionization energy cost	• Weak acids and weak bases only partially dissociate in water. • They require energy input to break bonds and release H^+ and OH^-, respectively. • Energy consumed in ionization reduces the net heat released during neutralization.
Net heat of neutralization	Heat of neutralization of weak acids and weak bases is the sum of endothermic ionization (energy absorbed to form H^+ or OH^-) and exothermic neutralization.

4. Slightly excessive of base concentration was used in the experiment rather than acid. Explain your answer.

Reasons	Explanations
Ensure complete neutralization	Guarantees all H^+ from the acid react completely
Experimental precision	Compensates for potential measurement errors in volume or concentration
Clear endpoint identification	• Makes the solution slightly basic at completion • Helps verify neutralization when using pH indicators
Accurate heat measurement	• Ensures all acid contributes to the heat measured • Prevents underestimation of heat of neutralization due to unreacted acid

5. If you repeated the experiment with 100 mL of 0.5 mol L^{-1} hydrochloric acid and 100 mL of 0.55 mol L^{-1} sodium hydroxide solution, how would the temperature change differ? Explain your answer.

Change	Explanation
Heat of neutralization remains unchanged	Because it is an intensive property for strong acid and strong base reaction.

Conclusion

This experiment determined the heat of neutralization between hydrochloric acid and sodium hydroxide solution to be -57.1 kJ mol^{-1}, matching theoretical expectations for strong acid-base reactions.

Experiment 12 Factors Affecting Rate of Reaction

Hypothesis

If a higher concentration of acid, a higher temperature, and smaller calcium carbonate sizes are used, then they will all lead to a faster rate of reaction.

Variables

Manipulated variable: Concentration of hydrochloric acid, temperature of water bath and size of calcium carbonate used

Responding variable: Rate of reaction

Controlled variable: Mass of calcium carbonate and volume of hydrochloric acid

Discussion

1. a) What does the slope (gradient of the tangent to the curve at any point represent?

Definition
The slope (gradient) of the tangent to the curve at any specific point represents the instantaneous reaction rate at that exact moment in time.

b) Why is the initial slope considered the most reliable measure of the initial reaction rate?

Reasons	Explanations
Known concentrations	At the start, reactant concentrations are exactly as prepared and at their maximum, ensuring a direct link between the measured rate and the initial conditions.
No complications	It avoids confounding factors that develop later in the reaction (e.g. Decreasing reactant concentrations, buildup of products that might interfere and potential side reactions).
Fair comparison	Using the initial slope allows for a direct and valid comparison between different experiments (e.g., different concentrations), as it captures the rate when the only intended variable differs.

c) Why do both graph of high acid concentration and low acid concentration eventually level off at the same final volume?

Reasons	Explanations
Limiting reactant is the same	Both reactions stop when the limiting reactant—usually the calcium carbonate is completely used up. Since the same mass of calcium carbonate is used in both experiments, it produces the same theoretical maximum amount of carbon dioxide gas. The acid concentration affects only the rate of the reaction, not the total amount of product formed once the limiting reactant is exhausted.
Mass of calcium carbonate is the same	
Total amount of product is determined by the limiting reactant	

2. What were the main sources of error or uncertainty in this experiment? Classify them into random errors or systematic errors.

Random errors

Random errors	Main sources
Timing inconsistencies	Human reaction time in starting/stopping the stopwatch
	Slight variations in reading the gas syringe at exact time intervals
Gas syringe friction	Fluctuating friction in the syringe plunger, causing "sticking" and uneven movement
Marble chip surface irregularities	Natural variations in marble chip porosity and surface texture affecting reactivity
Minor temperature fluctuations	Ambient temperature changes during the experiment influencing reaction rate

Systematic errors

Systematic errors	Main sources
Gas syringe calibration	A zero error (e.g., syringe not starting at exactly 0 mL) affecting all volume readings
Acid concentration drift	Evaporation of water from hydrochloric acid over time, gradually increasing concentration

Gas leaks	Small leaks in the apparatus (bung, tubing) causing consistent under-recording of gas volume
Incomplete reaction	Reactions stopping before full completion due to practical constraints (e.g., ending measurements too early)

3. Based on your experimental data for the reaction between calcium carbonate and hydrochloric acid, compare the effects of acid concentration and particle size on the reaction rate. Which factor demonstrated a more significant influence? Justify your answer.

Answer	Justification
Particle size	• This is because decreasing particle size exponentially increases the surface area available for collisions. • Powder provides millions of active sites for H^+ ions to attack simultaneously. • While increasing concentration also increases collision frequency, it only adds more particles to react with a still-limited surface area.

4. Suggest two specific improvements to the experimental procedure or apparatus that would increase the reliability or accuracy of the results. For each, explain how it would lead to better data.

Improvements	Explanations
Use a water bath for temperature control	• Maintaining a constant temperature ensures that the kinetic energy of the particles is identical for every repeat of an experiment. • This eliminates a major source of random error, leading to more consistent and reproducible results when testing the other variables.
Use a magnetic stirrer for consistent mixing	• Constant stirring maintains a uniform concentration of H^+ ions around the solid reactant and rapidly removes the product layer. • This ensures the reaction proceeds at a rate that is truly dependent on the bulk concentration and temperature, not on diffusion limitations, leading to more consistent and reliable rate data.
Use a more concentrated acid stock solution for precise dilutions	• Using precise volumetric glassware for dilution ensures that the acid concentrations are known with high accuracy. • This guarantees that any differences in reaction rate are truly due to the intended change in concentration, not from preparation errors, making the conclusions about the effect of concentration more valid.

Conclusion

This experiment confirmed that particle size (surface area) has the most dramatic effect on increasing the reaction rate between marble and hydrochloric acid, followed by temperature and then acid concentration.

Experiment 13 Determination of the Percentage of Acetic Acid in Vinegar

Hypothesis

If a standardized sodium hydroxide solution is titrated against a sample of vinegar, then the percentage of acetic acid in the vinegar can be accurately determined because the reaction between standardized sodium hydroxide solution and acetic acid is stoichiometric (1:1 ratio), allowing for precise calculation based on the volume of standardized sodium hydroxide solution required to reach neutralization.

Variables

Manipulated variable: Volume of standardized sodium hydroxide solution used

Responding variable: Percentage of acetic acid in vinegar

Controlled variable: Concentrations of the materials used

Discussion

1. Why using phenolphthalein instead of methyl orange as indicator when titrating sodium hydroxide solution to vinegar?

Reasons	Explanations	Indicator pH range
Nature of the reaction (weak acid vs strong base)	The titration procedures a basic salt (sodium acetate), making the solution slightly alkaline at the equivalence point.	Phenolphthalein pH range is 8.2 – 10.0 (more to alkaline) Methyl orange pH range is 3.1 – 4.4 (more to acidic)

2. How does the endpoint differ from the equivalence point in this titration?

Features	Endpoint	Equivalence point
Definition	The observed point where the indicator (phenolphthalein) changes color (Experimental)	The exact stoichiometric point where moles of alkaline = moles of acid (Theoretical)
Determination	Visually detected with indicator	Calculated from stoichiometry
Error source	None or ideal	• Slight overshooting due to indicator lag • Mis-judgement by human or random error

3. Why the mass of sodium hydroxide solution can be used in calculating the mass of acetic acid?

Reasons	Explanations
Molarity	The mass of sodium hydroxide solution in the solution is fixed by its molarity
Stoichiometry	The reaction is stoichiometric (1:1), the mass of sodium hydroxide solution directly relates to the mass of acetic acid

4. How do dilution errors impact the final percentage of acetic acid and what strategies can be used to prevent them?

Dilution errors	Explanations
Over-diluted vinegar or under-diluted sodium hydroxide solution	Underestimation of acetic acid percentage in vinegar as less sodium hydroxide solution needed for titration
Under-diluted vinegar or over-diluted sodium hydroxide solution	Overestimation of acetic acid percentage in vinegar as more sodium hydroxide solution needed for titration

Preventions	Explanations
Precise measurement of solutions	• Use a volumetric pipette to measure the vinegar instead of a graduated cylinder for better accuracy • If dilution is required, use a volumetric flask to prepare the diluted solution • Standardize the sodium hydroxide solution before to confirm its exact molarity • Store sodium hydroxide solution in a tightly sealed container to prevent absorption of atmospheric carbon dioxide which lowers its concentration over time
Equipment calibration	• Calibrate pipettes, burettes and volumetric flasks regularly to ensure volume accuracy • Use a digital balance for preparing sodium hydroxide solution if working with solid pellets

Technique improvements	• Avoid parallax errors when reading burette or pipette volumes (read at eye level at the meniscus bottom) • Rinse equipment properly [pipette with the analyte (vinegar) before use; Burette with the titrant (sodium hydroxide solution) before filling] • Add sodium hydroxide solution dropwise and swirl constantly near the endpoint to avoid overshooting • Prime the burette tip before titration if burette contains air bubbles
Verification steps	• Perform multiple trials (at least 3 times) to check the consistency • Compare results with known standards • Use a pH meter alongside phenolphthalein for a more precise endpoint

5. Explain how would the percentage change if impure vinegar (with citric acid) was used?

Percentage change	Explanation
Overestimate the true percentage of acetic acid	• Sodium hydroxide solution reacts with all acidic protons not just acetic acid • 1 mol of citric acid consumes 3 mols of sodium hydroxide solution $C_6H_8O_7 + 3NaOH \rightarrow Na_3C_6H_5O_7 + 3H_2O$

6. Wine (contains acetic acid) is commonly used in cooking to tenderize meat and add flavour. Besides these two purposes, what is another objective of using wine in this process?

Objectives	Explanations
Microbial stability	The alcohol (ethanol) and acidic components (e.g., acetic acid) in wine create an environment that mildly suppresses microbial growth. While not a substitute for proper food safety practices, this contributes to extending the meat’s freshness during preparation.
Acidic denaturation	Wine contains organic acids (e.g., acetic acid) that disrupt hydrogen bonds in muscle proteins. This acid-induced unfolding (denaturation) weakens the meat's connective tissue.

Collagen solubilization	Acidic environment helps convert tough collagen into soluble gelatine. This process is particularly effective in tough cuts with high connective tissue content.

Conclusion

This experiment successfully determined the acetic acid content in vinegar to be X% through **s**odium hydroxide solution titration, demonstrating the reliability of acid-base titration for quantitative analysis while highlighting the importance of precise measurements and standardized solutions to minimize errors. *X is the percentage of acetic acid in vinegar*.

Experiment 14 Electrolysis and Electroplating

Hypothesis

If an electric current is passed through (1) a concentrated saline solution between inert electrodes, then hydrogen and chlorine gas will form at the cathode and anode, respectively; and (2) a zinc sulfate solution with an iron nail cathode and zinc anode, then the nail will develop a protective zinc coating, demonstrating the principles of redox reactions in both processes.

Variables

Manipulated variable: Type of electrode used in electrolysis and electroplating

Responding variable: Volume of gases produced in electrolysis and zinc coating on iron nail in electroplating

Controlled variable: Voltage applied or concentration of solutions used in electrolysis and electroplating

Discussion

1. What gases are produced at the cathode and anode during the electrolysis of brine? Identify each gas and explain its formation.

Electrodes	Gases produced	Explanations
Cathode	Hydrogen gas	A “pop” sound was heard when ignited and burned, the phenolphthalein indicator turned pink at cathode
Anode	Chlorine gas	The yellowish-green gas turned the starch-potassium iodide test solution into blue

2. Why brine is needed during the electrolysis? Explain your answer.

Reasons	Explanations
High conductivity	• Brine contains a high concentration of Na^+ and Cl^-, which significantly increases electrical conductivity • Allows more efficient current flow, speeding up the electrolysis process

Selective discharge of Cl^-	- In brine, high Cl^- concentration favors chlorine gas production - In dilute brine, water molecules oxidize more easily than Cl^-, producing oxygen
Minimizes competing reactions	Brine reduces unwanted side reactions (like oxygen formation), ensuring maximum yield of hydrogen and chlorine gases

3. Why concentrated sodium hydroxide solution is used to wash iron nail before electroplating started? Explain your answer.

Reasons	Explanations
Removes surface contaminants	- Sodium hydroxide solution acts as a strong degreaser, breaking down organic residues (e.g., fingerprints, machine oils) that could interfere with plating adhesion - Sodium hydroxide doesn't dissolve rust directly, it helps loosen and lift corrosion products for easier removal by subsequent rinsing or scrubbing
Activates the metal surfaces	- Sodium hydroxide solution etches the iron nail surface microscopically, creating a rougher texture for better zinc adhesion during electroplating - Removes the passive oxide layer that might prevent uniform plating

4. How did electrode cleanliness (e.g., sanding the nail) impact the electroplating results? Explain your answer.

Impacts	Explanations
Improve adhesion of zinc coating	- Removes rust, oxide layer and dirt, exposing a fresh, active iron surface - Zn^{2+} bond more strongly to the clean metal, preventing peeling or flaking after plating
Uniform coating thickness	- Smooth surface ensures even current distribution during electroplating

Reduce contamination of electrolyte	- Impurities from dirty nail causes dull, rough plating or side reactions
Fasten the plating process	- Clean surfaces have low electrical resistance, allowing efficient electron transfer

5. Copper and gold could be tested for electroplating, and how might results differ in terms of their appearance and conductivity?

Other metals	Appearance	Conductivity
Copper	Bright reddish-gold coating	Better than zinc plated iron nail
Gold	Yellow shinning coating	Better than zinc plated iron nail

Conclusion

This experiment successfully demonstrated the principles of electrolysis by decomposing brine into hydrogen and chlorine gas, while electroplating effectively coated an iron nail with zinc, confirming the role of redox reactions.

Experiment 15 Chemical Properties of Aldehydes and Ketones

Hypothesis

Aldehydes have a hydrogen atom directly attached to their carbonyl carbon, while ketones have two alkyl groups, aldehydes will show greater reactivity in oxidation tests compare to ketones.

Variables

Manipulated variable: Carbonyl compounds used

Responding variable: Observations of different tests

Controlled variable: Concentration **OR** Volume of solutions and reagents used

Discussion

1. Write out the chemical equations for all reactions (except Schiff's reagent test) that produced positive test results in the experiment.

Tests	Ethanal
Silver mirror test	$CH_3CHO + 2Ag(NH_3)_2OH \rightarrow CH_3COONH_4 + 2Ag + H_2O + NH_3$
Fehling's solution test	$CH_3CHO + 2Cu^{2+} + NaOH + H_2O \overset{\Delta}{\rightarrow} CH_3COONa + Cu_2O + 4H^+$
Bromine water test	$CH_3CHO + Br_2 + H_2O \rightarrow CH_3COOH + 2HBr$
Acidified potassium permanganate solution test	$CH_3COH \xrightarrow{\text{Acidified } KMnO_4} CH_3COOH + CO_2 + H_2O$

2. Why different observations were recorded between ethanal and propanone in the experiment? Explain your answer.

Reasons	Explanations	
	Ethanal	Propanone
Structural differences	• C-H bond allows oxidation to carboxylic acids	• C-C bond cleavage only occurs at mild conditions
	• Low steric hindrance	• High steric hindrance (bulky CH_3 group)
	• Carbonyl carbon is more δ^+	• Alkyl groups donate electrons, reducing δ^+

3. If methanoic acid, fructose and sucrose are used in the experiment, predict their observations (except for bromine water test and Schiff's reagent test) and explain your answer.

Observations

Compounds	Tests		
	Silver mirror test	Fehling's solution test	Acidified potassium permanganate solution test
Methanoic acid	Silver mirror form	Red precipitate form	Purple solution decolourize
Fructose	Silver mirror form	Red precipitate form	Purple solution decolourize
Sucrose	No change	No change	No change

Explanations

Compounds	Tests		
	Silver mirror test	Fehling's solution test	Acidified potassium permanganate solution test
Methanoic acid	Structurally similar to an aldehyde contains -CHO group		
Fructose	Fructose tautomerizes to glucose (an aldose) exposing an aldehyde contains -CHO group which reacts		
Sucrose	• The glycosidic bond blocks the reactive carbonyl groups of its glucose and fructose components • No free aldehyde available for reactions		

4. Suggest why some ketones might give false positives in certain carbonyl tests if the reagent concentrations are not carefully controlled.

Reasons	Explanations
Overuse of oxidizing agents	Excessively strong or concentrated oxidants can forcefully oxidize ketones by breaking C-C bonds
Over-alkaline conditions	Excess base can convert ketones to enediols, which tautomerize to aldehydes and react

5. Compare the oxidation products of aldehydes and ketones. Why can aldehydes be oxidized to carboxylic acids, while ketones require stronger conditions?

Reasons	Explanations
Presence of C-H bond on the carbonyl carbon	• Aldehydes have a hydrogen atom directly attached to the carbonyl carbon, which can be removed as a proton during oxidation, forming a carboxylic acid • Ketones have two alkyl groups attached to the carbonyl carbon, so oxidation requires C-C bond cleavage (much harder than C-H cleavage)
No strong C-C bonds need to break	• Only the C-H bond is cleaved, which is weaker than C-C bonds • Strong oxidants are required to break C-C bonds, often leading to mixtures of smaller acids or ketones and CO_2

Conclusion

This experiment and theoretical analysis demonstrate that aldehydes (R-CHO) and ketones (R-CO-R’) exhibit distinct chemical behaviors due to their structural differences, particularly the presence or absence of a hydrogen atom on the carbonyl carbon.

Experiment 16 Preparation of Ethyl Ethanoate

Hypothesis

Ethyl ethanoate can be synthesized using acid-catalyzed esterification of ethanol and ethanoic acid in the presence of concentrated sulfuric acid.

Variables

Manipulated variable: -

Responding variable: Percentage yield of ethyl ethanoate

Controlled variable: Volume of chemicals used, Concentration of chemicals used **OR** Temperature applied in collecting ethyl ethanoate vapour

Discussion

1. The reaction mechanism is described as acid-catalyzed. Using the formation of ethyl ethanoate as an example, write the balanced chemical reaction.

Chemical reaction
$CH_3COOH + C_2H_5OH \underset{\Delta}{\overset{H^+}{\rightleftharpoons}} CH_3COOC_2H_5 + H_2O$

2. The collected ethyl ethanoate was channeled to the saturated sodium carbonate solution. Explain the purpose of this step, write out the balanced chemical equation that occurs and the observations expected.

Explanation	Chemical equation	Observations
To neutralize/remove the acidic impurities	$2CH_3COOH + Na_2CO_3 \rightarrow 2CH_3COONa + CO_2 + H_2O$	• Vigorous effervescence bubbling • Formation of two distinct layers

3. The experimental yield of ethyl ethanoate is often less than the theoretical yield. Discuss two possible reasons of loss that could account for this and suggest improvements to the procedure to minimize each one.

Possible reasons of loss	Discussion	Suggestions for improvements
Reversible reaction and equilibrium limitations	The esterification reaction is reversible. Even with a dehydrating agent (concentrated sulfuric acid), some water remains, shifting equilibrium backward to reactants and reducing ester yield.	• Use excess ethanol (the cheaper reactant). This pushes equilibrium toward ester formation (Le Chatelier's Principle). • Alternatively, employ a more efficient dehydrating agent (e.g., molecular sieves) to absorb water during the reaction.
Loss of volatile compounds during experiment	Ethyl ethanoate and ethanol are highly volatile. During distillation, transfer, or washing steps, vapors may escape, or droplets may adhere to apparatus surfaces.	• Use precise distillation techniques: Collect distillate only within the correct boiling range (77 – 78°C) to avoid co-distillation of reactants. • Rinse glassware with a volatile solvent (e.g., acetone) after transfer to recover adhered product. • Ensure airtight connections in the apparatus to prevent vapor leaks.
Incomplete reaction or side reactions	Side reactions, such as dehydration of ethanol to ethylene or diethyl ether consume reactants, reducing ester yield.	• Strictly control reaction temperature to avoid excessive heat consequently minimize side reactions. • Use anti-bumping granules to ensure smooth boiling.

4. Describe one chemical test (other than smell) that could be used to confirm the identity of the product as an ester. State their reagents used and the positive results you would observe.

Chemical test	Reagents used	Observation
Alkaline hydrolysis **OR** Saponification	Sodium hydroxide solution, hydrochloric acid and phenolphthalein indicator	• Loss of ester odor due to the formation of carboxylic acid. • The pink colour fades as ester hydrolyzes and consume base.

5. In the advance experiment, the ester is washed with sodium carbonate solution and then with water. Explain why the washing step with water is necessary after the sodium carbonate wash. Name one impurity is being removed?

Importance of water	Explanations	Impurities
To remove water-soluble ionic salts and residual base that remain in the organic layer after the initial neutralization	These ionic salts dissolve in the aqueous layer but may form emulsions or leave traces adsorbed to the organic layer.	• Sodium ethanoate (from ethanoic acid) • Sodium sulfate (from sulfuric acid) **OR** • Sodium carbonate

Conclusion

This experiment successfully synthesized ethyl ethanoate through acid-catalyzed esterification though the yield was lower than theoretical due to the reversible reaction, procedural losses and/or incomplete reaction.

Experiment 17 Preparation of Nylon-66

Hypothesis

Nylon-66 can be synthesized using condensation polymerization of 1, 6-hexanediamine and adipoyl chloride.

Variables

Manipulated variable: -

Responding variable: Mass of nylon-66 **OR** Percentage yield of nylon-66

Controlled variable: Weight of chemicals used

Discussion

1. Name the type of polymerization in the above experiment and explain your answer.

Type of polymerization	Explanation
Condensation polymerization	• A molecule of water is eliminated during the bond-forming. **OR** • The polymer chains build gradually through reactions between any species with compatible functional groups.

2. The synthesized nylon is often washed with dilute hydrochloric acid and then water. Explain the purpose of this acid wash?

Purpose	Explanation
To purify the polymer by dissolving and removing any unreacted hexamethylenediamine monomer	• The traces of the base starting monomer, 1, 6-hexanediamine, can remain trapped in the solid polymer. • Dilute hydrochloric acid neutralizes with this basic diamine, converting the water-insoluble diamine into a water-soluble salt. • The newly formed salt dissolves easily in the aqueous acid solution and is rinsed away.

3. Why dilute hydrochloric acid was used in the above experiment. Explain your answer.

Explanation
Dilute acid is strong enough to protonate the free amine monomers but not strong enough to attack the polymer backbone.

4. Nylon was one of the first synthetic polymers to be commercialized. Based on your knowledge, give two fundamental chemical principles make such a material so useful and versatile.

Fundamental chemical principles
• Strong intermolecular forces require more energy to overcome (also accept**:** high tensile strength**,** durability or high melting point) • Carbon-carbon and carbon-hydrogen single bonds allow for free rotation, giving molecular segments flexibility (also accept: high flexibility)

5. In industrial, a small amount of acetic acid is often added to the reaction mixture. Explain its function briefly.

Function	Explanation
Chain-terminating agent **OR** Molecular weight regulator	Acetic acid reacts with the terminal amine groups ($-NH_2$) of the growing polymer chains, prevents it from further reaction.

Conclusion

This experiment successfully synthesized nylon-66 through condensation polymerization though the yield was lower than theoretical due to the procedural losses and/or incomplete reaction.

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Name: __________________________ Date of completion: _____________

Class: __________________________ Teacher: _______________________

Experiment Title

A concise, descriptive and informative phrase used to clearly state the subject of an experiment.

Problem Statement

A clear and concise description of a problem or challenge that must be resolved to achieve the intended outcomes

Hypothesis

A testable prediction about an observed phenomenon, which serves as a preliminary answer to an experimental question and can be supported or refuted through experimentation.

Objectives

1. A clear and concise statement describing the intended goal of an experiment, designed to guide the experimental process and maintain project focus.

Variables

Manipulated variable: Factor that is deliberately altered or controlled in an experiment to observe its effects on other variables.

Responding variable: Factor that is measured and observed in an experiment to examine whether it is influenced by the manipulated variable(s).

Controlled variable: Factor that remains constant and unchanged throughout the experiment.

Materials and Apparatus

Materials/Apparatus	Concentrations/Specifications	Volumes/Quantities

Precautionary Steps (If there is)

1. Safety measures and error control strategies planned before and during the experiment, aimed at preventing harm to students, equipment or the environment while ensuring accurate and reliable results.

Procedures (Can be presented as a schematic diagram)

1. A detailed, step-by-step description of the procedure used to test a hypothesis. It typically includes the definition of variables, a list of materials, a sequential outline of operational steps and specifies data collection and analysis methods to ensure the experiment's reproducibility and reliability.
2. A detailed, step-by-step description of the procedure used to test a hypothesis. It typically includes the definition of variables, a list of materials, a sequential outline of operational steps and specifies data collection and analysis methods to ensure the experiment's reproducibility and reliability.
3. A detailed, step-by-step description of the procedure used to test a hypothesis. It typically includes the definition of variables, a list of materials, a sequential outline of operational steps and specifies data collection and analysis methods to ensure the experiment's reproducibility and reliability.

Results

X	Y

Discussion

1. Why is a Discussion section necessary in a lab report?
   <u>The Discussion section of a lab report requires an analysis of the experimental results: interpreting their meaning, comparing them with the initial hypothesis and existing knowledge, identifying sources of error and experimental limitations and proposing directions for future research.</u>

2. Why is a Discussion section necessary in a lab report?

   The Discussion section of a lab report requires an analysis of the experimental results: interpreting their meaning, comparing them with the initial hypothesis and existing knowledge, identifying sources of error and experimental limitations and proposing directions for future research.

3. Why is a Discussion section necessary in a lab report?

   The Discussion section of a lab report requires an analysis of the experimental results: interpreting their meaning, comparing them with the initial hypothesis and existing knowledge, identifying sources of error and experimental limitations and proposing directions for future research.

Conclusion

A concise summary of experimental outcomes, evaluation of hypothesis testing, analysis of potential error sources and limitations, along with the broader implications revealed by the study.