TO THE STUDENTS

The purpose of laboratory studies is to carry out operations and make measurements that will give us information about the physical world in which we live. This manual will serve you as the basis for chemistry laboratory experiences that establish an appreciation for the value and satisfactions of experimental science, and that help instil a qualitative, quantitative and critical approach to understanding the chemical world around us.

In this chemistry course, you will be working with your teacher and other students to develop the basic knowledge and skills which will help to understand more about chemistry and to apply in your daily life. You will learn how to demonstrate an interest in the creativity and innovation discover in chemistry. In some lessons, you will participate in group activities to develop skills in scientific methods of investigation based on the concepts, theory, terms, facts, laws and principles related to main themes.

Experiments have been grouped by topic to fit conveniently into most lecture schedules. This helps you feel more comfortable in the laboratory before the experiments begin to illustrate lecture material. It is essential that you understand what the experiment is about and what you are expected to do, when you start the laboratory period. This means that you must study each experiment before coming to the laboratory, referring, if necessary to your Grade 10 Chemistry Textbook for the theory and calculations involved in this experiment.

Your laboratory grade will be determined mainly by your systematic methodology of experimentation, the accuracy of your results, and your knowledge of what you are doing and why, as shown by your ability to answer questions included in the experiment.

After learning this experimental chemistry course you will develop and practice **higher order thinking skills**. Moreover, you will be able to participate actively in all experiments through the five C's, which are **collaboration**, **communication**, **critical thinking** and **problem solving**, **creativity and innovation**, and **citizenship**, as important **21st century skills for learning**.

A SPECIAL NOTE ON SAFETY

To avoid accidents in a laboratory, you must work carefully and thoughtfully. Every chemical must be regarded as potentially hazardous. Each experiment contains appropriate safety conditions in a **CAUTION** box after the **Procedure** section. The safety precautions for each experiment must be read and carefully considered before beginning the experiment.

Follow the laboratory safety rules which are mentioned in the Experiment 1. For your own safety, and that of your co-workers, please pay close attention to these regulations.

There are **eleven experiments** included in this Experimental Chemistry Course.

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EXPERIMENTS

Safety Rules and Report Writing

1. Laboratory Safety Rules and Writing Report

Separation Techniques

2. (a) Separation of a Mixture of Common Salt and Sand

(b) Separation of the Different Colouring Matters from In

Solubility

3. Determination of the Solubility of Common Salt in Water at Room Temperature

Formulae of Compounds

4. Determination of the Empirical Formula of Magnesium Oxide

Preparative Methods

5. Preparation of Calcium Oxide (Quicklime) and Study on Some of Its Chemical Properties

6. (a) Preparation of Carbon Dioxide Gas from Calcium Carbonate

(b) Study on the Physical Properties of Carbon Dioxide

Reactions

7. Test for Carbonate, Chloride, Bromide and Iodide Anions

Acids, Bases and Salts

8. (a) Neutralisation of Acids and Bases

(b) Preparation of Copper(II) Sulphate Crystals

Environment

9. Differentiation between Temporary and Permanent Hardness of Water

10. (a) To examine the soil alkalinity and acidity

(b) To determine the presence or absence of lime in the soil

Fossil Fuels

11. Relationship between Number of Carbon Atoms in Petroleum Products and Their Flow Time

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EXPERIMENT 1

Laboratory Safety Rules and Writing Report

Learning Objectives

- To be familiarised with the common chemicals and apparatus in the chemistry laboratory
- To understand the proper handling and uses of the different laboratory apparatus

Introduction

The students should avoid accidents while doing laboratory experiments. Laboratory safety is of first importance in any experimental work. The students have a responsibility to follow all the basic laboratory rules in experimental work.

(a) Basic Laboratory Safety Rules

No.	Do's	No.	Don'ts
1.	Follow instructions given by your teacher.	1.	Do not start the experiment without teacher's instruction.
2.	Know the general laboratory equipment, glassware etc. and other common items.	2.	Do not work alone in the laboratory.
3.	Keep the laboratory table (desk) clean and tidy.	3.	Do not touch any equipment, chemicals or other materials without the teacher's permission.
4.	Keep your workbooks and papers away from heating equipment, chemicals and flames.	4.	Do not eat or drink in the laboratory and never taste any chemicals.
5.	Know the locations and operating procedures of all safety equipment including the first aid kit, fire extinguisher and fire blanket.	5.	Do not smoke in the laboratory.
6.	Report all accidents and breakages to the teacher straight away.	6.	Do not smell gases or mixtures of chemicals directly.
7.	After doing the experiment, put all the apparatus and scattered stools in proper order in the laboratory.	7.	Do not dispose all chemical wastes in the sinks.
		8.	Burners must not be turned on whenever not in use: including electrical devices and gadgets.

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Personal protection

1. Wear safety goggles to protect the eyes from chemicals or pieces of broken glass.
2. Wear a laboratory coat to protect the body and clothes from the effects of chemical spills.
3. Wear plastic or rubber gloves to protect the hand from chemical spills.
4. Always wash your hands after handling laboratory materials.
5. Tie up long hair before doing an experiment.

(b) Handling the Common Apparatus and Chemicals

(i) Common apparatus

The following is the brief description of some common pieces of apparatus used in the chemistry laboratory.

Table 1.1 Some Common Apparatus Used in the Chemistry Laboratory

 balance	 beaker	 beehive shelf (porcelain)	 burette	 clay pipe
 conical flask	 delivery tubes	 desiccator	 filter paper	 flat-bottomed flask
 funnel	 gas jar	 glass capillary tubes	 glass rod	 glass tube

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gloves


goggle glass


litmus paper


measuring cylinder


petri dish


pipette


porcelain basin


porcelain crucible


round-bottomed flask


spatula


spirit burner


stand and clamp


test tube


test tube stand


thermometer


thistle funnel


tongs


tripod stand


wash bottle


watch glass

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(ii) Handling chemicals

1. All chemicals in the laboratory are to be considered dangerous. Read carefully the label on reagent bottle before use. Avoid handling chemicals with fingers.
2. Use a spatula to transfer a solid chemical to a test tube or beaker.

3. Never point the test tube towards people or look down into the test tube during heating.
4. Never pour water into a concentrated acid. Acid should be poured slowly into water.
5. Dispose the chemicals in the waste container provided in the laboratory.

(iii) Handling glassware and equipment

1. Examine glassware before each use. Never use chipped, cracked or dirty glassware.
2. Never handle broken glass with your bare hands.
3. Do not immerse hot glassware in cold water.

(iv) Handling of balance

Rules for Analytical Balances

The following rules summarise those procedures which must be followed in order to obtain accurate and reliable mass measurements with a single-pan analytical balance. Adherence to these rules will, at the same time, prevent damage to the balance.

1. Close the balance door, while weighing an object, in order to prevent air currents from disturbing the reading. When finished, the operator should close the balance door to prevent dust and dirt from entering the balance.
2. Only glass, ceramic, metal or plastic objects and containers should be placed in direct contact with the balance pan.
3. Do not handle objects to be weighed with bare hands. Moisture, grease and dirt on your fingers will affect the weight of the objects.
4. To be weighed accurately, all objects must be at room temperature. A warm object sets up convection currents inside the balance enclosure, which will make an object appear lighter than it really is. Also, warm air inside the enclosure is less dense than the air that it displaces and this also leads to a negative determinate error.
5. Never weigh chemicals directly in contact with the balance pan. Use containers such as beakers, flasks and weighing bottles.
6. All objects and materials that have recently been removed from a desiccator will absorb moisture and thereby gain weight. It is therefore good practice to record weights after identical time intervals. For example, if you are taking crucibles to constant weight, always record the weight of the crucible exactly 5 seconds after having placed the crucible on the balance pan. Using this technique, it is possible to minimise the effect of moisture absorption.
7. The use of weighing paper must be strictly avoided when using an analytical balance.
8. Do not spill chemicals inside the balance enclosure. If a spill occurs, clean it up immediately.

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(c) Writing Report

The following components should be included in the report on an experimental investigation.

(i) Title

Write the title of the experiment as a statement or question.

(ii) Aim

Write the aim of the experiment.

(iii) Introduction

The **Introduction** should:

- briefly explain relevant theory in sufficient detail;
- introduce any relevant laws, equations or theorems.

(iv) Apparatus and materials

A list of apparatus and materials used should be included.

(v) Procedure

Do not write a detailed procedure. Provide a brief outline only of what you did. Emphasise the important points. Mass, volume, temperature, other units, etc., should be cited.

(vi) Precautions

State the safety precautions taken in the investigation.

(vii) Observations, results and calculations

Record the measured data in a suitable format, e.g., a list or a table. Show clearly how calculations were done.

Experiment	Observation	Inference

(viii) Conclusion

Draw a conclusion which is consistent with results obtained and relate these results to the aim of the experiment.

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EXPERIMENT 2

Separation Techniques

Experiment 2 (a) Separation of a Mixture of Common Salt and Sand

Learning Objectives

- To learn the separation method for common salt and sand in a given mixture
- To know how to separate the solid mixtures

Introduction

Most substances are naturally found as mixture; therefore, separation methods indicate how the physical states of components in the mixture can be separated into pure substance. In this experiment, you will investigate how to separate the common salt and sand in a given mixture and which methods are used to achieve the successful separation. (Refer to Grade 10 Chemistry Textbook, Chapter 2)

Apparatus and materials

Beakers, glass rod, funnel, filter paper, evaporating basin (porcelain basin), common salt, sand, distilled water, conical flask, tripod stand, wire gauze, spirit burner/ hot plate, balance

Procedure

(1) Preparation of mixture

- Mix about 3 g of common salt with about 1 g of sand.

What physical state of substances is in this mixture?

Is it homogeneous mixture or heterogeneous mixture?

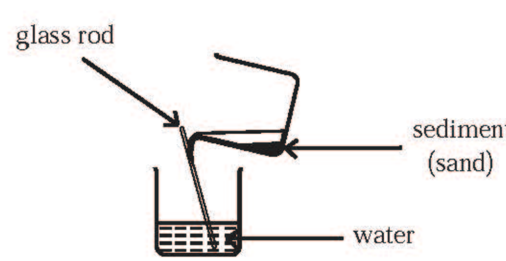
(2) Dissolution

- Put the mixture into a 250 mL beaker. Add about 50 mL of water to dissolve the salt. Stir the solution with a glass rod.

In dissolution process which substance remains undissolved and which substance dissolves in water?

(3) Decantation

- Allow the beaker to stand until all undissolved sand particles settle down in the beaker.
- Decant the clear liquid into a dry clean beaker using a glass rod to guide the clear solution, without disturbing the sand particles at the bottom, as shown in Figure (a).



(a) Decantation

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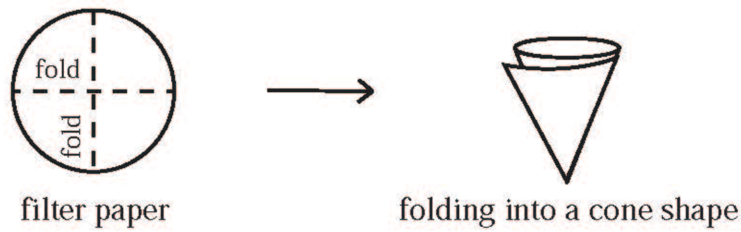
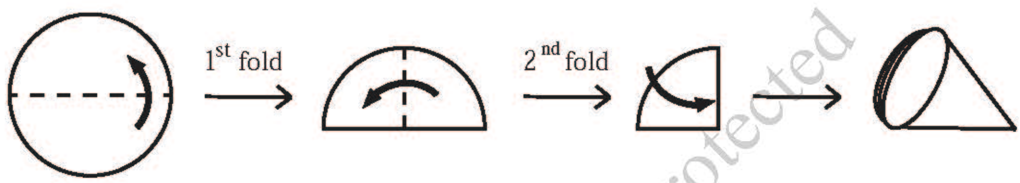
(4) Folding of filter paper

- Fold a filter paper in half. Then fold in half again, bringing the two corners together. Fold a filter paper into a cone shape as follows:

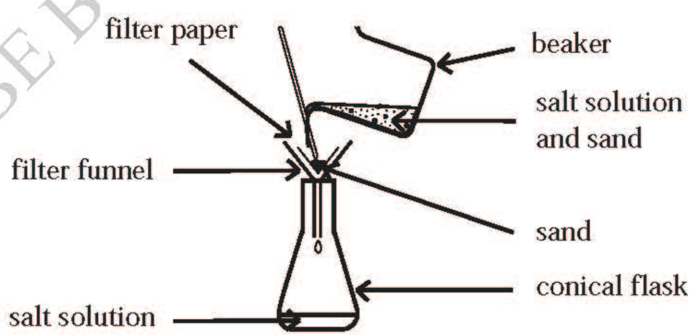




- Put the cone into a funnel and moisten the cone with small amount of water to make the filter paper adheres to the inner wall of the funnel.

(5) Filtration

- Add a small amount of water to the sand in the beaker to wash the common salt residue out of the sand.
- Filter the mixture into a conical flask through filter paper as shown in Figure (b).
- Repeat the washing, decantation and filtration at least three times.
- Pour the filtrate into an evaporating basin.



(b) Filtration

What remains on the filter paper?

Where is the salt after filtration? What type of mixture is observed in the conical flask?

What kind of substances can be separated by this filtration method?

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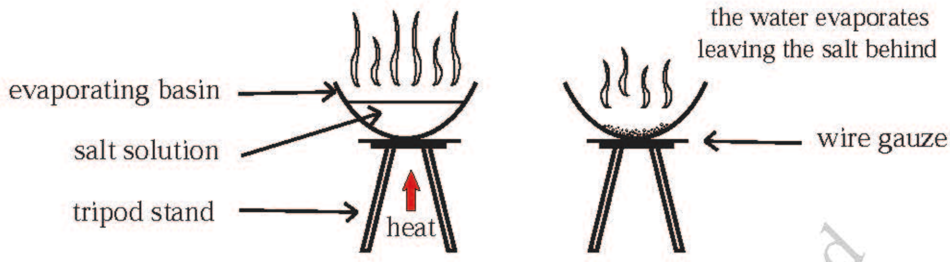
(6) Evaporation

- Heat gradually the evaporating basin on a wire gauze, until most of the water in the solution has evaporated as shown in Figure (c).
- Remove the flame and let the damp salt dry.



(c) Evaporation

CAUTION

- The filter paper should not extend above the brim of the funnel.
- The filter paper cone should be filled with the solution just below its brim.
- Run down the solution along a glass rod into the funnel as shown in Figure (a).
- Don't heat vigorously because water and salt burst out. It may be loss of salt.
- Beware of hot salt spitting when evaporation is almost complete.

  What will be lost due to this spurting?
  What will you do to avoid this loss?

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Observations, results and calculations

Experiment	Observation	Inference

Conclusion

Experiment 2 (b) Separation of the Different Colouring Matters from Ink

Learning Objectives

- To observe how chromatography can be used to separate mixtures of substances
- To separate colouring matters from blue ink, black ink and blue-black ink by using paper chromatography

Introduction

Chromatography is a method of separating and identifying various components in a mixture, which is present in small quantities. There are many types of chromatographic separation methods, one of which is paper chromatography. This uses a paper as stationary phase and a solvent as mobile phase. This method can be used to separate the different colouring matters in three samples of ink and find which samples of ink, if any, are the same.

Apparatus and materials

Filter paper, glass rod, capillary tube, pencil, glass hook, gas jar or large beaker (250 mL or larger), three water-soluble inks, solvents (distilled water, ethanol etc.)

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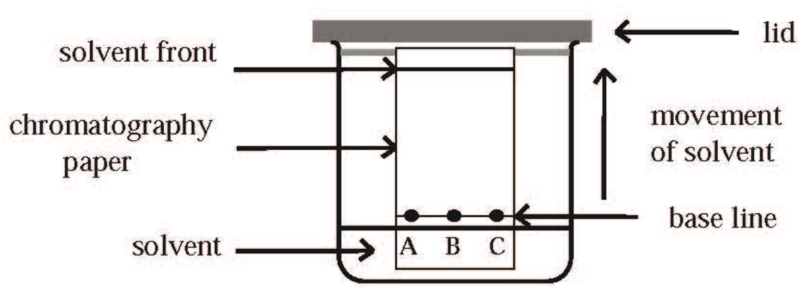
Procedure

- Cut out a rectangular piece (3 cm × 12 cm) of chromatography paper (filter paper can also be used).
- Draw a straight pencil line (base line) along one edge of the paper, about 1 cm from the edge. Lightly mark three equally-spaced dots with pencil along the line. Label the dots (A, B, C) as shown in Figure. Draw another straight pencil line along the other edge of the paper, about 1 cm from the edge. This is known as the solvent front.
- Put a small ink spot on each dot by means of thin capillary tube.
- Add solvent into a beaker to a depth of about 0.5 cm.
- Put the paper carefully into the beaker as shown in Figure. The spots should be slightly above the level of the solvent. Make it sure not to touch the inner wall of the beaker.
- Hang the strip of paper with the ink spots as shown in Figure.



- It will be found that the solvent rises up the paper, passing over the ink spot and carrying with it the different colour constituents at different speeds from the original spot on the base line, forming spots at different distances from the base line.
- When the solvent reaches solvent front, remove the paper and allow it to dry.
- The different spots at different distances from the pencil lines are the constituent substances of each ink. The pattern of paper strip obtained after paper chromatography is called a chromatogram.
- Count the coloured spots formed on the chromatogram.

How many different kinds of colouring matter are mixed up in ink which is used in the experiment?

Draw the picture of your chromatogram in the space below.

CAUTION

- Take care that the diameter of spot is not more than 5 mm in diameter.
- The spots should be slightly above the level of the solvent.
- Small concentrated spots give better results.

Note: The students can use different colours of soft pens instead of ink.

Conclusion

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EXPERIMENT 3

Determination of the Solubility of Common Salt in Water at Room Temperature

Learning Objective

- To determine the solubility of common salt in water at room temperature

Introduction

Solubility is the maximum quantity of a substance (solid, liquid, gas) that can be completely dissolved in a 100 g of water (solvent) to get a saturated solution at a given temperature.

Apparatus and materials

Evaporating basin, beakers, glass rods, measuring flask, tripod stand, spirit burner, tongs, thermometer, balance, desiccator, wire gauze, common salt, distilled water

Procedure

- Weigh a clean dry evaporating basin. Record the mass of the basin.

Given: density of water = 1 g mL^{-1}

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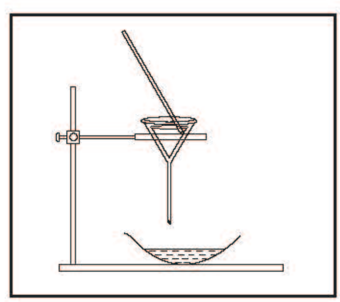
- Add 25 g of common salt into about 100 mL of distilled water in a beaker.
- Stir well until all salt dissolves. More salt is added until the solution becomes saturated.

  How do you know that the solution is saturated?



- Note the temperature of the solution. It should be the room temperature.

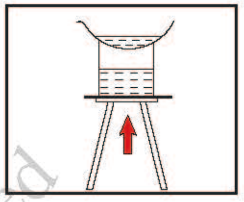
  Why do you use distilled water in this experiment to dissolve the solid?

- Filter half of the saturated solution into the basin by using clean dry filter paper and funnel.
- Weigh the basin containing the filtrate.
- Record the mass of the saturated solution from the calculation.

  Why are the dry filter paper and funnel used in this experiment?

- Heat the basin to evaporate the solution. When most of the water has evaporated, take a basin with a pair of tongs.
- Transfer the basin onto a beaker with half-full of boiling water which is being heated.
- Continue heating until all the water in the basin has evaporated.
- Allow the basin to cool in a desiccator.

  Why is it necessary to cool the basin in desiccator?



- Weigh the basin together with the solid.
- Repeat heating, cooling and weighing until the mass becomes constant.

  Why is it necessary to repeat heating, cooling and weighing until the mass becomes constant?

- Record the final mass of the basin and common salt.

  Calculate the solubility of common salt at room temperature.

CAUTION

- Don't heat vigorously because water and salt burst out. It may be loss of salt.
- Beware of hot salt spitting when evaporation is almost complete.

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Observation, results and calculation

temperature of the saturated solution (room temperature)	=	_____ °C		
mass of the basin	=	A g		
mass of the basin + saturated solution	=	B g		
mass of the basin + common salt	=	C g		
mass of the saturated solution	=	(B - A) g	=	W_1 g
mass of the common salt	=	(C – A) g	=	W_2 g
mass of the water	=	$(W_1 – W_2)$ g		

At _____ °C, $(W_1–W_2)$ g of water contains W_2 g of common salt.

∴ 100 g of water contains $[W_2/ (W_1–W_2)] \times 100 = \mathbf{X}$ g.

Conclusion

The solubility of common salt at _____ °C is found to be ____________.

EXPERIMENT 4

Determination of the Empirical Formula of Magnesium Oxide

Learning Objective

- To understand how to calculate the empirical formula of a compound

Introduction

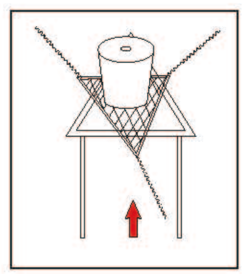
Empirical formula is the simplest type of chemical formula and it can determine the number of atoms present in the compound by using the ratio of mass composition and atomic weight of its elements. (Refer to Grade 10 Chemistry Textbook, Chapter 4)

Apparatus and materials

Crucible with lid, tripod stand, wire gauze, tongs, spirit burner, magnesium ribbon, sand paper

Procedure

- Take about 20 cm length of magnesium ribbon and clean it by rubbing two or three times with sand paper.
- Weigh a clean dry crucible with its lid and record the mass.
- Make a coil of the magnesium ribbon and place it in the crucible and put the lid on.
- Weigh the crucible with magnesium and its lid. Record the mass.
- Heat the crucible. During heating, lift the lid occasionally with a pair of tongs to let air in. When burning magnesium stops, remove the lid and heat the crucible strongly for two minutes to make sure that the combustion of magnesium is completed.



- Then remove the flame, put the lid on and allow the crucible to cool.
- Weigh the crucible and its lid. Record the results.
- Repeat heating, cooling and weighing the crucible and lid to obtain the constant mass.
- Record the last two constant masses.

  What substance is formed in the crucible?

  Are you sure that all magnesium is converted into its oxide?

CAUTION

- When magnesium begins to burn, lift the lid every minute taking care not to allow the white fumes (magnesium oxide) to escape.

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- To make sure that complete conversion of magnesium to its oxide has taken place, the heating, cooling and weighing should be carried on until there is no more increase in mass.

Observation, results and calculation

the mass of crucible and lid	=	A g
the mass of crucible and lid + magnesium	=	B g
[the mass of crucible and lid + magnesium oxide] (1st weighing)	=	W_1 g
[the mass of crucible and lid + magnesium oxide] (2nd weighing)	=	W_2 g
[the mass of crucible and lid + magnesium oxide] (3rd weighing) (constant weight)	=	W_2 g
the mass of magnesium	=	(B – A) g
the mass of magnesium oxide	=	(W_2 – A) g
the mass of oxygen, combined with magnesium	=	(W_2 – B) g

What is the mass of total number of magnesium atoms in magnesium oxide?

What is the mass of total number of oxygen atoms in magnesium oxide?

the relative atomic mass of magnesium = 24

the relative atomic mass of oxygen = 16

Symbol	Mg	O
The mass of each element	_____	_____
Divided by relative atomic mass	_____	_____
Empirical formula	_____	

Conclusion

The empirical formula of magnesium oxide is __________.

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EXPERIMENT 5

Preparation of Calcium Oxide (Quicklime) and Study on Some of Its Chemical Properties

Learning Objectives

- To prepare calcium oxide
- To study the interaction of calcium oxide with (i) water and (ii) dilute acid

Introduction

Calcium oxide (CaO), commonly known as quicklime, is a widely used chemical compound. It is a white crystalline solid at room temperature.

(i) Calcium oxide can be prepared by burning the hot calcium metal in oxygen. However, in this experiment it will be prepared by the decomposition of slaked lime (calcium hydroxide).

(ii) Calcium oxide is soluble in water to form calcium hydroxide.

(iii) Calcium oxide reacts with acid to produce metal salt and water.

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Apparatus and materials

Crucible, spirit burner, test tubes, watch glass, glass rods, glass tubes, beakers, calcium hydroxide, calcium oxide, blue and red litmus papers, dilute sulphuric acid or hydrochloric acid

(a) Preparation of Calcium Oxide (Quicklime)

Procedure

- Place a paste of slaked lime (calcium hydroxide) in a crucible.
- Heat the crucible strongly for about 30 minutes. Decomposition of slaked lime to calcium oxide should occur by then.

  Write a chemical equation for the above reaction.

(b) Study on Some Chemical Properties of Calcium Oxide

(i) Reaction of calcium oxide with water

- Take 0.5 g of quicklime (calcium oxide) in a dry test tube.

  Observe the state and colour of the substance.

- Then carefully add about 1-2 mL of water in the test tube.

  Observe what happens.

- Decant the clear solution into a beaker.

- Test this solution with both blue and red litmus papers.

  Observe what happens.

  Analyse the experiment with the following questions:

  What are the reactants?

  What is the product?

  Does this solution call alkali? Explain.

  Write the equation for this reaction.

(ii) Reaction of calcium oxide with acid

- Place a small amount of calcium oxide in a beaker.
- Then add 1-2 mL of 5 % (v/v) dilute sulphuric acid. (To get 5 % (v/v) sulphuric acid, 5 mL of concentrated sulphuric acid is slowly poured to 95 mL of distilled water).

  Observe what happens.

- Take a clear drop of the solution by a dropper and evaporate carefully on a watch glass.

  Is any residue left? If so, give the name of this residue.

  Is there any change in colour?

  Write an equation for this reaction.

Note: This experiment can be done with dilute hydrochloric acid instead of dilute sulphuric acid.

Conclusion

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EXPERIMENT 6

Preparation of Carbon Dioxide Gas from Calcium Carbonate and Study on Some of Its Physical Properties

Experiment 6 (a) Preparation of Carbon Dioxide Gas from Calcium Carbonate

Learning Objectives

- To prepare carbon dioxide gas from calcium carbonate
- To determine the physical properties of carbon dioxide gas

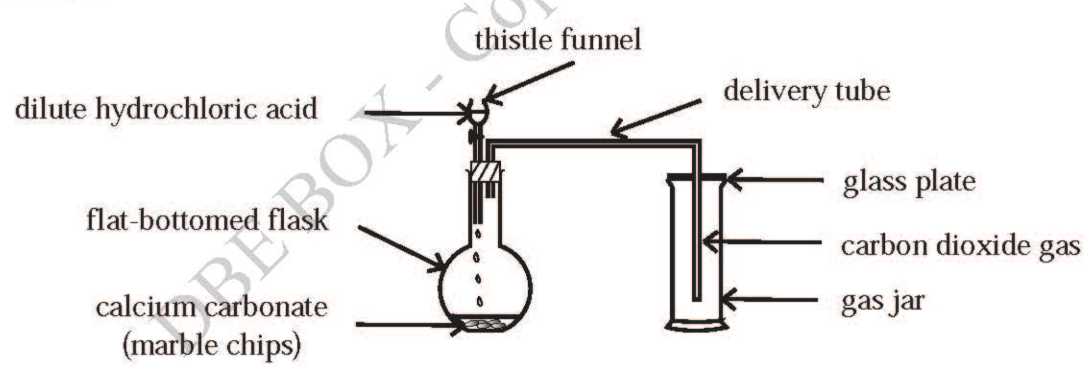
Introduction

Carbon dioxide is produced whenever an acid reacts with a carbonate. This means carbon dioxide is easy to prepare in the laboratory. Calcium carbonate and dilute hydrochloric acid are usually used in the preparation of carbon dioxide.
(Refer to Grade 10 Chemistry Textbook, Chapter 5)

Apparatus and materials

Flat-bottomed flask, thistle funnel, gas jar, delivery tube, candle, calcium carbonate (marble chips) and dilute hydrochloric acid, glass plate, grease

Procedure



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Preparation of carbon dioxide from calcium carbonate

- Put some marble chips in a flat-bottomed flask. Set up the apparatus as shown in Figure.
- Pour some dilute hydrochloric acid through the thistle funnel.

  Do you see something happening in the flask?

- Collect the carbon dioxide formed in the gas jar by upward displacement of air.

  What is the colour of carbon dioxide?

  How do you test whether the gas jar is full of the gas or not?

- Place a burning candle or burning splint at the mouth of the gas jar.
  What happens to the burning splint?
- Keep the gas jar with full of carbon dioxide by covering with a greased glass plate.

Observations and results

No.	Experiment	Observation	Inference
1.	Put some marble chips in a flat-bottomed flask.		
2.	Set up the apparatus as shown in Figure.		
3.	Pour some dilute hydrochloric acid through the thistle funnel.	***Do you see anything happening in the flask?***	
4.	Collect the carbon dioxide formed in the gas jar by upward displacement of air.	***What is the colour of carbon dioxide?***	
5.	You have to test whether the gas jar is full of the gas or not. Place a burning candle or splint at the mouth of the gas jar.	***What happens to the burning splint?***	
6.	Keep the gas jar with full of carbon dioxide by covering with a greased glass plate.		

Conclusion

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Experiment 6 (b) Study on the Physical Properties of Carbon Dioxide

Learning Objective

- To study the physical properties of carbon dioxide

Introduction

Carbon dioxide is a colourless, odourless and tasteless gas. It is sparingly soluble in water and heavier than air. Carbon dioxide solidified directly when cooled at atmospheric pressure to give dry ice, i.e., solid carbon dioxide sublimes at -78 °C.

Apparatus and materials

Gas jar with full of carbon dioxide gas, litmus papers, distilled water, a short piece of candle, deflagrating spoon, glass plate, grease

Procedure

(i) To study whether carbon dioxide gas will be acidic or basic

- Take a gas jar with full of carbon dioxide gas.

  What colour do you observe?

- Fan the mouth of the gas jar with a piece of paper, and smell the gas.

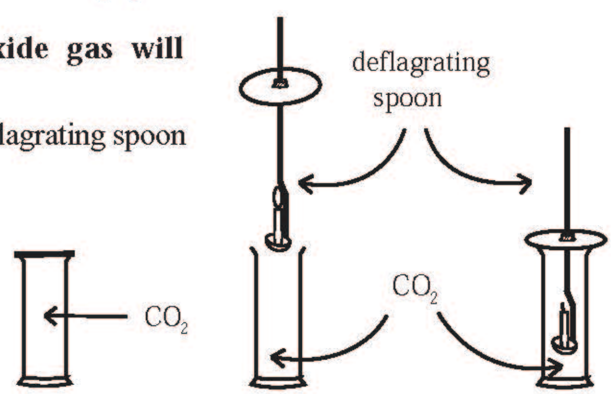
  What is the smell of carbon dioxide?

- Pour about 5 mL of distilled water into the gas jar containing full of carbon dioxide and shake it.
- Test this solution with a piece of blue litmus paper.

  What happens to the colour of blue litmus paper?

(ii) To study whether carbon dioxide gas will support combustion or not

- Fix a short piece of candle in the deflagrating spoon by means of melted wax.
- Light the candle.
- Place the deflagrating spoon with the burning candle in a gas jar full of carbon dioxide gas as shown in figure.



What happens to the burning candle?

Conclusion

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EXPERIMENT 7

Tests for Carbonate, Chloride, Bromide and Iodide Anions

Learning Objective

- To identify the unknown anions in a salt

Introduction

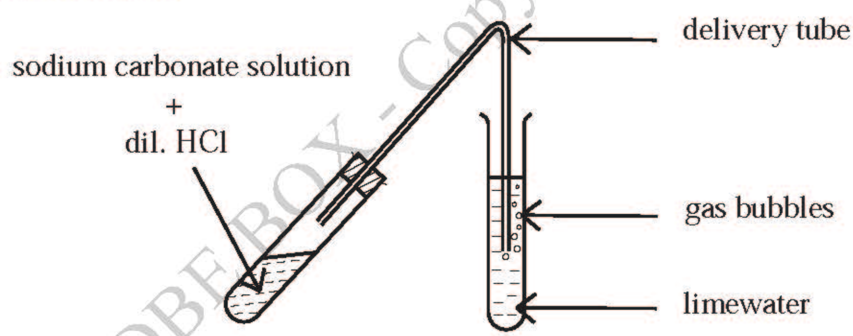
The knowledge of the precipitation reactions can be utilised to identify the unknown anions in the aqueous solution of their salts. For identification of unknown anions in solution, the unknown solution is treated with precipitating reagents. The precipitate formed or any gas evolved is observed and proper inference or conclusion can be drawn. This identification of the anions in solution forms a part of qualitative analysis.

Apparatus and materials

Test tubes, glass rod, glass tube, test tube holder, dilute nitric acid, dilute hydrochloric acid, sliver nitrate solution, sodium carbonate, sodium chloride, potassium bromide, potassium iodide, limewater, distilled water, delivery tube, stopper

Procedure

Test for carbonate ion



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- Add 1 % (w/v) sodium carbonate solution (0.5 g of sodium carbonate in 50 mL of distilled water) into a test tube to a depth about 1 cm.
- Add a few drops of dilute hydrochloric acid solution (5 mL of concentrated hydrochloric acid in 95 mL of distilled water).
- Attach tightly a delivery tube through stopper and pass the gas through limewater as shown in Figure.

Is the gas evolved? Name the gas.

Write down the chemical equation for this test.

What happens when the gas is passed through limewater?

Write down the chemical equation for this reaction.

Test for chloride, bromide and iodide ions

- Prepare salt solutions: sodium chloride, potassium bromide and potassium iodide as the same procedure of preparation of sodium carbonate solution.
- Put the salt solutions separately into each test tube to a depth about 1 cm.
- Add a few drops of dilute nitric acid solution and then add silver nitrate solution into each salt solution.

Why are a few drops of dilute nitric acid solution added?

Has a precipitate formed from each? If so, what colour is observed from each?

Compare your observations with each halide salt solution.

Write down the chemical equations for the chloride, bromide and iodide anions.

Record each result in the table.

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Observations and results

Anion (in solution)	**Test**	**Observation**	**Inferences**
carbonate (CO_3^{2-})			
chloride (Cl^-)			
bromide (Br^-)			
iodide (I^-)			

Conclusion

EXPERIMENT 8

Neutralisation of Acids and Bases and Preparation of Salt

Experiment 8 (a) Neutralisation of Acids and Bases

Learning Objectives

- To understand how to neutralise an acid and a base
- To identify the acid and base

Introduction

Neutralisation is the reaction of acid and base to obtain salt and water. It involves the combination of H^+ ion from acid and OH^- ion from base to generate water.

Apparatus and materials

Beaker, glass rod, glass tube, test tube, watch glass, spirit burner, 5 % (v/v) dilute sulphuric acid solution, distilled water, 1 % (w / v) sodium hydroxide (caustic soda) aqueous solution, blue and red litmus papers

Procedure

- Place 10 drops of 5 % (v/v) dilute sulphuric acid solution into a test tube and test with blue and red litmus papers.
  What do you observe the changes in colour of litmus papers?
- Add drop by drop of 1 % (w/v) sodium hydroxide solution (0.5 g of sodium hydroxide in 50 mL of distilled water) into the above dilute sulphuric acid solution.
- Test the solution with litmus papers after each addition, until both blue and red litmus papers do not change in colour.
  Is the solution neutral or acidic or basic?
- Take a few drops of the solution and evaporate on a watch glass.
  Is there any residue left on a watch glass? What is the colour of the residue?
  Write the chemical equation for this reaction.
  Record the results in the following table.

Observations and results

No.	Experiment	Observation	Inferences
1.			
2.			
3.			

Conclusion

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Experiment 8 (b) Preparation of Copper(II) Sulphate Crystals

Learning Objective

- To prepare copper(II) sulphate crystals

Introduction

Some metal oxides are insoluble in water. All metal oxides are bases and react with acids to give salts. The sulphates and chlorides of copper and lead can be prepared by the action of corresponding acids on the oxides of these metals. Soluble salts (e.g., copper(II) sulphate) are usually prepared by crystallisation method.

Apparatus and materials

Beaker, glass rod, filter paper, spirit burner, tripod stand, evaporating basin (porcelain basin), 5 % (v/v) sulphuric acid solution, copper(II) oxide and distilled water

Procedure

- Put 10 mL of 5 % (v/v) sulphuric acid solution into a beaker and warm it on a water bath.
- Add a little amount of black copper(II) oxide powder (0.5 g) into the above beaker and stir with a glass rod.

  What do you observe? Does the solid disappear or not?

  Is there any change in colour of the solution?

- Add further small amount of copper(II) oxide to the hot acid solution until no more solid dissolves. At this stage, all the acid is used up.
- Filter to remove the excess oxide.

  What substance is in the filtrate?

- Put the filtrate into an evaporating basin (porcelain basin) and heat it.
- Dip a glass rod into the hot solution and hold it up to cool.
- If small crystals form on the glass rod, stop heating, and leave it to stand.

  What is the colour of the crystal?

  What is the name of crystal obtained?

  Name the method used to obtain the copper(II) sulphate crystal.

  Write a chemical equation for this experiment.

Conclusion

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EXPERIMENT 9

Differentiation between Temporary and Permanent Hardness of Water

Learning Objectives

- To understand the chemical basis of water hardness
- To differentiate between temporary and permanent hard water

Introduction

Water is the most abundant compound on the Earth and because of its abundance, the importance of water is overlooked in daily life. The importance of water is due to its widespread occurrence and unique property. There is a shortage of fresh water in many parts of the world and most of the water is salty and hard. In this experiment you will have to study how to distinguish between temporary and permanent hard water.

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Apparatus and materials

Beaker, conical flasks, spirit burner, measuring cylinder, burette, test tubes, limewater (calcium hydroxide), calcium sulphate or magnesium sulphate (Epsom salt), 5 % (v/v) hydrochloric acid, solid soap, methylated spirit and distilled water

Procedure

(i) Preparation of soap solution

- Dissolve 1 g of pure solid soap in 100 mL of mixed solvent which is prepared by mixing equal volume of distilled water and methylated spirit.

(ii) Preparation of temporary hard water

- Take 150 mL of limewater and dilute with equal volume of distilled water.
- Pass carbon dioxide gas from the gas generator (for preparation of CO_2 gas, see Experiment 6) into limewater solution to form insoluble calcium carbonate.
- Continue the passage of carbon dioxide gas until all the calcium carbonate precipitate re-dissolves, resulting in a clear solution of calcium hydrogen carbonate.

(iii) Preparation of permanent hard water

- Prepare saturated calcium sulphate or magnesium sulphate solution and allow to settle for a few minutes.
- Decant the supernatant liquid into a beaker after 10 minutes and dilute with an equal volume of distilled water.

(iv) Differentiation between temporary and permanent hardness

- Take half of the temporary hard water into a beaker and boil for about 20 minutes (or until the white precipitate is formed).

  Name the precipitate.

- Pour the supernatant liquid into another beaker.
- Transfer the precipitate into a test tube and add 5 % (v/v) hydrochloric acid solution.

  What do you observe? Give the name of the gas.
- Take 10 mL each of the boiled and unboiled solutions and add into 50 mL conical flasks separately.
- Add solid soap solution from a burette to the boiled solution with vigorous shaking until a permanent lather is obtained.
- Repeat the experiment with the unboiled solution.

  What differences do you observe between the two solutions?

  What is the cause of the temporary hardness?
- Repeat the experiments (both the boiled and unboiled after adding solid soap solution) using permanent hard water (calcium sulphate or magnesium sulphate solution).

  What do you observe?
- Record the effect of boiling on its hardness.

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Observation and results

Experiment	Observation	Inferences

Conclusion

EXPERIMENT 10

Examination of Alkalinity and Acidity of Soil

Learning Objectives

- To examine the soil alkalinity and acidity
- To determine the presence or absence of lime

Introduction

Alkalinity and acidity of soil are important for the growing of plants. In this experiment, you will investigate whether the soil is alkaline or acidic by using litmus paper and also find out the presence of lime in soil.

Apparatus and materials

Beaker, evaporating basin (porcelain basin), test tube with stopper, glass rod, soil sample (garden soil), 5 % (v/v) hydrochloric acid solution, distilled water, litmus or pH papers, a sheet of paper and funnel

Experiment 10 (a) To examine the soil alkalinity and acidity

Procedure

- Take soil samples of garden from three different sites of the garden plot in your school compound or any other places to be tested.
- Collect the soil sample for each site as shown in Figure.

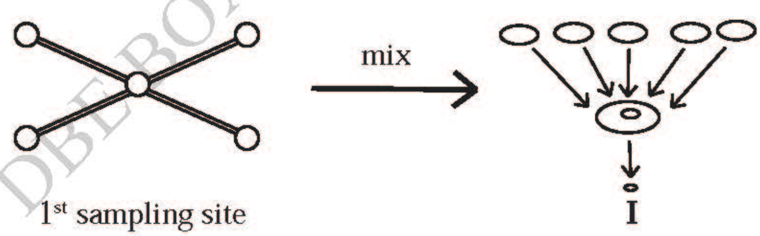

Figure Sampling design (grid and composite sampling for each point)

- Mix these five samples thoroughly from each sampling site on a sheet of paper.
- Denote **I** for 1st sampling site, **II** for 2nd and **III** for 3rd.
- Transfer each soil sample into separate beakers.

  Why is it necessary to take soil samples from different parts?

- Take the soil sample in the test tube to a depth of about 4 cm. Add water onto it to fill half of the test tube. Stopper it and shake well.
- Filter the contents and test the filtrate individually with blue and red litmus papers or pH paper.

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- To examine the soil acidity or alkalinity, observe the changes in colour of litmus or pH papers.

Experiment 10 (b) To determine the presence or absence of lime in the soil

- Take the soil sample in the test tube to a depth of about 4 cm. Pour dilute hydrochloric acid onto it to fill half of the test tube. If the liquid fizzes and bubbles freely, it indicates that the soil contains a sufficient quantity of lime. Lime in this sense means limestone, $CaCO_3$.
- If it only effervesces feebly, the lime content in the soil may be insufficient.
- Observe the results of your test and decide whether the soil contains sufficient quantity of lime or not.

  If any, write the chemical equation in words for the reaction between lime in soil sample and dilute hydrochloric acid.

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Observations and results

Test	Sampling site	Observation	Inferences
(a) Test the filtrates with litmus papers	I		
	II		
	III		
(b) Test the soil samples for the presence or absence of lime	I		
	II		
	III		

Conclusion

EXPERIMENT 11

Examination of Relationship between Number of Carbon Atoms in Petroleum Products and Their Flow Time

Learning Objective

- To compare the flow time of petroleum products related with number of carbon atoms

Introduction

Physical properties of materials such as colour, density, viscosity, melting point and boiling point are related with the molecules and atoms of the materials. Due to high intermolecular forces among long chain hydrocarbons, the materials with larger number of carbons become more viscous and more difficult to flow. Therefore, the number of carbon atoms in the liquids can be estimated by their flow time.

Apparatus and materials

Four samples of petroleum products [gasoline (petrol), diesel oil, lubricating oil (engine oil) and ethanol], 10 mL pipettes, beakers, stopwatch, suction rubber bulb

Procedure

- Place individually four samples of petroleum products in beakers.
- Take each sample with 10 mL pipette using suction rubber bulb.
- Record the time taken for drop off the liquid sample from the mark of the pipette using a stopwatch.
- Decide which petroleum product is more viscous.
- Repeat this procedure three times.
- Arrange these petroleum products in order of increasing carbon atoms. (Take the reference from Table 8.1 in Grade 10 Chemistry Textbook.)

CAUTION

- Keep away the petroleum products from naked flame.

Observation and results

No.	Sample (10 mL)	Flow Time (s)
1.		
2.		

Conclusion

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Useful Units Used In Laboratory

Symbol of unit	Name of unit	Quantity measured
$g\ mol^{-1}$	gram per mole	amount of substance in one mole
mol	mole	amount of substance (number of substance)
$mol\ dm^{-3}$	mole per cubic decimetre	concentration of solution
ppm	parts per million	concentration of solution
$mg\ L^{-1}$	milligram per litre	concentration of solution
$g\ cm^{-3}$	gram per cubic centimetre	density
$g\ dm^{-3}$	gram per cubic decimetre	density
cm	centimetre	dimensions: length, width, height
dm	decimetre	dimensions: length, width, height
g	gram	mass
amu	atomic mass unit	mass of small particle
$molecules\ mol^{-1}$	molecules per mole	number of molecules in one mole
atm	atmosphere	pressure
°C	degree celsius	temperature
K	kelvin	temperature
°F	degree fahrenheit	temperature
h	hour	time
min	minute	time
s	second	time
cm^3	cubic centimetre	volume
dm^3	cubic decimetre	volume
L	litre	volume
mL	millilitre	volume
$dm^3\ mol^{-1}$	cubic decimetre per mole	volume of gas in one mole

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