COMPETENCY BASED EDUCATION

KENYA CERTIFICATE OF BASIC EDUCATION SENIOR SCHOOL CURRICULUM

GRADE 11 CHEMISTRY NOTES



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STRAND 1.0: INORGANIC CHEMISTRY

SUB-STRAND 1.1: THE MOLE

1.0 INTRODUCTION TO THE MOLE

Chemistry involves the study of matter, its composition, properties and the changes it undergoes. Chemical reactions involve atoms, molecules and ions combining or separating in definite proportions.

Since atoms and molecules are extremely small, it is difficult to count them individually. Chemists therefore use a special counting unit known as the **mole** to express the amount of a substance.

The mole enables chemists to measure quantities of substances accurately during chemical reactions, laboratory experiments and industrial processes.

2.0 AMOUNT OF SUBSTANCE

2.1 Meaning of amount of substance

Amount of substance is the quantity of matter expressed in terms of the number of specified elementary particles present in a sample.

The elementary particles may include:

- Atoms.
- Molecules.
- Ions.
- Electrons.
- Formula units.

The SI unit for amount of substance is the **mole (mol).**

For example:

- Oxygen gas contains oxygen molecules (O_2).
- Sodium chloride contains sodium ions (Na^+) and chloride ions (Cl^-).
- Copper contains copper atoms (Cu).
- Water contains water molecules (H_2O).

2.2 The mole

A mole is the amount of substance that contains exactly $6.02214076 \times 10^{23}$ specified elementary entities.

This number is called the **Avogadro constant.**

For ordinary calculations, it is commonly approximated as:

$$N_A = 6.02 \times 10^{23}\ mol^{-1}$$

Therefore:

- 1 mole of carbon atoms contains 6.02×10^{23} carbon atoms.
- 1 mole of water contains 6.02×10^{23} water molecules.
- 1 mole of sodium ions contains 6.02×10^{23} sodium ions.

Important: Always identify the particles being counted. One mole of oxygen molecules is not the same as one mole of oxygen atoms.

2.3 Understanding the mole through everyday examples

The mole is a counting unit, just like a dozen.

Counting unit	Number of items
One pair	2
One dozen	12
One score	20
One gross	144
One mole	6.02×10^{23}

For example, when a trader in a Kenyan market buys one dozen eggs, the trader receives 12 eggs.

Similarly, when a chemist has one mole of water molecules, the chemist has 6.02×10^{23} water molecules.

The difference is that the mole represents an extremely large number of particles.

3.0 AVOGADRO'S CONSTANT

3.1 Meaning of Avogadro's constant

Avogadro's constant is the number of elementary particles contained in one mole of a substance.

It is represented by N_A.

$$N_A = 6.02 \times 10^{23}\ mol^{-1}$$

3.2 Importance of Avogadro's constant

Avogadro's constant is important because it:

1. Enables chemists to count extremely small particles indirectly.
2. Establishes the relationship between the mole and the number of particles.
3. Helps determine the number of atoms in a given mass of an element.
4. Helps calculate the number of molecules in a given mass of a compound.
5. Enables calculations involving chemical equations.
6. Helps determine the number of ions in ionic compounds.
7. Provides a connection between microscopic particles and measurable quantities of substances.

3.3 Relationship between moles and number of particles

The number of particles in a substance depends on the amount of substance in moles.

The relationship is:

$$N = n \times N_A$$

Where:

- N = Number of particles.
- n = Amount of substance in moles.
- N_A = Avogadro's constant.

To calculate the number of moles:

$$n = \frac{N}{N_A}$$

Worked example 1

Calculate the number of molecules in 2 moles of water.

Solution

Given:

$$n = 2\ mol$$

$$N_A = 6.02 \times 10^{23}$$

Using:

$$N = n \times N_A$$

$$N = 2 \times 6.02 \times 10^{23}$$

$$\boxed{N = 1.204 \times 10^{24}\ molecules}$$

Worked example 2

Calculate the number of atoms in 0.5 moles of copper.

Solution

$$N = n \times N_A$$

$$N = 0.5 \times 6.02 \times 10^{23}$$

$$\boxed{N = 3.01 \times 10^{23}\ atoms}$$

Worked example 3

Determine the number of moles represented by 3.01×10^{23} carbon atoms.

Solution

$$n = \frac{N}{N_A}$$

$$n = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$$

$$\boxed{n = 0.5\ mol}$$

3.4 Number of atoms in molecules

Some substances exist as molecules containing more than one atom.

For example:

- Hydrogen: H_2 contains 2 hydrogen atoms per molecule.
- Oxygen: O_2 contains 2 oxygen atoms per molecule.
- Water: H_2O contains 3 atoms per molecule.
- Carbon dioxide: CO_2 contains 3 atoms per molecule.
- Ammonia: NH_3 contains 4 atoms per molecule.

Worked example 4

Calculate the total number of atoms in 2 moles of water molecules.

One water molecule contains:

$$2 + 1 = 3\ atoms$$

Therefore:

$$N = 2 \times 6.02 \times 10^{23} \times 3$$

$$\boxed{N = 3.612 \times 10^{24}\ atoms}$$

4.0 RELATIVE ATOMIC MASS

4.1 Meaning of relative atomic mass

Relative atomic mass is the weighted average mass of the atoms of an element compared with one-twelfth of the mass of a carbon-12 atom.

It is represented by A_r.

Relative atomic mass has no units because it is a ratio.

For example:

Element	Symbol	Relative atomic mass
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Aluminium	Al	27
Sulphur	S	32
Chlorine	Cl	35.5
Potassium	K	39
Calcium	Ca	40
Iron	Fe	56
Copper	Cu	63.5
Zinc	Zn	65

These are commonly used approximate values for school-level chemical calculations.

4.2 Importance of relative atomic mass

Relative atomic mass is used to:

1. Calculate relative molecular mass.
2. Determine relative formula mass.
3. Calculate molar mass.
4. Convert mass into moles.
5. Determine empirical formulae.
6. Determine molecular formulae.
7. Carry out chemical equation calculations.

4.3 Relative molecular mass

Relative molecular mass is the sum of the relative atomic masses of all the atoms present in one molecule.

It is represented by M_r.

$$M_r = \text{Sum of relative atomic masses}$$

Worked example 5

Calculate the relative molecular mass of water (H_2O).

Given:

- H = 1
- O = 16

Solution:

$$M_r = (2 \times 1) + 16$$

$$M_r = 2 + 16$$

$$\boxed{M_r = 18}$$

Worked example 6

Calculate the relative molecular mass of sulphuric acid (H_2SO_4).

Given:

- H = 1
- S = 32
- O = 16

Solution:

$$M_r = (2 \times 1) + 32 + (4 \times 16)$$

$$= 2 + 32 + 64$$

$$\boxed{\boldsymbol{M_r} = 98}$$

4.4 Relative formula mass

Relative formula mass is the sum of the relative atomic masses of all the atoms represented in the formula unit of an ionic compound.

For example, sodium chloride (NaCl):

$$M_r = 23 + 35.5$$

$$\boxed{M_r = 58.5}$$

Other examples include:

- Magnesium oxide (MgO): 24 + 16 = 40.
- Calcium carbonate ($CaCO_3$): 40 + 12 + (3 × 16) = 100.
- Sodium carbonate (Na_2CO_3): (2 × 23) + 12 + (3 × 16) = 106.

4.5 Molar mass

Molar mass is the mass of one mole of a substance.

It is expressed in grams per mole (g mol^{-1}).

$$\text{Molar mass} = \frac{\text{Mass}}{\text{Number of moles}}$$

The numerical value of molar mass is equal to the relative atomic, molecular or formula mass, but it has units.

Examples:

- Carbon: 12 g mol^{-1}.
- Oxygen gas (O_2): 32 g mol^{-1}.
- Water (H_2O): 18 g mol^{-1}.
- Sodium chloride (NaCl): 58.5 g mol^{-1}.

5.0 CONVERSION OF MASS TO MOLES

5.1 Introduction

In a chemistry laboratory, substances are normally measured using a balance. However, chemical equations are interpreted using amounts of substances in moles.

Chemists therefore convert the measured mass of a substance into moles before carrying out many chemical calculations.

5.2 Relationship between mass, moles and molar mass

The relationship is:

$$\boxed{n = \frac{m}{M}}$$

Where:

- n = Number of moles (mol).
- m = Mass of substance (g).
- M = Molar mass (g mol^{-1}).

Rearranging:

$$m = n \times M$$

5.3 Steps followed when converting mass into moles

1. Identify the chemical formula of the substance.
2. Determine the relative atomic masses of the elements.
3. Calculate the relative molecular or formula mass.
4. Express the molar mass in g mol^{-1}.
5. Identify the given mass.
6. Substitute the values into the formula.
7. Calculate the number of moles.
8. Write the correct unit.

Worked example 7

Calculate the number of moles in 36 g of water.

Given:

- Mass = 36 g.
- Molar mass of water = 18 g mol^{-1}.

Using:

$$n = \frac{m}{M}$$

$$n = \frac{36}{18}$$

$$\boxed{n = 2\ mol}$$

Worked example 8

Calculate the number of moles in 49 g of sulphuric acid (H_2SO_4).

Given:

$$M = 98\ g\ mol^{-1}$$

$$n = \frac{49}{98}$$

$$\boxed{n = 0.5\ mol}$$

Worked example 9

Calculate the number of moles in 11.7 g of sodium chloride.

Molar mass:

$$M = 23 + 35.5 = 58.5\ g\ mol^{-1}$$

Therefore:

$$n = \frac{11.7}{58.5}$$

$$\boxed{n = 0.2\ mol}$$

6.0 CONVERSION OF MOLES TO MASS

6.1 Introduction

Sometimes a chemist knows the required number of moles but needs to determine the mass to measure using a laboratory balance.

The mass is calculated by multiplying the number of moles by the molar mass.

6.2 Formula

$$\boxed{m = n \times M}$$

Where:

- m = Mass in grams.
- n = Number of moles.
- M = Molar mass in g mol^{-1}.

Worked example 10

Calculate the mass of 2 moles of carbon dioxide.

Given:

- C = 12.
- O = 16.

Molar mass:

$$M = 12 + (2 \times 16)$$

$$M = 44\ g\ mol^{-1}$$

Mass:

$$m = n \times M$$

$$m = 2 \times 44$$

$$\boxed{m = 88\ g}$$

Worked example 11

Calculate the mass of 0.25 moles of calcium carbonate ($CaCO_3$).

Molar mass:

$$M = 40 + 12 + (3 \times 16)$$

$$M = 100\ g\ mol^{-1}$$

Therefore:

$$m = 0.25 \times 100$$

$$\boxed{m = 25\ g}$$

Worked example 12

What mass of magnesium oxide contains 0.5 moles of the compound?

Given:

- Mg = 24.
- O = 16.

$$M = 24 + 16 = 40$$

$$m = 0.5 \times 40$$

$$\boxed{m = 20\ g}$$

7.0 DETERMINATION OF CHEMICAL FORMULAE

A chemical formula represents the elements present in a substance and the relative number of atoms of each element.

There are two important types of formulae studied in this sub-strand:

1. Empirical formula.
2. Molecular formula.

7.1 Empirical formula

Meaning

An empirical formula is the simplest whole-number ratio of the atoms of different elements present in a compound.

It does not necessarily show the actual number of atoms in one molecule.

Examples:

Molecular formula	Empirical formula
H_2O_2	HO
$C_6H_{12}O_6$	CH_2O
C_2H_4	CH_2
C_6H_6	CH
N_2O_4	NO_2
CH_4	CH_4

7.2 Importance of empirical formulae

Empirical formulae are useful because they:

1. Show the simplest atomic ratio in a compound.
2. Help identify the composition of unknown compounds.
3. Assist in determining molecular formulae.
4. Help chemists interpret experimental results.
5. Are used in determining the composition of substances in chemical industries.

7.3 Methods of determining empirical formula

An empirical formula can be determined from:

- Percentage composition of elements.
- Mass composition of elements.
- Experimental combination of elements.
- Results obtained from chemical analysis.

7.4 General procedure for determining empirical formula

Follow these steps:

1. Write down the mass or percentage of each element.
2. If percentages are given, assume 100 g of the compound.
3. Convert the mass of each element into moles by dividing by its relative atomic mass.
4. Divide all mole values by the smallest mole value.
5. Obtain the simplest whole-number ratio.
6. Multiply the ratios where necessary to obtain whole numbers.
7. Write the empirical formula using the calculated ratios.

Worked example 13: Determining empirical formula from percentage composition

A compound contains 40% carbon, 6.67% hydrogen and 53.33% oxygen. Determine its empirical formula.

Given:

Element	Percentage	Relative atomic mass

Element	Percentage	Relative atomic mass
Carbon	40	12
Hydrogen	6.67	1
Oxygen	53.33	16

Step 1: Assume 100 g of the compound.

Therefore:

- Carbon = 40 g.
- Hydrogen = 6.67 g.
- Oxygen = 53.33 g.

Step 2: Convert masses into moles.

$$n = \frac{m}{A_r}$$

Carbon:

$$\frac{40}{12} = 3.33$$

Hydrogen:

$$\frac{6.67}{1} = 6.67$$

Oxygen:

$$\frac{53.33}{16} = 3.33$$

Step 3: Divide by the smallest mole value.

Element	Moles	Divide by 3.33	Ratio
C	3.33	3.33/3.33	1
H	6.67	6.67/3.33	2
O	3.33	3.33/3.33	1

Step 4: Write the empirical formula.

Ratio:

$$C:H:O = 1:2:1$$

Final answer

$$\boxed{CH_2O}$$

Worked example 14

A compound contains 24 g of carbon and 4 g of hydrogen. Determine its empirical formula.

Step 1: Calculate the number of moles.

Carbon:

$$n = \frac{24}{12} = 2$$

Hydrogen:

$$n = \frac{4}{1} = 4$$

Step 2: Divide by the smallest value.

Carbon:

$$\frac{2}{2} = 1$$

Hydrogen:

$$\frac{4}{2} = 2$$

Step 3: Write the formula.

$$\boxed{CH_2}$$

7.5 Experimental determination of empirical formula of magnesium oxide

Magnesium oxide is formed when magnesium reacts with oxygen.

The experiment can be used to determine the simplest ratio of magnesium atoms to oxygen atoms.

Aim

To determine the empirical formula of magnesium oxide by heating magnesium in air.

Requirements

- Magnesium ribbon.
- Crucible with lid.
- Crucible tongs.
- Bunsen burner.
- Electronic balance.
- Sandpaper.
- Pipe-clay triangle.
- Heat-resistant mat.

Procedure

1. Clean a piece of magnesium ribbon using sandpaper to remove the oxide layer.
2. Weigh a clean, dry crucible together with its lid.
3. Place the magnesium ribbon in the crucible.
4. Weigh the crucible, lid and magnesium ribbon.
5. Set up the crucible on a pipe-clay triangle.
6. Heat gently at first, then strongly, allowing oxygen to react with magnesium.
7. Keep the lid slightly open at intervals to allow oxygen into the crucible while preventing loss of the white product.
8. Heat until the magnesium has reacted completely.
9. Allow the crucible to cool.
10. Weigh the crucible, lid and magnesium oxide.
11. Repeat heating, cooling and weighing until a constant mass is obtained.

Observations

- Magnesium burns with a bright white flame.

- A white powder is formed.
- The white powder is magnesium oxide.

Relevant chemical equation

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Results table

Measurement	Mass
Empty crucible and lid	A g
Crucible, lid and magnesium	B g
Crucible, lid and magnesium oxide	C g

Mass of magnesium:

$$B - A$$

Mass of magnesium oxide:

$$C - A$$

Mass of oxygen:

$$(C - A) - (B - A) = C - B$$

Calculation

1. Moles of magnesium:

$$\frac{\text{Mass of magnesium}}{24}$$

2. Moles of oxygen:

$$\frac{\text{Mass of oxygen}}{16}$$

3. Divide both mole values by the smaller value.
4. Obtain the simplest whole-number ratio.

The expected empirical formula is:

$$\boxed{MgO}$$

Precautions

- Use clean magnesium ribbon.
- Avoid inhaling magnesium oxide fumes.
- Do not look directly at the intense white flame.
- Use crucible tongs when handling hot apparatus.
- Allow the crucible to cool before weighing.
- Heat carefully to avoid loss of the product.

7.6 Experimental determination of empirical formula of copper(II) oxide

Aim

To determine the empirical formula of copper(II) oxide using the reduction of copper(II) oxide by hydrogen gas.

Requirements

- Copper(II) oxide.
- Hydrogen gas.
- Hard glass tube.
- Delivery tubes.
- Bunsen burner.
- Balance.
- Heat-resistant apparatus.

Procedure

1. Weigh a clean, dry hard glass tube containing a known mass of copper(II) oxide.
2. Set up the apparatus to allow hydrogen gas to pass through the tube.
3. Allow hydrogen to flow through the apparatus for some time to displace air.
4. Heat the copper(II) oxide strongly while hydrogen continues to pass through it.
5. Observe the change in colour from black to reddish-brown.
6. Continue heating until the reaction is complete.
7. Allow the apparatus to cool while hydrogen is still flowing.
8. Stop the hydrogen supply only after the apparatus has cooled sufficiently.
9. Reweigh the tube and its contents.

Chemical equation

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Observations

- Black copper(II) oxide changes into reddish-brown copper.
- Water vapour is produced.

Calculations

Mass of oxygen removed:

$$\text{Initial mass of CuO} - \text{Final mass of Cu}$$

Moles of copper:

$$\frac{\text{Mass of copper}}{63.5}$$

Moles of oxygen:

$$\frac{\text{Mass of oxygen}}{16}$$

Determine the simplest ratio.

The expected empirical formula is:

$$\boxed{CuO}$$

Safety: Hydrogen is highly flammable. The apparatus must be checked for leaks, and hydrogen must be passed through the apparatus to displace air before heating. This experiment should only be conducted under appropriate teacher supervision and laboratory safety procedures.

8.0 MOLECULAR FORMULA

8.1 Meaning of molecular formula

A molecular formula shows the actual number of atoms of each element present in one molecule of a compound.

Unlike the empirical formula, the molecular formula gives the actual composition of a molecule.

Examples:

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Glucose	CH_2O	$C_6H_{12}O_6$
Ethene	CH_2	C_2H_4
Benzene	CH	C_6H_6
Dinitrogen tetroxide	NO_2	N_2O_4
Water	H_2O	H_2O

8.2 Relationship between empirical and molecular formulae

The molecular formula is a whole-number multiple of the empirical formula.

$$\boxed{\text{Molecular formula} = (\text{Empirical formula})_k}$$

Where:

$$k = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

8.3 Steps for determining molecular formula

1. Determine the empirical formula.
2. Calculate the empirical formula mass.
3. Obtain the relative molecular mass of the compound.
4. Divide the relative molecular mass by the empirical formula mass.
5. Multiply the subscripts in the empirical formula by the resulting whole number.
6. Write the molecular formula.

Worked example 15

A compound has an empirical formula of CH_2O and a relative molecular mass of 180. Determine its molecular formula.

Step 1: Calculate empirical formula mass.

$$12 + (2 \times 1) + 16 = 30$$

Step 2: Calculate the multiplier.

$$k = \frac{180}{30} = 6$$

Step 3: Multiply the subscripts by 6.

CH_2O becomes:

$$C_6H_{12}O_6$$

Final answer

$$\boxed{C_6H_{12}O_6}$$

Worked example 16

A compound has an empirical formula of CH_2 and a molecular mass of 56. Determine its molecular formula.

Empirical formula mass:

$$12 + (2 \times 1) = 14$$

Multiplier:

$$k = \frac{56}{14} = 4$$

Multiply the subscripts by 4:

$$\boxed{C_4H_8}$$

9.0 MOLAR SOLUTIONS

9.1 Meaning of a solution

A solution is a homogeneous mixture formed when one or more solutes dissolve in a solvent.

A solution has a uniform composition throughout.

Examples from everyday life in Kenya include:

- Salt dissolved in water.
- Sugar dissolved in tea.
- Glucose dissolved in water.
- Copper(II) sulphate dissolved in water.

9.2 Components of a solution

A solution has two main components.

1. Solute

A solute is the substance that dissolves in a solvent to form a solution.

Examples:

- Sodium chloride.
- Sugar.
- Copper(II) sulphate.
- Potassium nitrate.

2. Solvent

A solvent is the substance that dissolves a solute.

Water is the most commonly used solvent in school chemistry laboratories.

9.3 Meaning of a molar solution

A molar solution is a solution whose concentration is expressed in moles of solute per litre of solution.

For example, a 1 mol dm^{-3} sodium chloride solution contains one mole of sodium chloride dissolved to make one cubic decimetre (one litre) of solution.

9.4 Molarity

Molarity is the number of moles of solute dissolved in one cubic decimetre of solution.

It is represented by C.

$$\boxed{C = \frac{n}{V}}$$

Where:

- C = Concentration in mol dm^{-3}.
- n = Number of moles.
- V = Volume in dm^3.

Since:

$$n = \frac{m}{M}$$

Then:

$$C = \frac{m}{MV}$$

9.5 Conversion of volume units

Chemists commonly use cubic centimetres (cm^3) and cubic decimetres (dm^3) when measuring solution volumes.

The relationship is:

$$1\ dm^3 = 1000\ cm^3$$

Therefore:

$$1\ cm^3 = 0.001\ dm^3$$

To convert cm^3 into dm^3:

$$\boxed{V(dm^3) = \frac{V(cm^3)}{1000}}$$

Examples:

Volume in cm^3	Volume in dm^3
50	0.05
100	0.10
200	0.20
250	0.25
500	0.50
1000	1.00

Worked example 17

Calculate the concentration of a solution containing 0.5 moles of sodium hydroxide in 250 cm^3 of solution.

Given:

$$n = 0.5\ mol$$

$$V = \frac{250}{1000} = 0.25\ dm^3$$

Using:

$$C = \frac{n}{V}$$

$$C = \frac{0.5}{0.25}$$

$$\boxed{C = 2\ mol\ dm^{-3}}$$

Worked example 18

Calculate the number of moles in 500 cm^3 of a 0.2 mol dm^{-3} hydrochloric acid solution.

Given:

$$C = 0.2\ mol\ dm^{-3}$$

$$V = 0.5\ dm^3$$

Using:

$$n = C \times V$$

$$n = 0.2 \times 0.5$$

$$\boxed{n = 0.1\ mol}$$

10.0 PREPARATION OF STANDARD SOLUTIONS

10.1 Meaning of a standard solution

A standard solution is a solution whose concentration is accurately known.

Standard solutions are used in quantitative chemical analysis, particularly in titrations.

10.2 Importance of standard solutions

Standard solutions are important because they:

1. Help determine the concentration of unknown solutions.
2. Are used in acid-base titrations.
3. Are used in redox titrations.
4. Help determine the purity of substances.
5. Are used in quality control in chemical industries.
6. Help analyse substances in water treatment plants.
7. Are used in pharmaceutical industries.
8. Help determine the amount of a substance in a sample.

10.3 Apparatus used in preparing standard solutions

The following apparatus may be used:

1. Analytical balance – measures the mass of the solute accurately.
2. Weighing boat – holds the solid while weighing.
3. Spatula – transfers small quantities of solid chemicals.
4. Beaker – holds the solution during initial dissolving.

5. Glass rod – helps in stirring.
6. Funnel – transfers solution into a volumetric flask.
7. Volumetric flask – measures a fixed volume accurately.
8. Wash bottle – rinses apparatus using distilled water.
9. Dropper – helps adjust the solution level to the calibration mark.
10. Stopper – closes the volumetric flask.

10.4 Preparation of a standard solution from a solid substance

Example: Preparation of 250 cm^3 of 0.1 mol dm^{-3} sodium carbonate solution.

Step 1: Calculate the required mass.

Given:

$$C = 0.1\ mol\ dm^{-3}$$

$$V = 250\ cm^3 = 0.25\ dm^3$$

Molar mass of anhydrous sodium carbonate (Na_2CO_3):

$$M = (2 \times 23) + 12 + (3 \times 16)$$

$$M = 106\ g\ mol^{-1}$$

Calculate the number of moles:

$$n = C \times V$$

$$n = 0.1 \times 0.25$$

$$n = 0.025\ mol$$

Calculate the mass:

$$m = n \times M$$

$$m = 0.025 \times 106$$

$$\boxed{m = 2.65\ g}$$

Therefore, 2.65 g of anhydrous sodium carbonate is required.

Step 2: Weigh the solid.

- Place a clean, dry weighing boat on the analytical balance.
- Adjust the balance to zero.
- Use a spatula to add sodium carbonate until 2.65 g is obtained.

Step 3: Dissolve the solid.

- Transfer the sodium carbonate into a clean beaker.
- Add a small amount of distilled water.
- Stir using a glass rod until the solid dissolves completely.

Step 4: Transfer into the volumetric flask.

- Place a funnel in the neck of a 250 cm^3 volumetric flask.
- Carefully transfer the solution into the flask.
- Rinse the beaker and glass rod with distilled water.
- Transfer the rinsings into the volumetric flask to ensure complete transfer.

Step 5: Make up to the mark.

- Add distilled water until the solution approaches the calibration mark.
- Use a dropper to add the final drops.
- Ensure the bottom of the meniscus is at the calibration mark when viewed at eye level.

Step 6: Mix the solution.

- Stopper the flask.
- Invert it several times to mix the contents thoroughly.

Step 7: Label the solution.

The label should indicate:

- Name of the solution.
- Concentration.
- Date of preparation.
- Relevant safety information.

The prepared solution is 250 cm^3 of 0.1 mol dm^{-3} sodium carbonate.

10.5 Preparation of standard solutions from liquid substances

Some standard solutions are prepared by diluting concentrated solutions.

For example, hydrochloric acid may be supplied as a concentrated solution.

The concentration of the required solution is obtained by diluting the concentrated acid with distilled water.

The dilution relationship is:

$$\boxed{C_1V_1 = C_2V_2}$$

Where:

- C_1 = Initial concentration.
- V_1 = Volume of concentrated solution required.
- C_2 = Final concentration.
- V_2 = Final volume of diluted solution.

Worked example 19

Calculate the volume of 2 mol dm^{-3} hydrochloric acid required to prepare 250 cm^3 of 0.2 mol dm^{-3} hydrochloric acid.

Given:

$$C_1 = 2$$

$$C_2 = 0.2$$

$$V_2 = 250\ cm^3$$

Using:

$$C_1V_1 = C_2V_2$$

$$2 \times V_1 = 0.2 \times 250$$

$$2V_1 = 50$$

$$V_1 = \frac{50}{2}$$

$$\boxed{V_1 = 25\ cm^3}$$

Therefore, 25 cm^3 of the concentrated hydrochloric acid is required.

Add it carefully to distilled water and make up to a final volume of 250 cm^3.

Important laboratory safety: Always add acid to water slowly, never water to concentrated acid. Wear appropriate eye protection, gloves and a laboratory coat. Preparation of concentrated acids should be conducted under teacher supervision.

11.0 CONCENTRATION OF SOLUTIONS

11.1 Meaning of concentration

Concentration is the amount of solute dissolved in a given volume of solution.

A concentrated solution contains a relatively large amount of solute per unit volume, while a dilute solution contains a relatively small amount.

11.2 Methods of expressing concentration

Concentration may be expressed in different ways.

1. Moles per cubic decimetre (mol dm^{-3}).
2. Grams per cubic decimetre (g dm^{-3}).
3. Percentage concentration.
4. Parts per million (ppm), particularly for very dilute solutions.

11.3 Concentration in grams per cubic decimetre

This expresses the mass of solute dissolved in one cubic decimetre of solution.

$$\boxed{\text{Concentration} = \frac{\text{Mass of solute (g)}}{\text{Volume of solution (dm}^3\text{)}}}$$

Worked example 20

Calculate the concentration in g dm^{-3} of a solution containing 10 g of sodium chloride in 500 cm^3 of solution.

Convert volume:

$$V = \frac{500}{1000} = 0.5\ dm^3$$

Therefore:

$$C = \frac{10}{0.5}$$

$$\boxed{C = 20\ g\ dm^{-3}}$$

11.4 Relationship between g dm^{-3} and mol dm^{-3}

The concentration in g dm^{-3} can be converted into mol dm^{-3} by dividing by the molar mass.

$$\boxed{C(mol\ dm^{-3}) = \frac{C(g\ dm^{-3})}{M}}$$

Worked example 21

A sodium hydroxide solution has a concentration of 8 g dm^{-3}. Calculate its molar concentration.

Molar mass of NaOH:

$$23 + 16 + 1 = 40$$

Therefore:

$$C = \frac{8}{40}$$

$$\boxed{C = 0.2\ mol\ dm^{-3}}$$

12.0 DILUTION OF SOLUTIONS

12.1 Meaning of dilution

Dilution is the process of reducing the concentration of a solution by adding more solvent without changing the amount of solute present.

Water is commonly used to dilute aqueous solutions.

For example, adding distilled water to concentrated hydrochloric acid solution reduces its concentration.

12.2 Factors affecting dilution

1. Initial concentration of the solution.
2. Volume of the original solution.
3. Volume of solvent added.
4. Final volume of the solution.

12.3 Dilution formula

$$\boxed{C_1V_1 = C_2V_2}$$

The relationship is based on the fact that the number of moles of solute remains constant during dilution.

Worked example 22

100 cm^3 of a 2 mol dm^{-3} solution is diluted to 500 cm^3. Calculate the final concentration.

Given:

$$C_1 = 2$$

$$V_1 = 100$$

$$V_2 = 500$$

$$C_2 = ?$$

Using:

$$C_1V_1 = C_2V_2$$

$$2 \times 100 = C_2 \times 500$$

$$C_2 = \frac{200}{500}$$

$$\boxed{C_2 = 0.4\ mol\ dm^{-3}}$$

12.4 Importance of dilution

Dilution is important in:

1. Preparing solutions of required concentrations.
2. Preparing laboratory reagents.
3. Conducting chemical analysis.
4. Preparing medicines and pharmaceutical solutions.
5. Water treatment.
6. Food processing.
7. Industrial chemical production.
8. Preparing solutions for titration experiments.

13.0 ACID-BASE TITRATIONS

13.1 Meaning of titration

Titration is a quantitative analytical technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration.

In an acid-base titration, an acid reacts with a base to form salt and water.

13.2 Meaning of acid-base titration

An acid-base titration is a laboratory procedure in which an acid of known concentration or a base of known concentration is used to determine the concentration of the other solution.

Examples:

- Hydrochloric acid against sodium hydroxide.
- Sulphuric acid against potassium hydroxide.
- Nitric acid against sodium hydroxide.

13.3 Important terms used in titration

1. Titrant

The titrant is the solution of known concentration placed in the burette.

2. Analyte

The analyte is the solution whose concentration is being determined.

3. Indicator

An indicator is a substance that changes colour depending on the pH of a solution.

Examples:

- Phenolphthalein.
- Methyl orange.
- Litmus.

4. End point

The end point is the stage during a titration when the indicator changes colour.

5. Equivalence point

The equivalence point is the stage at which the reacting substances have combined in exactly the stoichiometric proportions shown by the balanced chemical equation.

6. Titre

A titre is the volume of titrant delivered from the burette during a titration.

7. Concordant titres

Concordant titres are closely agreeing titre values obtained from repeated titrations.

In school laboratories, titres within approximately 0.20 cm^3 of one another are commonly considered concordant, depending on the practical instructions.

13.4 Apparatus used in acid-base titration

1. Burette.
2. Pipette.
3. Pipette filler.
4. Conical flask.
5. Retort stand.
6. Clamp.
7. White tile.
8. Funnel.
9. Wash bottle.
10. Indicator.
11. Beakers.

13.5 Procedure for carrying out an acid-base titration

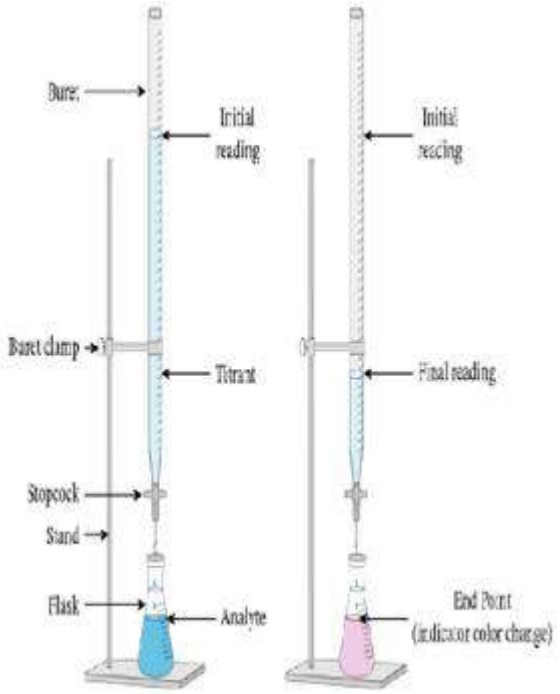
Example: Titration of hydrochloric acid against sodium hydroxide.

1. Rinse the burette with the hydrochloric acid solution.
2. Fill the burette with hydrochloric acid using a funnel.
3. Remove the funnel and ensure there are no air bubbles in the burette jet.
4. Record the initial burette reading at eye level.
5. Rinse a pipette with sodium hydroxide solution.
6. Use the pipette to transfer 25.0 cm^3 of sodium hydroxide into a clean conical flask.
7. Add two or three drops of a suitable indicator, such as methyl orange.
8. Place the conical flask on a white tile beneath the burette.
9. Slowly add hydrochloric acid while swirling the conical flask.
10. Near the end point, add the acid drop by drop.
11. Stop when the indicator shows a permanent colour change.
12. Record the final burette reading.
13. Calculate the titre by subtracting the initial reading from the final reading.
14. Repeat the titration until concordant titres are obtained.

Steps of Titration

1. The burette is filled with the acid. A burette is used to place the acid of unknown concentration.
2. The alkali is added to the conical flask using a pipette.
3. An indicator is added to the alkali. An indicator is added to the alkali in the conical flask, which causes a colour change to occur. When titrating a strong acid and strong alkali any commercial indicator is used, While when titrating a strong acid and a weak base methyl orange is used and finally when titrating a weak acid with a strong base phenolphthalein is used
4. The acid is added into the alkali. The acid from the burette is allowed to run into the alkali, swirling constantly to ensure the reactants mix well.
5. The acid Is added slowly when the indicator changes colour.
6. The acid stops being added when there is a permanent colour change
7. The titre is recorded in the table.

The titration is repeated three times;



Titration number	1	2	3
Final burette reading (cm^3)	20.0	20.0	20.0
Initial burette reading (cm^3)	0.0	0.0	0.0
Volume of solution used(cm^3)	20.0	20.0	20.0

Calculating Concentration of titrant

The known volume and concentration of the alkali enable the moles to be calculated.

The balanced symbol equation enable the moles ratio of acid: alkali to be determined. Hence the number of moles of acid can be calculated.

The mean titre is the volume of the acid. Using the moles and the volume, the concentration of the unknown acid can be determined.

13.6 Common indicators used in acid-base titrations

Indicator	Colour in acid	Colour in alkali
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink

The indicator selected should give a clear colour change close to the equivalence point.

13.7 Acid-base neutralisation

Neutralisation is the reaction between an acid and a base to form salt and water.

General equation:

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

Example:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Ionic equation:

$$H^+ + OH^- \rightarrow H_2O$$

13.8 Calculations involving acid-base titrations

For a reaction:

$$aA + bB \rightarrow \text{Products}$$

The mole relationship is:

$$\frac{n_A}{a} = \frac{n_B}{b}$$

Since:

$$n = CV$$

Then:

$$\boxed{\frac{C_A V_A}{a} = \frac{C_B V_B}{b}}$$

The volumes must be expressed in the same units.

Worked example 23

25 cm^3 of sodium hydroxide solution was completely neutralised by 20 cm^3 of 0.1 mol dm^{-3} hydrochloric acid. Calculate the concentration of the sodium hydroxide solution.

Chemical equation:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

The mole ratio is 1:1.

Step 1: Calculate moles of hydrochloric acid.

$$n = CV$$

Convert volume:

$$20\ cm^3 = 0.020\ dm^3$$

$$n = 0.1 \times 0.020$$

$$n = 0.002\ mol$$

Step 2: Determine moles of sodium hydroxide.

From the equation, the ratio is 1:1.

Therefore:

$$n(NaOH) = 0.002\ mol$$

Step 3: Calculate concentration of sodium hydroxide.

$$C = \frac{n}{V}$$

$$25\ cm^3 = 0.025\ dm^3$$

$$C = \frac{0.002}{0.025}$$

$$\boxed{C = 0.08\ mol\ dm^{-3}}$$

13.9 Applications of acid-base titrations

1. Determining the concentration of acids and alkalis.
2. Measuring acidity in agricultural soils.
3. Determining acidity in food products.
4. Quality control in pharmaceutical industries.
5. Testing the acidity of industrial waste.
6. Analysing water samples.
7. Determining the purity of chemicals.
8. Quality control in food-processing industries.

Aims

- To carry out an accurate titration using dilute hydrochloric acid, dilute sodium hydroxide solution, and phenolphthalein indicator.
- Titration is a practical technique used to determine the amount or concentration of a substance in a sample. It is an example of quantitative analysis. An acid-alkali titration can be used to find out what volume of acid (or alkali) of known concentration exactly neutralises a known volume of alkali (or acid) of unknown concentration. This concentration can then be calculated.
- To obtain valid results, it is important that measurements are precise and accurate. This can be achieved by using a standard procedure for carrying out a titration.

There are a number of ways that you could carry out a titration in Chemistry. This is an outline of the required steps to undertake one of these methods.

Practical activity 1 - Sample Titration table format

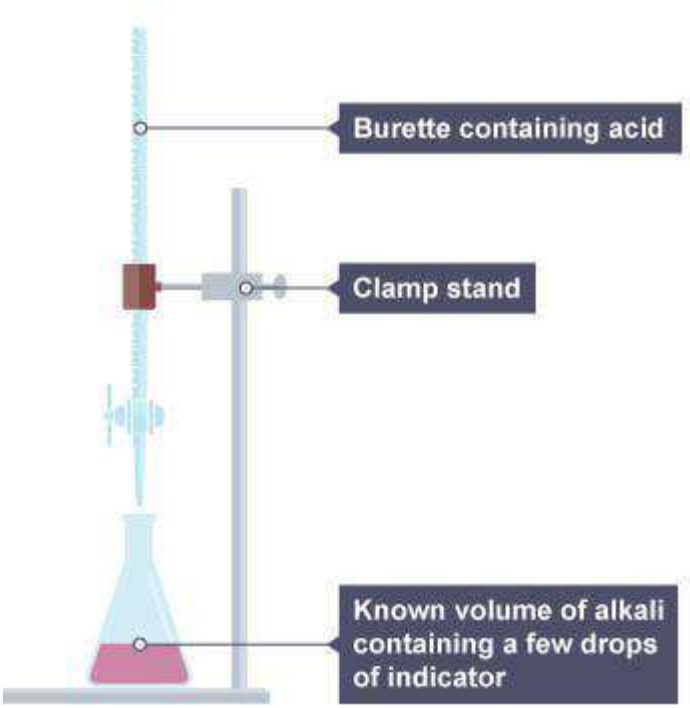
Final burette reading (cm^3)	24.0	24.0	24.0
Initial burette reading (cm^3)	0.0	0.0	0.0
Volume of solution used (cm^3)	24.0	24.0	24.0

Calculate the average volume of solution used

$$\frac{24.0 + 24.0 + 24.0}{3} = 24.0 \text{ cm}^3$$

Practical activity 2 - carrying out a titration

1. Use a pipette and pipette filler to add 25 cm^3 of alkali solution to a clean conical flask.
2. Add a few drops of a suitable indicator and put the conical flask on a white tile.
3. Fill the burette with dilute acid. Flush the tap through to remove any air bubbles. Ensure the burette is vertical.
4. Slowly add the acid from the burette to the conical flask, swirling to mix. (The mixture may at first change colour, and then back again when swirled.)
5. Stop adding the acid when the end-point is reached (when the colour first permanently changes). Note the final volume reading.
6. Repeat steps 1 to 5 until three results are repeatable (in close agreement). Ideally these should lie within 0.10 cm^3 of each other.



The solution of known concentration is placed in the burette and the solution of unknown concentration is added to the flask

Results

Record the results in a suitable table. The one here also shows some sample readings.

Run	Rough	1	2	3	4
End reading (cm^3)	25.45	24.80	47.90	23.70	23.90
Start reading (cm^3)	0.00	1.00	23.80	0.00	0.00
Titre (cm^3)	25.45	23.80	24.10	23.70	23.90

The titre is the volume added (the difference between the end and start readings). The burette is marked in 0.10 cm^3 graduations. It is possible to record results to the nearest 0.05 cm^3 by noting if the meniscus lies between markings.

Analysis

Example

Calculate the mean titre. In the calculation, ignore the rough run and any results that are not in close agreement (24.10 cm^3 in the table above).

Ignoring the rough run, and run 2 (because it is not concordant):

$$\text{mean titre} = \frac{23.80+23.70+23.90}{3}$$

$= 23.80\ cm^3$

Evaluation

Questions

1. Explain why a pipette is used to measure the acid, rather than a measuring cylinder.
2. Describe two steps needed to obtain accurate results.
3. Explain the importance of a suitable indicator in obtaining accurate results.

Answers

1. A pipette is more precise than a measuring cylinder. Adding slightly different volumes of alkali to the flask will result in a systematic error. The pipette allows the same volume of acid to be added each time, helping to make the results repeatable.
2. Take the readings from the bottom of the meniscus. Near to the end-point, rinse the inside of the flask with distilled water and add the acid drop by drop.
3. The indicator must change colour sharply when the solution in the flask is neutralised. This means the volume of acid measured is very close to the true value. The white tile makes it easier to see the colour change. Misjudging the colour change could result in a random error.

Hazards, risks and precautions

It is important in this practical activity to use appropriate apparatus and methods. This includes the safe use and careful handling of substances.

Evaluate the hazards and the precautions needed to reduce the risk of harm. For example:

Hazard	Possible harm	Possible precaution
Dilute sodium hydroxide solution	Causes skin and serious eye irritation	Wear gloves and eye protection, and use a pipette filler
Spilling hydrochloric acid while filling the burette	Causes eye irritation	Fill the burette slowly below eye level, using a

Note:

- As evidence of understanding the degree of accuracy of burettes, all readings must be recorded to a decimal point.
- As evidence of accuracy in carrying the out the titration, candidates value should be within 0.2 of the school value.
- The school value is the teachers readings presented to the examining body/council based on the concentrations of the solutions s/he presented to her/his candidates.
- Bonus mark is awarded for averaged reading within 0.1 school value as Final accuracy.
- Calculations involved after the titration require candidates thorough practice mastery on the:
  - ✓ relationship among the mole, molar mass, mole ratios, concentration, molarity.
  - ✓ mathematical application of 1^{st} principles.
- Very useful information which candidates forget appear usually in the beginning of the paper as: "You are provided with…"
- All calculation must be to the 4th decimal point unless they divide fully to a lesser decimal point.
- Never round off answers.

QUESTIONS

(1). Calculate the molarity of MCO_3 if 5.0cm^3 of MCO_3 react with 25.0cm^3 of 0.5M hydrochloric acid.(C = 12.0 ,O = 16.0)

(2). 2.0cm^3 of 0.5M hydrochloric acid react with 0.1M of M_2CO_3. Calculate the volume of 0.1M M_2CO_3 used.

(3). 5.0cm^3 of 0.1M sodium iodide react with 0.1M of Lead(II)nitrate(V). Calculate

a. The volume of Lead(II)nitrate (V) used.

b. The mass of Lead (II) Iodide formed.

Glossary of terms

- Titration - A quantitative procedure in which two solutions react in a known ratio, so if the concentration of one solution is known and the volumes of both are measured, the concentration of the other solution can be determined.
- Indicator - A substance that has different colours, depending upon the pH of the solution it is in.
- Quantitative - Results that can be expressed using numerical values.
- Acid - Corrosive substance which has a pH lower than 7. Acidity is caused by a high concentration of hydrogen ions.
- Alkali - A base which is soluble in water.
- Neutralises - To be made neutral by removing any acidic or alkaline nature.
- Valid - A measurement, test or observation that can be repeated or reproduced to get similar data to support a conclusion.
- precise - A measurement of the spread of numerical results. Data is precise if values are clustered closely together.
- Accurate - When the value of a measurement is close to the true value.
- Pipette - A piece of apparatus used to measure accurate and repeatable volumes of liquid. Also called a volumetric pipette.
- End-point - In a titration, the point at which the indicator first permanently changes colour.
- Repeatable - When a measurement gives a very similar result if it is taken many times over in the same experiment.
- Titre - Volume of one reactant needed to react completely with the other reactant in a titration.
- Meniscus - The curve in the upper surface of a liquid contained within a glass container. Where the container is narrow, like a burette, the meniscus is most noticeable.

14.0 REDOX TITRATIONS

14.1 Meaning of oxidation and reduction

Redox reactions are chemical reactions in which oxidation and reduction occur simultaneously.

Oxidation is defined as:

- Loss of electrons.
- Gain of oxygen.
- Loss of hydrogen.
- Increase in oxidation number.

Reduction is defined as:

- Gain of electrons.
- Loss of oxygen.
- Gain of hydrogen.
- Decrease in oxidation number.

14.2 Meaning of a redox titration

A redox titration is a quantitative analytical technique in which a solution of known concentration is used to determine the concentration of another solution through an oxidation-reduction reaction.

These titrations involve the transfer of electrons between reacting substances.

14.3 Common substances used in redox titrations

1. Potassium manganate(VII), $KMnO_4$.
2. Potassium dichromate(VI), $K_2Cr_2O_7$.
3. Iodine, I_2.
4. Sodium thiosulphate, $Na_2S_2O_3$.

14.4 Potassium manganate(VII) in redox titrations

Potassium manganate(VII) is a strong oxidising agent.

It is commonly used to determine the concentration of reducing agents such as iron(II) ions and oxalate ions.

It is purple in colour.

In acidic solution, manganate(VII) ions are reduced to nearly colourless manganese(II) ions.

The half-equation is:

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

14.5 Self-indicating nature of potassium manganate(VII)

Potassium manganate(VII) acts as its own indicator because it has a strong purple colour.

During titration:

- The purple colour disappears as manganate(VII) ions are reduced.
- At the end point, the first permanent faint pink colour indicates a slight excess of manganate(VII) ions.

Therefore, an additional indicator is usually not required.

14.6 Acid used in manganate(VII) titrations

Dilute sulphuric acid is commonly used to provide the acidic medium required for the reaction.

Hydrochloric acid is generally unsuitable because chloride ions may be oxidised under the titration conditions.

Nitric acid is unsuitable because it is an oxidising acid and may interfere with the reaction.

14.7 Example of a redox reaction

Reaction between acidified potassium manganate(VII) and iron(II) ions.

The ionic equation is:

$$MnO_4^- + 8H^+ + 5Fe^{2+}$$

$$\rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

From the equation:

$$1\ mol\ MnO_4^- : 5\ mol\ Fe^{2+}$$

14.8 Procedure for a redox titration

1. Rinse the burette with the potassium manganate(VII) solution.
2. Fill the burette with potassium manganate(VII) solution.
3. Record the initial reading.
4. Use a pipette to transfer a known volume of iron(II) solution into a conical flask.
5. Add a suitable measured amount of dilute sulphuric acid.
6. Place the flask beneath the burette.
7. Add potassium manganate(VII) slowly while swirling the flask.
8. Near the end point, add the solution drop by drop.
9. Stop when a faint pink colour persists for approximately 30 seconds.
10. Record the final burette reading.
11. Repeat to obtain concordant titres.
12. Use the balanced equation to calculate the concentration of the iron(II) solution.

14.9 Calculations involving redox titrations

Worked example 24

25 cm^3 of iron(II) sulphate solution was titrated with 20 cm^3 of 0.02 mol dm^{-3} potassium manganate(VII) solution in acidic conditions. Calculate the concentration of the iron(II) solution.

Ionic equation:

$$MnO_4^- + 8H^+ + 5Fe^{2+}$$

$$\rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

Step 1: Calculate moles of manganate(VII).

Convert volume:

$$20\ cm^3 = 0.020\ dm^3$$

$$n = CV$$

$$n = 0.02 \times 0.020$$

$$n = 0.0004\ mol$$

Step 2: Calculate moles of iron(II).

From the equation:

$$1\ mol\ MnO_4^- : 5\ mol\ Fe^{2+}$$

Therefore:

$$n(Fe^{2+}) = 5 \times 0.0004$$

$$= 0.002\ mol$$

Step 3: Calculate concentration of iron(II).

$$C = \frac{n}{V}$$

$$25\ cm^3 = 0.025\ dm^3$$

$$C = \frac{0.002}{0.025}$$

$$\boxed{C = 0.08\ mol\ dm^{-3}}$$

15.0 APPLICATIONS OF REDOX TITRATIONS IN INDUSTRY

Redox titrations have many applications in industrial production, quality control and chemical analysis.

15.1 Food-processing industries

Redox titrations are used to:

1. Determine vitamin C content in fruit juices.
2. Determine the concentration of reducing substances in food.
3. Assess the quality of food products.
4. Monitor changes in food composition during processing.
5. Determine the concentration of oxidising or reducing agents used in food preservation.

15.2 Pharmaceutical industries

They are used to:

1. Determine the concentration of active ingredients in medicines.
2. Assess the purity of pharmaceutical products.
3. Monitor the quality of chemical reagents.
4. Check the concentration of substances used in medicine production.
5. Ensure pharmaceutical products meet quality standards.

15.3 Water treatment industries

Redox titrations are used to:

1. Determine residual chlorine in treated water.
2. Monitor the concentration of oxidising agents.
3. Determine the amount of reducing substances in water.
4. Assess the quality of water used in industrial processes.
5. Monitor chemical treatment processes.

15.4 Mining and metallurgical industries

They are used to:

1. Determine the concentration of iron in mineral samples.
2. Analyse metallic ores.
3. Determine the purity of certain metal products.
4. Monitor oxidation-reduction processes during metal extraction.
5. Analyse industrial solutions used in mineral processing.

15.5 Agricultural industries

Redox titrations may be used to:

1. Analyse certain soil chemical properties.
2. Determine the concentration of iron in agricultural samples.
3. Analyse oxidising and reducing substances in fertiliser production.
4. Monitor chemical processes used in agrochemical industries.

15.6 Environmental monitoring

Redox titrations are used to:

1. Analyse water pollution.
2. Determine certain oxidising and reducing substances in wastewater.

3. Monitor industrial effluents.
4. Support environmental quality assessment.
5. Evaluate the effectiveness of wastewater treatment.

16.0 MOLE RATIOS IN CHEMICAL REACTIONS

16.1 Meaning of mole ratio

A mole ratio is the ratio of the amounts in moles of substances that react or are produced in a chemical reaction.

It is obtained from the coefficients in a balanced chemical equation.

16.2 Importance of mole ratios

Mole ratios are used to:

1. Determine the amount of reactants required in a reaction.
2. Calculate the amount of products formed.
3. Identify the limiting reactant.
4. Reduce wastage of chemicals.
5. Determine quantities of materials used in industries.
6. Predict the amounts of gases produced during chemical reactions.

16.3 Determining mole ratios from chemical equations

Consider:

$$2H_2 + O_2 \rightarrow 2H_2O$$

The mole ratio is:

$$H_2 : O_2 : H_2O = 2 : 1 : 2$$

This means:

- 2 moles of hydrogen react with 1 mole of oxygen.
- 2 moles of water are formed.

Other examples:

Formation of ammonia

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Mole ratio:

$$1 : 3 : 2$$

Formation of magnesium oxide

$$2Mg + O_2 \rightarrow 2MgO$$

Mole ratio:

$$2 : 1 : 2$$

Reaction between hydrochloric acid and sodium hydroxide

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Mole ratio:

$$1:1:1:1$$

16.4 Calculations involving mole ratios

Worked example 25

Calculate the number of moles of oxygen required to react completely with 4 moles of hydrogen.

Equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

From the equation:

2 moles of hydrogen react with 1 mole of oxygen.

Therefore:

$$4\ mol\ H_2 \rightarrow \frac{4}{2} = 2\ mol\ O_2$$

$$\boxed{2\ mol\ oxygen}$$

Worked example 26

Calculate the number of moles of ammonia produced when 6 moles of hydrogen react completely with excess nitrogen.

Equation:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Mole ratio:

$$3\ mol\ H_2 : 2\ mol\ NH_3$$

Therefore:

$$n(NH_3) = \frac{2}{3} \times 6$$

$$\boxed{4\ mol}$$

17.0 MOLAR GAS VOLUME

17.1 Meaning of molar gas volume

Molar gas volume is the volume occupied by one mole of a gas at a specified temperature and pressure.

The volume occupied by a mole of gas depends on the temperature and pressure.

At room temperature and pressure (RTP), a commonly used approximate molar gas volume is:

$$\boxed{24\ dm^3\ mol^{-1}}$$

At standard temperature and pressure (STP), defined as 273.15 K and 100 kPa, the ideal-gas molar volume is approximately:

$$\boxed{22.7\ dm^3\ mol^{-1}}$$

Always use the conditions specified in the question.

17.2 Relationship between gas volume and number of moles

At RTP:

$$\boxed{n = \frac{V}{24}}$$

Where:

- n = Number of moles.
- V = Volume of gas in dm^3.
- 24 = Approximate molar gas volume at RTP.

To calculate gas volume:

$$\boxed{V = n \times 24}$$

17.3 Gas volume in cubic centimetres

Since:

$$1\ dm^3 = 1000\ cm^3$$

The molar gas volume at RTP is approximately:

$$24{,}000\ cm^3\ mol^{-1}$$

Worked example 27

Calculate the number of moles in 48 dm^3 of oxygen gas at RTP.

$$n = \frac{V}{24}$$

$$n = \frac{48}{24}$$

$$\boxed{n = 2\ mol}$$

Worked example 28

Calculate the volume occupied by 0.5 moles of carbon dioxide at RTP.

$$V = n \times 24$$

$$V = 0.5 \times 24$$

$$\boxed{V = 12\ dm^3}$$

Worked example 29

Calculate the volume occupied by 3 moles of hydrogen gas at RTP.

$$V = 3 \times 24$$

$$\boxed{V = 72\ dm^3}$$

17.4 Calculations involving gas volumes and chemical equations

Consider:

$$2H_2 + O_2 \rightarrow 2H_2O$$

At the same temperature and pressure, gas volumes react in the ratio of their mole coefficients.

Therefore:

$$H_2: O_2 = 2:1$$

This means that 2 volumes of hydrogen react with 1 volume of oxygen, provided both gases are measured at the same temperature and pressure.

Worked example 30

Calculate the volume of oxygen required to react completely with 60 dm^3 of hydrogen at the same temperature and pressure.

Equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Volume ratio:

$$H_2: O_2 = 2:1$$

Therefore:

$$V(O_2) = \frac{60}{2}$$

$$\boxed{30\ dm^3}$$

18.0 LABORATORY SAFETY AND WASTE MANAGEMENT

Chemical experiments involving mole calculations, preparation of solutions and titrations require careful handling of chemicals and apparatus.

18.1 Laboratory safety precautions

1. Wear a laboratory coat, safety goggles and appropriate protective gloves.
2. Never taste chemicals.
3. Avoid directly inhaling chemical fumes.
4. Use a pipette filler instead of mouth pipetting.
5. Label all prepared solutions correctly.
6. Handle acids and alkalis carefully.
7. Always add acid to water when diluting concentrated acids.
8. Use clean apparatus to prevent contamination.
9. Handle hot apparatus using appropriate tongs.
10. Never return unused chemicals to their original containers.
11. Report chemical spills and accidents immediately.
12. Wash hands after handling chemicals.

18.2 Responsible disposal of chemical waste

Chemical waste should be disposed of according to laboratory procedures and environmental safety requirements.

Important practices include:

1. Separate chemical waste according to its nature.
2. Avoid pouring concentrated acids and alkalis into sinks.
3. Collect heavy-metal-containing waste in designated containers.
4. Do not dispose of potassium manganate(VII) waste carelessly.
5. Follow the teacher's instructions on neutralisation and disposal.

6. Avoid mixing incompatible chemicals.
7. Use designated waste containers.
8. Keep the laboratory clean.
9. Prevent chemical contamination of soil and water sources.
10. Follow institutional and environmental waste-management regulations.

SUMMARY OF IMPORTANT FORMULAE

The following formulae should be mastered when studying the mole.

Concept	Formula
Number of moles	$n = \frac{m}{M}$
Mass	$m = nM$
Number of particles	$N = nN_A$
Number of moles from particles	$n = \frac{N}{N_A}$
Concentration	$C = \frac{n}{V}$
Number of moles in solution	$n = CV$
Mass concentration	$C = \frac{m}{V}$
Moles from mass concentration	$n = \frac{CV}{M}$
Dilution	$C_1V_1 = C_2V_2$
Molecular formula multiplier	$k = \frac{M_r}{\text{Empirical formula mass}}$
Gas volume at RTP	$V = 24n$
Moles of gas at RTP	$n = \frac{V}{24}$

Important constants and conversions

- Avogadro's constant = $6.02 \times 10^{23}\ mol^{-1}$.
- Molar gas volume at RTP ≈ 24 $dm^3\ mol^{-1}$.
- 1 dm^3 = 1000 cm^3.
- 1 mole = 6.02×10^{23} specified particles.

REVISION QUESTIONS

STRAND 1.0: INORGANIC CHEMISTRY

SUB-STRAND 1.1: THE MOLE

SECTION A: MULTIPLE-CHOICE QUESTIONS

Choose the correct answer from the alternatives provided.

1. Which of the following is the SI unit for amount of substance?

A. Gram

B. Mole

C. Kilogram

D. Litre

2. One mole of a substance contains approximately:

A. 6.02×10^{20} particles.

B. 6.02×10^{21} particles.

C. 6.02×10^{23} particles.

D. 6.02×10^{25} particles.

3. The number of particles contained in one mole of a substance is known as:

A. Relative atomic mass.

B. Avogadro's constant.

C. Molecular mass.

D. Molar volume.

4. What is the relative molecular mass of water? (H = 1, O = 16)

A. 16 B. 17 C. 18 D. 20

5. Calculate the number of moles in 24 g of magnesium. (Mg = 24)

A. 0.5 mol

B. 1 mol

C. 2 mol

D. 4 mol

6. What is the mass of 2 moles of carbon dioxide? (C = 12, O = 16)

A. 22 g

B. 44 g

C. 66 g

D. 88 g

7. Which of the following represents the empirical formula of hydrogen peroxide (H_2O_2)?

A. H_2O_2

B. HO

C. H_2O

D. H_3O

8. The empirical formula of glucose, $C_6H_{12}O_6$, is:

A. CHO

B. CH_2O

C. $C_2H_4O_2$

D. $C_6H_{12}O_6$

9. Which apparatus is used to measure an accurate fixed volume of solution during titration?

A. Beaker

B. Conical flask

C. Pipette

D. Test tube

10. A standard solution is one whose:

A. Colour is known.

B. Concentration is accurately known.

C. Temperature is constant.

D. Volume is unknown.

11. Which of the following is used to prepare a solution to an accurate fixed volume?

A. Measuring cylinder B. Volumetric flask

C. Test tube D. Evaporating dish

12. What is the concentration of a solution containing 2 moles in 1 dm^3?

A. 0.5 mol dm^{-3}

B. 1 mol dm^{-3}

C. 2 mol dm^{-3}

D. 3 mol dm^{-3}

13. Which of the following is the correct relationship used in dilution calculations?

A. $C_1 + C_2 = V_1 + V_2$

B. $C_1V_1 = C_2V_2$

C. $C_1V_2 = C_2V_1$

D.

14. Which indicator is pink in alkaline solution and colourless in acidic solution?

A. Methyl orange

B. Litmus

C. Phenolphthalein

D. Universal indicator

15. The reaction between an acid and a base produces:

A. Salt and hydrogen.

B. Salt and water.

C. Oxygen and water.

D. Carbon dioxide and water.

16. Which substance is commonly used as a self-indicator in redox titrations?

A. Sodium chloride

B. Potassium manganate(VII)

C. Sodium hydroxide

D. Calcium carbonate

17. During a titration involving acidified potassium manganate(VII), the end point is indicated by:

A. A permanent faint pink colour.

B. A permanent blue colour.

C. A permanent green colour.

D. A permanent yellow colour.

18. Which acid is commonly used to provide an acidic medium in potassium manganate(VII) titrations?

A. Hydrochloric acid

B. Nitric acid

C. Sulphuric acid

D. Ethanoic acid

19. What is the mole ratio of nitrogen to hydrogen in the formation of ammonia?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

A. 1:1 B. 1:2

C. 1:3 D. 3:2

20. What is the approximate molar gas volume at room temperature and pressure?

A. 12 $dm^3\ mol^{-1}$

B. 18 $dm^3\ mol^{-1}$

C. 22 $dm^3\ mol^{-1}$

D. 24 $dm^3\ mol^{-1}$

21. Which of the following processes involves loss of electrons?

A. Reduction

B. Oxidation

C. Neutralisation

D. Precipitation

22. Which of the following is the correct ionic equation for neutralisation?

A. $H^+ + OH^- \rightarrow H_2O$

B. $H^+ + e^- \rightarrow H_2$

C. $Na^+ + Cl^- \rightarrow NaCl$

D. $O_2 + 2H_2 \rightarrow 2H_2O$

23. The simplest whole-number ratio of atoms in a compound is represented by its:

A. Molecular formula.

B. Structural formula.

C. Empirical formula.

D. Ionic equation.

24. Which of the following is the correct formula for calculating moles from mass?

A. $n = m \times M$

B. $n = \frac{m}{M}$

C. $n = \frac{M}{m}$

D. $n = m + M$

25. What is the mass of 0.5 moles of sodium chloride? (Na = 23, Cl = 35.5)

A. 14.625 g B. 29.25 g

C. 58.5 g D. 117 g

26. Which of the following substances is commonly used as a primary standard in acid-base titrations?

A. Sodium hydroxide

B. Hydrochloric acid

C. Anhydrous sodium carbonate

D. Ammonia solution

27. What is the main purpose of rinsing a burette with the solution to be used in it?

A. To increase the volume of solution.

B. To prevent dilution by residual water.

C. To change the colour of the solution.

D. To reduce the concentration.

28. Which of the following is a suitable method of disposing of heavy-metal-containing chemical waste?

A. Pouring it into a river.

B. Pouring it into the sink.

C. Collecting it in a designated waste container.

D. Pouring it onto the ground.

29. What volume of oxygen is required to react completely with 40 dm^3 of hydrogen at the same temperature and pressure?

$$2H_2 + O_2 \rightarrow 2H_2O$$

A. 10 dm^3

B. 20 dm^3

C. 40 dm^3

D. 80 dm^3

30. Which of the following best describes a redox reaction?

A. A reaction involving only heat transfer.

B. A reaction involving electron transfer.

C. A reaction involving only precipitation.

D. A reaction involving only neutralisation.

SECTION B: STRUCTURED QUESTIONS

31. The Mole and Avogadro's Constant

a) Define the term mole. (2 marks)

b) State the value of Avogadro's constant. (1 mark)

c) Explain the relationship between the mole and the number of particles. (2 marks)

d) Calculate the number of molecules in 0.25 moles of oxygen gas. (3 marks)

e) Determine the number of atoms present in 2 moles of helium gas. (2 marks)

32. Relative Atomic Mass and Molar Mass

a) Define relative atomic mass. (2 marks)

b) Distinguish between relative molecular mass and molar mass. (2 marks)

c) Calculate the relative molecular mass of:

d) Sulphuric acid, H_2SO_4. (2 marks)

ii) Calcium hydroxide, $Ca(OH)_2$. (2 marks)

iii) Calculate the number of moles in 20 g of calcium hydroxide. (2 marks)

33. Conversion of Mass and Moles

a) State the formula used to convert:

b) Mass into moles. (1 mark)

ii) Moles into mass. (1 mark)

b) Calculate the number of moles in 16 g of oxygen gas, O_2. (2 marks)

c) Calculate the mass of 0.4 moles of sodium carbonate, Na_2CO_3. (3 marks)

d) Calculate the number of molecules in 0.2 moles of carbon dioxide. (3 marks)

34. Empirical Formula

A compound contains 52.2% carbon, 13.0% hydrogen and 34.8% oxygen by mass.

a) Define empirical formula. (2 marks)

b) Explain the steps followed when determining an empirical formula from percentage composition. (3 marks)

c) Determine the empirical formula of the compound. (5 marks)

35. Molecular Formula

A compound has an empirical formula of CH_2O and a relative molecular mass of 180.

a) Define molecular formula. (2 marks)

b) Calculate the empirical formula mass. (2 marks)

c) Determine the molecular formula of the compound. (3 marks)

d) State two differences between empirical and molecular formulae. (2 marks)

36. Preparation of Standard Solutions

a) Define a standard solution. (2 marks)

b) Name four pieces of apparatus used in preparing a standard solution from a solid substance. (4 marks)

c) Calculate the mass of anhydrous sodium carbonate required to prepare 500 cm^3 of 0.2 mol dm^{-3} solution. (4 marks)

d) Describe the procedure followed when preparing a standard solution from a solid substance. (5 marks)

37. Concentration and Dilution

a) Define concentration of a solution. (2 marks)

b) Convert the following volumes into dm^3:

c) 250 cm^3. (1 mark)

ii) 750 cm^3. (1 mark)

iii) Calculate the concentration of a solution containing 0.5 moles of sodium hydroxide in 250 cm^3 of solution. (3 marks)

iv) 50 cm^3 of 2 mol dm^{-3} hydrochloric acid is diluted to 250 cm^3. Calculate the final concentration. (3 marks)

38. Acid-Base Titrations

a) Define acid-base titration. (2 marks)

b) State four pieces of apparatus used in carrying out an acid-base titration. (4 marks)

c) State three precautions observed during titration. (3 marks)

d) Explain the meaning of the following terms:

e) End point. (2 marks)

ii) Titre. (2 marks)

e) Write a balanced chemical equation for the reaction between hydrochloric acid and sodium hydroxide. (2 marks)

39. Calculations Involving Acid-Base Titrations

25 cm^3 of sodium hydroxide solution was neutralised by 20 cm^3 of 0.1 mol dm^{-3} hydrochloric acid.

a) Write the balanced chemical equation for the reaction. (2 marks)

b) Calculate the number of moles of hydrochloric acid used. (3 marks)

c) Determine the number of moles of sodium hydroxide that reacted. (2 marks)

d) Calculate the concentration of sodium hydroxide in mol dm^{-3}. (3 marks)

40. Redox Titrations

a) Define oxidation. (2 marks)

b) Define reduction. (2 marks)

c) Explain the meaning of a redox titration. (2 marks)

d) State three substances commonly used in redox titrations. (3 marks)

e) Explain why potassium manganate(VII) is described as a self-indicator. (3 marks)

f) State two safety precautions observed during redox titrations. (2 marks)

41. Applications of Redox Titrations

a) Explain four applications of redox titrations in industry. (4 marks)

b) State three applications of redox titrations in environmental monitoring. (3 marks)

c) Explain why redox titrations are important in quality control. (3 marks)

42. Mole Ratios in Chemical Reactions

Consider the reaction:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

a) State the mole ratio of nitrogen to hydrogen. (1 mark)

b) State the mole ratio of hydrogen to ammonia. (1 mark)

c) Calculate the number of moles of ammonia produced when 9 moles of hydrogen react completely with excess nitrogen. (3 marks)

d) Calculate the number of moles of nitrogen required to react completely with 12 moles of hydrogen. (3 marks)

43. Molar Gas Volume

a) Define molar gas volume. (2 marks)

b) State the approximate molar gas volume at RTP. (1 mark)

c) Calculate the number of moles in 48 dm^3 of hydrogen gas at RTP. (2 marks)

d) Calculate the volume occupied by 0.75 moles of oxygen gas at RTP. (2 marks)

e) Calculate the volume of oxygen required to react completely with 80 dm^3 of hydrogen at the same temperature and pressure.

$$2H_2 + O_2 \rightarrow 2H_2O$$

(3 marks)

44. Laboratory Safety and Waste Management

a) State five safety precautions observed when preparing standard solutions. (5 marks)

b) Explain four ways of disposing of chemical waste responsibly. (4 marks)

c) Explain why concentrated acids should be added to water rather than water being added to concentrated acids. (2 marks)

d) State two values developed through teamwork during laboratory activities. (2 marks)

21.0 PROJECT ACTIVITIES

The following practical activities can be used to develop practical skills, critical thinking, communication and collaboration.

Practical activity	Expected learning outcome
Preparation of sodium carbonate standard solution	Ability to prepare accurately known solutions
Determination of empirical formula of magnesium oxide	Ability to determine the simplest ratio of elements
Acid-base titration	Ability to determine unknown concentrations
Redox titration using potassium manganate(VII)	Ability to apply oxidation-reduction reactions
Calculations involving gas volumes	Ability to relate gas volume to amount of substance
Chemical waste segregation	Ability to practise responsible waste management

REVISION SUMMARY

THE MOLE

- A mole is defined as the mass of a substance that contains as many as particles(atoms, molecules, or ions) as the number of atoms in 12g of carbon-12.
- In the above definition, 12 is the mass number of carbon. So, one mole of carbon atoms has a mass of 12 grams. The relative atomic mass, which can be written as RAM, is the number just above the element in the periodic table
- This number is also called Avogadro's number. And is written as **6×10^{23}.**

1mole = 6.023×10^{23} particles

The mass of one mole of a substance is called **molar mass.**

The molar mass of:

(i) An **element** has mass equal to relative **atomic** mass /RAM(in grams) of the element e.g.

- Molar mass of carbon(C)= relative atomic mass = 12.0g
- **6.023×10^{23} particles of carbon = 1 mole = 12.0 g**

- Molar mass of sodium(Na) = relative atomic mass = 23.0g
- **6.023×10^{23} particles of sodium = 1 mole = 23.0 g**

- Molar mass of Iron (Fe) = relative atomic mass = 56.0g
- **6.023×10^{23} particles of iron = 1 mole = 56.0 g**

(ii)

- A **molecule** has mass equal to relative molecular mass/RMM (in grams)of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.
- The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic (e.g. O_2, H_2, N_2, F_2, Cl_2, Br_2, I_2)** noble gases are **monoatomic (e.g. He, Ar, Ne, Xe)**,Ozone gas(**O_3**) is **triatomic** e.g.

- Molar mass **Oxygen molecule(O_2)** = relative molecular mass = (16.0 x 2)g = 32.0
- **6.023×10^{23} particles of Oxygen molecule = 1 mole = 32.0 g**

- Molar mass **chlorine molecule(Cl_2)** =relative molecular mass = (35.5x 2)g =71.0g
- **6.023×10^{23} particles of chlorine molecule = 1 mole = 71.0 g**

- Molar mass **Nitrogen molecule(N_2)** = relative molecular mass =(14.0x 2)g =28.0g
- **6.023×10^{23} particles of Nitrogen molecule = 1 mole = 28.0 g**

- A **compound** has mass equal to relative formular mass /RFM (in grams)of the molecule.

- Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.

(i)

Molar mass **Water(H_2O)** = relative formular mass = [(1.0 x 2) + 16.0]g = 18.0g

6.023 x 10^{23} particles of Water molecule = 1 mole = 18.0 g

6.023 x10^{23} particles of Water molecule has:

- **-2** x 6.023 x10^{23} particles of Hydrogen atoms
- **-1** x 6.023 x10^{23} particles of Oxygen atoms

(ii)

Molar mass **sulphuric (VI) acid(H_2SO_4)** = relative formular mass

= [(1.0 x 2) + 32.0 + (16.0 x 4)]g = 98.0g

6.023 x10^{23} particles of sulphuric(VI)acid(H_2SO_4) = 1 mole = 98.0g

6.023 x10^{23} particles of sulphuric(VI)acid(H_2SO_4) has:

- **-2** x 6.023 x10^{23} particles of **H**ydrogen atoms
- **-1** x 6.023 x10^{23} particles of **S**ulphur atoms
- **-4** x 6.023 x10^{23} particles of **O**xygen atoms

QUESTIONS.

(Avogadro's constant L = 6.0 x 10^{23})

(1). Calculate the number of atoms present in 0.23 g of dilute sulphuric (VI) acid

(2). Calculate the number of atoms present in 0.23 g of sodium carbonate (IV) decahydrate

(3). How many molecules of Argon gas do we have when we say 3 moles of Argon gas?

(4). How many atoms of sodium do we have when we say we have 6 moles of sodium metal?

(5). Find the number of moles in 3.02 x 10^{24} atoms of Na.

(6). Find the number of moles make 1.510 x 10^{23} molecules of Ammonia.

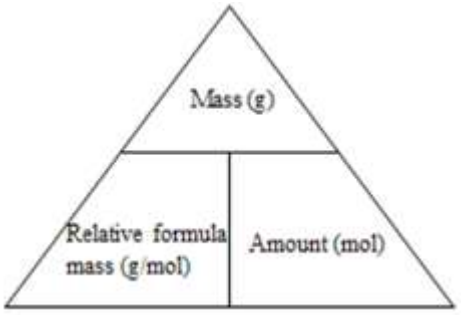
Interconversion of Moles, Mass & Molar Mass

- When the mass of an atom is compared to that of another, it is referred to as **relative atomic mass-RAM.**

- The relative atomic mass of an atom is defined as the ratio of the average mass of one atom of an element to one twelfth of the mass of an atom of carbon-12.
- The molar mass (Mw), also known as molecular weight (Mw), is the sum of the total mass in grams of all the atoms that make up a mole of a particular molecule. The unit used to measure is grams per mole (g/mol).

Molar mass answers the question, how many grams are in one mole of a given element?

The mole formula is: mole = mass (g) / molar mass (g/mol)



Questions;

(1). Calculate the number of moles in;

a. 0.23 g of Sodium atoms

b. 0.23 g of Chlorine atoms

c. 0.23 g of Chlorine molecules

d. 0.23 g of dilute sulphuric (VI)acid

(2). Calculate the mass of 0.25 moles of berylium.

(3). Calculate the molar mass of an element if ala mass of 5.6g contains 0.1 moles of the element.

Empirical and molecular formula

1. The empirical formula of a compound is its simplest formula. It is the simplest whole number ratios in which atoms of elements combine to form the compound.
2. It is mathematically the lowest common multiple (LCM) of the atoms of the elements in the compound.
3. Practically the empirical formula of a compound can be determined as in the following examples.

To determine the empirical formula of copper oxide

(a). Method 1: From copper to copper(II)oxide

Procedure.

Weigh a clean dry covered crucible **(M_1)**. Put two spatula full of copper **powder** into the crucible. Weigh again (M_2). Heat the crucible on a strong Bunsen flame for five minutes. Lift the lid, and swirl the crucible carefully using a pair of tong. Cover the crucible and continue heating for another five minutes. Remove the lid and stop heating. Allow the crucible to cool. When cool replace the lid and weigh the contents again **(M_3)**.

Sample results

Mass of crucible(M_1)	15.6g
Mass of crucible + copper before heating(M_2)	18.4
Mass of crucible + copper after heating(M_3)	19.1

Sample questions

1. Calculate the mass of copper powder used.

Mass of crucible + copper before heating (**M_2**) = 18.4

Less Mass of crucible (**M_1**) = 15.6g

18.4g - 15.6g = 2.8g

Mass of copper = 2.8 g

2. Calculate the mass of Oxygen used to react with copper.

Method I

Mass of crucible + copper after heating(**M_3**) = 19.1g

Mass of crucible + copper before heating(**M_2**) = 18.4g

19.1g - 18.4g = 0.7g

Mass of Oxygen = 0.7 g

Method II

Mass of crucible + copper after heating(**M_3**) = 19.1g

Mass of crucible = -15.6g

Mass of copper (II) Oxide = 3.5 g

Mass of copper (II) Oxide = 3.5 g

Mass of copper = - 2.8 g

Mass of Oxygen = **0.7g**

3.Calculate the number of moles of:

(i)

Copper used (Cu = 63.5)

number of moles of copper;

= mass used/Molar mass =

= 2.8/63.5 = **0.0441moles**

(ii)

Oxygen used (O = 16.0)

number of moles of oxygen = mass used/Molar mas

= 0.7/16.0 = **0.0441moles**

4. Determine the mole ratio of the reactants

Moles of copper = 0.0441moles

Moles of oxygen = 0.0441moles

= 0.0441moles/0.0441moles

Mole ratio Cu: O = **1:1**

5. What is the empirical, formula of copper oxide formed.

- CuO (copper (II) oxide)

6. State and explain the observations made during the experiment.

Observation

- Colour change from **brown** to **black**

Explanation

- Copper powder is brown. On heating it reacts with oxygen from the air to form black copper (II) oxide

7. Explain why magnesium ribbon/shavings would be unsuitable in a similar experiment as the one above.

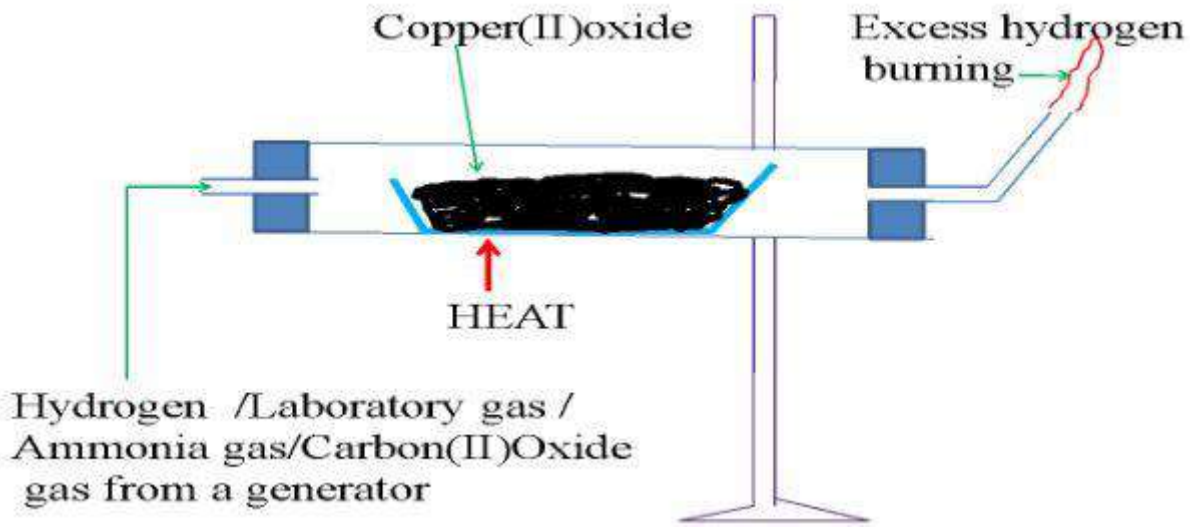
- Hot magnesium generates enough heat energy to react with **both** Oxygen and Nitrogen in the air forming a white solid mixture of Magnesium oxide and magnesium nitride. This causes experimental mass errors.

(b). <u>Method 2: From copper(II)oxide to copper</u>

Procedure.

- Weigh a clean dry porcelain boat (**M_1**). Put two spatula full of copper (II) oxide **powder** into the crucible. Reweigh the porcelain boat (**M_2**).Put the porcelain boat in a glass tube and set up the apparatus as below;

Determining the empirical formular from copper (ii) oxide to copper;



- Pass slowly(to prevent copper(II)oxide from being blown away)a stream of either dry Hydrogen /ammonia/laboratory gas/ carbon(II)oxide gas for about two minutes from a suitable generator.
- When all the in the apparatus set up is driven out ,heat the copper (II) oxide strongly for about five minutes until there is no further change. Stop heating.
- Continue passing the gases until the glass tube is cool.
- Turn off the gas generator.
- Carefully remove the porcelain boat form the combustion tube.
- Reweigh (M_3).

Sample results

Mass of boat(M_1)	15.6g
Mass of boat before heating(M_2)	19.1
Mass of boat after heating(M_3)	18.4

Sample questions

1. Calculate the mass of copper(II)oxide used.

Mass of boat before heating (M_2) = 19.1

Mass of empty boat (M_1) = 15.6g

M_2 - M_1 = Mass of copper (II) Oxide

19.1g - 15.6G = 3.5g

Mass of copper(II)Oxide = 3.5 g

2. Calculate the mass of

(i) Oxygen.

Mass of boat before heating(M_2) = 19.1

Mass of boat after heating (M_3) = 18.4g

M_2 - M_3 = Mass of oxygen

19.1 - 18.4 = 0.7g

Mass of oxygen = 0.7 g

(ii)Copper

Mass of copper (II) Oxide = 3.5 g

Mass of oxygen = 0.7 g

3.5 - 0.7 = 0.7g

Mass of oxygen = 2.8 g

3. Calculate the number of moles of:

(i) Copper used (Cu = 63.5)

number of moles of copper = mass used/Molar mass

2.8/63.5 = **0.0441moles**

(ii) Oxygen used (O = 16.0)

number of moles of oxygen = mass used/Molar mass

0.7/16.0 = **0.0441moles**

4. Determine the mole ratio of the reactants

Moles of copper/Moles of oxygen

= 0.0441moles/0.0441moles

= 1/1

Mole ratio Cu:O = **1:1**

5.What is the empirical, formula of copper oxide formed.

- CuO (copper (II) oxide)

6. State and explain the observations made during the experiment.

Observation;

- Colour change from **black** to **brown**

Explanation;

- Copper (II) oxide powder is black. On heating it is reduced by a suitable reducing agent to brown copper metal.

QUESTIONS

(1). A oxide of copper contain 80% by mass of copper. Determine its empirical formula. (Cu = 63.5, 16.0)

(2). 1.60g of an oxide of Magnesium contain 0.84g by mass of Magnesium. Determine its empirical formula(Mg = 24.0, 16.0)

(3). An oxide of Silicon contain 47% by mass of Silicon. What is its empirical formula(Si = 28.0, 16.0)

(4). A compound contain 70% by mass of Iron and 30% Oxygen. What is its empirical formula(Fe = 56.0, 16.0)

(5). During heating of a hydrated copper (II) sulphate (VI) crystals, the following readings were obtained:

Mass of evaporating dish = 300.0g

Mass of evaporating dish + hydrated salt = 305.0g

Mass of evaporating dish + anhydrous salt = 303.2g

Calculate the number of water of crystallization molecules in hydrated copper (II) sulphate (VI)

(Cu = 64.5, S = 32.0,O = 16.0, H = 1.0)

Calculating the Molecular Formula of a Compound -

- The molecular formula is the actual number of each kind of atoms present in a molecule of a compound.
- The empirical formula of an ionic compound is the same as the chemical formula but for simple molecular structured compounds, the empirical formula may not be the same as the chemical formula.
- The molecular formula is a multiple of empirical formula .It is determined from the relationship:

(i)

n = Relative formular mass/Relative empirical formula

where **n** is a whole number.

(ii)

Relative empirical formula x **n** = Relative formular mass, where **n** is a whole number.

Practice sample examples

(a). A hydrocarbon was found to contain 92.3% carbon and the remaining Hydrogen.

If the molecular mass of the compound is 78, determine the molecular formula(C=12.0, H=1.0)

Mass of Hydrogen = 100 – 92.3 => 7.7% of Oxygen

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = $\frac{\%\ \text{composition}}{\text{Molar mass}}$	$\frac{92.3}{12}$	$\frac{7.7}{1}$
Divide by the smallest value	$\frac{7.7}{7.7}$	$\frac{7.7}{7.7}$
Mole ratios	1	1

Empirical formula is CH

The molecular formular is thus determined:

n = Relative formular mass/Relative empirical formula

n = 78/13 = 6

The molecular formula is (C H) x 6 = C_6H_6

QUESTIONS

(1). On complete combustion of a sample of hydrocarbon, 3.52 g of carbon (IV) oxide and 1.44g of water were formed. Determine the molecular formula of the hydrocarbon. (Relative molecular mass of hydrocarbon is 56) (Carbon (10) oxide = 44) water=18) (H=1.0) (C=12.0)

(2). An alkanol has the following composition by mass: Hydrogen 13.5%, oxygen 21.6% and carbon 64. . 9%.

a. Determine the empirical formula of the alkanol (C=12.0; H=1.0; 0=16.0) (2 mks)

b. Given that empirical formula and the molecular formula of the alkanol are the same, draw the structure of the alkanol

(3). A hydrated salt has the following composition By mass: iron 20.2% Oxygen 23.0% sulphur 11.5% and 45.3% water of = crystallization. If RMM = 278

Determine the formula of the hydrate salt.

(4). A compound of carbon, hydrogen and oxygen contain 54.55% carbon, 9.09% and remaining 36.36% oxygen. If its relative molecular mass is 88, determine its molecular formula(C=12.0, H=1.0, O=16.0)

(5). A hydrocarbon burns completely in excess air to form 5.28 g of carbon (IV) oxide and 2,16g of water. If the molecular mass of the hydrocarbon is 84, draw and name its molecular structure.

(6). Compound A contain 5.2% by mass of Nitrogen .The other elements present are Carbon, hydrogen and Oxygen. On combustion of 0.085g of A in excess Oxygen,0.224g of carbon(IV)oxide and 0.0372g of water was formed. Determine the empirical formula of A (N=14.0, O=16.0 , C=12.0 , H=1.0

Concentration and Molarity of Solutions

- Concentration refers to the amount of a substance per defined space.
- Another definition of concentration is that concentration is the ratio of solute in a solution to either solvent or total solution.
- Molarity is a unit of concentration, defined as the number of moles of solute divided by the number of litres of solution.
- Molarity is the most common measure of concentration used in chemistry.

- Molarity is given as capital M for the units mol/L (or mol/dm^3 & 1L = 1dm^3 = 1000cm^3)

Concentration is affected by the following factors:

1. Mass of solute,
2. Dilution with volume of solvent and
3. Evaporation to remove some solvent by reducing its volume.

- Concentration given in g/L similar to how density is expressed in g/cm^3. The mass may be converted to moles and given as mol/L.
- Other units used are molality, normality, mass percent, mole fraction, parts per million, parts per billion e.t.c.

M = n(moles of solute)/v(volume of solution)

EXAMPLE;

(a). If 4.00g of sodium hydroxide were dissolved in 400cm^3 of distilled water. Determine the concentration in molarity of the resulting basic solution.

(Na = 22.99, O = 16.00, H = 1.01).

(i)

By first principles method

Chemical Formula: NaOH

Molar mass of NaOH: 23 + 16 + 1 = 40 g Moles in 4.0 g: 4/40 = 0.1 moles

Volume used: 400cm^3.

If 400 cm^3 contains 0.1 moles

Then 1000 cm^3 will have (1000 x 0.1)/400

Concentration = 0.250 mol/L (0.250 M)

(ii)

By formula

Chemical Formula: NaOH

Molar mass of NaOH: 23 + 16 + 1 = 40 g Moles in 4.0 g: 4/40 = 0.1 moles

Volume used: 400/1000 = 0.4 L

Concentration = 0.1/0.4

Concentration = 0.250 mol/L (0. 250 M)

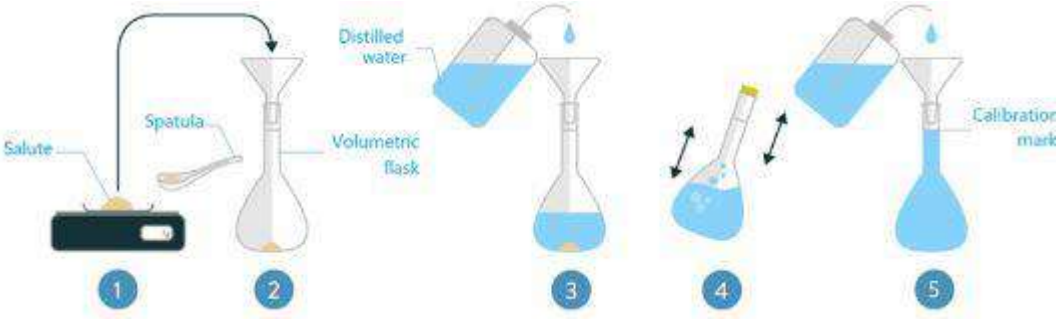
QUESTIONS

1. Calculate the molarity of a solution if;
2. 0.5 grams of sodium chloride is dissolved to make 0.05 litres of solution.
3. 734 grams of lithium sulfate are dissolved to make 2500ml of solution.
4. 6.7×10^{-2} grams of $Pb(C_2H_3O_2)_4$ are dissolved to make 3.5ml of solution.
5. Calculate the mass of : $25cm^3$ of 0.2^3 sodium hydroxide solution(Na =23.0, O =16.0, H=1.0)
6. The density of sulphuric (VI) is $1.84gcm^3$ Calculate the molarity of the acid.

Preparation of molar solution

Procedure

- Weigh accurately 4.0 g of sodium hydroxide pellets into a 250 cm^3 volumetric flask.
- Using a wash bottle add about 200 cm^3 of distilled water.
- Stopper the flask.
- Shake vigorously for three minutes.
- Remove the stopper for a second then continue to shake for about another two minutes until **all** the solid has dissolved.
- Add more water slowly up to **exactly** the $250cm^3$ **mark**.





- For accurate concentrations, one must make careful use of measuring apparatus such as: volumetric flasks, pipettes, and burettes.
- Accurate weighing is an essential part of chemistry since an extra 0.001 g might result in a different concentration. Use measurements and calculations to the appropriate number of significant figures or decimal places (usually 4).
- One of the core principles of science, and an important moral rule, is to describe experimental procedures completely and carefully so that another person can follow and reproduce the results.
- To avoid contamination, ensure that the spatula and all other apparatus are extremely clean.
- Always use volumetric flasks to prepare molar solutions and nothing else.

QUESTIONS

(1). What mass of lithium phosphate (Li_3PO_4) would you need to make?

a. 2.5litre of 1.06 M lithium phosphate solution?

b. 250 cm^3 of 2M solution. (Li=6.94, P=30.97, O= 16.00)

(2). A chemist has 4.0 g of silver nitrate, ($AgNO_3$) and needs to prepare 2.0 L of a 0.010 M solution.

a. Will there be enough silver nitrate? If so, how much silver nitrate will be left over?

b. What would be the concentration of he dissolved only 2.0 g in 500cm^3 of water? (Ag = 107.87, N = 14.01, O = 16.00)

(3). A chalice contains 36.45 grams ammonium chlorite, NH_4ClO_2, in 2.36 liters of solution - calculate the molarity.

(4). What mass would be needed to make 250 cm^3 of 0.20M solution?(N=14.01, Cl=35.45, O=16.00, H=1.01)

(5). 15.0cm^3 ethanoic acid (CH_3COOH) was dissolved in water to make 500 cm^3 of solution. Calculate the concentration of the solution in moles per litre. (C=12.0; H=1.0; O=16.0) density of ethanoic is 1.05g/cm^3

Dilution of Solutions

- **A Stock Solution** is a concentrated solution that will be diluted to some lower concentrated for actual use.
- Stock solutions save preparation time, conserve materials, reduce storage space, and improve the accuracy with which working solutions are prepared.
- **A serial dilution** is where the diluted material of the previous step is used to make the subsequent dilution.
- **A step dilution** involves using the same concentration to make a dilution.
- To dilute a solution means to add more solvent without adding more solute.
- Preparing a molar solution by dilution of a more concentrated stock solution is easier and quicker than by weighing the salt/solute and adding to a solvent each time you need to prepare the solution.
- The fact that the amount of solute does not change on dilution allows us to develop a calculation techniques.

We use an accurate volume V1 cm^3 of a stock (original) solution having concentration/ molarity M1, (mol/L) to prepare volume V2 cm^3 of dilute solution with lesser concentration of M2 (mol/L).

Measure accurately 25 cm^3 of 2M HCl. Transfer the solution into a 250 cm^3 volumetric flask. Add distilled water to the acid and make up to 250 cm^3 of solution.

Determine the number of moles in the 25 cm^3 of 2M HCl

Determine the molarity of the solution contained in the 250 cm^3 flask

Solution;

M = moles/volume

Moles = M x Volume

2moles - 1000 cm^3/25 cm^3

= 0.05moles

These moles are the ones that are still in the 250cm^3 of solution after water is added.

So;

0.05 moles = 250 cm^3/1000 cm^3

= 0.2M

So;

(2M x 25cm^3) = 50moles

(0.2M x 250cm^3) = 50 moles

M1xV1= M2xV2

The Dilution equation,

We know that the mass hence moles of the solute remain the same, therefore moles of solute before dilution = moles of solute after dilution, and that

Moles = Molarity x Volume [moles = M x V] for a given volume of any solution whose concentration/molarity is known. Therefore; moles in the stock/original solution is M1V1 , and also moles in the final dilute solution is M2V2,

Dilution equation equates the two moles as below:

M1V1 = M2V2.

QUESTIONS

1. 53.4 mL of a 1.50 M solution of NaCl is on hand, but you need some 0.800 M solution. How many mL of 0.800 M can you make?
2. 100.0 mL of 2.500 M KBr solution is on hand. You need 0.5500 M. What is the final volume of solution which results?
3. A stock solution of 1.00 M NaCl is available. How many milliliters are needed to make 100.0 mL of 0.750 M
4. What volume of 0.250 M KCl is needed to make 100.0 mL of 0.100 M solution?

Stoichiometry

A balanced chemical equation is the recipe for a reaction: it contains a list of all the reactants (the ingredients) and products (the cookies) as well as their relative proportions.

- Using a balanced chemical equation to calculate amounts of reactants and products is called **stoichiometry**.
- The **coefficients** from the balanced chemical reaction tell us the proportions of the reactants and products.
- We can use ratios of the coefficients to convert between amounts of reactants and products in our reaction.
- **A mole ratio** is the ratio between the amounts in moles of any two compounds involved in a chemical reaction.
- Mole ratios are used as conversion factors between products and reactants in many chemistry problems.

- **The mole ratio** may be determined by examining the coefficients in front of formulas in a balanced chemical equation.

Key things to consider while doing stoichiometry is:

1. Reactants and products must be known.
2. Write a well balanced chemical equation.
3. Derive the mole ratio from the coefficient used in balancing.
4. Convert mass given to moles.
5. Use the mole ratio to find the moles of the unknown.
6. Use the moles of the unknown and its rmm to find the mass of the unknown.

Example:

If 4.0 g of calcium was burnt in unlimited supply of oxygen, how much Calcium Oxide was produced?

$2Ca + 1O_2 \rightarrow 2CaO$

Moles of calcium 4/40 = 0.1 moles

Mole ratio,

2 : 1

0.1 --> ?

(0.1 x 1)/2 = 0.05 moles

Mass = moles x MM

= 0.05 x 40

= 2.0 g

- **Theoretical yield** is the amount of product that can be produced/ synthesized from reactants under ideal conditions.
- Experimental yield is the amount of product actually synthesized in the experiment.

- The percentage yield is determined by calculation:

(Actual yield x 100%)/Theoretical yield

- It indicates the percent of theoretical yield that was obtained from the final product in an experiment.
- If we are not told the reagent that is in excess, then we have to find the limiting reagent:
- **A limiting reagent** is the reactant that is used up first leading to a halt in the reaction.
- The other reactant is then referred to as the reagent **in excess.**

QUESTIONS

(1). Given the following reaction:

$Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$

How many moles of Ag will be produced from 16.0g of Cu, assuming AgNO3 is available in excess.

NB; moles(round to three significant figures)

(2).

$1Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCL(aq) + 1CO_2(g) + 1H_2O\ (l)$

a. If we had reacted 2.5g of sodium carbonate reacting with an excess of the acid, how much NaCl will be produced?

b. If this reaction was to be performed in the lab, will the mass of NaCl retrieved be exact as per your calculations? Explain.

Ionic equations

- An ionic equation is a chemical statement showing the movement of ions (cations and anions) from reactants to products.
- **Solids**, **gases** and **liquids** do not ionize/dissociate into free ions. **Only** ionic compounds in **aqueous/solution** or **molten** state ionize/dissociate into free cations and anions (**ions**)
- An ionic equation is usually derived from a stoichiometric equation by using the following guidelines;

Guidelines for writing ionic equations

(1). Write the balanced stoichiometric equation

(2). Indicate the state symbols of the reactants and products

(3). **Split** into cations and anions all the reactants and products that exist in **aqueous** state.

(4). **Cancel out** any cation and anion that appear on **both** the product and reactant side.

(5). Rewrite the chemical equation. It is an ionic equation.

Practice;

(**a**) Precipitation of an insoluble salt

- All insoluble salts are prepared in the laboratory from double decomposition /precipitation. This involves mixing **two soluble** salts to form **one soluble** and **one insoluble** salt

1. When silver nitrate(V) solution is added to sodium chloride solution, sodium nitrate (V) solution and a white precipitate of silver chloride are formed.

Balanced stoichiometric equation;

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Split reactants product existing in aqueous state as cation/anion;

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na+(aq)+ NO_3-(aq)$

Cancel out ions appearing on reactant and product side;

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq)+ NO_3^-(aq)$

Rewrite the equation

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$ (ionic equation)

2. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

Balanced stoichiometric equation

$Ba(NO_3)_2(aq) + CuSO_4(aq) \rightarrow BaSO_4(s) + Cu(NO_3)_2(aq)$

Split reactants product existing in aqueous state as cation/anion

$Ba^{2+}(aq) + 2NO_3^-(aq) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2NO_3^-(aq)+ Cu^{2+}(aq)$

Cancel out ions appearing on reactant and product side

$Ba^{2+}(aq) + 2NO_3^-(aq) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2NO_3^-(aq) + Cu^{2+}(aq)$

Rewrite the equation

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ (ionic equation)

3.A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.

Balanced stoichiometric equation

$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

Split reactants product existing in aqueous state as cation/anion;

$Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2NO_3^-(aq) + 2K^+(aq)$

Cancel out ions appearing on reactant and product side

$Pb^{2+}(aq) + 2NO3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2NO3^-(aq) + 2K^+(aq)$

Rewrite the equation

$Pb^{2+}(aq) + 2I^- (aq) \rightarrow PbI_2 (s)$ **(ionic equation)**

Neutralization

- Neutralization is the reaction of an acid with a **soluble** base/alkali or **insoluble** base.

(i)Reaction of alkalis with acids

(1). Reaction of nitric(V)acid with potassium hydroxide

Balanced stoichiometric equation

$HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$

Split reactants product existing in aqueous state as cation/anion

$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \rightarrow H_2O(l) + NO_3^-(aq) + K^+(aq)$

Cancel out ions appearing on reactant and product side

$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^- (aq) \rightarrow H_2O (l) + NO_3^-(aq) + K^+(aq)$

Rewrite the equation;

$H^+(aq) + OH^-(aq) \rightarrow H_2O (l)$ (ionic equation)

(2). Reaction of sulphuric(VI)acid with ammonia solution

Balanced stoichiometric equation

$H_2SO_4(aq) + 2NH_4OH(aq) \rightarrow H_2O(l) + (NH_4)\ 2SO_4(aq)$

Split reactants product existing in aqueous state as cation/anion

$2H^+(aq) + SO_4^{2-}(aq) + 2NH_4^+(aq) + 2OH^- (aq) \rightarrow 2H_2O(l) + SO_4^{2-} (aq) + 2NH_4^+(aq)$

Cancel out ions appearing on reactant and product side

$2H^+(aq) + SO_4^{2-}(aq) + 2NH_4^+(aq) + 2OH^-(aq) \rightarrow 2H_2O (l) + SO_4^{2-}(aq) + 2NH_4^+(aq)$

Rewrite the equation;

$2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ **(ionic equation)**

(3). Reaction of hydrochloric acid with Zinc hydroxide

Balanced stoichiometric equation

$2HCl(aq) + Zn(OH)_2(s) \rightarrow 2H_2(l) + ZnCl_2(aq)$

Split reactants product existing in aqueous state as cation/anion;

$2H^+(aq) + 2Cl^- (aq) + Zn(OH)_2(s) \rightarrow 2H_2O(l) + 2Cl^-(aq) + Zn^{2+}(aq)$

Cancel out ions appearing on reactant and product side

$2H^+(aq) + 2Cl^-(aq) + Zn(OH)_2(s) \rightarrow 2H_2O(l) + 2Cl^-(aq) + Zn^{2+}(aq)$

Rewrite the equation;

$2H+(aq) + Zn(OH)_2 (s)$

QUESTIONS

1. Write a the ionic equation for;
2. Reaction of hydrochloride acid with zinc metal
3. Reaction of sodium carbonate with zinc nitrate
4. Reaction of magnesium hydroxide with sulfuric acid
5. Reaction of potassium sulphate with lead nitrate
6. Reaction of sodium hydroxide with sulfuric acid.

Titration

1. **Titration** is an experimental technique used to determine the concentration of an unknown solution.
2. It usually involves acids and bases, but may be extended to any addition of one solution to another to find out the amounts required for reaction.

3. Since volume measurements play a key role in titration, it is also known as **volumetric analysis.**
4. A reagent, called the **titrant or titrator** is prepared as a standard solution.
5. A known concentration and volume of titrant/titrator reacts with a solution of **analyte or titrand,** whose concentration is unknown, in order to determine concentration

➔The procedure involves adding one solution to another solution, in the presence of an indicator that shows when the two solutions have completely reacted.

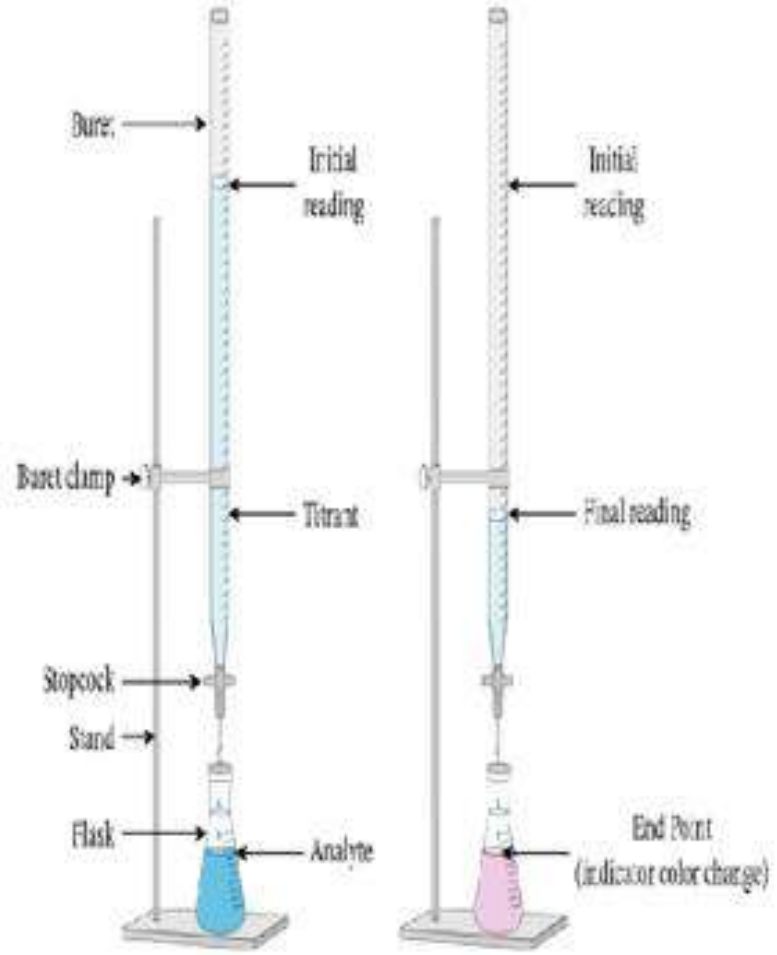
➔If the concentration and volume of one of the solutions is known and the volume of the other solution is known, then the unknown concentration can be calculated providing the stoichiometry of the reaction is known.

During volumetric analysis, the solution whose exact concentration is known is called a **standard solution.**

Steps of Titration

1. The burette is filled with the acid. A burette is used to place the acid of unknown concentration.
2. The alkali is added to the conical flask using a pipette.
3. An indicator is added to the alkali. An indicator is added to the alkali in the conical flask, which causes a colour change to occur. When titrating a strong acid and strong alkali any commercial indicator is used, While when titrating a strong acid and a weak base methyl orange is used and finally when titrating a weak acid with a strong base phenolphthalein is used
4. The acid is added into the alkali. The acid from the burette is allowed to run into the alkali, swirling constantly to ensure the reactants mix well.
5. The acid Is added slowly when the indicator changes colour.
6. The acid stops being added when there is a permanent colour change
7. The titre is recorded in the table.

The titration is repeated three times;



Titration number	1	2	3
Final burette reading (cm^3)	20.0	20.0	20.0
Initial burette reading (cm^3)	0.0	0.0	0.0
Volume of solution used(cm^3)	20.0	20.0	20.0

Calculating Concentration of titrant

The known volume and concentration of the alkali enable the moles to be calculated.

The balanced symbol equation enable the moles ratio of acid: alkali to be determined. Hence the number of moles of acid can be calculated.

The mean titre is the volume of the acid. Using the moles and the volume, the concentration of the unknown acid can be determined.

QUESTIONS

(1). Calculate the molarity of MCO_3 if $5.0cm^3$ of MCO_3 react with $25.0cm^3$ of 0.5M hydrochloric acid.(C = 12.0 ,O = 16.0)

(2). $2.0cm^3$ of 0.5M hydrochloric acid react with 0.1M of M_2CO_3. Calculate the volume of 0.1M M_2CO_3 used.

(3). $5.0cm^3$ of 0.1M sodium iodide react with 0.1M of Lead(II)nitrate(V). Calculate

a. The volume of Lead(II)nitrate (V) used.

b. The mass of Lead (II) Iodide formed.

DIRECT TITRATION;

- A direct titration is the basic titration method that involves the reaction between the unknown compound and the compound with known concentration. Here, the addition of excess reagents is not done as in back titrations. The unknown compound is directly reacted with the known compound. Therefore, the end point of the titration indicates the end of the reaction. By using that endpoint, the amount of unknown compound present in the sample solution can be determined.

You are provided with:

- 0.1M sodium hydroxide solution A
- Hydrochloric acid solution B

You are required to determine the concentration of solution B in moles per litre.

Procedure

Fill the burette with solution B. Pipette $25.0cm^3$ of solution A into a conical flask. Titrate solution A with solution B using phenolphthalein indicator to complete the titration table 1.

Sample results Titration table 1

Titration number	1	2	3
Final burette reading (cm^3)	20.0	20.0	20.0
Initial burette reading (cm^3)	0.0	0.0	0.0
Volume of solution B used(cm^3)	20.0	20.0	20.0

Sample worked questions

1. Calculate the average volume of solution B used

Average titre = (Titre 1 + Titre 2 +Titre 3)/3 => (20.0 +20.0 +20.0)/3 = **$20.0cm^3$**

2. How many moles of:

(i)solution A were present in $25cm^3$ solution.

Moles of solution A = (Molarity x volume)/1000 = (0.1 x 25)/1000 = **2.5×10^{-3}** moles

(ii)solution B were present in the average volume.

Chemical equation: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Mole ratio 1:1 → Moles of A = Moles of B = **2.5×10^{-3} moles**

(iii) solution B in moles per litre.

Moles of B per litre = (moles x 1000)/Volume = ($2.5 \times 10^{-3} \times 1000$)/20 = **0.1M**

Questions

(1). You are provided with;

Aqueous sulphuric (VI) acid labelled A.

Solution B containing 8.48g per litre of sodium carbonate

You're required to determine the:

Concentration of solution A.

Procedure:

(i) Using a pipette, place 25.0cm^3 of solution A into a 250 ml volumetric flask.

(ii) Add distilled water to make 250cm^3 solution. Label this solution M. Place solution M in a burette.

(iii) Clean the pipette and use it to place 25.0cm^3 of solution B into a conical flask.

(iv) Add 2 drops of methyl orange indicator provided and titrate with solution M.

(v) Record your results in the table below.

(vi) Repeat the titration two more times and complete the table.

Titration number	1	2	3
Final burette reading (cm3)	25.0	25.0	25.0
Initial burette reading (cm3)	0.0	0.0	0.0
Volume of solution M used(cm3)	25.0	25.0	25.0

Calculate the:

1. Average volume of solution M used [1 mark]
2. Concentration of sodium carbonate in solution B (2 marks) (Na = 23.0, O = 16.0, C = 12.0
3. Concentration of sulphuric (VI) acid in solution M. (2 marks)

Standard

A standard solution is one whose concentration is accurately known.

- For some substances a standard solution can be prepared by **weighing accurately** a given mass of pure substance and dissolving in water to give a **specific volume.** Such substances are called **primary Standards** like; sodium Carbonate and sodium ethanedioa.
- They must be stable.
- They should not react with. Components of the atmosphere
- **Most acids and bases** react or absorb components of the atmosphere. They are made by dissolving an **approximate mass** of the substance in water to a specific volume
- Their actual concentration is the determined by titrating them with a solution of a Primary Standard

- **Standardization** is the process in which one determines the concentration of a solution by accurately measuring the volume of it required to react with a known volume of a primary Standard
- The Standardization solution is called a **secondary Standard.**

QUESTION

(1). You are provided with;

- 0.125M HCl labelled solution **M**
- Anhydrous sodium carbonate labelled solid **P**
- Methyl orange indicator.

You required to prepare a solution of sodium carbonate and standardize it using hydrochloric acid solution **M.**

Procedure;

- Transfer all solid **P** into a 250cm^3 volumetric flask. Add 100cm^3 of distilled water and shake well until all solid dissolves. Add more distilled water and top up the solution to 250cm^3 mark. Label this as solution **T.** Using a measuring cylinder measure 50cm^3 of solution **T** and put it in a 250cm^3 beaker. Dilute it by adding 50cm^3 of distilled water. Shake well to form uniform solution, Label this solution as **R.**
- Pipette 25.0cm^3 of R into a conical flask. Add two drops of methyl orange indicator. Fill the burette with solution **M.** Carry out titration.
- Record your results in the table below. Repeat the titration twice and complete the table.
- Volume of pipette used ..... cm^3

Titration	1	2	3
Final burette readings (cm^3)			
Initial burette reading (cm^3)			
Volume of solution M used cm^3			

a. Calculate;

(i) Average volume of solution M used. (1mrk)

(ii) Moles of solution M used. (1mrk)

b. Write a balanced equation for the reaction between solution R and M.

c. Calculate;

(i) Moles of Na_2CO_3 in 25cm^3 of solution R (1mrk)

(ii) Number of moles Na_2CO_3 in 100cm^3 of solution R. (1mrk)

(iii) Number of moles Na_2CO_3 in 50cm^3 of solution T.

d. Given that Na = 23, C = 12, O = 16

Calculate;

(i) Mass of solid Na_2CO_3 that was dissolved to make solution T. (2mrks)

(ii) The concentration of Na_2CO_3 in solution I in moles per litre. (1mrk)

Redox Titrations

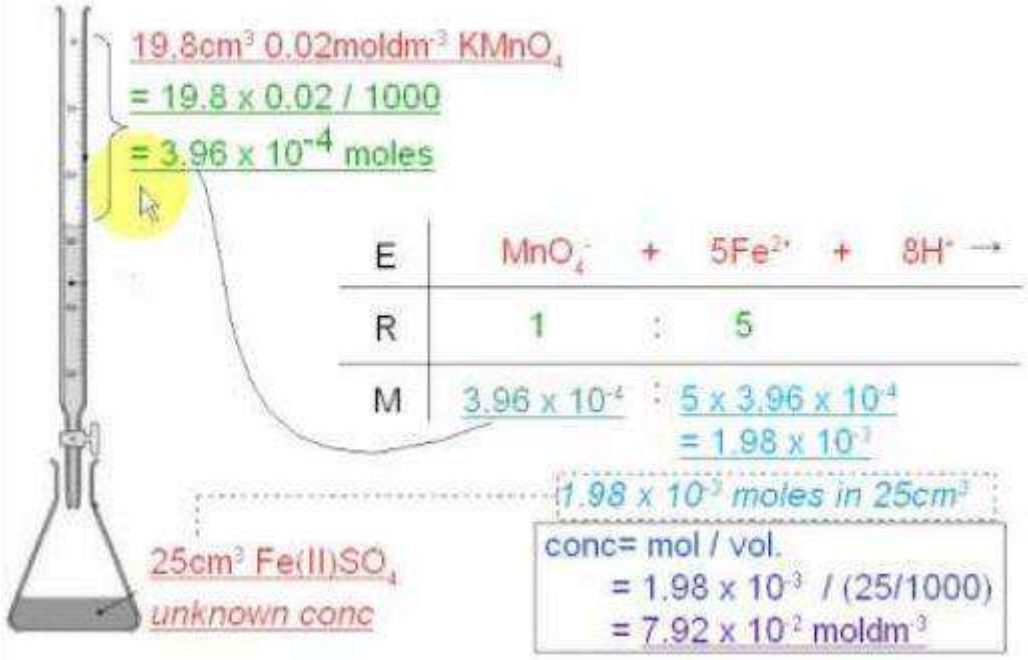
A redox titration is the process of determining the quantity of a sample by monitoring a redox process until the endpoint is reached, when all of the sample has reacted.

Most common oxidizing agents;

- Potassium manganate (VII)
- Potassium dichromate (VI)

EXAMPLE;

Find the concentration of an iron (ii) sulphate solution, given that 25.0cm³ of the solution, when acidified, required 19.8cm³ of 0.0200moldm⁻³ potassium manganate (vii) for oxidation.



QUESTIONS

(1). Find the concentration of Iron (II) sulphate solution, given that 25.0cm³ of the acidified solution required 19.8cm³ of 0.02M potassium manganate (VII) for complete oxidation.

(2). 5.0cm³ of 0.1M sodium iodide react with 0.1M of Lead (II)nitrate(V).

Calculate

a. The volume of Lead (II) nitrate(V) used.

b. The mass of Lead (II) Iodide formed.

(3). A solution of 25.0cm³ of Iron (II) ammonium sulphate needed exactly 20.0cm³ of a solution of potassium dichromate (VI) of concentration 0.05M for complete oxidation. State the color of the solution at the endpoint.

Given the equation: $6Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 6Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

Calculate the concentration of Iron (II) ions in the Iron (II) ammonium sulphate. [10 marks]

Back titration

A back titration is a titration method used to determine the concentration of an unknown using an excess amount of a compound with a known concentration. There is a chemical reaction between these

compounds. Since the amount of the compound with a known concentration added is known already, we can determine the amount of the compound that has reacted with the unknown compound by doing a back titration.

1. A back titration may be thought of as a normal titration but done in reverse.
2. A back titration is also called **indirect titration.**

Two steps are followed in a back titration.

1)First, the analyte is permitted to react with an excess reagent.

Next, a titration is conducted on the remaining quantity of the known solution. This is a way to measure the amount that remained unreacted i.e. the excess quantity

EXAMPLE;

(1). In an experiment, the following procedure was followed:

1 g of egg shells were reacted with 25 ml of a 3M HCl acid.

After the reaction stopped, the resulting solution was filtered into a 250 ml volumetric flask. Distilled water was used to top the solution to 250 ml mark and solution labelled Y. 25 ml of this solution was pipetted into a clean conical flask and a few drops of phenolphthalein indicator added. This solution was titrated against a standardized 2M NaOH solution to obtain four values.

The data below was recorded below.

	Trial	1	2	3
Final burette reading/cm^3	29.60	29.75	29.65	29.70
Initial burette reading/cm^3	0.00	0.00	0.00	0.00
Volume of NaOH used/cm^3	29.60	29.75	29.65	29.70

Steps:

Get average titre - Volume of NaOH required for neutralization.

Get moles of this NaOH required for neutralization.

Write a ba;anced chemical equation for reaction that took place.

Use equation get mole ratio and hence moles of HCl in 25 ml aliquot/ flask.

Work out moles of HCl in 250 ml volumetric flask. This will be the moles that remained after the reaction.

Work out original moles of the 25 ml of 3M HCl.

Subtract moles remaining from the original moles to get the moles of HCl that reacted.

Using equation of $CaCO_3$ and HCl, get the mole ratio and hence the moles of $CaCO_3$.

Using molar mass and moles of $CaCO_3$ calculate the mass of $CaCO_3$

Using the mass, convert the mass to percentage of $CaCO_3$ in 1g of the egg shells.

Note:

The resultant mixture was diluted to form 250 ml solution then a small volume, 25 ml, taken from the diluted solution to be analyzed by titration.

QUESTIONS;

(1). When 4g of a metal carbonate MCO_3 was dissolved in 160cm^3 of 1M hydrochloric acid and then the resultant solution diluted to one litre, 25.0cm^3 of this solution required 20.0cm^3 of 0.1M sodium hydroxide solution for complete neutralization. Calculate

a. The number of moles per litre of excess hydrochloric acid that reacted with sodium hydroxide, NaOH.

b. The number of moles per litre of acid that reacted with the carbonate.

c. The number of moles of carbonate that reacted with the acid.

d. The formula mass of the carbonate, MCO_3

e. The atomic mass of the metal M. (C = 12.0, O = 16.0)

(2). A 0.4108 gram sample of $CaCO_3$ (100.087 g/mol) is added to a flask along with 15.00 mL of 2.000 M HCl.) Enough water is then added to make 250.0 mL of solution A. A 20.00 mL aliquot of solution A is taken and titrated with 0.1160 M NaOH. How many mL of NaOH are used?

(3). A 0.9030 gram sample of $M(OH)_2$ was mixed with 20.00 mL of 2.000 M HCl and enough water added to make 100.0 mL of solution A. A 10.00 mL aliquot of solution A was taken and titrated with 17.64 mL of 0.05121 M NaOH. What is the metal, M?

Molar Gas Volume

Atomicity is the number of atoms of constituent elements combined to form a molecule.

Example;

- Noble gases (helium, neon, argon, krypton, etc.) have one atom each in their molecules such as He, Ne, Ar and Kr. So, the atomicity of noble gases is 1.

Example.

- Hydrogen gas (H_2), nitrogen gas (N_2), oxygen gas (O_2) have two atoms in each molecule. So their atomicity is 2.

- The terms, monoatomic, diatomic, triatomic e.t.c are used to describe the atomicity
- **AVOGADRO'S LAW - EQUAL VOLUMES OF GASES UNDER THE SAME CONDITIONS OF TEMPERATURE & PRESSURE CONTAIN THE SAME NUMBER OF MOLECULES/MOLES.**
- The molar volume, symbol Vm, is the volume occupied by one mole of a substance (chemical element or chemical compound) at a given temperature and pressure.
- It is equal to the molar mass (M) divided by the mass density (ρ).
- It has the SI unit cubic metres per mole (m^3/mol).

MOLAR GAS VOLUME

- One mole of any gas has a volume of 24000cm^3 (24 Litres = 24dm^3) at rtp (room temperature, 25^{0}C and pressure, 1 atmosphere).

- At stp (standard temperature, 0^0C and Pressure, 1 atmosphere) the molar gas volume (gas would occupy) is $22.4cm^3$ ($22.4L=24dm^3$)

EXAMPLE;

(1). Calculate the volume of carbon(IV)oxide at r.t.p produced when 5.0 g of calcium carbonate is strongly heated.(Ca = 40.0, C = 12.0,O = 16.0,1 mole of gas = 22.4 at r.t.p)

Chemical equation

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Mole ratios 1: 1: 1

Molar Mass $CaCO_3$ =100g

Method 1

100g $CaCO_3(s) \rightarrow 24dm^3$ $CO_2(g)$ at r.t.p

5.0 g $CaCO_3(s) \rightarrow (5.0g \times 24dm^3)/100g = 1.2dm^3/1200cm^3$

Method 2

Moles of 5.0 g $CaCO_3(s)$ = 5.0 g/100g = 0.05 moles

Mole ratio 1:1

Moles of $CO_2(g)$ = 0.05moles

Volume of $CO_2(g) = 0.05 \times 24000cm^3 = 1200cm^3$ /$1.2dm^3$

QUESTIONS

(1). 1.0g of an alloy of aluminium and copper were reacted with excess hydrochloric acid. If $840cm^3$ of hydrogen at s.t.p was produced, calculate the % of copper in the alloy. (Al = 27.0,one mole of a gas at s.t.p = $22.4dm^3$)

(2). When excess dilute hydrochloric acid was added to sodium sulphite, $960cm^3$ of sulphur (IV) oxide gas was produced. Calculate the mass of sodium sulphite that was used. (Molar mass of sodium sulphite = 126g: and molar gas volume = $24000cm^3$)

(3). Calculate the volume of carbon (iv) oxide measured at S.T.P that is evolved when 1 mole of copper (II) carbonate is heated to a constant mass

(4). How many molecules are there in $360cm^3$ of nitrogen as r.t.p

(5). Define the following terms

a. Monatomic gas

b. Diatomic gas

c. Atomicity of an element

Gay Lussac's Law

Gay Lussacs law states that "when gases combine/react they do so in simple volume ratios to each other and to their gaseous products at constant/same temperature and pressure"

- Gay Lussacs law thus only apply to gases
- Given the volume of one gas reactant, the other gaseous reactants can be deduced thus:

Examples

(1). Calculate the volume of Oxygen required to completely react with $50cm^3$ of Hydrogen.

Chemical equation:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Volume ratios 2 : 1 : 0

Reacting volumes $50cm^3 : 25cm^3$

= $50cm^3$ of Oxygen is used

QUESTIONS

(1). Calculate the volume of air required to completely reacts with $50cm^3$ of Hydrogen.(assume Oxygen is 21% by volume of air

(2). If $5cm^3$ of a hydrocarbon CxHy burn in $15cm^3$ of Oxygen to form $10cm^3$ of Carbon(IV)oxide and $10cm^3$ of water vapour/steam, obtain the equation for the reaction and hence find the value of x and y in CxHy.

(3). $100cm^3$ of nitrogen (II) oxide NO combine with $50cm^3$ of Oxygen to form $100cm^3$ of a single gaseous compound of nitrogen. All volumes measured at the same temperature and pressure. Obtain the equation for the reaction and name the gaseous product.

(4). When $15cm^3$ of a gaseous hydrocarbon was burnt in $100cm^3$ of Oxygen ,the resulting gaseous mixture occupied $70cm^3$ at room temperature and pressure. When the gaseous mixture was passed through, potassium hydroxide its volume decreased to $25cm^3$.

a. What volume of Oxygen was used during the reaction

b. Determine the molecular formula of the hydrocarbon

(5). Given the following reaction:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

If $20cm^3$ of Hydrogen gas reacts with $10cm^3$ of oxygen gas, how much steam ($H_2O(9)$will be evolved?

TOPICAL REVISION QUESTIONS

1. In an experiment magnesium ribbon was heated in air. The product formed was found to be heavier than the original ribbon. Potassium manganate (VII) was on the other hand, heated

in air and product formed was found to be lighter. Explain the differences on the observation made

2. In a filtration experiment 25cm^3 of a solution of Sodium Hydroxide containing 8g per litre was required for complete neutralization of 0.245g of a dibasic acid. Calculate the relative molecular mass of the acid (Na = 23.0, O = 16, H= 1)

3. D grams of Potassium hydroxide were dissolved is distilled water to make 100 cm^3 of solution. 50cm^3 of the solution required 50cm^3 of 2.0M nitric acid for complete neutralization. Calculate the mass D of Potassium hydroxide (RFM of KOH = 56)
$KOH_{(aq)} + HNO_{3(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$

4. When excess dilute hydrochloric acid was added to sodium sulphite, 960cm^3 of sulphuric (IV) Oxide gas was produced. Calculate the mass of sodium sulphate that was used.
(Molar gas volume = 24000cm^3 and Molar mass of sulphite = 126g)

5. The equation of the formation of iron (III) chloride is
$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_3$
Calculate the volume of chlorine which will react with iron to form 0.5g of Iron (III) chloride. (Fe = 56 Cl=35.5). Molar gas volume at 298K = 24dm^3)

6. 15.0cm^3 of ethanoic acid (CH_3COOH) was dissolved in water to make 500cm^3 of solution. Calculate the concentration of the solution in moles per litre [C=12, H = 1, O = 16, density of ethanoic acid is 1.05g/cm^3]

7. When 1.675g of hydrated sodium carbonate was reacted with excess hydrochloric acid, the volume carbon (IV) oxide gas obtained at room temperature and pressure was 150cm^3. Calculate the number of moles of water of crystallization in one mole of hydrated sodium carbonate:-
(Na=23, H =1, C=12, O=16, MGV at R.T.P = 24000cm^3)

8. How many chloride ions are present in 1.7g of magnesium chloride crystals? (Avogadro's constant = 6.0 x 10^{23}, Mg = 24, Cl = 35.5)

9. Calculate the volume of oxygen gas used during the burning of magnesium (O = 16, molar gas volume = 24,000cm^3 at room temperature)

10. A hydrated salt has the following composition by mass. Iron 20.2 %, oxygen 23.0%, sulphur 11.5%, water 45.3%

    i. Determine the formula of the hydrated salt (Fe=56, S=32, O=16, H=11)

    ii. 6.95g of the hydrated salt in (i) above were dissolved in distilled water and the total volume made to 250cm^3 of solution. Calculate the concentration of the resulting salt solution in moles per litre. (Given that the molecula mass of the salt is 278)

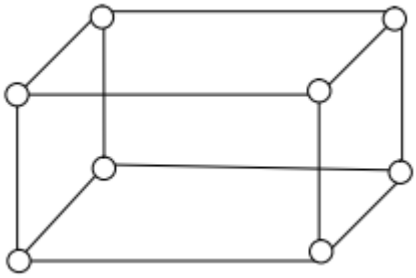
11.

    i. Lead (II) ions react with iodide ions according to the equation;
    $Pb^{2+}_{(aq)} + 2I^-_{(aq)} \rightarrow PbI_{2(s)}$
    300cm^3 of a 0.1m solution of iodide ions was added to a solution containing excess lead (II) ions. Calculate the mass in grams of lead II iodide formed

    ii. Identify the colour of the product formed in (i)

12.

i. The diagram below represents part of the structure of sodium chloride crystal



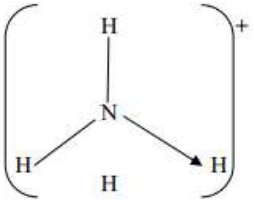
The position of one of the sodium ions in the crystal is shown as;

i. On the diagram, mark the positions of the other three sodium ions

ii. The melting and boiling points of sodium chloride are 801°C and 1413°C respectively. Explain why sodium chloride does not conduct electricity at 25°C, but does not at temperatures between 801°C and 1413°C

ii. Give a reason why ammonia gas is highly soluble in water

iii. The structure of ammonium ion is shown below;



Name the type of bond represented in the diagram by N → H

iv. Carbon exists in different crystalline forms. Some of these forms were recently discovered in soot and are called fullerenes

i. What name is given to different crystalline forms of the same element

ii. Fullerenes dissolve in methylbenzene while the other forms of carbon do not. Given that soot is a mixture of fullerenes and other solid forms of carbon, describe how crystals of fullerenes can be obtained from soot

iii. The relative molecular mass of one of the fullerenes is 720. What is the molecular mass of this fullerene

13. Calculate the volume of oxygen gas used during the burning of magnesium (O = 16, molar gas volume = 24,000cm^3 at room temperature)

14. Study the information in the table below and answer the questions that follow

Number of carbon atoms per molecule	Relative molecular mass of the hydrocarbon
2	28
3	42
4	56

i. Write the general formula of the hydrocarbons in the table

ii. Predict the relative atomic mass of the hydrocarbons with 5 carbon atoms

iii. Determine the relative atomic mass of the hydrocarbon in (ii) above and draw its structural formula (H=1.0, C=12.0)

15. Galvanized iron sheets are made by dipping the sheets in molten Zinc.

i. Explain how zinc protects iron from rusting

ii. Name the process applied in galvanization of iron with zinc

16. A factory uses nitric acid and ammonia gas as the only reactant for the preparation of the fertilizer if the daily production of the fertilizer is 4800kg. Calculate the mass of ammonia gas used daily (N = 14.0, O= 16.0, H = 1.0)

17. Calculate the volume of sulphur (VI) oxide gas that would be required to produce 178kg of oleum in step 3 molar gas volume at s.t.p = 22.4 litres (H = 1, O = 16, S = 32)

18. A sample of biogas contains 35.2% by mass of methane. A biogas cylinder contains 5.0kg of the gas. Calculate:

    i. Number of moles of methane in the cylinder (Molar mass of methane = 16)

    ii. Total volume of carbon (IV) oxide produced by the combustion of methane in the cylinder (Molar gas volume = 24.0dm^3 at room temperature and pressure)

19. 0.84g of aluminium were reacted completely with chlorine gas. Calculate the volume of chlorine gas used. (Molar gas volume is 24dm^3, Al = 27)

20. 3.52g of Carbon (IV) Oxide and 1.40g of water are produced when a mass of a hydrocarbon is completely burnt in oxygen. Determine the empirical formula of the hydrocarbon; (H = 1 , C= 12, O = 16)

21. Calculate the number of water molecules when 34.8g $Na_2CO_3.xH2O$ is heated and 15.9g of anhydrous Na_2CO_3 obtained (H=1, O=16, Na= 23, C = 12)

22. A weighed sample of crystallined sodium carbonate ($Na_2CO_3.nH2O$) was heated in a crucible until there was no further change in mass. The mass of the sample reduced by 14.5%. Calculate the number of moles (n) of water of crystallization (Na = 23, O = 16, C = 12, H = 1)

23. In a reaction 20cm^3 of 0.1 M Sodium Carbonate completely reacted with 13cm^3 of dilute sulphuric acid. Find the molarity of the sulphuric acid used.

24. An organic compound P contains 68.9% carbon, 13.5% hydrogen and 21.6% oxygen. The relative formula mass of p is 74. Determine its molecular formula. [C=12, H=1, O=16]

25. Campers GAZ cylinder contains about 1.12dm^3 of butane measured at 0^o and 1atm. Given that 25% of heat is lost, what is the maximum volume of water at room temperature which can be boiled to 100°C in order to make some coffee?
$C_4H_{10(g)} + 6½O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(l)}$; $\Delta H^\theta = -3{,}000KJmol^{-1}$
(Specific heat capacity of water = 4.2J g^{-1}°C, density of water 1gcm^{-3} Molar gas volume 22.41dm^3 at s.t.p)

26. An aqueous solution containing anhydrous sodium carbonate was prepared by dissolving 19.6g of the salt in 250cm^3 of distilled. Calculate the volume of 2M of magnesium chloride solution required to precipitate all the carbonate ions in the solution. (Na=23, C= 12; O = 16; Mg = 24; Cl =35.5)

27. 10.08g of ethanedioic acid ($H_2C_2O_4.xH_2O$) crystals were dissolved in water and made to 1dm^3 solution. 25.0 cm^3 of this solution was completely neutralized by 20cm^3 of 0.2M sodium hydroxide solution.
Calculate

    i. Molarity of the acid

    ii. the value of x in $H_2C_2O_4.xH_2O$ acid

28. 1.6g of magnesium metal is reacted with excess hydrochloric acid. Calculate the volume of hydrogen gas produced (Molar gas volume at stp = 22.4dm^3 Mg=24)

29. 60 litres of sulphur(IV) oxide were made to react with 40 litres of oxygen.

i. Which reactant was in excess and by how much?

ii. What is the volume of the product?

30. During welding of cracked railway lines by thermite 12.0g of oxide of iron is reduced by aluminium to 8.40g of iron. Determine the empirical formula of the oxide (Fe= 56.0, O= 16.0)

31. 0.84g of aluminium reacted completely with chlorine gas. Calculate the volume of chlorine gas used (Molar gas volume is $24dm^3$, Al = 27)

Answers

1. When a magnesium ribbon is heated in air it combines with oxygen forming magnesium oxide. When potassium manganate (VII) is heated it decomposes giving off oxygen which escapes in air

2. RFM of NaOH = 40
   Moles of NaOH = $^{8}/_{40}$ = 0.2M ✓
   Moles of NaOH in $25cm^3$
   $\frac{25 \times 0.2}{1000} = 0.005$
   Mole ratio 1:2
   Moles of acid = $^{0.005}/_{2}$
   = 0.0025
   $\frac{1 \times 0.245}{0.0025} = 98$

3. No. Of moles of HNO_3 acid
   $\frac{50 \times 2}{1000} = 0.1$moles
   Mole ratio 1:1
   The KOH will have 0.1moles; $\frac{0.1 \times 100}{50} = 0.2$moles
   Then D grams is 0.2 × 56
   = 11.2g

4. Number of moles of Q = $\frac{960\ cm^3 \times 1\ mole}{24000\ cm^3}$
   = 0.04moles
   Equation:
   $Na_2SO_{3(s)} + 2HCl(aq) \rightarrow 2NaCl_{(aq)} + SO_{2(g)} + H_2O_{(l)}$
   Mole ratio $Na_2SO_3 : SO_2$ is 1:1
   ∴No. of moles of Na_2SO_3 = 0.04moles
   Mass of $Na_2SO_3 = 126gmol^{-1} \times 0.04$
   = 5.04g

5. From the equation
   - (3x24) litres of chlorine react with iron to produce [(56 x 2) + (35.5 X3)] g of $FeCl_3$.
   325 g of Fecl3 is produced by 72 litres of Cl_2
   Then 0.5g of $Fecl_3$ is produced by:
   $\frac{0.5 \times 72}{325}$ =0.11078 litres
   = 110.78 cm^3

6. RMM (CH_3OOH) = 60
   Mass of $15cm^3$ and = 1.05 x 15 = 15.75g
   Moles in $500cm^3$ solution = $\frac{15.75}{60} = 0.2625$
   Molarity = $\frac{1000 \times 0.2625}{5000}$
   = 0.525M

7. If 24000 cm^3 = 1mole
   $150cm^3$ = ?
   $\frac{150 \times 1}{24000}$
   = 0.00625 moles of CO_2
   Since the ratio of Na_2CO_3; O_2 produced is 1:1 the mass of Na_2CO_3 = 0.00625 x 106 = 0.6625g

Na_2CO_3	H_2O
Mass 0.6625g RFM 106 Mole $^{0.6625}/_{106} = 0.00625$ Ratio $^{0.00625}/_{0.00625}$ $= 1$ $Na_2CO_3.9H_2O$	1.0125g 18 $^{1.0125}/_{18} = 0.5625$ $^{0.05625}/_{0.00625}$ $= 9$

8. $MgCl_2 \rightarrow Mg^{2+}(s) + 2Cl$
R.F.M of $MgCl_2 = 24 + 71$
$= 95$
Moles of Mass $= \frac{1.7}{95}$
R.F.M
$= 0.01789$moles
1 mole of $MgCl_2$ = 2moles of Cl^- ions
0.01789moles of $MgCl_2 = 0.01789 \times 2$
$= 0.03478$ moles of Cl^- ions
1mole = 6.0×10^{23} ions
0.03578moles = $\frac{0.03578 \times 6.0 \times 10^{23}}{1}$
$= 2.1468 \times 10^{22}$ ions of Cl^-

9. Mass of $O_2 = (4.0 - 2.4) = 1.6g$
Moles of $O_2 = {}^{1.6}/_{16} = 0.1$
If 1 mol O_2 _____ $24000cm^3$
0.1 Mol Mg = 0.5 mol $O_2 = 1200cm^3$
OR
2mg : O_2
2(24) 24000
$^{2.4}/_{2(24)} = {}^{x}/240000$
$x = \frac{2.4 \times 24000}{2(2.4)} = 1200cm^3$

10.

i.

Fe	**S**	**O**	**H_2O**
$^{20.2}/_{56}$	$^{11.5}/_{32}$	$^{23.0}/_{16}$	$^{45.3}/_{18}$
$^{0.36}/_{0.36}$	$^{0.36}/_{0.36}$	$^{1.44}/_{0.36}$	$^{2.52}/_{0.36}$
1	**1**	**4**	**7**

Empirical formula: $FeSO_4.7H_2O$

ii. 6.95g = $^{6.95}/_{278} = 0.025$
∴ 0.05 moles in $250cm^3 = 0.025 \times {}^{1000}/_{250} = 0.1$

11.

i. R.F.M of $PbI_2 = 207 + (127 \times 2) = 461$
2 moles of I^- ions produces 1 mole of PbI_2
Moles of I-ions $= \frac{0.1 \times 300}{1000} = 0.03$ mole
Mole ratio $PbI_2 : I^-$ mole of PbI_2 formed = $^{0.03}/_{2} = 0.05$
1 : 2
Mass of PbI_2 formed = 0.015 mole × 461
= 6.915 g

ii. Yellow precipitateMocks Topical Analysis www.easyelimu.com 20

12.

i.

i.

ii. At 25° C, sodium chloride is in solid form. Ions cannot move. Between 801°C and 1413°C sodium chloride is in liquid state, ions are mobile

ii. Both ammonia and water are polar moleculer and hydrogen bonds are formed

iii. N → H // co-ordinate bond / Dative bond

iv.

i. Allotrope

ii. Add methylbenzene to soot in a beaker. Shake and filter. Warm the filtrate to concentrate it. Allow the concentrate to cool for crystals to form. Filter to obtain crystals of fullerene

iii. $^{720}/_{12} = 60$

13. Mass of O_2 = (4.0 – 2.4)= 1.6g
Moles of O_2 = $^{1.6}/_{16} = 0.1$
If 1 mol $O_2 \rightarrow 24000cm^3$
0.1 Mol Mg = 0.5 mol $O_2 = 1200cm^3$
OR
2mg : O_2
2(24) 24000
$^{2.4}/_{2(24)} = {}^{x}/240000$
$x = \frac{2.4 \times 24000}{2(2.4)} = 1200cm^3$

14.

i. C_nH_{2n}, where n = No. of carbon atoms

ii. 70

iii. C_5H_{10}, $CH_3CH{=}CHCH_2CH_3$
OR $CH_3CH_2CHCH_2 = CH_2$

15.

i. Zinc is more reactive// higher reduction potential than copper it will react with// get oxidized in preference to iron oxygen to form Zinc Oxide coat which protects iron from rusting

ii. Sacrificial protection or cathodic protection

16. $NH_{3(g)} + HNO_{3(g)} \rightarrow NH_4NO_{3(s)}$
RMM of NH_4NO_3 = 80
Moles of NH_4NO_3 = $^{4800}/_{80}$ = 60moles
RMM of NH_3 = 17
Mass of NH_3 = 60 x 17 = 1020KJ

17. From the equation of step 3
$SO_{3(g)} + H_2SO_{4(l)} \rightarrow H_2S_2O_{7(l)}$
RFM of $H_2S_2O_7$ = 2 + (2 × 32) + (7 × 16) = 178 ✓ ½ mark
178g of Oleum are produced by 22.4 liters of SO3 ✓ ½ mark
178 kg of Oleum will produce
$\frac{178 \times 1000 \times 22.4}{}$l ✓ 1½ mark

178g
= 22,4000 liters ✓ ½ mark
(Total 13 marks)

18.

i. $\frac{35.2 \times 1000}{100 \times 16}$
= 10Moles
Or mass of CH4 = $\frac{35.2 \times 5}{1000}$ = 1.76g
Mass in g = 1.76 x 1000 = 1760kg
Moles of methane = $^{1760}/_{16}$
= 110 Moles

ii. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ – (ignore states)
Volume = 110 x 24.0
= 2640 dm^3
Mark consequential from equation and b(ii) (Without equation max *TZM*)

19. Volume of Cl_2 used
= 0.047 x 24
= 1.128dm^3

20. Mass due Carbon in CO2 = $^{12}/_{4}$ x 35.2
= 0.96
Moles carbon = $^{0.96}/_{12}$ = 0.08
Mass due Hydrogen in H_2O = $^{2}/_{18}$ x 1.40
= 0.156
Moles hydrogen = $^{0.156}/_{1}$ = 0.156
Mole ratio C:H = 1:1.95
E.F = CH_2

21. $Na_2CO_{3.x}H2O \rightarrow Na_2CO_3 + H_2O$ √1

34.8g	$\frac{15.9g}{106}$	$\frac{18.9g}{18}$
	$\frac{0.15}{0.15}$ √1	$\frac{1.15}{0.15}$

x = 7 √1
Total 3 marks

22. % of H_2O lost = 14.5%^
5 of anhydrous Na_2CO_3 = 85.5% (½mk)
R.F.M of Na_2CO_3 = 106 (½mk)
RMM of H_2O = 18 (½mk)

Na_2CO_3	**H_2O**
$\frac{85.5}{106}$	$\frac{14.5}{18}$
$\frac{0.8066}{0.8055}$	$\frac{0.8066}{0.8055}$

23. n = 1 ($Na_2CO_3.H_2O$) (½mk)

24. Moles of Na_2CO_3 = $\frac{20 \times 0.1}{1000}$ = 0.002 moles
$Na_2CO_3 + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

Mole ratio 1 : 1
Moles of H_2SO_4 = Moles of Na_2CO_3
= 0.002 moles
Molarity of H2SO4 = $\frac{10000 \times 0.002}{13}$ = 0.154 moles

25.

Element	C	H	O
%	68.9	13.5	21.6
Molar mass	12	1	16
Moles	$^{68.9}/_{12}$ 5.403	$^{13.5}/_{1}$ 13.5	$^{216}/_{16}$ 1.35
MR	$^{5.403}/_{1.33}$ 4	$^{13.5}/_{1.35}$ 10	$^{1.35}/_{1.35}$ 1
Ratio	4	10	1

26.
h ($C_4H_{10}O$) = 74
h (12 x 4) + (10 x 1) +16 = 74
74h = 74
h= 1
Formula $C_4H_{10}O$

27. Moles C_4H_{10} = $^{1.12}/_{22.4}$ = 0.05 mol
Heat produced + 0.05 × (3000) = 150 kJ
Usefull heat = $\frac{75 \times 150}{100}$ = 112.5 kJ
Let volume of water = V
Room temperature = 25°C
Boiling point = 100°C
Change in temperature, ΔT = 100 − 25 = 75°C ½ mk
Q = ΔT × mass × C
= 75 × V × 4.2 =112500
315V = 112500
V = $^{112500}/_{315}$ ½ mk
V = 357cm^3 ½ mk

28. RFM Na_2CO_3 = 43 + 12 + 48 = 106
Mol. Na_2CO_3 = $^{19.6}/_{106}$ = 0.8149057
Molarity of Na_2CO_3 = $^{0.1849057}/_{0.25}$ = 0.73962M
$Na_2CO_{3(aq)} + MgCl_{2(aq)} \rightarrow MgCO_{3(s)} + 2NaCl_{(aq)}$
Mole ratio $NaCO_3$: $MgCl_2$ is 1:1
∴ mol. $MgCl_2$ Reacted = 0.1849
If 2.0 mol. = 1000cm^3 solution $MgCl_2$
= 0.1849mol = $\frac{0.1849 \times 1000}{2}$
= 92.45 or 92.5 cm^3

29.

i. ACID BASE
1 2

½ ×0.004 $\quad\quad$ $\underline{20cm^3}$ × 0.2 moles
= 0.002 moles ✓½ $\quad$ $1000cm^3$ = 0.004 moles
$25cm^3$ _____ 0.002 moles ✓ ½
$1000cm^3$ _____ ?
$1000cm^3$ × 0.002 moles = 0.08M✓ ½

ii. 0.08 moles _____ 10.08g $H_2C_2O_4.xH2O$ ✓ ½
1 mole _____ ?
$\underline{1\ mole}$ × 10.08 = 126 ✓ ½
0.08 moles
126 _____ $H_2C_2O_4.xH_2O$
18x = 126 – 90 ✓ ½
18x = 36
x = 2 ✓ ½

30. $Mg_{(g)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
24g _____$22.4dm^3$
16g _____?
$\underline{1.6g}$ x $22.4dm^3$ ✓ ½ = 1.4933 dm^3
24g

31.

i. $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$, $SO_2 : O_2$
2 $\quad$ 1 $\quad$ 2 $\quad$ 60 : 30 ✓½
60l $\quad$ 40l
Oxygen ✓½ by 10 litres

ii. 60 litres

32. Mass of Oxygen = 12 – 8.4 = 3.5g

Element	Fe	O
Mass	8.4	3.6
RAM	56	16
No. of moles	$^{8.4}/_{56}$ 0.15	$^{3.6}/_{16}$ 0.225
Mole ration	$^{0.15}/_{0.15}$ 1 2	$^{0.225}/_{0.15}$ 1.5 × 2 3

33.
The empirical formula is Fe_2O_3

34. 2AL $3CL_2$ = $2ALCL_3$
Mole of alumminium $\underline{0.84}$ = 0.0311.bt accrdin 2 d equation 2 mole of AL react with 3mole of Cl.
27
1mole of Al required 1.5mole of cl
0.0311mole of Al required 0.0311 ×1.5mole of Cl = 0.04665mole.
1 mole of a gas occupied $22.4dm^3$,0.04665 mole will occupy 0.04665 × 22.4 = 1.0450dm^3 of Cl

STRAND 1.0: INORGANIC CHEMISTRY

SUB-STRAND 1.2: NON-METALS

1.0 Introduction to non-metals

Non-metals are elements that generally do not possess the physical and chemical properties associated with metals. They are found on the right-hand side of the Periodic Table, except hydrogen, which is placed on the left-hand side.

Examples of non-metals include:

- Hydrogen
- Oxygen
- Nitrogen
- Carbon
- Sulphur
- Chlorine
- Phosphorus

In this sub-strand, we focus mainly on hydrogen, oxygen and nitrogen.

1.1 General characteristics of non-metals

Most non-metals have the following characteristics:

1. **Poor conductors of heat and electricity:** Most non-metals do not conduct heat and electricity effectively. Graphite is an exception because it conducts electricity.
2. **Low melting and boiling points:** Many non-metals have relatively low melting and boiling points compared with metals.
3. **Dull appearance:** Most solid non-metals are not shiny.
4. **Brittleness:** Solid non-metals are generally brittle and break easily when hammered.
5. **Formation of acidic oxides:** Many non-metals react with oxygen to form acidic oxides.
6. **Formation of covalent compounds:** Non-metals commonly combine by sharing electrons.
7. **Different physical states:** Non-metals may exist as solids, liquids or gases at room temperature.

1.2 Importance of non-metals

Non-metals are important in everyday life because they:

1. Support respiration in living organisms.
2. Are used in the manufacture of fertilisers.
3. Are used in the production of fuels.
4. Support industrial manufacturing processes.
5. Are used in water treatment.
6. Are essential in food production and preservation.
7. Are used in medical and scientific applications.
8. Form important compounds found in living organisms and the environment.

2.0 SOURCES OF SELECTED NON-METALS

Elements such as hydrogen, oxygen and nitrogen occur naturally in the environment. They can be obtained from different natural sources and industrial processes.

2.1 Sources of hydrogen

Hydrogen is the lightest chemical element. It is a colourless, odourless gas that exists naturally in combined form in many substances.

Natural sources of hydrogen

1. **Water**
   - ✓ Water contains hydrogen chemically combined with oxygen.
   - ✓ Its chemical formula is H_2O.
   - ✓ Hydrogen can be obtained from water through electrolysis.
2. **Natural gas**
   - ✓ Natural gas contains methane, CH_4, which is a major source of hydrogen for industrial production.
   - ✓ Hydrogen is obtained through processes such as steam reforming.
3. **Petroleum**
   - ✓ Petroleum contains hydrocarbons made up of hydrogen and carbon.
   - ✓ Hydrogen can be produced during petroleum refining.
4. **Organic matter**
   - ✓ Plants and animals contain hydrogen in substances such as carbohydrates, proteins and fats.
   - ✓ Hydrogen is released through suitable chemical and biological processes.
5. **Acids**
   - ✓ Hydrogen can be produced when certain metals react with dilute acids.
   - ✓ For example, zinc reacts with dilute hydrochloric acid to produce hydrogen gas.

Industrial sources of hydrogen

Hydrogen is produced industrially through:

- Steam reforming of natural gas.
- Electrolysis of water.
- Gasification of carbon-containing materials.
- Other industrial chemical processes.

Example: Electrolysis of water

When an electric current is passed through acidified water, hydrogen and oxygen gases are produced.

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

Hydrogen is produced at the cathode (negative electrode), while oxygen is produced at the anode (positive electrode).

2.2 Sources of oxygen

Oxygen is an essential non-metal that occurs naturally in the atmosphere, water and rocks.

Natural sources of oxygen

1. **Atmosphere**
   - ✓ Oxygen makes up approximately 21% of the volume of dry air.
   - ✓ It is available for respiration and combustion.
2. **Water**
   - ✓ Water contains oxygen chemically combined with hydrogen.
   - ✓ Its chemical formula is H_2O.
   - ✓ Oxygen can be obtained through electrolysis.

3. **Photosynthesis**
    - ✓ Green plants release oxygen during photosynthesis.
    - ✓ They use sunlight to manufacture food from carbon dioxide and water.
    - ✓ Oxygen is released as a by-product.
4. **Oxides**
    - ✓ Oxygen occurs in many metal and non-metal oxides.
    - ✓ Examples include iron(III) oxide, calcium oxide and carbon dioxide.
5. **Minerals and rocks**
    - ✓ Oxygen is chemically combined in minerals such as silicates and carbonates.
    - ✓ Many rocks in the Earth's crust contain oxygen.

Industrial sources of oxygen

Oxygen is mainly obtained industrially through fractional distillation of liquid air.

It can also be produced through electrolysis of water and certain industrial chemical processes.

2.3 Sources of nitrogen

Nitrogen is a colourless gas that forms the largest proportion of the Earth's atmosphere.

Natural sources of nitrogen

1. **Atmosphere**
    - ✓ Nitrogen makes up approximately 78% of dry air by volume.
    - ✓ It exists mainly as diatomic nitrogen molecules, N_2.
2. **Soil**
    - ✓ Nitrogen occurs in soil in compounds such as nitrates and ammonium salts.
    - ✓ These compounds are important for plant growth.
3. **Organic matter**
    - ✓ Nitrogen occurs in proteins and other nitrogen-containing substances in plants and animals.
4. **Fertilisers**
    - ✓ Nitrogen is present in fertilisers such as ammonium nitrate, ammonium sulphate and urea.
5. **Biological nitrogen fixation**
    - ✓ Certain bacteria convert atmospheric nitrogen into nitrogen-containing compounds that plants can use.
    - ✓ Some nitrogen-fixing bacteria live in root nodules of leguminous plants such as beans, peas and groundnuts.

Industrial sources of nitrogen

Nitrogen is mainly obtained through fractional distillation of liquid air.

It can also be produced by removing oxygen and other gases from air using appropriate chemical or physical processes.

3.0 GENERAL METHODS OF GAS COLLECTION

Gases are collected in the laboratory using methods that depend on their physical properties.

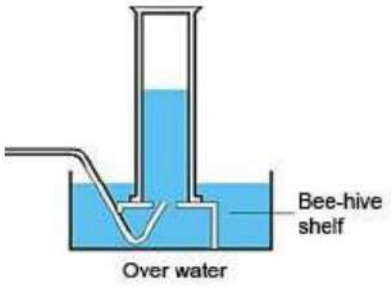
The main factors considered when selecting a method of gas collection include:

1. Solubility of the gas in water.
2. Density of the gas compared with air.
3. Chemical reactivity of the gas.

4. Whether the gas reacts with water.
5. Whether the gas reacts with substances in the collection apparatus.

3.1 Collection over water

This method is also called **downward displacement of water**.



It is used to collect gases that are insoluble or only slightly soluble in water and do not react with water.

Principle

The gas produced is passed through a delivery tube into an inverted gas jar filled with water.

The gas enters the jar and pushes water out. As more gas enters, the water level falls until the jar is filled with gas.

Examples of gases collected over water

- Hydrogen
- Oxygen
- Nitrogen

Advantages

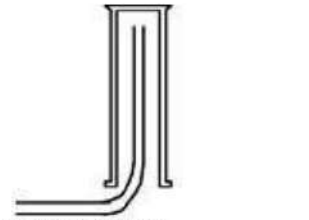
1. It is a simple method of collecting gases.
2. It prevents contamination of the gas by air when properly carried out.
3. It allows the volume of gas collected to be observed.
4. It is suitable for gases that are not very soluble in water.

Limitations

1. It cannot be used for gases that are highly soluble in water.
2. It cannot be used for gases that react with water.
3. The collected gas contains water vapour and is therefore moist.

3.2 Upward delivery

This method is also called **downward displacement of air**.



Upward delivery
(downward displacement of air)

It is used to collect gases that are denser than air.

Principle

The gas is introduced into an upright gas jar. Because the gas is denser than air, it settles at the bottom and pushes air upwards and out of the jar.

Examples

- Carbon dioxide
- Chlorine
- Sulphur dioxide

Advantages

1. It is useful for collecting gases that are highly soluble in water.
2. It does not require the gas to pass through water.
3. It is suitable for gases denser than air.

Limitations

1. Some air may remain mixed with the collected gas.
2. It is not suitable for gases lighter than air.
3. It may expose the surroundings to harmful gases if proper precautions are not taken.

3.3 Downward delivery

This method is also called **upward displacement of air**.



Downward delivery
(upward displacement of air)

It is used to collect gases that are less dense than air.

Principle

The gas is introduced into an inverted gas jar. Because it is less dense than air, it rises and pushes air downwards and out of the jar.

Examples

- Ammonia

- Hydrogen, where suitable apparatus is used

Advantages

1. It is useful for gases that are highly soluble in water.
2. It does not require water as a collecting medium.
3. It is suitable for gases lighter than air.

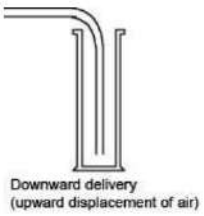
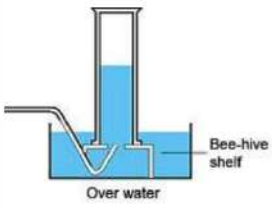
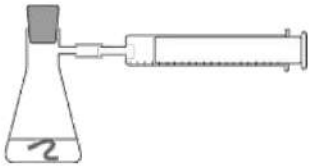
Limitations

1. It is not suitable for gases denser than air.
2. The gas may mix with air if the apparatus is not properly arranged.

3.4 Comparison of gas collection methods

Method	Principle	Suitable gases
Collection over water	Gas displaces water	Hydrogen, oxygen, nitrogen
Upward delivery	Dense gas pushes air upwards	Carbon dioxide, chlorine
Downward delivery	Light gas pushes air downwards	Ammonia

Important: The choice of collection method depends on the gas's solubility, density and chemical behaviour. Density alone should not be used to choose a collection method when the gas reacts with water.

Method	Upward displacement of air	Downward displacement of air	Over water	Gas syringe
Use when ...	gas is denser than air	gas is less dense than air	gas is insoluble or only slightly soluble in water	you want to measure the volume
Apparatus				
Gases collected by this method	Carbon dioxide (CO_2) Chlorine (Cl_2)	Hydrogen (H_2) Ammonia (NH_3)	Hydrogen (H_2) Oxygen (O_2) Carbon dioxide (CO_2)	All gases

4.0 DRYING OF GASES

A gas collected over water usually contains water vapour. Such a gas is described as a moist gas.

Drying is the process of removing water vapour from a gas without reacting with the gas itself.

A drying agent is a substance that absorbs or removes moisture from a gas.

4.1 Common drying agents

1. **Anhydrous calcium chloride**
   - ✓ Absorbs water vapour.
   - ✓ Can be used to dry gases such as hydrogen and oxygen under suitable conditions.
   - ✓ It should not be used to dry ammonia because it absorbs ammonia.

2. **Concentrated sulphuric acid**
   - ✓ Absorbs water vapour.
   - ✓ It can be used to dry gases such as oxygen and hydrogen.
   - ✓ It cannot be used for ammonia because ammonia reacts with sulphuric acid.
3. **Quicklime (calcium oxide)**
   - ✓ Removes moisture.
   - ✓ It can be used to dry ammonia.
   - ✓ It should not be used for acidic gases such as carbon dioxide and sulphur dioxide because they react with it.
4. **Anhydrous copper(II) sulphate**
   - ✓ Absorbs water and changes from white to blue.
   - ✓ It can be used to demonstrate the presence of water.

4.2 Importance of drying gases

1. To obtain pure, dry gases for experiments.
2. To prevent water vapour from interfering with chemical reactions.
3. To improve accuracy in gas volume measurements.
4. To prevent unwanted reactions involving moisture.
5. To obtain gases suitable for industrial processes.

5.0 LABORATORY PREPARATION OF OXYGEN

Oxygen can be prepared in the laboratory by decomposing hydrogen peroxide using manganese(IV) oxide as a catalyst.

5.1 Preparation using hydrogen peroxide

Requirements

- Hydrogen peroxide solution.
- Manganese(IV) oxide powder.
- Conical flask.
- Two-holed stopper.
- Delivery tube.
- Water trough.
- Gas jar.
- Beehive shelf.

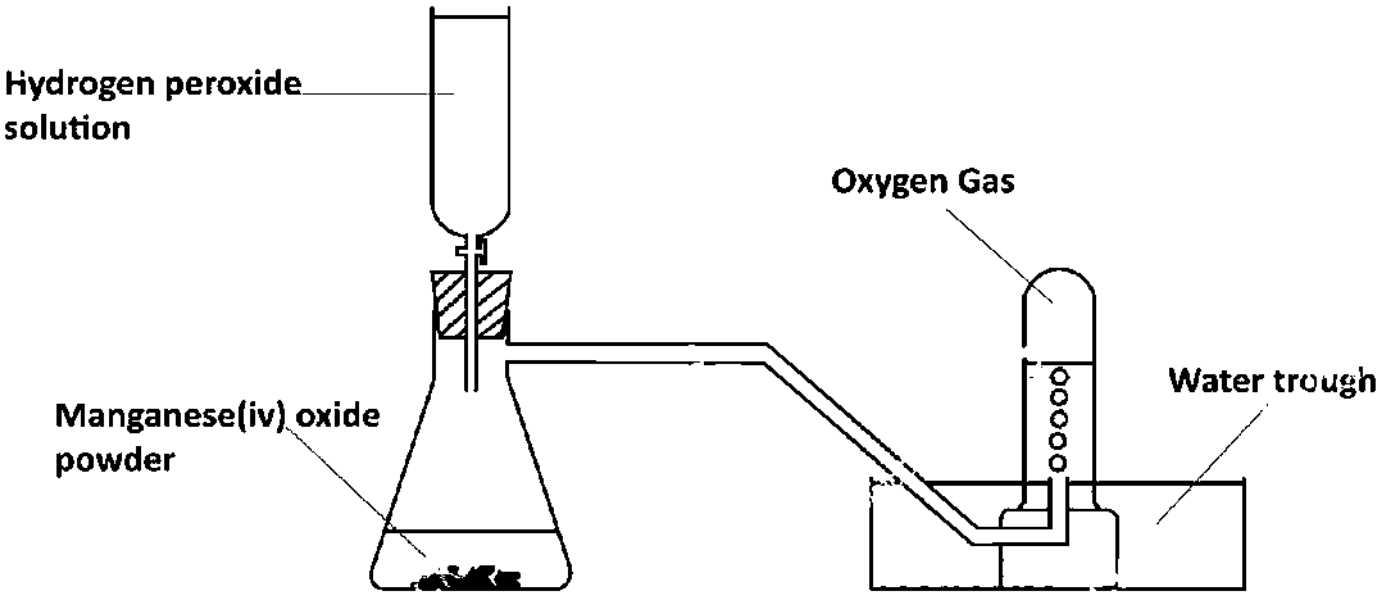
Procedure

1. Place a small quantity of manganese(IV) oxide in a conical flask.
2. Fit the flask with a stopper containing a delivery tube.
3. Connect the delivery tube to a water trough.
4. Fill a gas jar with water and invert it over a beehive shelf in the trough.
5. Add hydrogen peroxide solution carefully into the flask.
6. Observe the production of bubbles.
7. Allow the gas to pass through the delivery tube into the inverted gas jar.
8. Collect the oxygen by downward displacement of water.
9. Test the collected gas using a glowing wooden splint.

Observation

- Bubbles of gas are produced.
- A colourless gas collects in the gas jar.
- The glowing wooden splint relights when inserted into the gas jar.

Explanation



Equation

$$\text{Hydrogen peroxide} \xrightarrow[\text{Catalyst}]{MnO_{2(s)}} \text{Water + Oxygen}$$

$$2H_2O_{2(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)}$$

Manganese(IV) oxide acts as a catalyst. It increases the rate of decomposition of hydrogen peroxide without being used up in the reaction.

Chemical equation

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

Role of manganese(IV) oxide

1. It speeds up the decomposition of hydrogen peroxide.
2. It provides an alternative reaction pathway with lower activation energy.
3. It remains chemically unchanged at the end of the reaction.

Test for oxygen

A glowing wooden splint is inserted into a gas jar containing the gas.

Positive result: The glowing splint relights.

Conclusion: The gas is oxygen.

Precautions

1. Handle hydrogen peroxide carefully because concentrated solutions can cause burns.
2. Avoid inhaling manganese(IV) oxide dust.
3. Ensure the apparatus is airtight.
4. Do not allow the delivery tube to become blocked.
5. Keep the apparatus away from unnecessary sources of heat.

5.2 Alternative preparation by heating potassium chlorate

Oxygen can also be prepared by heating potassium chlorate in the presence of manganese(IV) oxide as a catalyst.

Chemical equation

$$2KClO_3(s) \overset{\Delta, MnO_2}{\rightarrow} 2KCl(s) + 3O_2(g)$$

Explanation

1. Potassium chlorate decomposes when heated.
2. Potassium chloride is formed as a solid product.
3. Oxygen gas is released.
4. Manganese(IV) oxide increases the rate of decomposition.

Safety: Heating potassium chlorate requires careful supervision because it is a strong oxidising agent. It should not be mixed with combustible materials.

6.0 LABORATORY PREPARATION OF HYDROGEN

Hydrogen can be prepared in the laboratory by reacting a moderately reactive metal, such as zinc, with a dilute acid.

6.1 Preparation using zinc and dilute hydrochloric acid

Requirements

- Zinc granules.
- Dilute hydrochloric acid.
- Conical flask.
- Thistle funnel.
- Two-holed stopper.
- Delivery tube.
- Water trough.
- Gas jar.
- Beehive shelf.

Procedure

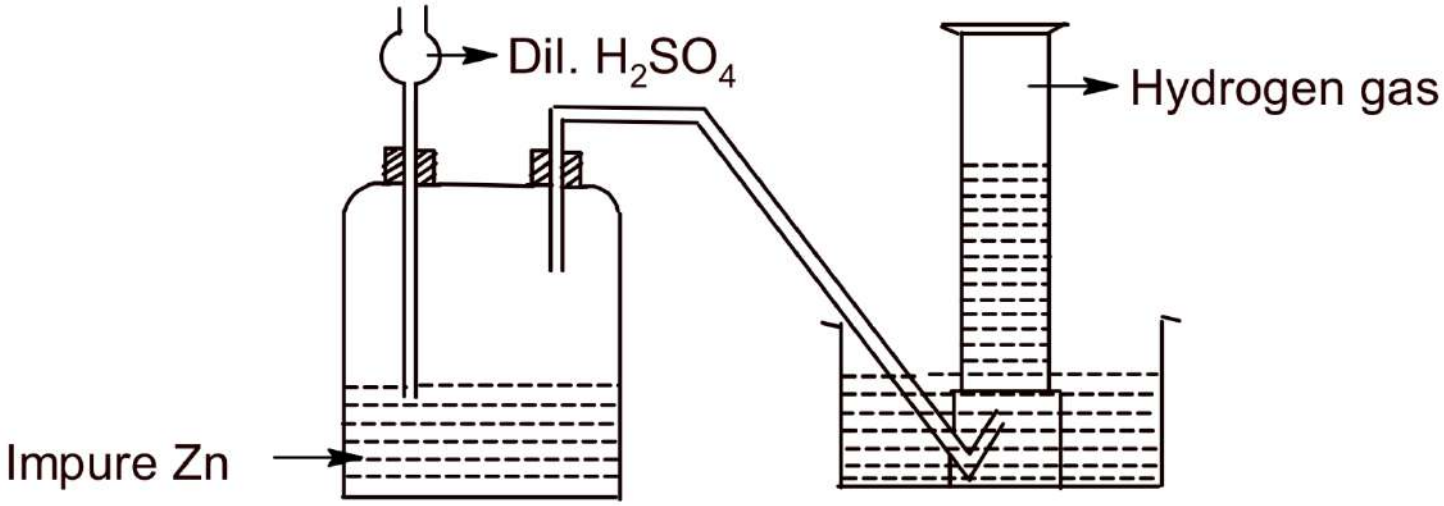
1. Place zinc granules in a conical flask.
2. Fit the flask with a stopper containing a thistle funnel and delivery tube.
3. Ensure the end of the thistle funnel is below the level of the acid that will be added.
4. Connect the delivery tube to a water trough.
5. Fill a gas jar with water and invert it over the beehive shelf.
6. Add dilute hydrochloric acid through the thistle funnel.
7. Observe bubbles forming around the zinc granules.
8. Allow the first portion of gas to escape to remove air from the apparatus.
9. Collect the hydrogen over water.
10. Test a small sample of the gas using a burning splint.

Observation

- Bubbles of gas are produced.
- Zinc gradually dissolves.
- A colourless gas collects in the gas jar.
- The gas burns with a characteristic squeaky pop when a small sample is tested.

Chemical equation

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$



Explanation

1. Zinc reacts with dilute hydrochloric acid.
2. Zinc displaces hydrogen from the acid.
3. Zinc chloride is formed.
4. Hydrogen gas is released.

6.2 Preparation using zinc and dilute sulphuric acid

Zinc also reacts with dilute sulphuric acid to produce hydrogen.

Chemical equation

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

Observations

- Effervescence occurs.
- Zinc gradually dissolves.
- Hydrogen gas is produced.

6.3 Test for hydrogen

A small sample of the gas is tested using a burning splint.

Positive result: The gas burns with a squeaky pop sound.

Conclusion: Hydrogen is present.

The squeaky pop is caused by rapid combustion of hydrogen with oxygen in the air.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Precautions

1. Hydrogen is highly flammable.
2. Keep flames away from the gas-generating apparatus.
3. Allow the first gas produced to escape because it may contain air.
4. Test only a small sample of the gas.

5. Never bring a burning splint close to the gas generator.
6. Use dilute acid carefully.

7.0 LABORATORY PREPARATION OF NITROGEN

Nitrogen can be prepared in the laboratory by producing ammonium nitrite in solution and gently warming it.

Ammonium nitrite is unstable and decomposes to form nitrogen gas and water.

7.1 Preparation using ammonium nitrite

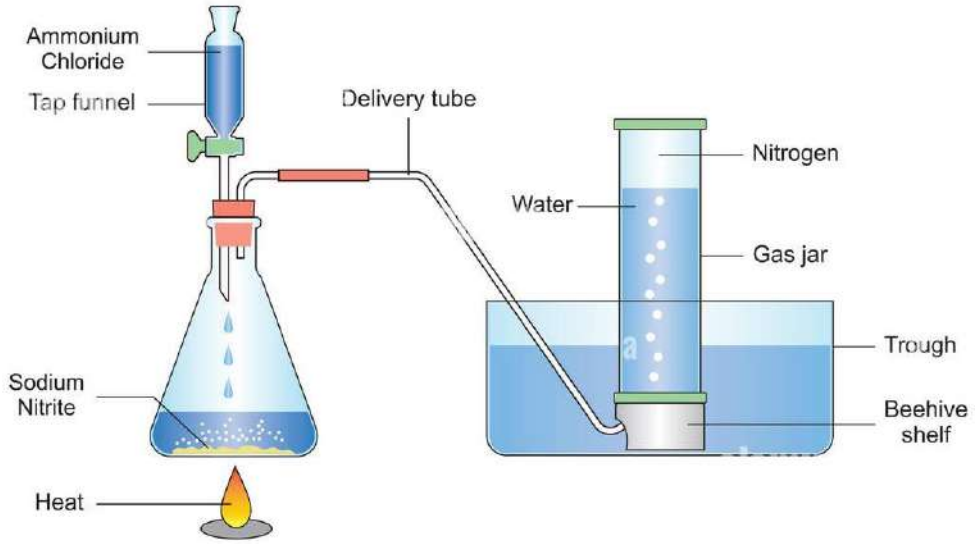
Requirements

- Sodium nitrite solution.
- Ammonium chloride solution.
- Conical flask.
- Delivery tube.
- Heat source.
- Water trough.
- Gas jar.
- Beehive shelf.

Procedure

1. Place sodium nitrite solution in a suitable flask.
2. Add ammonium chloride solution.
3. Fit the flask with a stopper and delivery tube.
4. Connect the delivery tube to a water trough.
5. Fill a gas jar with water and invert it over a beehive shelf.
6. Warm the mixture gently under teacher supervision.
7. Allow the gas produced to pass through the delivery tube.
8. Collect the nitrogen over water.

Chemical reactions



First, ammonium nitrite is formed:

$$NaNO_2 + NH_4Cl \rightarrow NH_4NO_2 + NaCl$$

Ammonium nitrite then decomposes:

$$NH_4NO_2 \rightarrow N_2 + 2H_2O$$

Overall equation

$$NaNO_2 + NH_4Cl \rightarrow N_2 + NaCl + 2H_2O$$

Observations

1. Bubbles of gas are produced.
2. A colourless gas collects in the gas jar.
3. The gas does not support combustion.
4. A burning splint placed in the gas is extinguished.

Test for nitrogen

Introduce a burning or glowing splint into a gas jar containing the gas.

Observation: The flame is extinguished.

Conclusion: The gas does not support combustion. This observation is consistent with nitrogen, but it is not by itself a unique test because other gases can also extinguish a flame.

Precautions

1. Warm the mixture gently.
2. Avoid excessive heating.
3. Wear protective goggles and laboratory coats.
4. Carry out the preparation under teacher supervision.
5. Ensure the apparatus is properly fitted.

7.2 Preparation of nitrogen by removing oxygen from air

Nitrogen can also be obtained by passing air over heated copper turnings.

Procedure

1. Place copper turnings in a hard-glass tube.
2. Connect the tube to a source of air.
3. Heat the copper turnings strongly.
4. Pass air slowly over the heated copper.
5. Allow oxygen in the air to react with copper.
6. Collect the gas that remains.

Chemical equation

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Explanation

- Oxygen in the air reacts with hot copper to form black copper(II) oxide.
- Most of the nitrogen passes through without reacting.
- The remaining gas is mainly nitrogen, together with argon and small amounts of other gases.

Important: This method does not produce completely pure nitrogen because air contains other gases besides oxygen and nitrogen.

8.0 INDUSTRIAL ISOLATION OF OXYGEN AND NITROGEN FROM LIQUID AIR

Oxygen and nitrogen are obtained industrially through a process known as **fractional distillation of liquid air**.

Air is a mixture of gases. The main components of dry air are approximately:

Gas	Percentage by volume
Nitrogen	78%
Oxygen	21%
Argon	0.93%
Carbon dioxide	About 0.04%
Other gases	Trace amounts

Fractional distillation separates components of a mixture based on differences in their boiling points.

8.1 Principle of fractional distillation of liquid air

Nitrogen and oxygen have different boiling points.

- Nitrogen boils at approximately −196°C.
- Oxygen boils at approximately −183°C.

Nitrogen has a lower boiling point than oxygen. Therefore, nitrogen is more volatile and boils off before oxygen.

8.2 Steps involved in industrial isolation

Step 1: Filtration of air

Atmospheric air is passed through filters to remove dust and other suspended particles.

Importance:

- Prevents blockage of equipment.
- Protects machinery from damage.
- Improves the quality of the final products.

Step 2: Removal of water vapour

Air is cooled or passed through suitable drying systems to remove moisture.

Importance:

- Prevents water from freezing inside pipes.
- Prevents blockage of equipment during cooling.

Step 3: Removal of carbon dioxide

Carbon dioxide is removed using suitable chemical or adsorption processes.

Importance:

- Prevents carbon dioxide from freezing and blocking the equipment.
- Improves the purity of the gases obtained.

Step 4: Compression of air

The purified air is compressed using compressors.

Compression increases the pressure and helps prepare the air for cooling.

Step 5: Cooling of compressed air

The compressed air is cooled in heat exchangers.

Further cooling is achieved through expansion and refrigeration processes until the air becomes liquid.

Step 6: Fractional distillation

Liquid air is introduced into a fractionating column.

As the liquid air is warmed:

1. Nitrogen, which has a lower boiling point, vaporises first.
2. Nitrogen rises towards the top of the column and is collected.
3. Oxygen remains in the liquid phase at a higher temperature than nitrogen.
4. Oxygen is collected from the lower part of the column.

Argon can also be separated as a useful by-product through additional separation stages.

8.3 Summary of fractional distillation

Property	Nitrogen	Oxygen
Formula	N_2	O_2
Approximate percentage in air	78%	21%
Boiling point	−196°C	−183°C
Relative volatility	More volatile	Less volatile
Position of collection	Near the top	Lower part of column

8.4 Advantages of fractional distillation of liquid air

1. It produces large quantities of oxygen and nitrogen.
2. It allows the separation of gases from atmospheric air.
3. It can produce gases of high purity.
4. It provides argon as a useful by-product.
5. It supports medical, agricultural and industrial activities.

8.5 Environmental and safety considerations

1. Air separation requires considerable energy.
2. Energy use may contribute to greenhouse gas emissions depending on the source of electricity.
3. Industrial plants should use energy-efficient equipment.
4. Liquid nitrogen and liquid oxygen are extremely cold and can cause severe cold burns.
5. Liquid oxygen supports vigorous combustion, so it must be kept away from oils, grease and combustible materials.
6. Nitrogen can cause suffocation by displacing oxygen in enclosed spaces.

9.0 PHYSICAL PROPERTIES OF OXYGEN, HYDROGEN AND NITROGEN

Physical properties are characteristics of substances that can be observed or measured without changing their chemical composition.

The physical properties of gases include:

- Colour.
- Odour.
- Density.
- Solubility in water.
- Effect on moist litmus paper.
- Behaviour towards burning or glowing splints.

9.1 Physical properties of oxygen

1. **Colour**
   - ✓ Oxygen is a colourless gas at room temperature.
   - ✓ It does not have a visible colour in ordinary gas jars.
2. **Odour**
   - ✓ Oxygen is odourless.
   - ✓ It has no characteristic smell.
3. **Solubility in water**
   - ✓ Oxygen is slightly soluble in water.
   - ✓ Dissolved oxygen is important for aquatic organisms such as fish.
4. **Effect on moist litmus paper**
   - ✓ Oxygen does not change the colour of moist blue or red litmus paper.
   - ✓ This shows that oxygen is neutral under ordinary conditions.
5. **Effect on a glowing wooden splint**
   - ✓ Oxygen relights a glowing wooden splint.
   - ✓ This is the standard laboratory test for oxygen.
6. **Density**
   - ✓ Oxygen is denser than air.
   - ✓ Its relative molecular mass is 32, while the average relative molecular mass of air is about 29.
7. **State**
   - ✓ Oxygen exists as a gas at room temperature.
   - ✓ Its molecules contain two oxygen atoms, O_2.

9.2 Physical properties of hydrogen

1. **Colour**
   - ✓ Hydrogen is a colourless gas.
2. **Odour**
   - ✓ Hydrogen is odourless.
3. **Solubility in water**
   - ✓ Hydrogen is very slightly soluble in water.
   - ✓ It can therefore be collected over water.
4. **Effect on moist litmus paper**
   - ✓ Hydrogen does not change the colour of moist red or blue litmus paper.
   - ✓ It is neutral under ordinary conditions.
5. **Effect on a burning splint**
   - ✓ Hydrogen burns with a pale blue flame in air.
   - ✓ A small sample produces a characteristic squeaky pop when tested with a burning splint.
6. **Density**
   - ✓ Hydrogen is the least dense gas.
   - ✓ It is much lighter than air.
7. **State**

- ✓ Hydrogen exists as a gas at room temperature.
- ✓ Its molecules contain two hydrogen atoms, H_2.

9.3 Physical properties of nitrogen

1. **Colour**
   - ✓ Nitrogen is a colourless gas.
2. **Odour**
   - ✓ Nitrogen is odourless.
3. **Solubility in water**
   - ✓ Nitrogen is only slightly soluble in water.
   - ✓ It can be collected over water.
4. **Effect on moist litmus paper**
   - ✓ Nitrogen does not change the colour of moist red or blue litmus paper.
   - ✓ It is neutral under ordinary conditions.
5. **Effect on a burning splint**
   - ✓ Nitrogen does not support combustion under ordinary conditions.
   - ✓ A burning splint is extinguished when introduced into a gas jar containing nitrogen.
6. **Density**
   - ✓ Nitrogen is slightly less dense than air.
   - ✓ Its relative molecular mass is 28.
7. **State**
   - ✓ Nitrogen exists as a gas at room temperature.
   - ✓ Its molecules contain two nitrogen atoms, N_2.

9.4 Comparison of physical properties

Property	Oxygen	Hydrogen	Nitrogen
Formula	O_2	H_2	N_2
Colour	Colourless	Colourless	Colourless
Odour	Odourless	Odourless	Odourless
Solubility in water	Slightly soluble	Very slightly soluble	Slightly soluble
Effect on litmus	Neutral	Neutral	Neutral
Glowing splint	Relights	Does not relight; burns if ignited	Extinguishes
Density compared with air	Denser	Much lighter	Slightly lighter
Combustion behaviour	Supports combustion	Burns readily	Does not support combustion

10.0 CHEMICAL PROPERTIES OF OXYGEN

Chemical properties describe how a substance reacts with other substances to form new products.

Oxygen is chemically reactive and supports combustion. It combines with many elements to form oxides.

10.1 Reaction of oxygen with metals

Oxygen reacts with many metals to form metal oxides.

This process is called oxidation.

A. Reaction with magnesium

When magnesium ribbon is heated in air or oxygen, it burns with a dazzling white flame.

A white powder of magnesium oxide is formed.

Chemical equation:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Observations:

1. Magnesium burns with a bright white flame.
2. A white solid is formed.
3. The product is magnesium oxide.

Importance: Magnesium is used in some pyrotechnic applications because it burns with a brilliant white light.

B. Reaction with copper

When copper is heated strongly in oxygen, it reacts to form black copper(II) oxide.

Chemical equation:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Observations:

1. The reddish-brown copper changes colour.
2. A black solid is formed.
3. The product is copper(II) oxide.

C. Reaction with iron

Iron reacts with oxygen, especially in the presence of moisture, to form rust.

Rust is mainly hydrated iron(III) oxide.

A simplified equation for iron reacting with oxygen is:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

Rusting is a slow process that requires oxygen and water.

Examples of rusting in Kenya:

- Iron roofing sheets.
- Metal gates.
- Water tanks.
- Agricultural tools.
- Iron fences.

10.2 Reaction of oxygen with carbon

Carbon burns in oxygen to produce carbon dioxide when oxygen is sufficient.

Chemical equation:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Observations:

1. Carbon glows or burns.
2. A colourless gas is produced.
3. The gas turns limewater milky.

Reaction in limited oxygen

When oxygen is insufficient, carbon may form carbon monoxide.

$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

Carbon monoxide is a poisonous gas. It is dangerous because it reduces the ability of blood to transport oxygen.

Safety: Poorly ventilated charcoal stoves and generators can produce dangerous concentrations of carbon monoxide.

10.3 Reaction of oxygen with sulphur

Sulphur burns in oxygen with a blue flame to form sulphur dioxide.

Chemical equation:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Observations:

1. Sulphur burns with a blue flame.
2. A colourless gas is produced.
3. The gas has a choking smell.
4. Sulphur dioxide dissolves in water to form an acidic solution.

Chemical equation for reaction with water:

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

Sulphur dioxide contributes to acid rain when released into the atmosphere.

10.4 Summary of reactions of oxygen

Substance	Reaction	Product
Magnesium	$2Mg + O_2$	$2MgO$
Copper	$2Cu + O_2$	$2CuO$
Carbon	$C + O_2$	CO_2
Carbon in limited oxygen	$2C + O_2$	$2CO$
Sulphur	$S + O_2$	SO_2

11.0 CHEMICAL PROPERTIES OF NITROGEN

Nitrogen is relatively unreactive at room temperature.

This is because nitrogen molecules contain a very strong triple covalent bond between the two nitrogen atoms.

The bond is represented as:

$$N \equiv N$$

A large amount of energy is required to break this bond before nitrogen can react.

However, nitrogen reacts with certain substances under suitable conditions.

11.1 Reaction of nitrogen with magnesium

Magnesium reacts with nitrogen when heated strongly to form magnesium nitride.

Procedure

1. Place magnesium ribbon in a suitable combustion spoon.
2. Heat the magnesium strongly.
3. Allow it to burn in a nitrogen atmosphere under controlled laboratory conditions.
4. Observe the formation of a solid product.

Chemical equation

$$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

Observations

1. Magnesium burns.
2. A solid product is formed.
3. The product is magnesium nitride.

Reaction of magnesium nitride with water

Magnesium nitride reacts with water to produce magnesium hydroxide and ammonia gas.

$$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(s) + 2NH_3(g)$$

Importance

This reaction demonstrates that nitrogen can combine with metals to form nitrides.

11.2 Reaction of nitrogen with hydrogen

Nitrogen reacts with hydrogen to form ammonia.

This reaction is used industrially in the Haber process.

Chemical equation

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Conditions required

- Temperature: approximately 400–500°C.
- Pressure: approximately 150–250 atmospheres in many industrial operations.
- Catalyst: finely divided iron.
- Nitrogen and hydrogen are supplied in a ratio of 1:3 by volume.

Explanation

1. Nitrogen and hydrogen are mixed in the correct proportions.
2. The mixture is compressed to high pressure.
3. The gases are heated to the required temperature.

4. The mixture is passed over an iron catalyst.
5. Ammonia is formed.
6. Ammonia is cooled and separated.
7. Unreacted nitrogen and hydrogen are recycled.

Importance of the Haber process

1. It produces ammonia on a large scale.
2. Ammonia is used in the manufacture of nitrogen fertilisers.
3. Fertilisers improve agricultural productivity.
4. Ammonia is used in the manufacture of nitric acid.
5. It supports food production by providing nitrogen compounds for crops.

11.3 Reaction of nitrogen with oxygen

Nitrogen reacts with oxygen at very high temperatures to form nitrogen monoxide.

This reaction can occur during lightning and in high-temperature industrial combustion processes.

Chemical equation

$$N_2(g) + O_2(g) \xrightarrow{\text{high temperature}} 2NO(g)$$

Nitrogen monoxide further reacts with oxygen to form nitrogen dioxide.

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Observations

1. Nitrogen monoxide is a colourless gas.
2. Nitrogen dioxide is a reddish-brown gas.
3. Nitrogen dioxide dissolves in water to form acids.

Importance and environmental effects

1. Nitrogen oxides are produced in vehicle engines.
2. They are released from some industrial processes.
3. They contribute to air pollution.
4. They contribute to the formation of acid rain.
5. They contribute to the formation of photochemical smog.

12.0 CHEMICAL PROPERTIES OF HYDROGEN

Hydrogen is a reactive non-metal that burns readily in oxygen.

It also acts as a reducing agent because it can remove oxygen from some metal oxides.

12.1 Reaction of hydrogen with oxygen

Hydrogen burns in oxygen to form water.

Chemical equation

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Observations

1. Hydrogen burns with a pale blue flame in air.
2. Water is formed.
3. The reaction releases a large amount of heat.

4. A mixture of hydrogen and oxygen can explode if ignited.

Explanation

Hydrogen combines with oxygen to form water.

The reaction is highly exothermic, meaning it releases energy in the form of heat.

Importance

1. Hydrogen can be used as a fuel.
2. It is used in fuel cells to generate electricity.
3. It is used in some rocket propulsion systems.
4. It produces water as the main product when it burns completely in oxygen.

Safety: Hydrogen and oxygen must not be mixed and ignited in an uncontrolled manner. Their mixture can be explosive.

12.2 Reaction of hydrogen with metal oxides

Hydrogen reacts with certain heated metal oxides by removing oxygen from them.

This process is called reduction.

Hydrogen acts as a reducing agent.

A. Reaction with copper(II) oxide

When hydrogen is passed over heated copper(II) oxide, it reduces copper(II) oxide to copper.

Chemical equation

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Procedure

1. Place black copper(II) oxide in a suitable combustion tube.
2. Pass dry hydrogen through the tube.
3. Heat the copper(II) oxide strongly.
4. Observe the changes taking place.
5. Allow the apparatus to cool while maintaining a safe flow of hydrogen as directed by the teacher.

Observations

1. The black copper(II) oxide changes to reddish-brown copper.
2. Water droplets may form in the cooler part of the apparatus.
3. Hydrogen is oxidised to water.
4. Copper(II) oxide is reduced to copper.

Explanation

- Copper(II) oxide loses oxygen.
- Hydrogen gains oxygen to form water.
- Hydrogen acts as the reducing agent.
- Copper(II) oxide acts as the oxidising agent.

B. Reaction with lead(II) oxide

Hydrogen can also reduce heated lead(II) oxide to lead.

$$PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(g)$$

Importance of reduction by hydrogen

1. It is useful in the extraction of certain metals.
2. It demonstrates the reducing properties of hydrogen.
3. It is used in some industrial chemical processes.
4. It helps learners understand oxidation and reduction reactions.

13.0 USES OF OXYGEN, HYDROGEN AND NITROGEN

These gases are important in medicine, agriculture, manufacturing, transport, food preservation and other areas of daily life.

13.1 Uses of oxygen

1. **Medical treatment**
   - ✓ Oxygen is administered to patients who require additional oxygen support.
   - ✓ It is used in hospitals, ambulances and emergency medical care.
   - ✓ Medical oxygen is administered under professional supervision.
2. **Welding and metal cutting**
   - ✓ Oxygen is used with fuel gases in oxy-fuel welding and cutting.
   - ✓ It helps produce high-temperature flames.
3. **Steel manufacturing**
   - ✓ Oxygen is used in steelmaking to remove excess carbon and other impurities from molten iron.
4. **Water treatment**
   - ✓ Oxygen is used in aeration processes.
   - ✓ It supports microorganisms that break down organic waste in wastewater treatment plants.
5. **Respiration**
   - ✓ Oxygen is needed by most living organisms for aerobic respiration.
   - ✓ It helps release energy from food.
6. **Rocket propulsion**
   - ✓ Liquid oxygen is used as an oxidiser in rocket engines.
   - ✓ It supports the burning of rocket fuel.
7. **Industrial chemical production**
   - ✓ Oxygen is used in the manufacture of chemicals such as sulphuric acid and nitric acid through relevant industrial processes.
8. **Aquaculture**
   - ✓ Oxygen is supplied to fish ponds and fish tanks through aeration systems.
   - ✓ It helps maintain dissolved oxygen levels required by aquatic organisms.

13.2 Uses of hydrogen

1. **Manufacture of ammonia**
   - ✓ Hydrogen combines with nitrogen in the Haber process to produce ammonia.
   - ✓ Ammonia is used in fertiliser manufacture.
2. **Hydrogenation of vegetable oils**
   - ✓ Hydrogen is used to convert some unsaturated oils into more saturated fats.
   - ✓ This process changes the physical properties of the oils.
3. **Fuel cells**
   - ✓ Hydrogen is used in fuel cells to generate electricity.
   - ✓ Water is produced as a major product in hydrogen-oxygen fuel cells.
4. **Rocket fuel**

- ✓ Liquid hydrogen is used as a fuel in some rocket engines.
- ✓ It reacts with liquid oxygen to release large amounts of energy.

5. **Manufacture of hydrochloric acid**
   - ✓ Hydrogen combines with chlorine to form hydrogen chloride gas.
   - ✓ Hydrogen chloride dissolves in water to form hydrochloric acid.

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

6. **Reduction of metal oxides**
   - ✓ Hydrogen is used to remove oxygen from certain metal oxides.
   - ✓ This can produce metals such as copper.
7. **Production of methanol**
   - ✓ Hydrogen is used with carbon monoxide or carbon dioxide in the industrial manufacture of methanol.
8. **Petroleum refining**
   - ✓ Hydrogen is used in processes such as hydrocracking and desulphurisation.

13.3 Uses of nitrogen

1. **Manufacture of ammonia**
   - ✓ Nitrogen reacts with hydrogen to produce ammonia.
   - ✓ Ammonia is used to manufacture fertilisers.
2. **Food preservation and packaging**
   - ✓ Nitrogen is used in food packaging to reduce oxidation.
   - ✓ It helps maintain the quality of some packaged foods.
3. **Liquid nitrogen for cooling**
   - ✓ Liquid nitrogen is used for rapid freezing.
   - ✓ It is used in scientific laboratories, medical research and food processing.
4. **Preservation of biological materials**
   - ✓ Liquid nitrogen is used to preserve selected biological materials at very low temperatures.
   - ✓ It is used in specialised medical and research facilities.
5. **Inert atmosphere**
   - ✓ Nitrogen is used to provide an inert atmosphere in certain industrial processes.
   - ✓ It reduces unwanted reactions with oxygen.
6. **Electronics manufacture**
   - ✓ Nitrogen is used in some electronics manufacturing processes to reduce oxidation and contamination.
7. **Tyre inflation**
   - ✓ Nitrogen is used to inflate certain vehicle and aircraft tyres.
   - ✓ It is particularly useful in specialised applications requiring controlled gas properties.
8. **Chemical industries**
   - ✓ Nitrogen is used in the manufacture of chemicals such as ammonia and other nitrogen-containing compounds.

13.4 Summary of uses

Oxygen	Hydrogen	Nitrogen
Medical oxygen	Manufacture of ammonia	Manufacture of ammonia
Welding and cutting	Fuel cells	Food packaging

Oxygen	Hydrogen	Nitrogen
Steel manufacture	Rocket fuel	Liquid nitrogen cooling
Water treatment	Hydrogenation of oils	Preservation of biological materials
Respiration	Metal oxide reduction	Inert industrial atmosphere
Aquaculture	Petroleum refining	Electronics manufacture

14.0 IMPORTANCE OF GASES IN DAY-TO-DAY LIFE

Gases play an important role in supporting life and economic activities.

14.1 Importance in human health

1. Oxygen supports respiration.
2. Medical oxygen helps patients with breathing difficulties when clinically indicated.
3. Oxygen is used in emergency medical services.
4. Nitrogen is used in specialised medical and biological preservation processes.

14.2 Importance in agriculture

1. Nitrogen is essential for the growth of plants.
2. Nitrogen fertilisers improve crop production when used correctly.
3. Oxygen supports respiration in plant roots and soil organisms.
4. Nitrogen-fixing bacteria help make atmospheric nitrogen available to plants.
5. Hydrogen is used in the manufacture of ammonia, which is an important raw material for fertilisers.

14.3 Importance in industry

1. Oxygen is used in steel manufacture.
2. Hydrogen is used in petroleum refining.
3. Nitrogen is used to create inert atmospheres.
4. Oxygen supports metal cutting and welding.
5. Hydrogen and nitrogen are used in chemical production.

14.4 Importance in environmental management

1. Oxygen supports aquatic life through dissolved oxygen.
2. Oxygen is used in wastewater treatment.
3. Nitrogen is part of the natural nitrogen cycle.
4. Responsible industrial gas use can reduce waste and improve efficiency.
5. Monitoring nitrogen oxides and other gaseous pollutants helps protect air quality.

15.0 ENVIRONMENTAL EFFECTS OF GASES AND THEIR COMPOUNDS

Some gases are beneficial, while others may cause environmental and health problems when released in excessive amounts.

15.1 Effects of nitrogen oxides

Nitrogen monoxide and nitrogen dioxide are commonly grouped as nitrogen oxides (NO_x).

Their effects include:

1. Contributing to acid rain.

2. Irritating the respiratory system.
3. Contributing to the formation of photochemical smog.
4. Affecting plant growth.
5. Reducing air quality in urban and industrial areas.

15.2 Effects of sulphur dioxide

Sulphur dioxide is formed when sulphur-containing fuels are burned.

Its effects include:

1. Irritation of the respiratory system.
2. Formation of acid rain.
3. Damage to crops and vegetation.
4. Corrosion of buildings and monuments.
5. Acidification of water bodies and soils.

15.3 Effects of carbon monoxide

Carbon monoxide is produced during incomplete combustion of carbon-containing fuels.

Its effects include:

1. It is poisonous when inhaled.
2. It combines with haemoglobin in the blood.
3. It reduces the ability of blood to transport oxygen.
4. It can cause headaches, dizziness and unconsciousness.
5. High exposure can cause death.

15.4 Ways of reducing harmful gaseous emissions

1. Use cleaner energy sources.
2. Maintain vehicle engines properly.
3. Install suitable industrial emission-control equipment.
4. Avoid burning waste unnecessarily.
5. Use efficient cooking stoves.
6. Ensure proper ventilation when using fuel-burning appliances.
7. Plant and protect trees.
8. Monitor emissions from industries and vehicles.
9. Use appropriate technology to reduce sulphur dioxide and nitrogen oxide emissions.

16.0 LABORATORY SAFETY DURING THE PREPARATION AND INVESTIGATION OF GASES

Laboratory activities involving gases require proper handling because some gases are flammable, poisonous, irritating or capable of displacing oxygen.

16.1 General safety precautions

1. Wear a laboratory coat and safety goggles.
2. Follow the teacher's instructions.
3. Use small quantities of chemicals.
4. Avoid inhaling gases directly.
5. Conduct experiments involving harmful gases in a fume cupboard.
6. Ensure the apparatus is properly fitted.

7. Keep flammable substances away from naked flames.
8. Do not taste laboratory chemicals.
9. Wash hands after experiments.
10. Dispose of chemical waste according to laboratory guidelines.

16.2 Specific precautions

Gas	Hazard	Safety measure
Oxygen	Supports vigorous combustion	Keep away from oils, grease and combustible materials
Hydrogen	Highly flammable and forms explosive mixtures with air	Keep away from flames and sparks
Nitrogen	Can cause suffocation in enclosed spaces	Ensure good ventilation
Sulphur dioxide	Toxic and irritating	Use a fume cupboard
Nitrogen dioxide	Toxic and irritating	Avoid inhalation; use a fume cupboard
Carbon monoxide	Highly poisonous	Prevent incomplete combustion and ensure ventilation

PRACTICAL-BASED LEARNING ACTIVITIES

Activity 1: Identifying sources of selected gases

Task: In groups, discuss the natural and industrial sources of oxygen, hydrogen and nitrogen.

Expected outcomes:

- Identify natural sources of the gases.
- Explain how the gases are obtained.
- Relate the gases to everyday activities.

Activity 2: Investigating the preparation of oxygen

Task: Under teacher supervision, prepare oxygen using hydrogen peroxide and manganese(IV) oxide.

Expected outcomes:

- Assemble gas-preparation apparatus.
- Prepare oxygen.
- Collect oxygen over water.
- Test oxygen using a glowing splint.
- Record observations.

Activity 3: Investigating the preparation of hydrogen

Task: Prepare hydrogen by reacting zinc with dilute hydrochloric acid.

Expected outcomes:

- Identify the reactants.
- Assemble the apparatus.
- Collect hydrogen.
- Test a small sample safely.
- Write the balanced chemical equation.

Activity 4: Investigating the physical properties of gases

Task: Compare the physical properties of oxygen, hydrogen and nitrogen using prepared samples or teacher-provided demonstrations.

Expected outcomes:

- Compare colour and odour.
- Investigate solubility.
- Investigate effects on moist litmus paper.
- Carry out appropriate gas tests.
- Present observations in a table.

Activity 5: Discussing industrial isolation

Task: Use diagrams, textbooks or digital media to explain how oxygen and nitrogen are separated from liquid air.

Expected outcomes:

- Explain the preparation of liquid air.
- Describe fractional distillation.
- Compare boiling points.
- Explain why nitrogen is collected before oxygen.

Activity 6: Environmental effects of gases

Task: Discuss the effects of industrial gaseous emissions on the environment and suggest ways of reducing pollution.

Expected outcomes:

- Identify harmful gaseous pollutants.
- Explain their effects.
- Suggest practical pollution-control measures.

KEY TERMS AND THEIR MEANINGS

Term	Meaning
Non-metal	An element that generally lacks metallic properties
Catalyst	A substance that changes the rate of a chemical reaction without being consumed
Combustion	A chemical reaction involving burning, usually in oxygen
Oxidation	A process involving gain of oxygen or loss of electrons
Reduction	A process involving loss of oxygen or gain of electrons
Oxide	A compound formed when an element combines with oxygen
Gas collection	The process of gathering a gas produced during an experiment
Drying agent	A substance used to remove moisture from a gas
Fractional distillation	Separation of a mixture based on differences in boiling points
Liquefaction	Conversion of a gas into a liquid
Solubility	The ability of a substance to dissolve in another substance

Inert atmosphere	An environment containing a gas that does not readily react under the conditions used
Haber process	Industrial process used to manufacture ammonia from nitrogen and hydrogen
Nitrogen fixation	Conversion of atmospheric nitrogen into nitrogen-containing compounds
Exothermic reaction	A chemical reaction that releases heat energy

'SUMMARY OF IMPORTANT CHEMICAL EQUATIONS

Reaction	Balanced chemical equation
Preparation of oxygen	$2H_2O_2 \rightarrow 2H_2O + O_2$
Preparation of hydrogen	$Zn + 2HCl \rightarrow ZnCl_2 + H_2$
Preparation of hydrogen using sulphuric acid	$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
Preparation of nitrogen	$NH_4NO_2 \rightarrow N_2 + 2H_2O$
Oxygen with magnesium	$2Mg + O_2 \rightarrow 2MgO$
Oxygen with copper	$2Cu + O_2 \rightarrow 2CuO$
Oxygen with carbon	$C + O_2 \rightarrow CO_2$
Oxygen with sulphur	$S + O_2 \rightarrow SO_2$
Nitrogen with magnesium	$3Mg + N_2 \rightarrow Mg_3N_2$
Nitrogen with hydrogen	$N_2 + 3H_2 \rightleftharpoons 2NH_3$
Nitrogen with oxygen	$N_2 + O_2 \rightarrow 2NO$
Nitrogen monoxide with oxygen	$2NO + O_2 \rightarrow 2NO_2$
Hydrogen with oxygen	$2H_2 + O_2 \rightarrow 2H_2O$
Hydrogen with copper(II) oxide	$CuO + H_2 \rightarrow Cu + H_2O$
Hydrogen with lead(II) oxide	$PbO + H_2 \rightarrow Pb + H_2O$

REVISION QUESTIONS

Section A: Multiple-choice questions

Choose the correct answer from the alternatives given.

1. Which of the following gases forms the largest percentage of atmospheric air?

A. Oxygen
B. Nitrogen
C. Hydrogen
D. Carbon dioxide

2. Which gas relights a glowing wooden splint?

A. Nitrogen
B. Hydrogen
C. Oxygen
D. Carbon dioxide

3. Which substance is used as a catalyst in the laboratory preparation of oxygen from hydrogen peroxide?

A. Sodium chloride
B. Manganese(IV) oxide
C. Calcium carbonate
D. Copper(II) oxide

4. The chemical formula of hydrogen gas is:

A. H
B. H_2
C. H_2O
D. H_3

5. Which gas produces a squeaky pop when a small sample is tested with a burning splint?

A. Nitrogen
B. Oxygen
C. Hydrogen
D. Carbon dioxide

6. Which of the following methods is suitable for collecting oxygen in the laboratory?

A. Collection over water
B. Collection over concentrated sulphuric acid
C. Collection over sodium hydroxide
D. Collection over kerosene

7. Nitrogen makes up approximately what percentage of dry atmospheric air?

A. 21%
B. 50%
C. 78%
D. 93%

8. Which gas is used in the manufacture of ammonia through the Haber process?

A. Oxygen
B. Nitrogen
C. Carbon dioxide
D. Sulphur dioxide

9. What is the role of manganese(IV) oxide in the preparation of oxygen?

A. It acts as a fuel.
B. It acts as a catalyst.
C. It acts as a drying agent.
D. It acts as a reactant.

10. Which gas is highly flammable?

A. Nitrogen
B. Oxygen
C. Hydrogen
D. Carbon dioxide

11. What is formed when magnesium reacts with oxygen?

A. Magnesium chloride
B. Magnesium oxide

C. Magnesium sulphate
D. Magnesium hydroxide

12. Which of the following is the correct equation for the reaction between hydrogen and oxygen?

A. $H_2 + O_2 \rightarrow H_2O$

B. $2H_2 + O_2 \rightarrow 2H_2O$

C. $H_2 + 2O_2 \rightarrow H_2O$

D. $2H_2 + 2O_2 \rightarrow H_2O$

13. Which gas is used in food packaging to reduce oxidation?

A. Nitrogen
B. Chlorine
C. Sulphur dioxide
D. Hydrogen sulphide

14. The boiling point of nitrogen is approximately:

A. −183°C
B. −196°C
C. 0°C
D. 100°C

15. During fractional distillation of liquid air, nitrogen is collected before oxygen because nitrogen:

A. Is more reactive.
B. Has a higher boiling point.
C. Has a lower boiling point.
D. Is denser than oxygen.

16. Which product is formed when nitrogen reacts with magnesium?

A. Magnesium oxide
B. Magnesium nitride
C. Magnesium chloride
D. Magnesium sulphide

17. Which gas does not support combustion under ordinary conditions?

A. Oxygen
B. Hydrogen
C. Nitrogen
D. Both oxygen and hydrogen

18. Which substance can be used to dry ammonia gas?

A. Concentrated sulphuric acid
B. Quicklime
C. Water
D. Dilute hydrochloric acid

19. What is the product formed when carbon burns in sufficient oxygen?

A. Carbon monoxide
B. Carbon dioxide
C. Methane
D. Ethane

20. Which industrial process is used to manufacture ammonia?

A. Contact process
B. Haber process
C. Fractional distillation
D. Electrolysis

21. Which gas is used in oxy-fuel welding and cutting?

A. Oxygen
B. Nitrogen
C. Carbon dioxide only
D. Hydrogen chloride

22. Which of the following gases is slightly denser than air?

A. Hydrogen
B. Nitrogen
C. Oxygen
D. Helium

23. What happens when hydrogen is passed over heated copper(II) oxide?

A. Copper is oxidised.
B. Copper(II) oxide is reduced.
C. Hydrogen is reduced to copper.
D. Copper(II) oxide forms magnesium oxide.

24. Which gas is released during photosynthesis?

A. Nitrogen
B. Hydrogen
C. Oxygen
D. Sulphur dioxide

25. Which of the following is a major environmental effect of nitrogen dioxide?

A. Formation of acid rain
B. Formation of pure oxygen
C. Improvement of air quality
D. Prevention of smog

26. Which gas is used in fuel cells to generate electricity?

A. Hydrogen
B. Nitrogen
C. Chlorine
D. Sulphur dioxide

27. What is the main reason nitrogen is relatively unreactive at room temperature?

A. It is highly soluble in water.
B. It has a strong triple bond.
C. It is a metal.
D. It has a low density.

28. Which gas is produced when zinc reacts with dilute hydrochloric acid?

A. Oxygen
B. Nitrogen
C. Hydrogen
D. Carbon dioxide

29. Which of the following is a correct use of liquid nitrogen?

A. Welding metals
B. Rapid freezing and cooling
C. Supporting combustion
D. Manufacture of oxygen

30. Which of the following statements about oxygen is correct?

A. Oxygen is itself flammable.
B. Oxygen supports combustion.
C. Oxygen is highly soluble in water.
D. Oxygen turns moist blue litmus red.

Section B: Structured questions

1.
a) Define a non-metal.
b) Name three non-metals that exist as gases at room temperature.
c) State four general characteristics of non-metals.
d) Explain three ways in which non-metals are important in everyday life.

2.
a) State three natural sources of hydrogen.
b) Explain how hydrogen can be obtained from water.
c) Name two industrial sources of hydrogen.
d) State three uses of hydrogen.

3.
a) State four natural sources of oxygen.
b) Explain how plants release oxygen into the atmosphere.
c) State four uses of oxygen in industry and medicine.

4.
a) State the percentage of nitrogen in dry atmospheric air.
b) Name three natural sources of nitrogen.
c) Explain the importance of nitrogen in agriculture.
d) State four uses of nitrogen.

5.
a) Explain the principle of collecting gases over water.
b) State three gases that can be collected using this method.
c) Give two limitations of collecting gases over water.
d) Explain why ammonia cannot be collected over water.

6.
a) Name three methods used to collect gases in the laboratory.
b) Explain the principle of upward delivery.

c) Explain the principle of downward delivery.
d) State two factors considered when selecting a gas collection method.

7.
Describe the laboratory preparation of oxygen from hydrogen peroxide under the following headings:

a) Requirements.

b) Procedure.

c) Observations.

d) Chemical equation.

e) Test for the gas.

f) Precautions.

8.
a) Describe how hydrogen is prepared in the laboratory using zinc and dilute hydrochloric acid.
b) Write the balanced chemical equation.
c) State three observations.
d) Describe the test for hydrogen.
e) State three safety precautions.

9.
a) Describe the laboratory preparation of nitrogen using ammonium nitrite.
b) Write the chemical equation for the decomposition of ammonium nitrite.
c) State two observations.
d) Explain why nitrogen does not support combustion.

10.
a) What is fractional distillation?
b) Explain how oxygen and nitrogen are separated from liquid air.
c) State the boiling points of oxygen and nitrogen.
d) Explain why nitrogen is collected before oxygen.
e) State three advantages of fractional distillation of liquid air.

11.
Compare the physical properties of oxygen, hydrogen and nitrogen under the following headings:

a) Colour.

b) Odour.

c) Solubility in water.

d) Effect on moist litmus paper.

e) Effect on a glowing or burning splint.

f) Density compared with air.

12.
a) Describe the reaction between magnesium and oxygen.
b) Write the balanced chemical equation.

c) State the observations when magnesium burns in oxygen.
d) Explain why the reaction is classified as oxidation.

13.
a) Write the chemical equation for the reaction between carbon and sufficient oxygen.
b) State the product formed when carbon burns in limited oxygen.
c) Explain why carbon monoxide is dangerous to human beings.

14.
a) Describe the reaction between sulphur and oxygen.
b) Write the balanced chemical equation.
c) State two environmental effects of sulphur dioxide.

15.
a) Explain why nitrogen is relatively unreactive at room temperature.
b) Describe the reaction between nitrogen and magnesium.
c) Write the balanced chemical equation.
d) Name the product formed.
e) State what happens when the product reacts with water.

16.
a) Write the balanced chemical equation for the reaction between nitrogen and hydrogen.
b) Name the industrial process involved.
c) State four conditions required in the process.
d) Explain three uses of ammonia.

17.
a) Write the chemical equation for the reaction between nitrogen and oxygen at high temperatures.
b) Name the product formed.
c) Explain how nitrogen dioxide is formed.
d) State three environmental effects of nitrogen oxides.

18.
a) Explain the reaction between hydrogen and oxygen.
b) Write the balanced chemical equation.
c) State three uses of hydrogen as a fuel.
d) Explain why mixtures of hydrogen and oxygen must be handled carefully.

19.
a) Explain how hydrogen reduces copper(II) oxide.
b) Write the balanced chemical equation.
c) State the observations made during the reaction.
d) Identify the reducing agent.
e) Identify the substance that is reduced.

20.
A learner collected three colourless gases labelled A, B and C. The learner carried out the following tests:

- Gas A relit a glowing wooden splint.
- Gas B produced a squeaky pop when tested with a burning splint.
- Gas C extinguished a burning splint.

a) Identify gases A, B and C.

b) State the property demonstrated by each test.

c) Give one use of each gas.

d) Explain why the three gases are important in everyday life.

STRAND 2.0: PHYSICAL CHEMISTRY

SUB-STRAND 2.1: SALTS

'1.0 INTRODUCTION TO SALTS

1.1 Meaning of salts

A **salt** is an ionic compound formed when the hydrogen ions of an acid are wholly or partly replaced by metal ions or ammonium ions.

Salts are made up of positively charged ions called cations and negatively charged ions called anions.

Examples include:

- Sodium chloride, NaCl.
- Potassium nitrate, KNO_3.
- Calcium carbonate, $CaCO_3$.
- Copper(II) sulphate, $CuSO_4$.
- Ammonium chloride, NH_4Cl.
- Magnesium sulphate, $MgSO_4$.

Salts are important in agriculture, medicine, food preparation, construction, water treatment and manufacturing industries.

1.2 Formation of salts

Salts can be prepared through different chemical reactions.

A. Reaction between an acid and a metal

A suitable metal reacts with a dilute acid to produce a salt and hydrogen gas.

Example:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

Magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen gas.

B. Reaction between an acid and a base

An acid reacts with a base to form a salt and water. This is called neutralisation.

Example:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Hydrochloric acid reacts with sodium hydroxide to form sodium chloride and water.

C. Reaction between an acid and a metal carbonate

An acid reacts with a metal carbonate to produce a salt, water and carbon dioxide.

Example:

$$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$$

D. Reaction between an acid and a metal oxide

An acid reacts with a metal oxide to produce a salt and water.

Example:

$$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$$

E. Precipitation reaction

Two soluble salts react in solution to form an insoluble salt called a precipitate.

Example:

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Silver chloride is formed as a white precipitate.

1.3 Classification of salts

Salts may be classified according to their composition and behaviour.

A. Normal salts

Normal salts are formed when all replaceable hydrogen ions of an acid have been replaced by metal or ammonium ions.

Examples:

- Sodium chloride, NaCl.
- Potassium sulphate, K_2SO_4.

- Calcium nitrate, $Ca(NO_3)_2$.
- Ammonium chloride, NH_4Cl.

B. Acid salts

Acid salts are formed when only some of the replaceable hydrogen ions of a polybasic acid have been replaced.

Examples:

- Sodium hydrogen carbonate, $NaHCO_3$.
- Potassium hydrogen sulphate, $KHSO_4$.
- Sodium hydrogen sulphate, $NaHSO_4$.

C. Basic salts

Basic salts contain hydroxide ions together with other anions in the same compound. They may be formed through incomplete neutralisation of a base.

An example is basic copper(II) carbonate, commonly represented as:

$$Cu_2(OH)_2CO_3$$

1.4 Importance of salts in everyday life

1. **Food preparation:** Sodium chloride is used for seasoning and preserving food.
2. **Agriculture:** Ammonium nitrate, potassium nitrate and other salts supply essential nutrients to crops.
3. **Medicine:** Some salts are used in oral rehydration solutions and other medical preparations.
4. **Construction:** Calcium carbonate is used in the manufacture of cement and lime.
5. **Water treatment:** Certain salts are used in water purification processes.
6. **Industry:** Sodium carbonate is used in glass and detergent manufacture.
7. **Farming:** Copper(II) sulphate is used in some agricultural and laboratory applications.
8. **Household activities:** Sodium hydrogen carbonate is used in baking and cleaning.

2.0 ACTION OF HEAT ON SALTS

2.1 Meaning of thermal decomposition

Thermal decomposition is the process in which a chemical compound breaks down into two or more simpler substances when heated.

Some salts are stable when heated, while others decompose to form new substances.

The products formed depend on the type of salt and the conditions under which it is heated.

2.2 General factors affecting thermal decomposition

1. **Nature of the salt:** Different salts have different thermal stabilities.
2. **Position of the metal in the reactivity series:** The thermal stability of many metal salts is influenced by the nature of the metal ion.
3. **Temperature:** Some salts require higher temperatures than others to decompose.
4. **Presence of impurities:** Impurities may affect the decomposition process.
5. **Particle size:** Smaller particles may heat more uniformly and react faster.

2.3 Action of heat on metal carbonates

Metal carbonates are salts containing the carbonate ion, CO_3^{2-}.

When many metal carbonates are heated, they decompose to form a metal oxide and carbon dioxide gas.

General equation

$$MCO_3(s) \xrightarrow{\Delta} MO(s) + CO_2(g)$$

Here, M represents a suitable divalent metal.

A. Heating calcium carbonate

Calcium carbonate decomposes on strong heating to form calcium oxide and carbon dioxide.

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

Observations:

1. A white solid is heated.
2. A colourless gas is released.
3. The gas turns limewater milky.
4. Calcium oxide remains as a white solid.

Explanation:

- Calcium carbonate breaks down due to heat.
- Calcium oxide is the solid product.
- Carbon dioxide is released as a gas.

Industrial application:

This reaction is used in the manufacture of quicklime in lime kilns.

Calcium oxide is used in:

- Cement manufacture.
- Soil treatment.
- Steel production.
- Water treatment.
- Manufacture of calcium hydroxide.

B. Heating copper(II) carbonate

Copper(II) carbonate decomposes on heating to form black copper(II) oxide and carbon dioxide.

$$CuCO_3(s) \xrightarrow{\Delta} CuO(s) + CO_2(g)$$

Observations:

1. Green copper(II) carbonate is heated.
2. The green solid changes to black.
3. A colourless gas is released.
4. The gas turns limewater milky.

Explanation:

The green copper(II) carbonate breaks down to form black copper(II) oxide and carbon dioxide.

C. Heating zinc carbonate

Zinc carbonate decomposes on heating to form zinc oxide and carbon dioxide.

$$ZnCO_3(s) \xrightarrow{\Delta} ZnO(s) + CO_2(g)$$

Observations:

1. Zinc carbonate is heated.
2. A colourless gas is produced.
3. The solid residue is zinc oxide.
4. Zinc oxide is yellow when hot and white when cold.

2.4 Summary of thermal decomposition of carbonates

Carbonate	Solid product	Gaseous product
Calcium carbonate	Calcium oxide	Carbon dioxide
Copper(II) carbonate	Copper(II) oxide	Carbon dioxide
Zinc carbonate	Zinc oxide	Carbon dioxide

Important: Not all carbonates decompose at the same temperature. Sodium carbonate and potassium carbonate are relatively stable under ordinary laboratory heating conditions.

2.5 Action of heat on hydrogen carbonates

Hydrogen carbonates are salts containing the hydrogen carbonate ion, HCO_3^-.

When many metal hydrogen carbonates are heated, they decompose to form a carbonate, water and carbon dioxide.

General equation

$$2MHCO_3(s) \xrightarrow{\Delta} M_2CO_3(s) + H_2O(g) + CO_2(g)$$

This general equation applies to hydrogen carbonates of monovalent metals.

A. Heating sodium hydrogen carbonate

Sodium hydrogen carbonate decomposes when heated to form sodium carbonate, water and carbon dioxide.

$$2NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Observations:

1. White sodium hydrogen carbonate is heated.
2. A colourless gas is produced.
3. Water vapour is formed.
4. A white solid residue remains.
5. The gas turns limewater milky.

Applications:

Sodium hydrogen carbonate is used in:

- Baking powder formulations.
- Some fire extinguishers.
- Household cleaning.
- Certain medical preparations.

B. Heating calcium hydrogen carbonate

Calcium hydrogen carbonate is unstable and decomposes to form calcium carbonate, water and carbon dioxide.

$$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

This reaction occurs when water containing dissolved calcium hydrogen carbonate is heated.

Importance:

It explains the formation of temporary hardness deposits when hard water is boiled.

2.6 Difference between carbonates and hydrogen carbonates on heating

Carbonates	Hydrogen carbonates
Contain CO_3^{2-} ions	Contain HCO_3^- ions
Many decompose to form metal oxides and CO_2	Many decompose to form carbonates, water and CO_2
Example: $CaCO_3$	Example: $NaHCO_3$
Calcium carbonate forms calcium oxide and CO_2	Sodium hydrogen carbonate forms sodium carbonate, water and CO_2

3.0 ACTION OF HEAT ON NITRATES

Nitrates are salts containing the nitrate ion, NO_3^-.

The thermal decomposition of nitrates depends on the metal present in the salt.

Different groups of nitrates produce different decomposition products.

3.1 Thermal decomposition of nitrates of potassium and sodium

Nitrates of sodium and potassium decompose on heating to form the corresponding nitrite and oxygen gas.

General equation:

$$2MNO_3(s) \xrightarrow{\Delta} 2MNO_2(s) + O_2(g)$$

A. Heating potassium nitrate

$$2KNO_3(s) \xrightarrow{\Delta} 2KNO_2(s) + O_2(g)$$

Observations:

1. Potassium nitrate is a white solid.
2. On strong heating, it decomposes.
3. Oxygen gas is released.
4. A white solid residue of potassium nitrite remains.

Test for oxygen: A glowing splint relights.

B. Heating sodium nitrate

$$2NaNO_3(s) \xrightarrow{\Delta} 2NaNO_2(s) + O_2(g)$$

Sodium nitrate decomposes to form sodium nitrite and oxygen.

3.2 Thermal decomposition of nitrates of moderately reactive metals

Many nitrates of metals such as lead, copper and zinc decompose on heating to form the metal oxide, nitrogen dioxide and oxygen.

General equation

$$2M(NO_3)_2(s) \xrightarrow{\Delta} 2MO(s) + 4NO_2(g) + O_2(g)$$

This equation applies to many divalent metal nitrates.

A. Heating lead(II) nitrate

$$2Pb(NO_3)_2(s) \xrightarrow{\Delta} 2PbO(s) + 4NO_2(g) + O_2(g)$$

Observations:

1. White lead(II) nitrate crystals are heated.
2. Brown fumes of nitrogen dioxide are released.
3. Oxygen gas is also produced.
4. A yellow residue of lead(II) oxide remains.

Explanation:

Lead(II) nitrate decomposes to produce lead(II) oxide, nitrogen dioxide and oxygen.

Safety: Nitrogen dioxide is toxic. This experiment must only be demonstrated under suitable fume-cupboard conditions by a trained teacher.

B. Heating copper(II) nitrate

$$2Cu(NO_3)_2(s) \xrightarrow{\Delta} 2CuO(s) + 4NO_2(g) + O_2(g)$$

Observations:

1. Blue copper(II) nitrate crystals are heated.
2. Brown fumes of nitrogen dioxide are released.
3. Black copper(II) oxide remains.
4. Oxygen gas is also produced.

3.3 Thermal decomposition of nitrates of very reactive metals

Nitrates of very reactive metals such as lithium, sodium and potassium require different considerations.

- Lithium nitrate decomposes to lithium oxide, nitrogen dioxide and oxygen.
- Sodium and potassium nitrates decompose to nitrites and oxygen.

For example:

$$4LiNO_3(s) \xrightarrow{\Delta} 2Li_2O(s) + 4NO_2(g) + O_2(g)$$

3.4 Summary of decomposition of selected nitrates

Nitrate	Solid residue	Gases produced
Potassium nitrate	Potassium nitrite	Oxygen
Sodium nitrate	Sodium nitrite	Oxygen
Lead(II) nitrate	Lead(II) oxide	Nitrogen dioxide and oxygen

Nitrate	Solid residue	Gases produced
Copper(II) nitrate	Copper(II) oxide	Nitrogen dioxide and oxygen
Lithium nitrate	Lithium oxide	Nitrogen dioxide and oxygen

4.0 ACTION OF HEAT ON SULPHATES

Sulphates are salts containing the sulphate ion, SO_4^{2-}.

Many sulphates are thermally stable, but some decompose when strongly heated.

The decomposition products depend on the metal present.

4.1 Heating copper(II) sulphate pentahydrate

Copper(II) sulphate pentahydrate is a blue crystalline salt containing five molecules of water of crystallisation.

Its formula is:

$$CuSO_4 \cdot 5H_2O$$

When heated, it loses water of crystallisation and forms white anhydrous copper(II) sulphate.

Chemical equation

$$CuSO_4 \cdot 5H_2O(s) \xrightarrow{\Delta} CuSO_4(s) + 5H_2O(g)$$

Observations:

1. Blue crystals are heated.
2. Water vapour is released.
3. The blue crystals turn white.
4. Water droplets may condense on the cooler part of the apparatus.

Explanation:

The salt loses its water of crystallisation and becomes anhydrous copper(II) sulphate.

Rehydration

When water is added to anhydrous copper(II) sulphate, it becomes blue again.

$$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4 \cdot 5H_2O(s)$$

Application: Anhydrous copper(II) sulphate can be used to test for the presence of water.

4.2 Strong heating of copper(II) sulphate

On very strong heating, anhydrous copper(II) sulphate decomposes to form copper(II) oxide and sulphur trioxide.

$$CuSO_4(s) \xrightarrow{\Delta} CuO(s) + SO_3(g)$$

Sulphur trioxide may further decompose at high temperatures:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

4.3 Heating calcium sulphate

Calcium sulphate is relatively thermally stable. On strong heating, it can decompose to form calcium oxide and sulphur oxides.

A simplified representation under suitable high-temperature conditions is:

$$CaSO_4(s) \xrightarrow{\Delta} CaO(s) + SO_3(g)$$

The exact decomposition pathway depends on the temperature and atmosphere.

4.4 Heating sodium sulphate

Sodium sulphate is relatively stable under ordinary laboratory heating conditions and does not readily decompose.

4.5 Summary of action of heat on selected sulphates

Sulphate	Effect of heating
Copper(II) sulphate pentahydrate	Loses water and turns white
Anhydrous copper(II) sulphate	Decomposes on very strong heating
Calcium sulphate	Relatively stable; decomposes under suitable high-temperature conditions
Sodium sulphate	Relatively stable under ordinary heating

'5.0 PRACTICAL INVESTIGATION: ACTION OF HEAT ON SALTS

Activity 1: Heating selected salts

Learning intention: Investigate the effect of heat on selected salts and identify the products formed.

Requirements:

- Copper(II) carbonate.
- Sodium hydrogen carbonate.
- Copper(II) sulphate pentahydrate.
- Potassium nitrate.
- Test tubes or suitable hard-glass apparatus.
- Test-tube holder.
- Bunsen burner.
- Delivery tubes and suitable gas-testing equipment.
- Limewater.
- Appropriate safety equipment.

Procedure:

1. Place a small quantity of a selected salt in a suitable dry test tube.
2. Hold the test tube using a test-tube holder.
3. Heat gently, then more strongly where appropriate.
4. Observe any changes in colour or appearance.
5. Test any gas produced using an appropriate method.
6. Record the observations.
7. Repeat for the other salts.

Expected observations

Salt heated	Observation

Salt heated	Observation
Copper(II) carbonate	Green solid changes to black; CO_2 released
Sodium hydrogen carbonate	Colourless gases and water vapour released; white residue remains
Copper(II) sulphate pentahydrate	Blue crystals turn white; water released
Potassium nitrate	Oxygen released; white potassium nitrite remains
Lead(II) nitrate	Brown nitrogen dioxide fumes; yellow lead(II) oxide remains

Discussion questions

1. Which salts released carbon dioxide?
2. Which salt changed from blue to white?
3. Which salt released oxygen?
4. What evidence showed that a chemical change had occurred?
5. Why should gases produced during heating be tested carefully?

Safety precautions

- Use small quantities of salts.
- Wear protective goggles and a laboratory coat.
- Point the mouth of the test tube away from people.
- Do not inhale gases directly.
- Do not heat nitrate salts in the presence of combustible substances.
- Toxic gases such as nitrogen dioxide must only be generated in a properly functioning fume cupboard.
- Allow heated apparatus to cool before handling.

6.0 SOLUBILITY OF SALTS IN WATER

6.1 Meaning of solubility

Solubility is the maximum mass of a solute that dissolves in a specified mass of solvent at a particular temperature to form a saturated solution.

In this sub-strand, water is the solvent and salts are the solutes.

Solubility is commonly expressed as:

$$\text{Solubility} = \frac{\text{Mass of solute}}{\text{Mass of water}} \times 100$$

The unit commonly used is grams of salt per 100 g of water.

For example, if 36 g of sodium chloride dissolves in 100 g of water at a particular temperature, its solubility is 36 g per 100 g of water.

6.2 Meaning of a solution

A **solution** is a homogeneous mixture formed when a solute dissolves in a solvent.

Examples:

- Sodium chloride dissolved in water.
- Copper(II) sulphate dissolved in water.
- Potassium nitrate dissolved in water.

A salt dissolves when its particles separate into ions that become surrounded by water molecules.

For example:

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

The symbol (aq) indicates that the ions are dissolved in water.

For copper(II) sulphate:

$$CuSO_4(s) \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$$

6.3 Types of solutions based on the amount of dissolved solute

A. Unsaturated solution

An unsaturated solution contains less dissolved solute than the maximum amount it can dissolve at a given temperature.

Characteristics:

1. More solute can dissolve at the same temperature.
2. No undissolved solute remains if sufficient solvent is present.
3. It can become saturated when more solute is added.

B. Saturated solution

A saturated solution contains the maximum amount of dissolved solute possible at a particular temperature.

Characteristics:

1. It cannot dissolve more solute at that temperature under the same conditions.
2. Excess solute remains undissolved if added.
3. It is in dynamic equilibrium with undissolved solute when both are present.

C. Supersaturated solution

A supersaturated solution contains more dissolved solute than would normally be present in a saturated solution at the same temperature.

Characteristics:

1. It is usually prepared by dissolving excess solute at a higher temperature.
2. It is then cooled carefully without allowing crystals to form.
3. It is unstable.
4. Adding a small crystal of the solute may trigger rapid crystallisation.

6.4 Factors affecting the solubility of salts in water

1. Temperature

Temperature affects the amount of salt that can dissolve in a fixed amount of water.

- The solubility of many salts increases as temperature increases.
- Some salts show very little change in solubility with temperature.
- A few salts become less soluble as temperature increases.

Example: Potassium nitrate becomes much more soluble in water as temperature increases.

2. Nature of the solute

Different salts have different solubilities because of differences in their ionic structures and interactions with water molecules.

For example:

- Sodium chloride is readily soluble in water.
- Potassium nitrate is soluble in water.
- Calcium sulphate is only slightly soluble in water.
- Many carbonates, such as calcium carbonate, are poorly soluble in water.

3. Nature of the solvent

The solubility of a substance depends on the solvent used.

Water dissolves many ionic compounds because its molecules are polar and can attract positive and negative ions.

However, not all salts dissolve readily in water.

4. Presence of other dissolved ions

The presence of other ions in a solution can affect the solubility of a salt.

For example, adding sodium chloride to a solution of silver chloride may reduce the solubility of silver chloride because both substances provide chloride ions.

This is an example of the common-ion effect.

5. Pressure

Pressure has very little effect on the solubility of solid salts in liquids.

Pressure has a much greater effect on the solubility of gases in liquids.

Important distinction: Stirring and crushing a solid salt into smaller particles increase the rate of dissolving, but they do not normally change the equilibrium solubility at a fixed temperature.

7.0 SOLUBILITY OF CHLORIDES, NITRATES, SULPHATES AND CARBONATES

Different salts have different solubilities in water. The general solubility rules help predict whether a salt will dissolve or form a precipitate.

7.1 Solubility of chlorides

Chlorides are salts containing the chloride ion, Cl^-.

Examples:

- Sodium chloride, NaCl.
- Potassium chloride, KCl.
- Calcium chloride, $CaCl_2$.
- Silver chloride, AgCl.
- Lead(II) chloride, $PbCl_2$.

General solubility rules for chlorides

1. Most chlorides are soluble in water.
2. Sodium chloride and potassium chloride are soluble.

3. Most ammonium chlorides are soluble.
4. Silver chloride is insoluble or very slightly soluble in water.
5. Lead(II) chloride is sparingly soluble in cold water but more soluble in hot water.
6. Mercury(I) chloride is insoluble in water.

Examples

Chloride	Solubility in water
Sodium chloride, NaCl	Soluble
Potassium chloride, KCl	Soluble
Calcium chloride, $CaCl_2$	Soluble
Ammonium chloride, NH_4Cl	Soluble
Silver chloride, AgCl	Insoluble
Lead(II) chloride, $PbCl_2$	Sparingly soluble in cold water

Application: The low solubility of silver chloride makes it useful in testing for chloride ions.

7.2 Solubility of nitrates

Nitrates are salts containing the nitrate ion, NO_3^-.

Examples:

- Sodium nitrate, $NaNO_3$.
- Potassium nitrate, KNO_3.
- Calcium nitrate, $Ca(NO_3)_2$.
- Copper(II) nitrate, $Cu(NO_3)_2$.
- Lead(II) nitrate, $Pb(NO_3)_2$.

General solubility rule

All common nitrates are soluble in water.

Examples:

Nitrate	Solubility
Sodium nitrate	Soluble
Potassium nitrate	Soluble
Calcium nitrate	Soluble
Copper(II) nitrate	Soluble
Lead(II) nitrate	Soluble
Silver nitrate	Soluble

Nitrates are therefore commonly used to prepare solutions of metal ions for qualitative analysis.

7.3 Solubility of sulphates

Sulphates are salts containing the sulphate ion, SO_4^{2-}.

Examples:

- Sodium sulphate, Na_2SO_4.
- Potassium sulphate, K_2SO_4.
- Magnesium sulphate, $MgSO_4$.
- Copper(II) sulphate, $CuSO_4$.
- Barium sulphate, $BaSO_4$.
- Calcium sulphate, $CaSO_4$.
- Lead(II) sulphate, $PbSO_4$.

General solubility rules for sulphates

1. Most sulphates are soluble in water.
2. Sodium sulphate and potassium sulphate are soluble.
3. Magnesium sulphate is soluble.
4. Barium sulphate is insoluble.
5. Lead(II) sulphate is insoluble.
6. Calcium sulphate is sparingly soluble.

Examples

Sulphate	Solubility
Sodium sulphate	Soluble
Potassium sulphate	Soluble
Magnesium sulphate	Soluble
Copper(II) sulphate	Soluble
Barium sulphate	Insoluble
Lead(II) sulphate	Insoluble
Calcium sulphate	Sparingly soluble

Application: Barium sulphate is used in medical imaging because it is highly insoluble in water and can be used as a contrast medium under medical supervision.

7.4 Solubility of carbonates

Carbonates are salts containing the carbonate ion, CO_3^{2-}.

Examples:

- Sodium carbonate, Na_2CO_3.
- Potassium carbonate, K_2CO_3.
- Calcium carbonate, $CaCO_3$.
- Magnesium carbonate, $MgCO_3$.
- Copper(II) carbonate, $CuCO_3$.

General solubility rules for carbonates

1. Most carbonates are insoluble in water.

2. Sodium carbonate is soluble.
3. Potassium carbonate is soluble.
4. Ammonium carbonate is soluble.
5. Calcium carbonate is insoluble.
6. Magnesium carbonate is sparingly soluble.
7. Copper(II) carbonate is insoluble.

Examples

Carbonate	Solubility
Sodium carbonate	Soluble
Potassium carbonate	Soluble
Ammonium carbonate	Soluble
Calcium carbonate	Insoluble
Magnesium carbonate	Sparingly soluble
Copper(II) carbonate	Insoluble

7.5 Summary of general solubility rules

Salt group	General rule	Important exceptions
Chlorides	Most are soluble	AgCl insoluble; $PbCl_2$ sparingly soluble
Nitrates	All common nitrates are soluble	No common school-level exception
Sulphates	Most are soluble	$BaSO_4$ and $PbSO_4$ insoluble; $CaSO_4$ sparingly soluble
Carbonates	Most are insoluble	Sodium, potassium and ammonium carbonates soluble

7.6 Importance of solubility rules

Solubility rules are useful because they help learners and chemists to:

1. Predict whether a precipitate will form.
2. Identify unknown ions in solution.
3. Select suitable reagents for qualitative analysis.
4. Prepare soluble salts.
5. Separate ions from mixtures.
6. Understand the formation of mineral deposits.
7. Explain the formation of scale in kettles, pipes and boilers.

8.0 SOLUBILITY CURVES

8.1 Meaning of a solubility curve

A **solubility curve** is a graph showing how the solubility of a substance changes with temperature.

The graph usually has:

- Temperature on the horizontal axis (x-axis).
- Solubility on the vertical axis (y-axis).

Solubility is commonly expressed in grams of solute per 100 g of water.

8.2 Importance of solubility curves

Solubility curves are used to:

1. Compare the solubilities of different salts.
2. Determine the solubility of a salt at a particular temperature.
3. Predict the effect of heating and cooling on solutions.
4. Determine the mass of salt that crystallises when a saturated solution is cooled.
5. Identify temperatures at which two salts have equal solubility.
6. Select suitable temperatures for fractional crystallisation.
7. Estimate the amount of water required to dissolve a given mass of salt.

8.3 Shape and interpretation of solubility curves

Different salts have different solubility curves.

A. Potassium nitrate, KNO_3

- Its solubility increases rapidly with temperature.
- It is much more soluble in hot water than in cold water.
- It is suitable for demonstrating crystallisation by cooling.

B. Sodium chloride, NaCl

- Its solubility changes only slightly with temperature.
- Its curve is nearly horizontal compared with that of potassium nitrate.
- Cooling a hot saturated sodium chloride solution produces a relatively small amount of crystals.

C. Potassium chlorate, $KClO_3$

- Its solubility increases with temperature.
- The increase is more gradual than that of potassium nitrate over much of the usual school-level graph range.

D. Calcium sulphate, $CaSO_4$

- It is only slightly soluble in water.
- Its solubility generally decreases as temperature increases over the commonly considered range.
- Its curve is relatively low on a solubility graph.

E. Sodium carbonate, Na_2CO_3

- Its solubility varies considerably with temperature.
- Its curve is affected by the formation and changes of hydrated crystals.
- The curve is not necessarily a simple steadily rising line over the entire temperature range.

8.4 Illustrative solubility data

The following rounded figures are for learning how to plot and interpret solubility curves. Actual values vary slightly with experimental conditions, the form of the salt and the data source. Sodium carbonate has additional complications due to changes in hydration state.

Approximate solubility in grams of salt per 100 g of water.

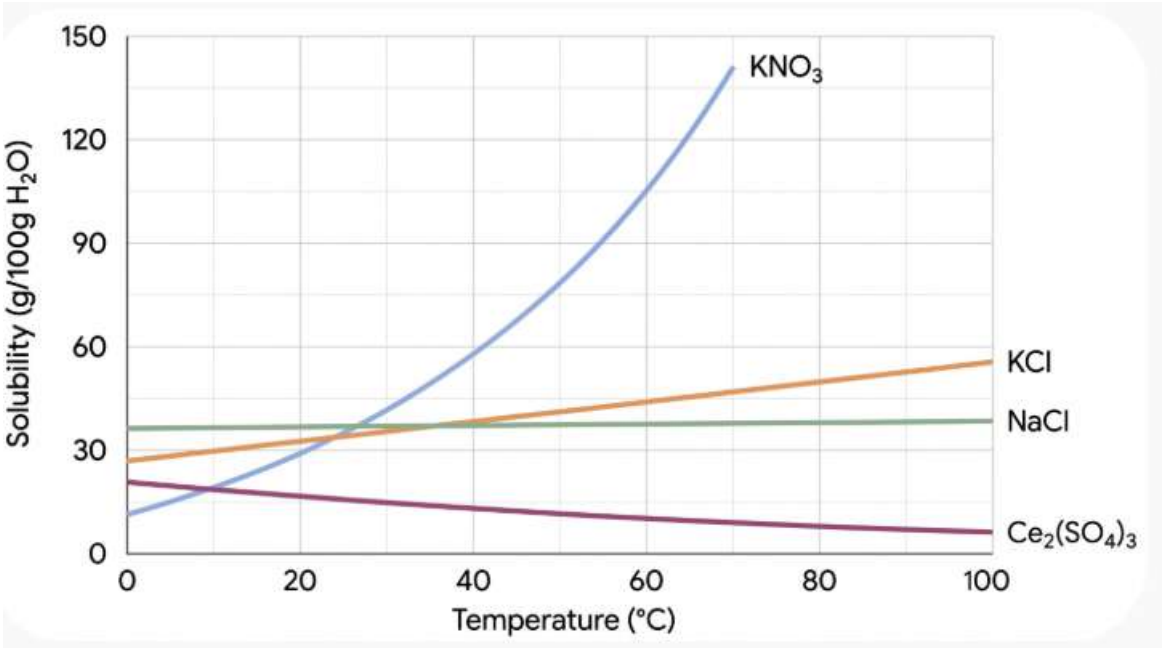
Temperature (°C)	NaCl	KNO_3	$KClO_3$	$CaSO_4$
0	35.7	13.3	3.3	0.20
20	36.0	31.6	7.4	0.21
40	36.6	63.9	13.9	0.20
60	37.3	110	22.4	0.19
80	38.4	169	31.5	0.17
100	39.8	246	44.5	0.16

8.5 How to draw a solubility curve

Procedure:

1. Draw two perpendicular axes.
2. Label the horizontal axis temperature in °C.
3. Label the vertical axis solubility in grams per 100 g of water.
4. Choose suitable scales for both axes.
5. Plot the solubility values for each salt.
6. Join the points using smooth curves.
7. Label each curve clearly.
8. Give the graph a suitable title.

Solubility curves for salts



8.6 Interpreting solubility curves

A learner should be able to use a solubility curve to answer questions such as:

- What is the solubility of potassium nitrate at 40°C?
- Which salt is most soluble at 60°C?
- Which salt shows the greatest increase in solubility with temperature?
- How much potassium nitrate crystallises when a saturated solution is cooled?
- At what temperature do two salts have equal solubility?

8.7 Worked example 1: Reading a solubility curve

A solubility curve shows that 64 g of potassium nitrate dissolves in 100 g of water at 40°C.

Question: How much potassium nitrate can dissolve in 200 g of water at 40°C?

Solution:

100 g of water dissolves 64 g of potassium nitrate.

200 g of water dissolves:

$$\frac{64 \times 200}{100} = 128g$$

Answer

128 g of potassium nitrate.

8.8 Worked example 2: Calculating the mass of crystals formed

A saturated solution contains 110 g of potassium nitrate in 100 g of water at 60°C. The solution is cooled to 20°C, where the solubility is 31.6 g per 100 g of water.

Calculate the mass of potassium nitrate that crystallises out.

Step 1: Identify the mass initially dissolved.

$$110g$$

Step 2: Identify the mass that remains dissolved at 20°C.

$$31.6g$$

Step 3: Calculate the mass crystallised.

$$\text{Mass crystallised} = \text{Initial dissolved mass} - \text{Final dissolved mass}$$

$$= 110 - 31.6$$

$$= 78.4g$$

Answer

78.4 g of potassium nitrate crystallises.

8.9 Worked example 3: Crystallisation from 250 g of water

A saturated potassium nitrate solution contains 110 g of potassium nitrate per 100 g of water at 60°C. It is cooled to 20°C, where the solubility is 31.6 g per 100 g of water.

Calculate the mass of crystals formed when the solution contains 250 g of water.

Step 1: Calculate the initial mass of dissolved salt.

$$\frac{110 \times 250}{100} = 275g$$

Step 2: Calculate the final mass that remains dissolved.

$$\frac{31.6 \times 250}{100} = 79g$$

Step 3: Calculate the mass crystallised.

$$275 - 79 = 196g$$

Answer

196 g of potassium nitrate crystallises.

8.10 Worked example 4: Calculating the mass of water needed

The solubility of potassium nitrate at 40°C is 63.9 g per 100 g of water.

Calculate the mass of water required to dissolve 127.8 g of potassium nitrate at 40°C.

Solution:

$$63.9g \text{ salt} \rightarrow 100g \text{ water}$$

$$127.8g \text{ salt} \rightarrow x$$

$$x = \frac{127.8 \times 100}{63.9}$$

$$x = 200g$$

Answer: 200 g of water is required.

'9.0 CRYSTALLISATION OF SALTS

9.1 Meaning of crystallisation

Crystallisation is the process of obtaining solid crystals of a substance from its solution.

It is commonly used to purify soluble salts.

9.2 Principle of crystallisation

Crystallisation is based on the difference in solubility of a substance at different temperatures.

For many salts:

1. More salt dissolves in hot water than in cold water.
2. A hot saturated solution is prepared.
3. The solution is cooled.
4. The solubility decreases.
5. Excess dissolved salt separates out as crystals.

9.3 Steps involved in crystallisation

Step 1: Preparation of a solution

- Place water in a suitable container.
- Add the required salt.
- Heat and stir until the salt dissolves.

Step 2: Formation of a saturated solution

- Add more salt gradually until no more dissolves at the selected temperature.
- Ensure the solution is saturated.

Step 3: Filtration

- Filter the hot solution if insoluble impurities are present.
- Collect the clear filtrate.

Step 4: Concentration

- Heat the filtrate gently to evaporate some water.
- Stop heating when the solution approaches the crystallisation point.

Step 5: Cooling

- Allow the concentrated solution to cool slowly.
- Crystals begin to form as the solubility decreases.

Step 6: Separation of crystals

- Filter the mixture to separate the crystals from the remaining solution.

Step 7: Washing

- Wash the crystals with a small quantity of cold distilled water where appropriate.
- This removes traces of the mother liquor and soluble impurities.

Step 8: Drying

- Dry the crystals between clean filter papers or using a suitable drying method.
- Avoid excessive heating where the crystals contain water of crystallisation.

9.4 Factors affecting crystal formation

1. **Rate of cooling:** Slow cooling often produces larger, well-formed crystals.
2. **Concentration:** A more concentrated solution is more likely to produce crystals on cooling.
3. **Temperature:** Solubility changes with temperature.
4. **Purity:** Impurities may interfere with crystal formation.

5. **Stirring:** Excessive disturbance may affect the size and formation of crystals.
6. **Time:** Allowing sufficient time may improve crystal growth.

9.5 Importance of crystallisation

1. Purification of salts.
2. Manufacture of crystalline chemicals.
3. Production of fertiliser crystals.
4. Preparation of laboratory reagents.
5. Recovery of dissolved substances from solutions.
6. Production of crystals for industrial and commercial use.

10.0 FRACTIONAL CRYSTALLISATION

10.1 Meaning of fractional crystallisation

Fractional crystallisation is a method of separating two or more dissolved substances based on differences in their solubilities at different temperatures.

It is useful when a solution contains two or more soluble salts.

10.2 Principle of fractional crystallisation

Different salts have different solubilities in water at the same temperature and their solubilities change by different amounts as temperature changes.

When a hot solution containing two or more salts is cooled:

1. The less soluble salt may begin to crystallise first.
2. The crystals are separated by filtration.
3. The remaining solution contains a greater proportion of the more soluble salt.
4. Further cooling or concentration can produce crystals of another salt.

10.3 Example: Separation of potassium nitrate and sodium chloride

Suppose a solution contains potassium nitrate and sodium chloride.

At high temperatures, both salts dissolve in water.

When the solution is cooled:

- The solubility of potassium nitrate decreases greatly.
- The solubility of sodium chloride decreases only slightly.
- A relatively large proportion of potassium nitrate crystallises.
- Most of the sodium chloride remains in solution.

The potassium nitrate crystals can be separated by filtration.

The remaining solution can then be concentrated to recover more sodium chloride.

10.4 Procedure for fractional crystallisation

1. Prepare a hot concentrated solution containing the two salts.
2. Heat until both salts dissolve.
3. Allow the solution to cool slowly.
4. Observe which salt crystallises first.
5. Separate the crystals by filtration.
6. Wash the crystals with a small quantity of cold suitable solvent.
7. Dry the crystals.
8. Further concentrate the filtrate to recover the other salt.

9. Compare the crystals obtained.

10.5 Industrial applications of fractional crystallisation

1. **Fertiliser production:** Used in the purification and separation of selected fertiliser salts.
2. **Chemical manufacturing:** Used to separate and purify soluble chemical compounds.
3. **Food processing:** Used in some processes involving recovery and purification of dissolved substances.
4. **Pharmaceutical production:** Used to purify certain chemical ingredients.
5. **Salt production:** Used in the purification and recovery of selected salts from brines.

10.6 Advantages of fractional crystallisation

1. It can separate substances that are both soluble in water.
2. It can produce relatively pure crystals.
3. It is useful in large-scale industrial processes.
4. It allows recovery of valuable dissolved materials.
5. It reduces wastage of useful chemical substances.

10.7 Limitations of fractional crystallisation

1. It may require large quantities of water.
2. It may take a long time.
3. Some salts have similar solubility curves, making separation difficult.
4. Some salts may form mixed crystals.
5. Repeated crystallisation may be required to obtain high purity.

11.0 PRECIPITATION REACTIONS AND PREPARATION OF INSOLUBLE SALTS

11.1 Meaning of precipitation

Precipitation is the formation of an insoluble solid when two suitable solutions are mixed.

The insoluble solid formed is called a **precipitate**.

A precipitation reaction occurs when ions in solution combine to form a compound that is insoluble or only slightly soluble in water.

11.2 Conditions necessary for precipitation

1. Two solutions containing suitable ions must be mixed.
2. The ions must be able to form an insoluble or sparingly soluble compound.
3. The ionic product must exceed the solubility limit of the compound under the conditions used.

11.3 Examples of precipitation reactions

A. Preparation of silver chloride

When silver nitrate solution is mixed with sodium chloride solution, a white precipitate of silver chloride forms.

Balanced chemical equation:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Net ionic equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Observation: A white precipitate forms.

B. Preparation of barium sulphate

When barium chloride solution is mixed with sodium sulphate solution, a white precipitate of barium sulphate forms.

Balanced chemical equation:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

Net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

Observation: A white precipitate forms.

C. Preparation of calcium carbonate

When calcium chloride solution is mixed with sodium carbonate solution, a white precipitate of calcium carbonate forms.

Balanced chemical equation:

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$$

Net ionic equation:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

D. Preparation of copper(II) hydroxide

When copper(II) sulphate solution is mixed with sodium hydroxide solution, a blue precipitate of copper(II) hydroxide forms.

Balanced chemical equation:

$$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

Net ionic equation:

$$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$$

E. Preparation of lead(II) iodide

When lead(II) nitrate solution is mixed with potassium iodide solution, a bright yellow precipitate of lead(II) iodide forms.

Balanced chemical equation:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Net ionic equation:

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

11.4 Steps involved in preparing an insoluble salt by precipitation

1. Identify the insoluble salt required.
2. Select two soluble salts that provide the required ions.
3. Prepare solutions of the two soluble salts.
4. Mix the solutions in a clean beaker.
5. Stir the mixture to allow the reaction to occur.
6. Allow the precipitate to settle if necessary.
7. Filter the mixture using filter paper and a funnel.
8. Wash the precipitate with distilled water to remove soluble impurities.

9. Dry the precipitate using a suitable method.
10. Store the dry salt in a clean, labelled container.

11.5 Importance of washing a precipitate

Washing removes soluble impurities that may remain attached to the precipitate.

For example, when preparing barium sulphate from barium chloride and sodium sulphate, the filtrate contains sodium chloride.

Some sodium chloride solution may remain trapped around the barium sulphate particles.

Washing with distilled water helps remove these soluble impurities.

11.6 Importance of precipitation reactions

1. Preparation of insoluble salts.
2. Identification of ions during qualitative analysis.
3. Removal of unwanted ions from water.
4. Treatment of industrial wastewater.
5. Purification of chemical solutions.
6. Recovery of useful metals and compounds.
7. Manufacture of pigments and specialised chemicals.

11.7 Worked example: Writing a precipitation equation

Question: Write the balanced chemical equation for the reaction between calcium nitrate and sodium carbonate.

Step 1: Identify the ions.

Calcium nitrate contains:

$$Ca^{2+} \text{ and } NO_3^-$$

Sodium carbonate contains:

$$Na^+ \text{ and } CO_3^{2-}$$

Step 2: Exchange the ions.

The products are calcium carbonate and sodium nitrate.

Step 3: Write the formulae.

$$CaCO_3 \quad \text{and} \quad NaNO_3$$

Step 4: Balance the equation.

$$Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaNO_3(aq)$$

Step 5: Write the net ionic equation.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

12.0 COMPLEX IONS

12.1 Meaning of a complex ion

A **complex ion** is a charged particle formed when a central metal ion combines with surrounding molecules or ions called ligands through coordinate bonds.

A ligand is an ion or molecule that donates a pair of electrons to a central metal ion to form a coordinate bond.

Examples of ligands include:

- Ammonia, NH_3.
- Water, H_2O.
- Chloride ions, Cl^-.
- Hydroxide ions, OH^-.
- Cyanide ions, CN^-.

A complex ion normally contains a central metal ion surrounded by a specific number of ligands.

12.2 Formation of complex ions

A complex ion forms when a ligand donates a lone pair of electrons to a metal ion.

For example, copper(II) ions react with excess ammonia to form a deep-blue complex ion.

$$[Cu(NH_3)_4(H_2O)_2]^{2+}$$

The deep-blue colour is characteristic of the copper(II)-ammonia complex in aqueous solution.

Example 1: Copper(II) ions and ammonia

When a small amount of aqueous ammonia is added to a solution containing copper(II) ions, a pale-blue precipitate of copper(II) hydroxide forms.

$$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$$

When excess aqueous ammonia is added, the precipitate dissolves to form a deep-blue solution containing a copper-ammonia complex.

Observations

1. Initially, a pale-blue precipitate forms.
2. On adding excess ammonia, the precipitate dissolves.
3. A deep-blue solution is formed.

Example 2: Zinc ions and excess ammonia

Zinc ions react with a small quantity of aqueous ammonia to form a white precipitate of zinc hydroxide.

$$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$$

When excess ammonia is added, the white precipitate dissolves to form a colourless solution containing a zinc-ammonia complex.

$$[Zn(NH_3)_4]^{2+}$$

Example 3: Silver ions and ammonia

Silver ions react with chloride ions to form a white precipitate of silver chloride.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Silver chloride dissolves in dilute ammonia solution due to the formation of a colourless diamminesilver(I) complex.

$$AgCl(s) + 2NH_3(aq) \rightleftharpoons [Ag(NH_3)_2]^+(aq) + Cl^-(aq)$$

12.3 Importance of complex ions

1. **Identification of metal ions:** Formation of characteristic coloured complexes helps identify certain metal ions.

2. **Qualitative analysis:** Complex formation is used to distinguish between ions such as copper(II) and zinc(II).
3. **Industrial extraction:** Complex formation is used in some metal extraction and purification processes.
4. **Biological processes:** Complex ions occur in biological substances, including metal-containing proteins.
5. **Water treatment:** Complex formation influences the behaviour and removal of some metal ions.
6. **Chemical manufacturing:** Complex ions are useful in specialised chemical reactions.

13.0 QUALITATIVE ANALYSIS OF SALTS

13.1 Meaning of qualitative analysis

Qualitative analysis is the process of identifying the substances or ions present in a sample without determining their exact quantities.

In salt analysis, qualitative analysis is used to identify:

- Positive ions (cations).
- Negative ions (anions).
- Gases produced during chemical reactions.

For example, a learner may be given an unknown salt and asked to determine whether it contains copper(II) ions, chloride ions or sulphate ions.

13.2 Meaning of cations and anions

A. Cations

Cations are positively charged ions formed when atoms or groups of atoms lose electrons.

Examples:

Cation	Name
Na^+	Sodium ion
K^+	Potassium ion
Ca^{2+}	Calcium ion
Mg^{2+}	Magnesium ion
Ba^{2+}	Barium ion
Zn^{2+}	Zinc ion
Pb^{2+}	Lead(II) ion
Cu^{2+}	Copper(II) ion
Fe^{2+}	Iron(II) ion
Fe^{3+}	Iron(III) ion
Al^{3+}	Aluminium ion

B. Anions

Anions are negatively charged ions formed when atoms or groups of atoms gain electrons.

Examples:

Anion	Name
Cl^-	Chloride ion
SO_4^{2-}	Sulphate ion
SO_3^{2-}	Sulphite ion
NO_3^-	Nitrate ion
CO_3^{2-}	Carbonate ion
HCO_3^-	Hydrogen carbonate ion
OH^-	Hydroxide ion

13.3 Reagents used in qualitative analysis

A reagent is a substance added to a sample to bring about a chemical reaction that helps identify the substances present.

Common reagents include:

1. **Sodium hydroxide solution**
   - ✓ Used to identify several metal cations.
   - ✓ Different ions form precipitates of different colours and solubilities.
2. **Aqueous ammonia**
   - ✓ Used to identify metal ions through precipitation and complex formation.
   - ✓ Some precipitates dissolve in excess ammonia.
3. **Dilute hydrochloric acid**
   - ✓ Used to test for carbonate ions and some other anions.
   - ✓ Carbonates release carbon dioxide.
4. **Dilute nitric acid**
   - ✓ Used to acidify samples before certain anion tests.
   - ✓ It is preferred in the silver nitrate test for halides because it does not introduce chloride ions.
5. **Silver nitrate solution**
   - ✓ Used to test for halide ions, especially chloride ions.
6. **Barium chloride solution**
   - ✓ Used to test for sulphate ions after acidifying the sample.
7. **Lead(II) nitrate solution**
   - ✓ Can be used to demonstrate the presence of certain anions through formation of insoluble lead salts.
8. **Iron(II) sulphate solution and concentrated sulphuric acid**
   - ✓ Used in the brown-ring test for nitrate ions.

14.0 TESTS FOR CATIONS

A cation can be identified by adding suitable reagents and observing the formation, colour and solubility of any precipitate.

The main reagents used in this section are:

- Sodium hydroxide solution, NaOH.
- Aqueous ammonia, $NH_3(aq)$.

14.1 Test for lead(II) ions, Pb^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise to a solution containing lead(II) ions.

Observation: A white precipitate forms.

The precipitate dissolves in excess sodium hydroxide to form a colourless solution.

Ionic equation:

$$Pb^{2+}(aq) + 2OH^{-}(aq) \rightarrow Pb(OH)_2(s)$$

In excess hydroxide ions, the precipitate dissolves to form a soluble hydroxo complex, commonly represented as:

$$Pb(OH)_2(s) + 2OH^{-}(aq) \rightleftharpoons [Pb(OH)_4]^{2-}(aq)$$

Using aqueous ammonia

Add aqueous ammonia dropwise.

Observation: A white precipitate forms. It does not dissolve in excess aqueous ammonia.

Inference: The observations are consistent with the presence of Pb^{2+} ions.

14.2 Test for calcium ions, Ca^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A white precipitate may form, especially at higher concentrations. It does not dissolve in excess sodium hydroxide.

Using aqueous ammonia

Add aqueous ammonia.

Observation: Usually no precipitate forms, or only a very slight precipitate may be observed in sufficiently concentrated solutions.

Inference: The results can help distinguish calcium ions from other cations, but a confirmatory test should be used.

14.3 Test for magnesium ions, Mg^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A white precipitate forms.

The precipitate does not dissolve in excess sodium hydroxide.

Ionic equation:

$$Mg^{2+}(aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2(s)$$

Using aqueous ammonia

A white precipitate may form. It is insoluble in excess ammonia.

Inference: The observations are consistent with Mg^{2+} ions.

14.4 Test for barium ions, Ba^{2+}

Using sodium hydroxide

Add sodium hydroxide solution.

Observation: Usually no visible precipitate forms under ordinary test conditions.

Using aqueous ammonia

Add aqueous ammonia.

Observation: No visible precipitate normally forms.

Confirmatory test

Add dilute sulphuric acid or a suitable soluble sulphate solution.

Observation: A white precipitate of barium sulphate forms.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The white precipitate is insoluble in dilute acids.

14.5 Test for zinc ions, Zn^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A white precipitate forms.

The precipitate dissolves in excess sodium hydroxide to form a colourless solution.

$$Zn^{2+}(aq) + 2OH^{-}(aq) \rightarrow Zn(OH)_2(s)$$

In excess sodium hydroxide:

$$Zn(OH)_2(s) + 2OH^{-}(aq) \rightleftharpoons [Zn(OH)_4]^{2-}(aq)$$

Using aqueous ammonia

Add aqueous ammonia dropwise.

Observation: A white precipitate forms.

In excess aqueous ammonia, the precipitate dissolves to form a colourless solution.

A zinc-ammonia complex is formed:

$$[Zn(NH_3)_4]^{2+}$$

Inference: The formation of a white precipitate that dissolves in excess of both reagents is consistent with zinc ions.

14.6 Test for aluminium ions, Al^{3+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A white gelatinous precipitate forms.

The precipitate dissolves in excess sodium hydroxide to form a colourless solution.

$$Al^{3+}(aq) + 3OH^{-}(aq) \rightarrow Al(OH)_3(s)$$

In excess sodium hydroxide:

$$Al(OH)_3(s) + OH^-(aq) \rightleftharpoons [Al(OH)_4]^-(aq)$$

Using aqueous ammonia

Add aqueous ammonia.

Observation: A white gelatinous precipitate forms.

The precipitate does not dissolve in excess ammonia.

Inference: These observations are consistent with aluminium ions.

14.7 Test for copper(II) ions, Cu^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A pale-blue precipitate forms.

The precipitate does not dissolve in excess sodium hydroxide.

$$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$$

Using aqueous ammonia

Add aqueous ammonia dropwise.

Observation: A pale-blue precipitate forms.

When excess ammonia is added, the precipitate dissolves to form a deep-blue solution.

Inference: This deep-blue solution is characteristic of the copper(II)-ammonia complex.

14.8 Test for iron(II) ions, Fe^{2+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A green precipitate forms.

The precipitate is iron(II) hydroxide.

$$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$$

The green precipitate gradually turns brown on exposure to air because iron(II) hydroxide is oxidised to iron(III) compounds.

Using aqueous ammonia

A green precipitate forms and does not dissolve in excess ammonia.

Inference: The initial green precipitate is consistent with Fe^{2+} ions.

14.9 Test for iron(III) ions, Fe^{3+}

Using sodium hydroxide

Add sodium hydroxide solution dropwise.

Observation: A reddish-brown precipitate forms.

The precipitate is iron(III) hydroxide.

$$Fe^{3+}(aq) + 3OH^{-}(aq) \rightarrow Fe(OH)_3(s)$$

The precipitate does not dissolve in excess sodium hydroxide.

Using aqueous ammonia

A reddish-brown precipitate forms and remains insoluble in excess ammonia.

Inference: The observations are consistent with Fe^{3+} ions.

14.10 Summary table of cation tests

Cation	NaOH added dropwise	In excess NaOH	Aqueous ammonia
Pb^{2+}	White precipitate	Dissolves	White precipitate, insoluble in excess
Ca^{2+}	White precipitate may form	Insoluble	Usually no precipitate
Mg^{2+}	White precipitate	Insoluble	White precipitate, insoluble in excess
Ba^{2+}	Usually no precipitate	—	Usually no precipitate
Zn^{2+}	White precipitate	Dissolves	White precipitate dissolves in excess
Al^{3+}	White gelatinous precipitate	Dissolves	White precipitate, insoluble in excess
Cu^{2+}	Pale-blue precipitate	Insoluble	Deep-blue solution in excess
Fe^{2+}	Green precipitate	Insoluble	Green precipitate, insoluble in excess
Fe^{3+}	Reddish-brown precipitate	Insoluble	Reddish-brown precipitate, insoluble in excess

Note: The exact appearance of a precipitate may depend on concentration, pH and experimental conditions. Confirmatory tests should be used where necessary.

15.0 TESTS FOR ANIONS

Anions can be identified by reacting their solutions with appropriate reagents and observing the products formed.

The anions covered include:

- Chloride, Cl^{-}.
- Sulphite, SO_3^{2-}.
- Sulphate, SO_4^{2-}.
- Nitrate, NO_3^{-}.
- Carbonate, CO_3^{2-}.

15.1 Test for chloride ions, Cl^{-}

Reagents

- Dilute nitric acid.
- Silver nitrate solution.

Procedure

1. Place a small quantity of the salt solution in a test tube.
2. Add a few drops of dilute nitric acid.
3. Add silver nitrate solution.
4. Observe the reaction.

Observation

A white precipitate forms.

The precipitate is silver chloride.

Ionic equation

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Confirmatory test

Add dilute aqueous ammonia to the white precipitate.

Observation: The precipitate dissolves.

Inference: Chloride ions are present.

Explanation: Silver chloride dissolves in dilute ammonia through the formation of a soluble silver-ammonia complex.

15.2 Test for sulphate ions, SO_4^{2-}

Reagents

- Dilute hydrochloric acid.
- Barium chloride solution.

Procedure

1. Place a small quantity of the salt solution in a test tube.
2. Add a few drops of dilute hydrochloric acid.
3. Add barium chloride solution.
4. Observe the reaction.

Observation

A white precipitate forms.

The precipitate is barium sulphate.

Ionic equation

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

Explanation

Barium ions combine with sulphate ions to form insoluble barium sulphate.

The acidification step helps prevent interference from some other anions, such as carbonate and sulphite.

15.3 Test for sulphite ions, SO_3^{2-}

Reagents

Dilute hydrochloric acid or dilute sulphuric acid may be used in a suitable test.

Procedure

1. Place a small quantity of the suspected sulphite salt in a test tube.
2. Add dilute acid.
3. Observe the gas released.
4. Pass the gas into acidified potassium manganate(VII) solution or use another appropriate confirmatory test.

Observation

A colourless gas with a choking smell is produced.

The gas is sulphur dioxide.

Chemical equation

For sodium sulphite:

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2$$

Confirmatory test

Sulphur dioxide decolourises acidified potassium manganate(VII) solution.

Observation: The purple solution becomes colourless.

Explanation: Sulphur dioxide acts as a reducing agent.

Safety: Sulphur dioxide is toxic and irritating. Never smell it directly. Use a fume cupboard or an approved microscale demonstration.

15.4 Test for nitrate ions, NO_3^-

One common confirmatory method is the brown-ring test.

Reagents

- Fresh iron(II) sulphate solution.
- Concentrated sulphuric acid.

Procedure

1. Place the nitrate solution in a test tube.
2. Add freshly prepared iron(II) sulphate solution.
3. Carefully run concentrated sulphuric acid down the side of the inclined test tube so that it forms a separate lower layer.
4. Do not shake the test tube.
5. Observe the boundary between the two layers.

Observation

A brown ring forms at the junction of the two liquid layers.

Explanation

Nitrate ions are reduced to nitric oxide under the strongly acidic conditions. The nitric oxide forms a brown complex with iron(II) ions.

Inference: The formation of a brown ring indicates nitrate ions.

Safety: Concentrated sulphuric acid is highly corrosive. This test should be conducted only by a trained teacher using appropriate protective equipment.

15.5 Test for carbonate ions, CO_3^{2-}

Reagents

Dilute hydrochloric acid and limewater.

Procedure

1. Place a small amount of the suspected carbonate in a test tube.
2. Add dilute hydrochloric acid.
3. Observe any effervescence.
4. Pass the gas produced into limewater.

Observation

1. Effervescence occurs.
2. A colourless gas is produced.
3. Limewater turns milky.

Chemical equation

For calcium carbonate:

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

The carbon dioxide reacts with limewater:

$$CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$$

The calcium carbonate formed is an insoluble white precipitate that makes limewater appear milky.

Inference: The results indicate the presence of carbonate ions.

15.6 Summary of anion tests

Anion	Reagent	Observation	Inference
Cl^-	Dilute nitric acid, then silver nitrate	White precipitate; dissolves in dilute ammonia	Chloride present
SO_4^{2-}	Dilute hydrochloric acid, then barium chloride	White precipitate	Sulphate present
SO_3^{2-}	Dilute acid	Colourless choking gas; decolourises acidified manganate(VII)	Sulphite present
NO_3^-	Iron(II) sulphate and carefully layered concentrated sulphuric acid	Brown ring at the interface	Nitrate present
CO_3^{2-}	Dilute acid; gas passed into limewater	Effervescence; limewater turns milky	Carbonate present

16.0 TESTS FOR GASES EVOLVED DURING SALT ANALYSIS

Some salts release gases when they react with acids or are heated.

The gases can be identified using characteristic tests.

16.1 Test for carbon dioxide

Test: Bubble the gas through limewater.

Observation: Limewater turns milky.

Equation:

$$CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$$

Inference: Carbon dioxide is present.

16.2 Test for sulphur dioxide

Test: Pass the gas through acidified potassium manganate(VII) solution.

Observation: Purple potassium manganate(VII) solution becomes colourless.

Inference: Sulphur dioxide is present.

16.3 Test for ammonia

Ammonia is released when ammonium salts are warmed with an alkali such as sodium hydroxide.

Procedure:

1. Add sodium hydroxide solution to the suspected ammonium salt.
2. Warm gently.
3. Test the gas using damp red litmus paper.

Observation: Damp red litmus paper turns blue.

Chemical equation:

$$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$$

Inference: Ammonia gas is present.

16.4 Test for oxygen

Test: Insert a glowing wooden splint into the gas.

Observation: The splint relights.

Inference: Oxygen is present.

16.5 Test for hydrogen

Test: Apply a burning splint to a small sample of the gas.

Observation: A squeaky pop is produced.

Inference: Hydrogen is present.

16.6 Test for hydrogen sulphide

Hydrogen sulphide may be produced when certain sulphides react with dilute acids.

Test: Use moist lead(II) ethanoate paper.

Observation: The paper turns black due to formation of lead(II) sulphide.

Inference: Hydrogen sulphide is present.

Safety: Hydrogen sulphide is highly toxic. It must not be deliberately smelled or generated in an ordinary open laboratory.

16.7 Summary of gas tests

Gas	Test	Positive observation

Gas	Test	Positive observation
CO_2	Bubble through limewater	Limewater turns milky
SO_2	Acidified potassium manganate(VII)	Purple solution becomes colourless
NH_3	Damp red litmus paper	Turns blue
O_2	Glowing splint	Relights
H_2	Burning splint applied to a small sample	Squeaky pop
H_2S	Moist lead(II) ethanoate paper	Turns black

17.0 FLAME TESTS FOR METALLIC IONS

17.1 Meaning of a flame test

A **flame test** is a qualitative analytical method used to identify certain metal ions by observing the characteristic colour they produce when introduced into a non-luminous flame.

When a metal ion is heated:

1. Energy from the flame is absorbed by electrons.
2. Some electrons move to higher energy levels.
3. When the electrons return to lower energy levels, energy is released.
4. Some of this energy is released as visible light.
5. The light produces a characteristic flame colour.

17.2 Apparatus and materials

- Bunsen burner.
- Clean nichrome or platinum wire loop.
- Dilute hydrochloric acid for cleaning the wire.
- Samples of suitable salts.
- Safety goggles.
- Laboratory coat.

17.3 Procedure for carrying out a flame test

1. Light the Bunsen burner and obtain a non-luminous blue flame.
2. Clean the wire loop using dilute hydrochloric acid.
3. Heat the wire loop until it produces no characteristic flame colour.
4. Dip the clean wire loop into the salt sample.
5. Introduce the sample into the edge of the blue flame.
6. Observe the flame colour.
7. Record the result.
8. Clean the wire loop thoroughly before testing another salt.

17.4 Flame colours of selected metal ions

A. Lithium ions, Li^+

Flame colour: Crimson red.

Lithium salts produce a crimson-red flame.

B. Sodium ions, Na^+

Flame colour: Golden yellow.

Sodium produces an intense yellow flame.

Even small amounts of sodium contamination can interfere with other flame colours.

C. Potassium ions, K^+

Flame colour: Lilac.

Potassium produces a pale lilac flame.

A cobalt-blue glass may be used to reduce interference from sodium's strong yellow emission.

D. Calcium ions, Ca^{2+}

Flame colour: Brick red or orange-red.

Calcium salts produce a characteristic orange-red flame.

E. Barium ions, Ba^{2+}

Flame colour: Apple green.

Barium salts produce a pale green or apple-green flame.

17.5 Summary of flame tests

Metal ion	Flame colour
Lithium, Li^+	Crimson red
Sodium, Na^+	Golden yellow
Potassium, K^+	Lilac
Calcium, Ca^{2+}	Brick red/orange-red
Barium, Ba^{2+}	Apple green

17.6 Importance of flame tests

1. Identification of selected metal ions.
2. Qualitative analysis of salts.
3. Demonstration of electronic energy changes.
4. Introduction to atomic emission spectra.
5. Identification of certain metallic elements in laboratory samples.

17.7 Limitations of flame tests

1. Some metal ions produce similar flame colours.
2. Sodium contamination can mask other colours.
3. Not all metal ions produce easily visible flame colours.
4. Flame colours may be affected by the concentration of the sample.
5. Flame tests alone may not provide a definite identification.

18.0 APPLICATIONS OF SALTS IN KENYA

Salts have numerous applications in Kenyan households, farms, schools, hospitals and industries.

18.1 Agriculture

1. Ammonium nitrate supplies nitrogen to crops.
2. Potassium nitrate supplies nitrogen and potassium.
3. Calcium salts are used in some soil improvement practices.
4. Magnesium salts provide magnesium needed for chlorophyll formation.
5. Phosphate salts supply phosphorus for plant growth.

18.2 Food industry

1. Sodium chloride is used in seasoning and food preservation.
2. Sodium hydrogen carbonate is used in baking.
3. Selected salts are used as food additives under approved standards.
4. Salts are used in food processing and manufacturing.

18.3 Construction industry

1. Calcium carbonate is used in cement manufacture.
2. Calcium sulphate is used in the manufacture of plaster and related building materials.
3. Calcium oxide is used in lime production.
4. Sodium carbonate is used in glass manufacture.

18.4 Medicine and healthcare

1. Sodium chloride is used in the preparation of saline solutions.
2. Certain salts are used in oral rehydration formulations.
3. Calcium salts are used in selected medical preparations.
4. Barium sulphate is used as a contrast medium in certain medical imaging procedures.

18.5 Water treatment

1. Some aluminium salts are used as coagulants in water treatment.
2. Calcium compounds are used in water treatment and pH adjustment.
3. Certain salts are used to remove unwanted ions from water.
4. Chemical precipitation is used in some wastewater treatment processes.

18.6 Industrial manufacturing

1. Sodium carbonate is used in glass manufacture.
2. Sodium chloride is an important raw material in the chlor-alkali industry.
3. Copper(II) sulphate is used in selected industrial and agricultural processes.
4. Potassium nitrate is used in fertiliser manufacture.
5. Calcium carbonate is used in the manufacture of cement and lime.

19.0 LABORATORY SAFETY AND ENVIRONMENTAL CONSIDERATIONS

The investigation of salts involves heating chemicals, preparing solutions, carrying out precipitation reactions and testing gases. Learners must observe laboratory safety at all times.

19.1 General laboratory safety precautions

1. Wear a laboratory coat, safety goggles and appropriate protective clothing.
2. Follow instructions given by the teacher.
3. Use small quantities of chemicals during experiments.
4. Never taste laboratory chemicals.
5. Avoid direct inhalation of gases.
6. Do not touch chemicals with bare hands.
7. Use clean apparatus to avoid contamination.

8. Label all prepared solutions clearly.
9. Keep flammable materials away from flames.
10. Wash hands thoroughly after practical activities.
11. Dispose of chemical waste according to laboratory procedures.
12. Report spills and accidents immediately.

19.2 Safety when heating salts

1. Use a suitable heat-resistant test tube.
2. Hold the test tube with a test-tube holder.
3. Point the mouth of the test tube away from people.
4. Heat small quantities of chemicals.
5. Never look directly into a heated test tube.
6. Allow hot apparatus to cool before handling.
7. Do not inhale gases produced during heating.
8. Perform experiments that produce toxic gases only in a suitable fume cupboard.

19.3 Environmental effects of salts

Some salts are beneficial, while excessive quantities may cause environmental problems.

A. Effects of excessive fertiliser salts

1. Excess nitrates can be washed into rivers, lakes and dams.
2. High nitrate levels may contaminate drinking water.
3. Excess nutrients can cause eutrophication.
4. Eutrophication may lead to excessive growth of algae.
5. Decomposition of dead algae can reduce dissolved oxygen in water.
6. Aquatic organisms may die due to reduced oxygen levels.

B. Effects of excessive sodium salts in soil

7. Excess sodium can damage soil structure.
8. It may reduce water infiltration.
9. It can make it difficult for plant roots to absorb water.
10. It may reduce agricultural productivity.
11. Proper soil and water management helps control salinity and sodicity.

C. Effects of industrial salt waste

1. Some industrial salts may contaminate rivers and groundwater.
2. Certain metal salts may be toxic to aquatic organisms.
3. Improper disposal may change the chemical composition of soil.
4. Some dissolved salts may interfere with water treatment.
5. Industrial wastewater should be treated before discharge.

19.4 Responsible management of salt waste

1. Avoid releasing concentrated chemical solutions into drainage systems.
2. Collect heavy-metal-containing waste separately.
3. Reuse and recycle chemicals where appropriate.
4. Treat industrial wastewater before discharge.
5. Follow environmental regulations.
6. Use the smallest practical quantity of chemicals during experiments.

20.0 COMPETENCY-BASED LEARNING ACTIVITIES

The following activities support the development of scientific knowledge, practical skills, values and core competencies.

Activity 1: Investigating the action of heat on salts

Task: In groups, investigate the effect of heat on selected carbonates, hydrogen carbonates, nitrates and sulphates.

Learners should:

1. Identify the salts being heated.
2. Observe changes in colour.
3. Identify gases produced.
4. Record the solid residues.
5. Write balanced chemical equations.
6. Explain the environmental effects of gases released.

Activity 2: Investigating solubility of salts

Task: Investigate the solubility of sodium chloride, potassium nitrate and other selected salts in water.

Learners should:

1. Measure equal masses of water.
2. Add known masses of selected salts.
3. Stir until the salt dissolves or no more dissolves.
4. Record the maximum mass dissolved.
5. Repeat at different temperatures.
6. Compare the results.
7. Explain the differences in solubility.

Activity 3: Drawing solubility curves

Task: Use experimental data or supplied solubility data to draw solubility curves.

Learners should:

1. Prepare a suitable data table.
2. Select appropriate graph scales.
3. Plot temperature against solubility.
4. Draw smooth curves.
5. Interpret the curves.
6. Estimate the amount of salt crystallised when a solution is cooled.

Activity 4: Preparing an insoluble salt

Task: Prepare barium sulphate or calcium carbonate through a precipitation reaction under teacher supervision.

Learners should:

1. Select suitable soluble reactants.
2. Prepare the required solutions.
3. Mix the solutions.
4. Observe the precipitate.
5. Filter the mixture.
6. Wash and dry the precipitate.

7. Write the balanced chemical and ionic equations.

Activity 5: Qualitative analysis of salts

Task: Use appropriate reagents to identify selected cations and anions in unknown salt solutions.

Learners should:

1. Carry out preliminary observations.
2. Add sodium hydroxide solution.
3. Add aqueous ammonia.
4. Test for anions using suitable reagents.
5. Record observations.
6. Draw conclusions.
7. Present their findings.

Activity 6: Flame tests

Task: Carry out flame tests on selected salts containing lithium, sodium, potassium, calcium and barium ions.

Learners should:

1. Clean the wire loop.
2. Introduce the salt sample into the flame.
3. Observe the flame colour.
4. Record the results.
5. Compare the flame colours.
6. Explain how flame tests help identify metal ions.

Activity 7: Investigating the importance of salts in Kenya

Task: In groups, identify examples of salts used in agriculture, food production, construction, medicine and water treatment in Kenya.

Learners should:

1. Identify common salts.
2. State their chemical formulae.
3. Explain their uses.
4. Discuss environmental effects of improper disposal.
5. Present recommendations for responsible use.

21.0 KEY FORMULAE AND EQUATIONS TO REMEMBER

Reaction	Chemical equation
Heating calcium carbonate	$CaCO_3 \rightarrow CaO + CO_2$
Heating copper(II) carbonate	$CuCO_3 \rightarrow CuO + CO_2$

Reaction	Chemical equation
Heating zinc carbonate	$ZnCO_3 \rightarrow ZnO + CO_2$
Heating sodium hydrogen carbonate	$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
Heating potassium nitrate	$2KNO_3 \rightarrow 2KNO_2 + O_2$
Heating lead(II) nitrate	$2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$
Dehydration of copper(II) sulphate	$CuSO_4 \cdot 5H_2O \rightarrow CuSO_4 + 5H_2O$
Formation of silver chloride	$Ag^+ + Cl^- \rightarrow AgCl$
Formation of barium sulphate	$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$
Formation of calcium carbonate	$Ca^{2+} + CO_3^{2-} \rightarrow CaCO_3$
Formation of copper(II) hydroxide	$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$
Formation of lead(II) iodide	$Pb^{2+} + 2I^- \rightarrow PbI_2$
Carbonate test	$CO_3^{2-} + 2H^+ \rightarrow CO_2 + H_2O$
Sulphite test	$SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$

Note: Heating symbols and state symbols may be added when writing complete chemical equations.

22.0 IMPORTANT TERMS AND DEFINITIONS

Term	Meaning
Salt	An ionic compound formed when replaceable hydrogen ions of an acid are replaced by metal or ammonium ions
Solubility	Maximum mass of solute that dissolves in a specified amount of solvent at a given temperature
Solute	Substance that dissolves in a solvent
Solvent	Substance that dissolves a solute
Solution	Homogeneous mixture of a solute and solvent
Saturated solution	Solution containing the maximum amount of dissolved solute at a given temperature
Unsaturated solution	Solution that can dissolve more solute at a given temperature
Supersaturated solution	Unstable solution containing more dissolved solute than normally possible at a given temperature
Thermal decomposition	Breakdown of a compound into simpler substances by heating
Crystallisation	Process of obtaining crystals from a solution
Fractional crystallisation	Separation of soluble substances based on differences in their solubilities
Precipitation	Formation of an insoluble solid from a solution
Precipitate	Insoluble solid formed during a chemical reaction in solution
Cation	Positively charged ion
Anion	Negatively charged ion
Qualitative analysis	Identification of substances present in a sample

Term	Meaning
	without measuring their quantities
Reagent	Substance used to bring about a chemical reaction during analysis
Complex ion	Charged species containing a central metal ion bonded to ligands
Ligand	Ion or molecule that donates an electron pair to a central metal ion
Flame test	Test that identifies selected metal ions using characteristic flame colours
Hydrated salt	Salt containing water molecules in its crystal structure
Anhydrous salt	Salt without water of crystallisation

23.0 REVISION QUESTIONS

SECTION A: MULTIPLE-CHOICE QUESTIONS

Choose the correct answer from the alternatives provided.

1. A salt is best described as an ionic compound formed when replaceable hydrogen ions of an acid are replaced by:

A. Oxygen atoms only
B. Metal or ammonium ions
C. Hydrogen molecules
D. Non-metal atoms only

2. Which of the following salts is formed when hydrochloric acid reacts with sodium hydroxide?

A. Sodium sulphate
B. Sodium chloride
C. Sodium nitrate
D. Sodium carbonate

3. What are the products formed when calcium carbonate is strongly heated?

A. Calcium and oxygen
B. Calcium oxide and carbon dioxide
C. Calcium hydroxide and oxygen
D. Calcium chloride and carbon dioxide

4. Which gas is released when sodium hydrogen carbonate is heated?

A. Hydrogen
B. Nitrogen
C. Carbon dioxide
D. Chlorine

5. When copper(II) carbonate is heated, the green solid changes to:

A. White copper
B. Black copper(II) oxide
C. Blue copper sulphate
D. Brown copper metal

6. What is the solid product formed when potassium nitrate is heated?

A. Potassium oxide
B. Potassium nitrite
C. Potassium hydroxide
D. Potassium carbonate

7. Which gas produces brown fumes during the thermal decomposition of lead(II) nitrate?

A. Oxygen
B. Hydrogen
C. Nitrogen dioxide
D. Carbon dioxide

8. What happens when blue copper(II) sulphate pentahydrate is heated?

A. It turns green.
B. It turns black.
C. It turns white.
D. It turns red.

9. Which of the following salts is generally insoluble in water?

A. Sodium chloride
B. Potassium nitrate
C. Barium sulphate
D. Ammonium chloride

10. Which group of salts is generally soluble in water?

A. Carbonates
B. Nitrates
C. Most metal hydroxides
D. Most metal sulphides

11. Which of the following chlorides is insoluble in water?

A. Sodium chloride
B. Potassium chloride
C. Silver chloride
D. Calcium chloride

12. Which salt is sparingly soluble in water?

A. Calcium sulphate
B. Sodium nitrate
C. Potassium chloride
D. Ammonium nitrate

13. A saturated solution is one that:

A. Contains no dissolved solute.
B. Contains the maximum amount of dissolved solute at a given temperature.
C. Contains only water.
D. Always contains an undissolved gas.

14. Which of the following factors does not normally affect the equilibrium solubility of a solid salt in water?

A. Temperature
B. Nature of the solute
C. Nature of the solvent
D. Stirring speed

15. A solubility curve is a graph showing the relationship between:

A. Pressure and volume
B. Temperature and solubility
C. Mass and density
D. Time and reaction rate

16. Which salt shows a particularly large increase in solubility as temperature increases?

A. Potassium nitrate
B. Calcium sulphate
C. Sodium chloride
D. Barium sulphate

17. The solubility of sodium chloride changes with temperature:

A. Very rapidly.
B. Only slightly.
C. By decreasing to zero.
D. By increasing without limit.

18. A solution containing more dissolved solute than normally possible at a given temperature is called:

A. Unsaturated solution
B. Dilute solution
C. Supersaturated solution
D. Suspension

19. Crystallisation is mainly used to:

A. Produce gases from liquids.
B. Purify soluble solid substances.
C. Separate liquids by boiling point.
D. Change metals into non-metals.

20. Fractional crystallisation separates substances based on differences in their:

A. Densities only
B. Solubilities at different temperatures
C. Colours only
D. Electrical conductivity

21. What is formed when silver nitrate solution reacts with sodium chloride solution?

A. Silver chloride precipitate
B. Sodium chloride precipitate
C. Silver metal
D. Sodium metal

22. Which of the following precipitates is white?

A. Copper(II) hydroxide
B. Iron(III) hydroxide
C. Silver chloride
D. Lead(II) iodide

23. Which reagent is commonly used to test for sulphate ions?

A. Silver nitrate solution
B. Barium chloride solution
C. Sodium chloride solution
D. Potassium iodide solution

24. Which gas is released when a carbonate reacts with dilute hydrochloric acid?

A. Oxygen
B. Nitrogen
C. Carbon dioxide
D. Hydrogen

25. A white precipitate that dissolves in excess sodium hydroxide and excess ammonia may indicate:

A. Zinc ions
B. Iron(III) ions
C. Barium ions
D. Sodium ions

26. Which cation produces a pale-blue precipitate with sodium hydroxide?

A. Fe^{3+}
B. Cu^{2+}
C. Zn^{2+}
D. Mg^{2+}

27. Which cation produces a green precipitate with sodium hydroxide that gradually turns brown on exposure to air?

A. Fe^{2+}
B. Fe^{3+}
C. Cu^{2+}
D. Al^{3+}

28. Which cation produces a reddish-brown precipitate with sodium hydroxide?

A. Calcium
B. Magnesium
C. Iron(III)
D. Zinc

29. A white precipitate formed when a solution is treated with silver nitrate after acidification with dilute nitric acid indicates the presence of:

A. Sulphate ions
B. Chloride ions
C. Nitrate ions
D. Carbonate ions

30. Which test is used to identify nitrate ions?

A. Limewater test
B. Brown-ring test
C. Glowing splint test
D. Flame test

31. Which metal ion produces a lilac flame?

A. Sodium
B. Potassium
C. Calcium
D. Lithium

32. Which metal ion produces an apple-green flame?

A. Barium
B. Sodium
C. Lithium
D. Calcium

33. Which metal ion produces a golden-yellow flame?

A. Potassium
B. Lithium
C. Sodium
D. Barium

34. What is the flame colour of lithium ions?

A. Lilac
B. Crimson red
C. Apple green
D. Golden yellow

35. A complex ion is formed when:

A. A metal ion combines with ligands.
B. Two gases are mixed.
C. A salt is heated.
D. A solid dissolves without forming ions.

36. When excess aqueous ammonia is added to copper(II) ions, the final solution becomes:

A. Green
B. Colourless
C. Deep blue
D. Yellow

37. Which reagent is used in the brown-ring test for nitrates together with concentrated sulphuric acid?

A. Sodium hydroxide
B. Iron(II) sulphate
C. Silver nitrate
D. Barium chloride

38. What happens when sulphur dioxide is passed through acidified potassium manganate(VII) solution?

A. It turns blue.
B. It becomes colourless.
C. It forms a yellow precipitate.
D. It produces a green flame.

39. Which of the following is an environmental effect of excessive nitrate fertiliser use?

A. Eutrophication
B. Increased oxygen in all water bodies
C. Elimination of all soil salts
D. Prevention of algal growth

40. Which statement about precipitation reactions is correct?

A. They always produce gases only.
B. They involve the formation of an insoluble solid from solutions.
C. They occur only between metals and oxygen.
D. They never involve ions.

SECTION B: STRUCTURED QUESTIONS

1. Introduction to salts
   g) Define a salt.

   h) Explain how salts are formed through neutralisation.

   i) Distinguish between normal salts and acid salts.

   j) Give four examples of salts used in everyday life.

   k) State five uses of salts in Kenya.

2. Thermal decomposition of carbonates
   l) Define thermal decomposition.

   m) State the products formed when calcium carbonate is strongly heated.

   n) Write a balanced chemical equation for the reaction.

   o) Describe how the gas produced can be identified.

   p) State three industrial uses of calcium oxide.

   q) Explain why calcium carbonate is important in the cement industry.

3. Thermal decomposition of hydrogen carbonates
   r) What is a hydrogen carbonate?

   s) Describe what happens when sodium hydrogen carbonate is heated.

   t) Write a balanced chemical equation for the reaction.

   u) State three observations made during the experiment.

   v) Explain one use of sodium hydrogen carbonate in everyday life.

4. Thermal decomposition of nitrates
   a) What is a nitrate?
   b) Describe the action of heat on potassium nitrate.
   c) Write a balanced chemical equation for the reaction.
   d) State the products formed when lead(II) nitrate is heated.
   e) Write a balanced equation for the decomposition of lead(II) nitrate.
   f) Explain why the heating of lead(II) nitrate must be carried out in a fume cupboard.
5. Thermal decomposition of sulphates
   a) State the formula of copper(II) sulphate pentahydrate.
   b) Describe what happens when blue copper(II) sulphate crystals are heated.
   c) Write a balanced chemical equation for the reaction.
   d) Explain how anhydrous copper(II) sulphate can be used to test for water.
   e) State two observations that may be made during the experiment.
6. Solubility of salts
   a) Define solubility.
   b) Distinguish between saturated and unsaturated solutions.
   c) Explain the meaning of a supersaturated solution.
   d) State four factors that affect the solubility of salts in water.
   e) Explain why stirring increases the rate of dissolving but does not normally change equilibrium solubility.
7. General solubility rules
   a) State the general solubility rule for nitrates.
   b) State two chlorides that are insoluble or sparingly soluble in water.
   c) Name three sulphates that are soluble in water.
   d) Name two carbonates that are soluble in water.
   e) Explain how solubility rules are used in qualitative analysis.
8. Solubility curves
   a) Define a solubility curve.
   b) State the quantities represented on the horizontal and vertical axes.

c) Explain why potassium nitrate has a steeply rising solubility curve.

d) Explain why sodium chloride has a relatively flat solubility curve.

e) State four uses of solubility curves.

9. Solubility calculations

The solubility of potassium nitrate is 110 g per 100 g of water at 60°C and 31.6 g per 100 g of water at 20°C.

a) Calculate the mass of potassium nitrate that can dissolve in 200 g of water at 60°C.

b) Calculate the mass of potassium nitrate that remains dissolved in 200 g of water at 20°C.

c) Calculate the mass of potassium nitrate that crystallises when the saturated solution is cooled from 60°C to 20°C.

10. Another solubility calculation

A saturated solution contains 169 g of potassium nitrate dissolved in 100 g of water at 80°C. The solution is cooled to 40°C, where the solubility is 63.9 g per 100 g of water.

a) Calculate the mass of potassium nitrate that remains dissolved at 40°C.

b) Calculate the mass of potassium nitrate that crystallises.

c) Explain why crystals form when the solution is cooled.

11. Crystallisation

a) Define crystallisation.

b) Explain the principle of crystallisation.

c) Describe the steps followed when preparing pure copper(II) sulphate crystals from a suitable solution.

d) State four factors that affect crystal formation.

e) Explain why crystals are washed with a small quantity of cold distilled water.

12. Fractional crystallisation

a) Define fractional crystallisation.

b) Explain how fractional crystallisation can be used to separate potassium nitrate and sodium chloride.

c) State four steps involved in fractional crystallisation.

d) Give three industrial applications of fractional crystallisation.

e) State three limitations of fractional crystallisation.

13. Precipitation reactions

a) Define a precipitation reaction.

b) Explain the meaning of a precipitate.

c) Describe how barium sulphate can be prepared from suitable soluble salts.

d) Write the balanced chemical equation for the reaction.

e) Write the net ionic equation.

f) Explain why the precipitate should be washed before drying.

14. Preparation of insoluble salts

A learner is required to prepare calcium carbonate in the laboratory.

a) Name two soluble salts that can be used.

b) Describe the procedure for preparing calcium carbonate.

c) Write the balanced chemical equation.

d) Write the net ionic equation.

e) State three observations.

f) Explain how the precipitate can be purified.

15. Complex ions

a) Define a complex ion.

b) Explain the meaning of a ligand.

c) Describe what happens when aqueous ammonia is added dropwise to a solution containing copper(II) ions.

d) Explain what happens when excess ammonia is added.

e) State two other examples of complex ions.

f) Give three applications of complex ions.

16. Tests for cations

a) Explain how sodium hydroxide solution can be used to identify metal ions.

b) State the observation when sodium hydroxide is added to a solution containing:

i. Copper(II) ions.

ii. Iron(II) ions.

iii. Iron(III) ions.

iv. Zinc ions.

v. Aluminium ions.

c) Explain how zinc ions can be distinguished from aluminium ions using aqueous ammonia.

d) State the observation when excess sodium hydroxide is added to a solution containing lead(II) ions.

17. Further cation analysis

A learner adds sodium hydroxide solution to four separate solutions labelled A, B, C and D.

- A produces a pale-blue precipitate.
- B produces a green precipitate that turns brown on standing.
- C produces a reddish-brown precipitate.
- D produces a white precipitate that dissolves in excess sodium hydroxide and excess ammonia.

a) Identify the likely cation in each solution.

b) Name the precipitate formed in A, B and C.

c) Explain the colour change observed in B.

d) Explain the behaviour of D in excess sodium hydroxide and ammonia.

18. Tests for anions

a) Describe the test for chloride ions.

b) State the expected observation.

c) Describe the test for sulphate ions.

d) State the expected observation.

e) Explain how carbonate ions can be identified.

f) Describe the test for nitrate ions.

19. Gas tests during salt analysis

a) State the test for carbon dioxide.

b) State the test for sulphur dioxide.

c) State the test for ammonia.

d) Explain how oxygen can be identified.

e) Explain how hydrogen can be identified.

f) State one safety precaution when testing gases produced during salt analysis.

20. Flame tests

a) Define a flame test.

b) Explain why metal ions produce different flame colours.

c) State the flame colours produced by:

i. Lithium ions.

ii. Sodium ions.

iii. Potassium ions.

iv. Calcium ions.

v. Barium ions.

d) Describe the procedure for carrying out a flame test.

e) State three limitations of flame tests.

21. Environmental effects of salts

a) Explain how excessive use of nitrate fertilisers can affect rivers and lakes.

b) Define eutrophication.

c) State four environmental effects of excessive salts in agricultural soils.

d) Explain three ways of managing salt-containing industrial waste.

e) Suggest four ways farmers can use fertilisers responsibly.

22. Application of salts in Kenya

a) Name two salts used in agriculture.

b) State three uses of sodium chloride.

c) Explain the importance of calcium carbonate in construction.

d) State three applications of salts in water treatment.

e) Explain why salts are important in medicine and healthcare.

23. Practical investigation

A learner is provided with an unknown white salt labelled X.

The learner carries out the following tests:

- When dilute hydrochloric acid is added, a colourless gas is released.
- The gas turns limewater milky.
- When a solution of X is treated with calcium chloride solution, a white precipitate forms.

a) Identify the likely anion in X.

b) Name the gas released.

c) Write a balanced ionic equation for the reaction with acid.

d) Write a balanced ionic equation for the formation of the precipitate with calcium ions.

e) State two other salts that may contain the same anion.

24. Integrated salt analysis

A learner is provided with a solution containing an unknown salt Y. The following observations are made:

- Addition of sodium hydroxide produces a white precipitate.
- The precipitate dissolves in excess sodium hydroxide.
- Addition of excess aqueous ammonia also dissolves the white precipitate.
- Addition of dilute nitric acid followed by silver nitrate produces a white precipitate.

a) Identify the likely cation in Y.

b) Identify the anion in Y.

c) Name the salt Y.

d) Write the ionic equation for the formation of the metal hydroxide precipitate.

e) Write the ionic equation for the formation of the white precipitate with silver nitrate.

f) State why the results of more than one test are important in identifying an unknown salt.

STRAND 2.0: PHYSICAL CHEMISTRY

SUB-STRAND 2.2: GAS LAWS

INTRODUCTION

Gases are an important state of matter found in everyday life, industries, agriculture, medicine, transport and environmental processes. Air, oxygen, nitrogen, carbon dioxide, hydrogen and natural gas are examples of gases.

Gases behave differently when their temperature, pressure or volume changes. These relationships are explained using scientific principles known as **gas laws**.

Gas laws help us understand why:

- A bicycle tyre becomes harder when more air is pumped into it.
- A balloon expands when heated.
- A balloon shrinks when placed in a refrigerator.
- A pressure cooker develops high pressure when heated.
- Perfume spreads through a room after being sprayed.
- Oxygen is compressed and stored in medical cylinders.
- Gases combine in definite volume ratios during chemical reactions.

The main gas laws studied in this sub-strand are:

1. Boyle's law.
2. Charles' law.
3. Combined gas law.
4. Gay-Lussac's law.
5. Graham's law of diffusion.
6. Avogadro's law.

3.0 KEY TERMS USED IN GAS LAWS

3.1 Gas

A gas is a state of matter in which particles move freely and randomly, occupying the entire volume of the container holding them.

Examples include:

- Oxygen.
- Hydrogen.
- Nitrogen.
- Carbon dioxide.
- Chlorine.
- Ammonia.
- Water vapour.

3.2 Pressure

Pressure is the force exerted per unit area.

$$P = \frac{F}{A}$$

Where:

- P = pressure.
- F = force.
- A = area.

The SI unit of pressure is the pascal (Pa).

Other units include:

- Kilopascal (kPa).
- Atmosphere (atm).
- Millimetres of mercury (mmHg).
- Bar.

Useful conversions:

$$1\,atm = 101325\,Pa$$

$$1\,atm = 101.325\,kPa$$

$$1\,atm = 760\,mmHg$$

$$1\,bar = 100000\,Pa$$

3.3 Volume

Volume is the amount of space occupied by a substance.

The SI unit is the cubic metre (m^3).

Other commonly used units include:

- Cubic centimetres (cm^3).
- Litres (L).
- Millilitres (mL).

Conversions:

$$1\,L = 1000\,cm^3$$

$$1\,mL = 1\,cm^3$$

$$1\,m^3 = 1000\,L$$

3.4 Temperature

Temperature is the measure of the degree of hotness or coldness of a substance.

Gas law calculations require absolute temperature measured in kelvin (K).

The relationship between Celsius and Kelvin is:

$$T(K) = t(^{\circ}C) + 273$$

For more precise calculations, use 273.15.

Examples:

1. Convert 27°C into kelvin.

$$T = 27 + 273$$

$$\boxed{T = 300\,K}$$

2. Convert 100°C into kelvin.

$$T = 100 + 273$$

$$\boxed{T = 373\,K}$$

3. Convert 25 K into degrees Celsius.

$$t = 25 - 273$$

$$\boxed{t = -248^{\circ}C}$$

Important: Never substitute Celsius temperatures directly into gas law equations involving absolute temperature.

3.5 Amount of gas

The amount of a substance is measured in moles (mol).

One mole contains approximately:

$$6.022 \times 10^{23}$$

particles.

This number is known as the Avogadro constant.

3.6 Fixed mass of gas

A fixed mass of gas refers to a definite quantity of gas whose amount does not change during an investigation.

For example, when investigating Boyle's law, the same quantity of air is trapped in a sealed syringe throughout the experiment.

4.0 KINETIC THEORY OF GASES

4.1 Meaning of kinetic theory

The kinetic theory of gases explains the behaviour of gases in terms of the movement and arrangement of their particles.

It describes how gas particles move, collide and interact with one another and with the walls of their container.

4.2 Main assumptions of the kinetic theory of gases

The kinetic theory is based on the following assumptions:

1. **Gases consist of tiny particles.**

   A gas is made up of very small particles called atoms or molecules.

2. **Gas particles are in continuous random motion.**

   The particles move freely and randomly in all directions.

3. **Gas particles move in straight lines between collisions.**

   Their direction changes when they collide with other particles or the walls of the container.

4. **The actual volume of gas particles is negligible.**

   The volume occupied by individual gas particles is very small compared with the volume of the container.

5. **There are negligible intermolecular forces in an ideal gas.**

   Gas particles do not significantly attract or repel one another except during collisions.

6. **Collisions are perfectly elastic.**

   No overall kinetic energy is lost during collisions between ideal gas particles.

7. **The average kinetic energy of gas particles is directly proportional to absolute temperature.**

   When temperature increases, the average kinetic energy of gas particles increases.

4.3 Explanation of gas pressure using kinetic theory

Gas pressure results from collisions of gas particles with the walls of their container.

The process occurs as follows:

1. Gas particles move continuously and randomly.
2. The particles collide with the walls of the container.
3. During collisions, they exert forces on the walls.
4. The combined force exerted by the particles per unit area produces gas pressure.

Factors affecting gas pressure

1. Number of gas particles.
2. Volume of the container.
3. Temperature of the gas.
4. Frequency and force of collisions with the container walls.

4.4 Explanation of the effect of temperature on gas particles

When a gas is heated:

1. Its particles gain kinetic energy.
2. They move faster.
3. Their collisions with the container walls become more frequent and energetic.
4. If the volume is fixed, pressure increases.

When a gas is cooled:

1. Its particles lose kinetic energy.
2. They move more slowly.
3. Their collisions with the container walls become less energetic.
4. At constant volume, pressure decreases.

4.5 Properties of gases

Gases have the following properties:

1. **They have no fixed shape.** They take the shape of their containers.
2. **They have no fixed volume.** They expand to fill available space.
3. **They are compressible.** Their volume can be reduced by applying pressure.
4. **They have low density.** Their particles are widely separated compared with particles in solids and liquids.
5. **They diffuse easily.** Their particles mix with particles of other gases.
6. **They exert pressure.** Their particles collide with the walls of containers.
7. **They expand when heated.** At constant pressure, their volume increases as temperature rises.

5.0 BOYLE'S LAW

5.1 Meaning of Boyle's law

Boyle's law states that:

The volume of a fixed mass of a gas is inversely proportional to its pressure, provided the temperature remains constant.

This means that when the pressure of a gas increases, its volume decreases, and when pressure decreases, its volume increases, provided temperature and the amount of gas remain constant.

Mathematically:

$$V \propto \frac{1}{P}$$

Therefore:

$$PV = k$$

Where:

- P = pressure of the gas.
- V = volume of the gas.
- k = constant at a fixed temperature.

For two different conditions:

$$\boxed{P_1V_1 = P_2V_2}$$

Where:

- P_1 = initial pressure.
- V_1 = initial volume.
- P_2 = final pressure.
- V_2 = final volume.

5.2 Explanation of Boyle's law using kinetic theory

Consider a fixed quantity of gas enclosed in a container with a movable piston.

When the piston is pushed down:

1. The volume occupied by the gas decreases.
2. The same number of particles occupies a smaller space.
3. The particles travel a shorter distance before colliding with the walls.
4. Collisions with the container walls occur more frequently.
5. Gas pressure increases.

When the piston is pulled upwards:

1. The volume increases.
2. Gas particles have more space to move around.
3. Collisions with the walls occur less frequently per unit area.
4. Gas pressure decreases.

The temperature must remain constant during this investigation.

5.3 Experimental verification of Boyle's law

Aim

To investigate how the volume of a fixed mass of gas changes when its pressure changes at constant temperature.

Apparatus

- Gas syringe or Boyle's law apparatus.
- Pressure gauge or pressure sensor.
- Connecting tube.
- Clamps.
- Fixed mass of trapped air.

Procedure

1. Trap a fixed amount of air in a gas syringe.
2. Connect the syringe to a pressure-measuring device.
3. Record the initial volume and pressure.
4. Gradually compress the trapped air.
5. Record the new volume and pressure.
6. Repeat the procedure using different volumes.
7. Maintain constant temperature throughout the experiment.
8. Compare the pressure and volume readings.

Expected results

Volume (cm^3)	Pressure (kPa)	Pressure × Volume

Volume (cm^3)	Pressure (kPa)	Pressure × Volume
60	50	3000
50	60	3000
40	75	3000
30	100	3000
20	150	3000

Illustrative experimental data; actual results depend on the apparatus and experimental conditions.

Observations

- As volume decreases, pressure increases.
- As volume increases, pressure decreases.
- The product of pressure and volume remains approximately constant.

Conclusion

The volume of a fixed mass of gas is inversely proportional to its pressure at constant temperature.

This verifies Boyle's law.

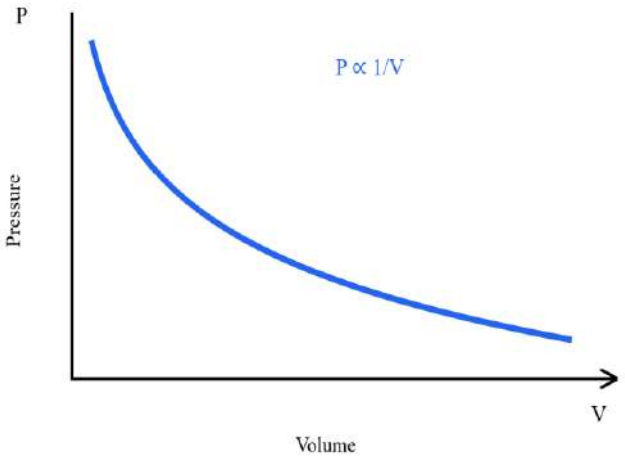
5.4 Graphs for Boyle's law

A. Graph of pressure against volume

When pressure is plotted against volume, a downward-curving graph is obtained.

The graph is a rectangular hyperbola.

Boyle's law: Pressure against volume



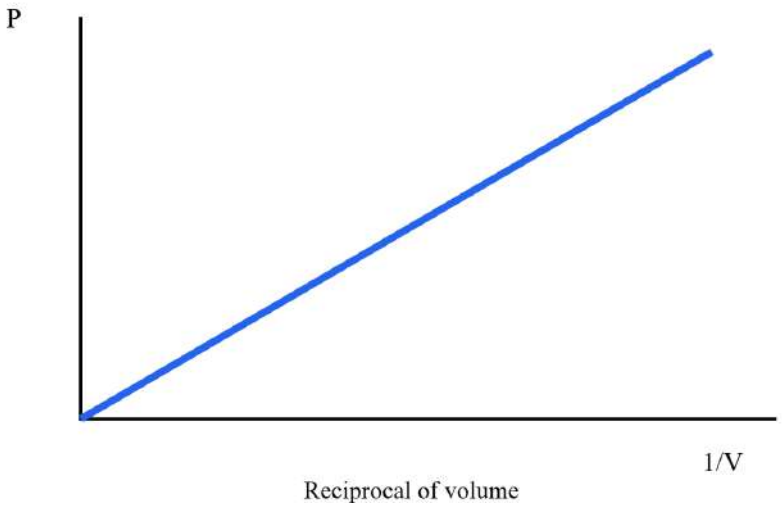
Pressure decreases as volume increases at constant temperature.

B. Graph of pressure against reciprocal volume

When pressure is plotted against reciprocal volume, a straight line passing through the origin is obtained.

$$P \propto \frac{1}{V}$$

Pressure against 1/Volume



A straight line demonstrates the direct proportionality between pressure and reciprocal volume.

5.5 Mathematical applications of Boyle's law

Use:

$$P_1V_1 = P_2V_2$$

Example 1

A gas occupies a volume of 300 cm³ at a pressure of 100 kPa. Calculate its volume when the pressure is increased to 150 kPa at constant temperature.

Solution

Given:

$$P_1 = 100\,kPa$$

$$V_1 = 300\,cm^3$$

$$P_2 = 150\,kPa$$

$$V_2 = ?$$

Using Boyle's law:

$$P_1V_1 = P_2V_2$$

Substituting:

$$100 \times 300 = 150 \times V_2$$

$$V_2 = \frac{100 \times 300}{150}$$

$$\boxed{V_2 = 200\,cm^3}$$

Answer: The final volume is 200 cm³.

Example 2

A gas occupies 500 cm^3 at 80 kPa. Calculate the pressure when its volume is reduced to 200 cm^3 at constant temperature.

Solution

$$P_1 = 80\,kPa$$

$$V_1 = 500\,cm^3$$

$$V_2 = 200\,cm^3$$

$$P_2 = ?$$

$$P_1V_1 = P_2V_2$$

$$80 \times 500 = P_2 \times 200$$

$$P_2 = \frac{80 \times 500}{200}$$

$$\boxed{P_2 = 200\,kPa}$$

Example 3

A gas occupies 240 cm^3 at 120 kPa. Find the pressure when the volume increases to 480 cm^3.

Solution

$$P_1V_1 = P_2V_2$$

$$120 \times 240 = P_2 \times 480$$

$$P_2 = \frac{120 \times 240}{480}$$

$$\boxed{P_2 = 60\,kPa}$$

5.6 Applications of Boyle's law

1. **Breathing in human beings:** When the diaphragm contracts, the volume of the chest cavity increases and pressure inside the lungs decreases. Atmospheric pressure then drives air into the lungs.
2. **Breathing out:** The volume of the chest cavity decreases, increasing pressure inside the lungs and causing air to move out.
3. **Medical syringes:** Pulling the plunger increases the volume inside the syringe and lowers its pressure, allowing liquid to enter.
4. **Bicycle pumps:** Compressing air in a pump decreases its volume and increases its pressure, enabling air to enter a tyre.
5. **Scuba diving:** Changes in pressure affect the volume of gas in diving equipment and air spaces.
6. **Air compressors:** Air is compressed into a smaller volume, increasing its pressure for industrial applications.
7. **Respiratory equipment:** Pressure-volume relationships are important in the design and operation of some breathing devices.

6.0 CHARLES' LAW

6.1 Meaning of Charles' law

Charles' law states that:

The volume of a fixed mass of gas is directly proportional to its absolute temperature, provided the pressure remains constant.

This means that when the temperature of a gas increases, its volume increases, provided pressure remains constant.

Similarly, when the temperature decreases, the volume decreases.

Mathematically:

$$V \propto T$$

Therefore:

$$\frac{V}{T} = k$$

For two different conditions:

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

Where:

- V_1 = initial volume.
- T_1 = initial temperature in kelvin.
- V_2 = final volume.
- T_2 = final temperature in kelvin.

6.2 Explanation of Charles' law using kinetic theory

When a gas is heated at constant pressure:

1. Its particles gain kinetic energy.
2. They move faster.
3. Collisions with the container walls become more energetic.
4. The gas expands.
5. The volume increases until the pressure returns to the required constant value.

When the gas is cooled:

1. Its particles lose kinetic energy.
2. They move more slowly.
3. The gas contracts.
4. Its volume decreases.

6.3 Experimental verification of Charles' law

Aim

To investigate the relationship between the volume and temperature of a fixed mass of gas at constant pressure.

Apparatus

- Gas syringe with a movable piston.

- Thermometer.
- Water baths.
- Beakers.
- Heating source.
- Clamp stand.
- Fixed mass of trapped air.

Procedure

1. Trap a fixed amount of air in a gas syringe.
2. Ensure that the piston can move freely.
3. Place the gas syringe in a water bath.
4. Record the temperature and gas volume.
5. Heat the water bath gradually.
6. Record the new volume at each temperature.
7. Ensure that pressure remains constant throughout the experiment.
8. Convert temperatures from degrees Celsius to kelvin.
9. Compare the volume and absolute temperature readings.

Expected results

Temperature (°C)	Temperature (K)	Volume (cm^3)
0	273	100
27	300	110
54	327	120
81	354	130
108	381	140

Illustrative data for demonstrating the relationship; actual experimental readings may differ.

Observations

- As temperature increases, volume increases.
- As temperature decreases, volume decreases.
- The ratio of volume to absolute temperature remains approximately constant.

Conclusion

The volume of a fixed mass of gas is directly proportional to its absolute temperature at constant pressure.

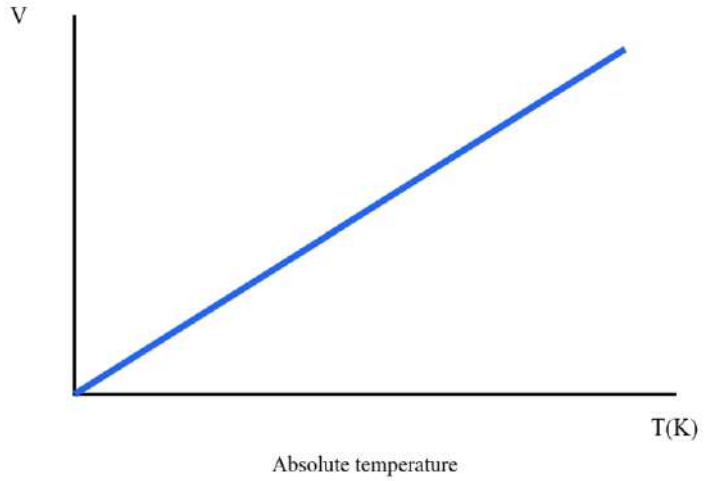
This verifies Charles' law.

6.4 Graphs for Charles' law

A. Volume against temperature in kelvin

A straight line passing through the origin is obtained when volume is plotted against absolute temperature.

Volume against absolute temperature

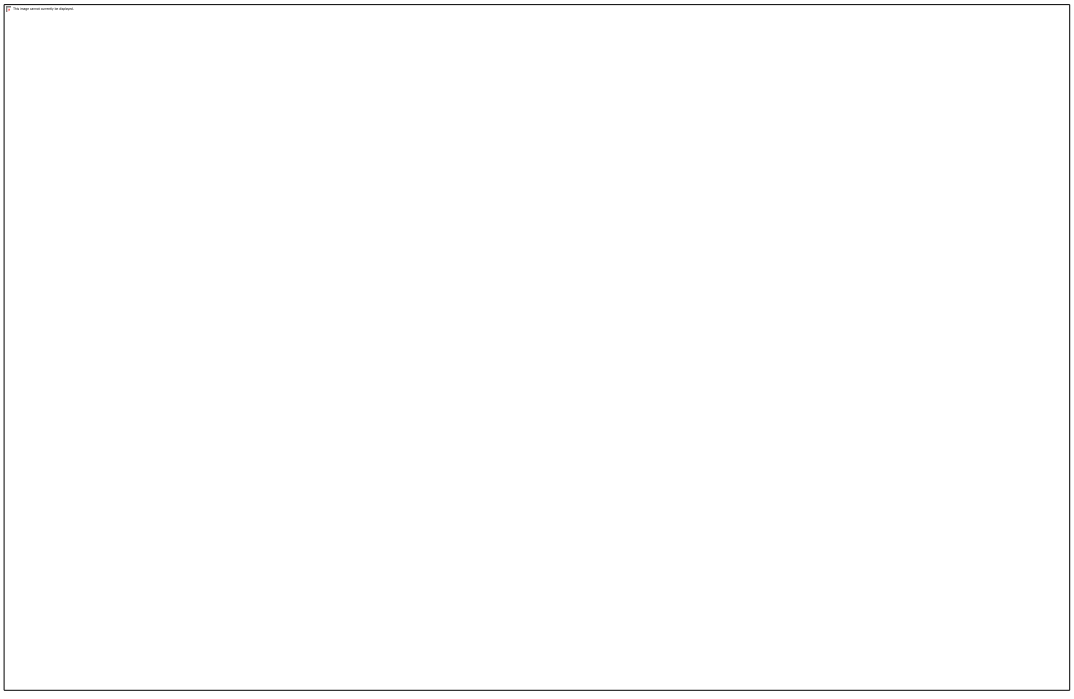


Volume increases directly with absolute temperature at constant pressure.

B. Volume against temperature in degrees Celsius

When volume is plotted against temperature in degrees Celsius, the graph is a straight line that, if extrapolated, approaches zero volume at approximately −273°C.

Volume against Celsius temperature



The extrapolated intercept corresponds to absolute zero.

Important: The zero-volume intercept is a theoretical extrapolation for an ideal gas. Real gases liquefy before reaching this condition.

6.5 Absolute zero

Absolute zero is the theoretical lowest possible temperature, corresponding to zero kelvin.

$$0\,K = -273.15°C$$

At absolute zero, the classical kinetic theory predicts that the average translational kinetic energy of an ideal gas would approach zero.

Real gases cannot be cooled as gases all the way to absolute zero because they undergo phase changes.

6.6 Mathematical applications of Charles' law

Use:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Example 1

A gas occupies 200 cm^3 at 27°C. Calculate its volume when heated to 127°C at constant pressure.

Solution

Given:

$$V_1 = 200\,cm^3$$

$$T_1 = 27 + 273 = 300\,K$$

$$T_2 = 127 + 273 = 400\,K$$

$$V_2 = ?$$

Using Charles' law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{200}{300} = \frac{V_2}{400}$$

$$V_2 = \frac{200 \times 400}{300}$$

$$\boxed{V_2 = 266.7\,cm^3}$$

Example 2

A gas occupies 500 cm^3 at 300 K. Calculate its volume at 450 K if pressure remains constant.

Solution

$$\frac{500}{300} = \frac{V_2}{450}$$

$$V_2 = \frac{500 \times 450}{300}$$

$$\boxed{V_2 = 750\,cm^3}$$

Example 3

A gas occupies 600 cm^3 at 100°C. Find its volume at 0°C at constant pressure.

Solution

$$T_1 = 100 + 273 = 373\,K$$

$$T_2 = 0 + 273 = 273\,K$$

$$V_1 = 600\,cm^3$$

$$\frac{600}{373} = \frac{V_2}{273}$$

$$V_2 = \frac{600 \times 273}{373}$$

$$\boxed{V_2 = 439.1\,cm^3}$$

6.7 Applications of Charles' law

1. **Hot-air balloons:** Heating the air inside a balloon causes it to expand at approximately constant pressure. The heated air becomes less dense than surrounding air, helping the balloon rise.
2. **Expansion of balloons:** A balloon expands when warmed because the gas inside increases in volume.
3. **Cold weather effects:** Balloons shrink when cooled because the gas inside occupies a smaller volume.
4. **Weather balloons:** Changes in atmospheric temperature influence the volume of gas in balloons as they rise.
5. **Industrial gas processing:** Temperature-volume relationships are considered when designing gas storage and processing equipment.
6. **Vehicle tyres:** Temperature changes affect the behaviour of gases inside tyres. However, a tyre has nearly fixed volume, so the pressure change is more directly explained by Gay-Lussac's law.

7.0 COMBINED GAS LAW

7.1 Meaning of the combined gas law

The combined gas law describes the relationship between pressure, volume and absolute temperature for a fixed mass of gas.

It combines Boyle's law and Charles' law.

It is used when pressure, volume and temperature all change while the amount of gas remains constant.

7.2 Derivation of the combined gas law

From Boyle's law:

$$V \propto \frac{1}{P}$$

From Charles' law:

$$V \propto T$$

Combining the two relationships:

$$V \propto \frac{T}{P}$$

Therefore:

$$\frac{PV}{T} = k$$

For two different conditions:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This is the combined gas law equation.

$$\boxed{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}$$

Where:

- P_1 = initial pressure.
- V_1 = initial volume.
- T_1 = initial absolute temperature.
- P_2 = final pressure.
- V_2 = final volume.
- T_2 = final absolute temperature.

7.3 Conditions for using the combined gas law

1. The amount of gas must remain constant.
2. The gas should behave approximately ideally.
3. Temperature must be expressed in kelvin.
4. Pressure units must be consistent.
5. Volume units must be consistent.

7.4 Mathematical applications of the combined gas law

Example 1

A gas occupies 400 cm³ at a pressure of 100 kPa and a temperature of 27°C. Calculate its volume at a pressure of 150 kPa and a temperature of 127°C.

Solution

Given:

$$P_1 = 100\,kPa$$

$$V_1 = 400\,cm^3$$

$$T_1 = 27 + 273 = 300\,K$$

$$P_2 = 150\,kPa$$

$$T_2 = 127 + 273 = 400\,K$$

$$V_2 = ?$$

Using the combined gas law:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Substituting:

$$\frac{100 \times 400}{300} = \frac{150 \times V_2}{400}$$

Rearranging:

$$V_2 = \frac{100 \times 400 \times 400}{300 \times 150}$$

$$\boxed{V_2 = 355.6\, cm^3}$$

Example 2

A gas occupies 2 litres at 100 kPa and 300 K. Find its pressure when the volume becomes 1.5 litres and temperature becomes 450 K.

Solution

$$P_1 = 100\, kPa$$

$$V_1 = 2\, L$$

$$T_1 = 300\, K$$

$$V_2 = 1.5\, L$$

$$T_2 = 450\, K$$

$$P_2 = ?$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{100 \times 2}{300} = \frac{P_2 \times 1.5}{450}$$

$$P_2 = \frac{100 \times 2 \times 450}{300 \times 1.5}$$

$$\boxed{P_2 = 200\, kPa}$$

Example 3

A gas occupies 600 cm³ at 120 kPa and 300 K. Calculate its temperature when its volume is 800 cm³ and its pressure is 90 kPa.

Solution

$$\frac{120 \times 600}{300} = \frac{90 \times 800}{T_2}$$

$$\frac{72000}{300} = \frac{72000}{T_2}$$

$$240 = \frac{72000}{T_2}$$

$$T_2 = \frac{72000}{240}$$

$$\boxed{T_2 = 300\, K}$$

8.0 GAY-LUSSAC'S LAW OF PRESSURE

8.1 Meaning of Gay-Lussac's law

Gay-Lussac's pressure law states that:

The pressure of a fixed mass of gas is directly proportional to its absolute temperature, provided the volume remains constant.

This means that when a gas in a rigid container is heated, its pressure increases.

When the gas is cooled, its pressure decreases.

Mathematically:

$$P \propto T$$

Therefore:

$$\frac{P}{T} = k$$

For two different conditions:

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$$

Where:

- P_1 = initial pressure.
- T_1 = initial temperature in kelvin.
- P_2 = final pressure.
- T_2 = final temperature in kelvin.

8.2 Explanation using kinetic theory

When a fixed volume of gas is heated:

1. The gas particles gain kinetic energy.
2. They move faster.
3. They collide more frequently with the container walls.
4. The collisions become more energetic.
5. The pressure increases.

When the gas is cooled, the particles move more slowly, reducing the pressure.

8.3 Experimental verification of Gay-Lussac's law

Aim

To investigate the relationship between the pressure and absolute temperature of a fixed mass of gas at constant volume.

Apparatus

- Rigid sealed gas container suitable for pressure experiments.
- Pressure sensor or gauge.
- Thermometer.
- Controlled water baths.
- Heating equipment.

Procedure

1. Trap a fixed quantity of gas in a rigid container.
2. Record the initial temperature and pressure.
3. Place the container in a controlled-temperature water bath.
4. Increase the temperature gradually.
5. Record the pressure at each temperature.
6. Repeat at different temperatures.
7. Convert temperatures to kelvin.
8. Compare the pressure-temperature ratios.

Observations

- Pressure increases as temperature increases.
- Pressure decreases as temperature decreases.
- The ratio of pressure to absolute temperature remains approximately constant.

Conclusion

The pressure of a fixed mass of gas is directly proportional to its absolute temperature at constant volume.

8.4 Graph for Gay-Lussac's law

A graph of pressure against absolute temperature is a straight line passing through the origin for an ideal gas.

Pressure against absolute temperature



Pressure increases directly with absolute temperature at constant volume.

8.5 Mathematical applications of Gay-Lussac's law

Example 1

A gas in a sealed rigid container has a pressure of 100 kPa at 27°C. Calculate its pressure at 127°C.

Solution

$$P_1 = 100\,kPa$$

$$T_1 = 27 + 273 = 300\,K$$

$$T_2 = 127 + 273 = 400\,K$$

$$P_2 = ?$$

Using:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{100}{300} = \frac{P_2}{400}$$

$$P_2 = \frac{100 \times 400}{300}$$

$$\boxed{P_2 = 133.3\,kPa}$$

Example 2

A gas has a pressure of 200 kPa at 400 K. Calculate its temperature when the pressure decreases to 150 kPa at constant volume.

Solution

$$\frac{200}{400} = \frac{150}{T_2}$$

$$T_2 = \frac{150 \times 400}{200}$$

$$\boxed{T_2 = 300\,K}$$

8.6 Applications of Gay-Lussac's law

1. **Aerosol containers:** Heating a sealed aerosol container increases the pressure of the gas inside. This is why aerosol containers must not be exposed to excessive heat.
2. **Vehicle tyres:** Tyre pressure usually increases as the air inside warms during driving, approximately at constant volume.
3. **Pressure cookers:** Heating a sealed pressure cooker increases the pressure of the steam inside, allowing food to cook at a higher temperature.
4. **Gas cylinders:** Temperature changes affect the pressure of gases stored in rigid cylinders.
5. **Industrial gas storage:** Engineers consider pressure changes caused by temperature variations when designing storage systems.
6. **Fire safety:** Sealed gas containers exposed to fire may experience a dangerous increase in internal pressure.

9.0 GRAHAM'S LAW OF DIFFUSION

9.1 Meaning of diffusion

Diffusion is the spontaneous movement of particles from a region of higher concentration to a region of lower concentration until they become evenly distributed.

Diffusion occurs in gases and liquids.

Examples of diffusion in everyday life

1. The smell of perfume spreads throughout a room.
2. The smell of cooking food spreads from the kitchen to other rooms.
3. Smoke spreads through the surrounding air.
4. Oxygen moves from the air in the lungs into the blood.
5. Carbon dioxide moves from the blood into the air spaces in the lungs.
6. Tea particles spread through hot water.
7. A drop of ink spreads through water.

9.2 Diffusion in gases

Gas particles move randomly and continuously.

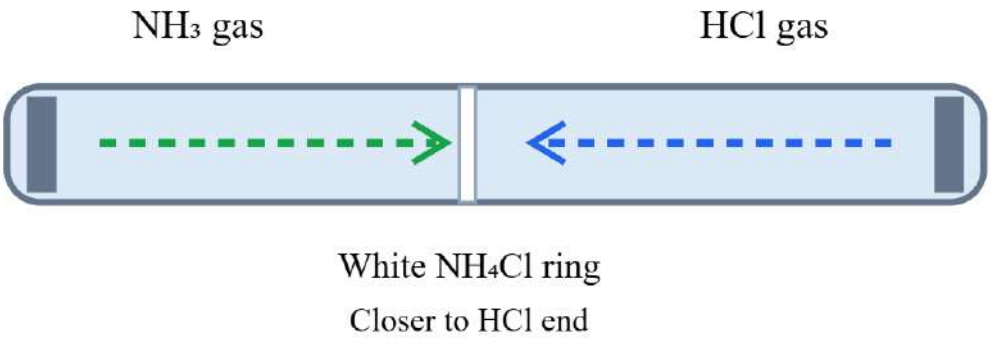
When two different gases are brought into contact, their particles mix gradually due to random motion.

For example, when ammonia gas and hydrogen chloride gas are introduced at opposite ends of a glass tube, they diffuse towards one another and react to form a white ring of ammonium chloride.

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

The white ring forms closer to the hydrogen chloride end because ammonia diffuses faster than hydrogen chloride.

Illustration: Diffusion of ammonia and hydrogen chloride



Ammonia travels further in the same time because it has a lower molar mass.

9.3 Factors affecting the rate of diffusion of gases

1. **Molecular mass**

   Lighter gas molecules generally diffuse faster than heavier gas molecules at the same temperature.

2. **Temperature**

   Increasing temperature increases the average kinetic energy of gas particles, making diffusion faster.

3. **Concentration gradient**

A greater difference in concentration between two regions generally produces faster net diffusion.

4. **Nature of the gas**

   Molecular size and interactions can influence the rate of diffusion.

5. **Pressure and density**

   Changes in pressure and gas density can influence diffusion under different conditions.

6. **Distance travelled**

   A gas takes longer to diffuse over a greater distance.

9.4 Graham's law of diffusion

Graham's law states that:

The rate of diffusion of a gas is inversely proportional to the square root of its molar mass, provided the gases are at the same temperature and pressure.

Mathematically:

$$r \propto \frac{1}{\sqrt{M}}$$

Where:

- r = rate of diffusion.
- M = molar mass of the gas.

For two gases A and B:

$$\boxed{\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}}$$

Where:

- r_A = rate of diffusion of gas A.
- r_B = rate of diffusion of gas B.
- M_A = molar mass of gas A.
- M_B = molar mass of gas B.

Remember: The lighter gas diffuses faster than the heavier gas under the same conditions.

9.5 Mathematical applications of Graham's law

Example 1: Comparing hydrogen and oxygen

Calculate the ratio of the rates of diffusion of hydrogen and oxygen.

Given:

Molar mass of hydrogen gas, H_2:

Molar mass of oxygen gas, O_2:

Using Graham's law:

$$\frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{32}{2}}$$

$$= \sqrt{16}$$

$$\boxed{\frac{r_{H_2}}{r_{O_2}} = 4}$$

Answer: Hydrogen diffuses four times faster than oxygen under the same conditions.

Example 2: Comparing ammonia and hydrogen chloride

Calculate the ratio of the rates of diffusion of ammonia and hydrogen chloride.

Given:

$$M_{NH_3} = 14 + (3 \times 1) = 17$$

$$M_{HCl} = 1 + 35.5 = 36.5$$

Using:

$$\frac{r_{NH_3}}{r_{HCl}} = \sqrt{\frac{36.5}{17}}$$

$$= \sqrt{2.147}$$

$$\boxed{\frac{r_{NH_3}}{r_{HCl}} \approx 1.47}$$

Answer: Ammonia diffuses approximately 1.47 times faster than hydrogen chloride.

Example 3: Finding an unknown molar mass

Gas A diffuses twice as fast as gas B. If the molar mass of gas A is 4 g/mol, calculate the molar mass of gas B.

Using:

$$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$$

$$2 = \sqrt{\frac{M_B}{4}}$$

Squaring both sides:

$$4 = \frac{M_B}{4}$$

$$M_B = 4 \times 4$$

Example 4: Comparing diffusion times

For the same distance under the same conditions, a gas with molar mass 4 g/mol takes 2 minutes to diffuse. How long will a gas with molar mass 64 g/mol take?

For a fixed distance, time is inversely proportional to rate.

Therefore:

$$\frac{t_A}{t_B} = \frac{r_B}{r_A}$$

Since rate is inversely proportional to the square root of molar mass:

$$\frac{t_A}{t_B} = \sqrt{\frac{M_A}{M_B}}$$

Alternatively:

$$\frac{t_B}{t_A} = \sqrt{\frac{M_B}{M_A}}$$

$$\frac{t_B}{2} = \sqrt{\frac{64}{4}}$$

$$\frac{t_B}{2} = 4$$

$$\boxed{t_B = 8\,minutes}$$

9.6 Applications of diffusion in everyday life

1. **Breathing:** Oxygen diffuses from the air spaces in the lungs into the blood, while carbon dioxide diffuses from the blood into the lungs.
2. **Perfumes:** Fragrance particles diffuse through air, allowing people to smell perfume from a distance.
3. **Cooking:** The smell of food spreads through the surrounding environment.
4. **Gas exchange in plants:** Carbon dioxide and oxygen diffuse through stomata during photosynthesis and respiration.
5. **Medical oxygen delivery:** Oxygen moves from areas of higher concentration to areas of lower concentration in respiratory systems.
6. **Industrial gas separation:** Differences in molecular movement are considered in certain gas separation processes.
7. **Environmental processes:** Gases diffuse between the atmosphere, soil and water.

10.0 AVOGADRO'S LAW

10.1 Meaning of Avogadro's law

Avogadro's law states that:

Equal volumes of all gases, at the same temperature and pressure, contain equal numbers of molecules.

It also states that the volume of a gas is directly proportional to the number of moles of gas at constant temperature and pressure.

Mathematically:

$$V \propto n$$

Therefore:

$$\frac{V}{n} = k$$

For two different conditions:

$$\boxed{\frac{V_1}{n_1} = \frac{V_2}{n_2}}$$

Where:

- V_1 = initial volume.
- n_1 = initial number of moles.
- V_2 = final volume.
- n_2 = final number of moles.

10.2 Explanation of Avogadro's law

Consider two containers of equal volume containing different gases at the same temperature and pressure.

According to Avogadro's law:

1. The containers contain equal numbers of gas molecules.
2. The gases may have different molecular masses.
3. Their total masses may therefore be different.
4. However, their volumes are equal because their temperature, pressure and number of particles are equal.

For example, equal volumes of hydrogen and oxygen at the same temperature and pressure contain equal numbers of molecules.

10.3 Experimental illustration of Avogadro's law

Aim

To demonstrate that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

Procedure

1. Prepare two suitable gas containers of equal volume.
2. Fill one container with hydrogen and the other with oxygen.
3. Ensure both gases are at the same temperature and pressure.
4. Compare the quantities in terms of the number of moles.
5. Discuss the relationship between volume and number of particles.

Conclusion

Equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules.

10.4 Molar volume of gases

Molar volume is the volume occupied by one mole of a gas at a specified temperature and pressure.

At a temperature of 0°C and pressure of 1 atmosphere, one mole of an ideal gas occupies approximately:

$$\boxed{22.4\,dm^3}$$

This condition is known as standard temperature and pressure (STP), using the conventional chemistry definition of 273 K and 1 atm.

At 25°C and 1 atmosphere, one mole of an ideal gas occupies approximately:

$$\boxed{24.5\,dm^3}$$

Important: Molar volume depends on temperature and pressure. Always use the conditions specified in a question.

10.5 Mathematical applications of Avogadro's law

Example 1

A gas occupies 2 litres when it contains 0.1 moles. Calculate the volume occupied by 0.3 moles at the same temperature and pressure.

Solution

$$V_1 = 2\,L$$

$$n_1 = 0.1\,mol$$

$$n_2 = 0.3\,mol$$

$$V_2 = ?$$

Using:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{2}{0.1} = \frac{V_2}{0.3}$$

$$V_2 = \frac{2 \times 0.3}{0.1}$$

$$\boxed{V_2 = 6\,L}$$

Example 2

A gas occupies 12 dm^3 when it contains 0.5 moles. Calculate the number of moles in 30 dm^3 at the same temperature and pressure.

Solution

$$\frac{12}{0.5} = \frac{30}{n_2}$$

$$n_2 = \frac{30 \times 0.5}{12}$$

$$\boxed{n_2 = 1.25\,mol}$$

10.6 Applications of Avogadro's law

1. Calculating the number of moles in a given gas volume.
2. Comparing equal volumes of different gases.
3. Determining reacting volumes of gases.
4. Calculating gas volumes during industrial chemical processes.
5. Understanding the relationship between the number of particles and gas volume.
6. Determining the amount of gas produced in laboratory experiments.

11.0 GAY-LUSSAC'S LAW OF COMBINING VOLUMES

11.1 Meaning of Gay-Lussac's law of combining volumes

Gay-Lussac's law of combining volumes states that:

When gases react, their volumes and the volumes of gaseous products bear simple whole-number ratios, provided all gas volumes are measured at the same temperature and pressure.

This law is different from Gay-Lussac's pressure law.

- Gay-Lussac's pressure law relates pressure and temperature at constant volume.
- Gay-Lussac's law of combining volumes relates the volumes of reacting gases and gaseous products.

11.2 Explanation of the law

During a chemical reaction involving gases, the volumes of reacting gases combine in simple whole-number ratios.

For example, hydrogen reacts with oxygen to form water vapour.

Balanced equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

The volume ratio is:

$$2:1:2$$

This means:

- 2 volumes of hydrogen react with 1 volume of oxygen.
- 2 volumes of water vapour are formed.

The gas volumes must be compared at the same temperature and pressure.

11.3 Experimental illustration using hydrogen and oxygen

Aim

To demonstrate the combining volume ratio of hydrogen and oxygen.

Procedure

A suitable demonstration can be carried out using an approved eudiometer or gas-volume apparatus under teacher supervision.

1. Measure suitable volumes of hydrogen and oxygen.
2. Mix the gases in the correct ratio.
3. Introduce a controlled electrical spark using suitable apparatus.
4. Allow the reaction to occur.
5. Cool the apparatus to the initial temperature.
6. Observe the change in gas volume.
7. Compare the initial and final gas volumes.

Observation

Two volumes of hydrogen react with one volume of oxygen to produce water.

After the water condenses, the remaining gas volume is reduced according to the reacting proportions.

Safety: Hydrogen-oxygen mixtures are explosive. This experiment should only be demonstrated using approved small-scale apparatus by a trained teacher.

11.4 Examples of Gay-Lussac's law of combining volumes

Example 1: Formation of water

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Volume ratio:

$$2:1:2$$

Therefore:

- 2 dm^3 of hydrogen react with 1 dm^3 of oxygen.
- 4 dm^3 of hydrogen react with 2 dm^3 of oxygen.
- 10 dm^3 of hydrogen react with 5 dm^3 of oxygen.

Example 2: Formation of ammonia

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Volume ratio:

$$1:3:2$$

Therefore:

- 1 volume of nitrogen reacts with 3 volumes of hydrogen.
- 2 volumes of nitrogen react with 6 volumes of hydrogen.
- 5 volumes of nitrogen react with 15 volumes of hydrogen.

Example 3: Formation of hydrogen chloride

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Volume ratio:

$$1:1:2$$

Therefore:

- 1 volume of hydrogen reacts with 1 volume of chlorine.
- 5 volumes of hydrogen react with 5 volumes of chlorine.
- 10 volumes of hydrogen react with 10 volumes of chlorine.

11.5 Mathematical applications of Gay-Lussac's law of combining volumes

Example 1

Calculate the volume of oxygen required to react completely with 60 cm^3 of hydrogen to form water vapour.

Equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Volume ratio:

$$2:1$$

Therefore:

$$\frac{V_{H_2}}{V_{O_2}} = \frac{2}{1}$$

$$\frac{60}{V_{O_2}} = \frac{2}{1}$$

$$V_{O_2} = \frac{60}{2}$$

$$\boxed{V_{O_2} = 30\ cm^3}$$

Example 2

Nitrogen reacts with hydrogen to produce ammonia.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Calculate the volume of ammonia formed when 20 dm^3 of nitrogen reacts completely with sufficient hydrogen.

Volume ratio:

$$1:3:2$$

1 volume nitrogen → 2 volumes ammonia

$$20 \times 2 = 40$$

$$\boxed{40\ dm^3}$$

Example 3

Hydrogen reacts with chlorine to form hydrogen chloride.

$$H_2 + Cl_2 \rightarrow 2HCl$$

Calculate the volume of hydrogen chloride formed when 15 cm^3 of chlorine reacts completely with sufficient hydrogen.

Volume ratio:

$$1:1:2$$

Therefore:

$$15 \times 2 = 30$$

$\boxed{30\ cm^3}$

11.6 Applications of Gay-Lussac's law of combining volumes

1. Calculating the volumes of gases required in chemical manufacturing.
2. Determining the quantities of hydrogen and nitrogen required in ammonia production.
3. Calculating the volumes of oxygen needed for combustion reactions.
4. Determining the amount of gaseous products formed during industrial reactions.
5. Understanding the composition of reacting gas mixtures.
6. Solving stoichiometric problems involving gaseous reactants and products.

12.0 RELATIONSHIP BETWEEN AVOGADRO'S LAW AND GAY-LUSSAC'S LAW OF COMBINING VOLUMES

The two laws are closely related.

Avogadro's law

Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

Gay-Lussac's law of combining volumes

Reacting gases combine in simple whole-number volume ratios when measured at the same temperature and pressure.

How the laws are related

1. Avogadro's law explains that gas volume is proportional to the number of molecules.
2. Chemical equations show the relative numbers of molecules involved in reactions.
3. Therefore, the coefficients in balanced gaseous chemical equations represent the volume ratios of reacting gases and gaseous products under the same conditions.

For example:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

The molecular ratio is:

$$1:3:2$$

The corresponding gas volume ratio is also:

$$1:3:2$$

This relationship is important in industrial chemical calculations.

13.0 SUMMARY OF THE SIX GAS LAWS

Gas law	Relationship	Condition	Formula
Boyle's law	Volume inversely proportional to pressure	Constant temperature	$P_1V_1 = P_2V_2$

Gas law	Relationship	Condition	Formula
Charles' law	Volume directly proportional to temperature	Constant pressure	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Combined gas law	Pressure, volume and temperature relationship	Fixed mass of gas	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Gay-Lussac's pressure law	Pressure directly proportional to temperature	Constant volume	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Graham's law	Rate inversely proportional to square root of molar mass	Same temperature and pressure	$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$
Avogadro's law	Volume directly proportional to number of moles	Constant temperature and pressure	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Gay-Lussac's combining volumes	Reacting gas volumes are in simple whole-number ratios	Same temperature and pressure	Use balanced equation coefficients

14.0 THE IDEAL GAS EQUATION

14.1 Meaning of an ideal gas

An ideal gas is a hypothetical gas whose particles obey the assumptions of kinetic theory perfectly under all conditions.

An ideal gas:

- Has particles of negligible volume.
- Has no intermolecular attractions or repulsions.
- Experiences perfectly elastic collisions.
- Obeys the gas laws exactly.

Real gases behave approximately like ideal gases at low pressure and high temperature.

14.2 Derivation of the ideal gas equation

From the combined gas law:

$$\frac{PV}{T} = k$$

Avogadro's law shows that:

$$V \propto n$$

Combining these relationships:

$$PV \propto nT$$

Introducing the universal gas constant, R:

$$\boxed{PV = nRT}$$

Where:

- P = pressure.
- V = volume.
- n = number of moles.
- R = universal gas constant.
- T = absolute temperature in kelvin.

The value of R depends on the units used.

Common values include:

$$R = 8.314\, J\, mol^{-1} K^{-1}$$

or

$$R = 0.0821\, L\, atm\, mol^{-1} K^{-1}$$

14.3 Applications of the ideal gas equation

The equation can be used to calculate:

1. Pressure of a gas.
2. Volume occupied by a gas.
3. Number of moles of a gas.
4. Temperature of a gas.
5. Molar mass of a gas when mass and volume information is available.

14.4 Worked examples

Example 1: Calculating the volume of a gas

Calculate the volume occupied by 2 moles of an ideal gas at 300 K and a pressure of 100 kPa.

Given:

$$n = 2\, mol$$

$$T = 300\, K$$

$$P = 100000\, Pa$$

$$R = 8.314$$

Using:

$$PV = nRT$$

Rearranging:

$$V = \frac{nRT}{P}$$

Substituting:

$$V = \frac{2 \times 8.314 \times 300}{100000}$$

$$V = 0.049884\, m^3$$

Converting to litres:

$$0.049884 \times 1000 = 49.884\, L$$

$$\boxed{V \approx 49.9\, L}$$

Example 2: Calculating the number of moles

Calculate the number of moles in 24.93 litres of a gas at 300 K and 1 atmosphere.

Given:

$$V = 24.93\,L$$

$$T = 300\,K$$

$$P = 1\,atm$$

$$R = 0.0821\,L\,atm\,mol^{-1}K^{-1}$$

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{1 \times 24.93}{0.0821 \times 300}$$

$$\boxed{n \approx 1.01\,mol}$$

15.0 COMPARISON OF THE DIFFERENT GAS LAWS

It is important to distinguish the different gas laws because they involve different conditions and relationships.

Law	What changes?	What remains constant?	Main idea
Boyle's law	Pressure and volume	Temperature and amount of gas	Increasing pressure decreases volume
Charles' law	Volume and temperature	Pressure and amount of gas	Heating increases volume
Gay-Lussac's pressure law	Pressure and temperature	Volume and amount of gas	Heating increases pressure
Combined gas law	Pressure, volume and temperature	Amount of gas	Relates all three variables
Graham's law	Diffusion rate and molar mass	Temperature and pressure for comparison	Lighter gases diffuse faster
Avogadro's law	Volume and number of moles	Temperature and pressure	More moles occupy greater volume
Gay-Lussac's combining volumes	Reacting gas volumes	Temperature and pressure	Gases combine in simple volume ratios

16.0 APPLICATIONS OF GAS LAWS IN EVERYDAY LIFE AND INDUSTRY

Gas laws have many applications in Kenya and throughout the world.

16.1 Transportation

1. **Vehicle tyres:** The pressure inside vehicle tyres changes with temperature. Drivers should check tyre pressure when tyres are cold and follow the manufacturer's recommended pressure.
2. **Bicycle tyres:** Air is compressed into bicycle tyres using pumps.
3. **Air brakes:** Compressed air is used in braking systems in some heavy vehicles.
4. **Air conditioning:** Gas pressure, temperature and volume relationships are important in refrigeration and air-conditioning systems.

16.2 Agriculture

1. **Greenhouse farming:** Gas diffusion influences the movement of carbon dioxide and oxygen around crops.
2. **Soil aeration:** Oxygen diffuses through soil pores to plant roots.
3. **Animal respiration:** Oxygen diffuses into the blood while carbon dioxide diffuses out.
4. **Grain storage:** Gas composition and movement are considered in controlled-atmosphere storage.

16.3 Medicine

1. Oxygen cylinders store compressed oxygen for medical use.
2. Medical oxygen regulators control gas delivery pressure.
3. Respiratory equipment depends on pressure and gas-flow principles.
4. Gas diffusion is important in the exchange of oxygen and carbon dioxide in the lungs.
5. Anaesthetic gases are administered using controlled gas delivery systems.

16.4 Industry

1. Compressed gases are used in welding.
2. Industrial gas cylinders store gases under pressure.
3. Gas laws help engineers design storage tanks and pipelines.
4. Ammonia production uses gas-volume relationships.
5. Gas behaviour is considered in chemical manufacturing.
6. Industrial furnaces depend on controlled gas supply and pressure.

16.5 Domestic applications

1. LPG cylinders store liquefied petroleum gas under pressure.
2. Pressure cookers use increased steam pressure to cook food faster.
3. Aerosol sprays use pressurised gases to release products.
4. Air fresheners spread through diffusion.
5. Gas burners use controlled gas flow during cooking.

16.6 Environmental applications

1. Diffusion explains the spread of gases in the atmosphere.
2. Gas exchange occurs between water bodies and the atmosphere.
3. Oxygen and carbon dioxide movement is important in ecosystems.
4. Gas behaviour is considered in pollution monitoring.
5. Diffusion helps explain the movement of atmospheric gases.

17.0 PRACTICAL INVESTIGATIONS AND EXPERIMENTS

Practical 1: Investigating Boyle's law

Aim: To establish the relationship between pressure and volume of a fixed mass of gas at constant temperature.

Materials: Gas syringe, pressure gauge, trapped air and connecting tubes.

Procedure:

1. Trap a fixed mass of air in the apparatus.
2. Record its initial volume and pressure.
3. Compress the gas gradually.
4. Record pressure and volume at each stage.
5. Repeat the measurements.
6. Calculate the product of pressure and volume.

Expected finding: The product of pressure and volume remains approximately constant.

Safety: Do not exceed the pressure rating of the apparatus.

Practical 2: Investigating Charles' law

Aim: To establish the relationship between gas volume and temperature at constant pressure.

Materials: Gas syringe with movable piston, thermometer, beakers and water baths.

Procedure:

1. Trap a fixed amount of air in the syringe.
2. Record the initial temperature and volume.
3. Place the syringe in a water bath.
4. Increase the temperature gradually.
5. Record the corresponding gas volume.
6. Convert temperatures into kelvin.
7. Plot volume against temperature.

Expected finding: The volume increases as absolute temperature increases.

Safety: Use a water bath rather than placing the syringe directly over a flame.

Practical 3: Demonstrating diffusion of gases

Aim: To demonstrate that gases diffuse and mix with one another.

Materials: Long glass tube, cotton wool, ammonia solution, hydrochloric acid and stoppers.

Procedure:

1. Place small pieces of cotton wool at opposite ends of a suitable glass tube.
2. Moisten one with ammonia solution and the other with concentrated hydrochloric acid under teacher supervision.
3. Insert the cotton wool simultaneously.
4. Close the tube with suitable stoppers.
5. Observe the formation of a white ring.

Observation: A white ring of ammonium chloride forms closer to the hydrochloric acid end.

Equation:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

Conclusion: Ammonia diffuses faster than hydrogen chloride because it has a lower molar mass.

Safety: Ammonia and hydrogen chloride vapours are irritating. Conduct the demonstration in a fume cupboard.

Practical 4: Demonstrating gas volume ratios

Aim: To investigate the volume relationships between reacting gases.

Materials: Approved gas-volume apparatus and suitable gas samples.

Procedure:

1. Use a suitable gas-volume apparatus.
2. Measure known volumes of reacting gases.
3. Carry out a teacher-supervised reaction.
4. Record the volume changes.
5. Compare the observed volume ratios with balanced chemical equations.

Conclusion: Reacting gas volumes combine in simple whole-number ratios when measured at the same temperature and pressure.

18.0 COMMON ERRORS IN GAS LAW CALCULATIONS

Learners should avoid the following mistakes:

1. **Using Celsius instead of kelvin.**

   Always convert temperature to kelvin before substituting into gas law equations.

2. **Mixing pressure units.**

   Do not substitute one pressure in kPa and another in Pa without converting them to the same unit.

3. **Mixing volume units.**

   Use consistent volume units throughout the calculation.

4. **Using the wrong formula.**

   Identify which quantities remain constant before selecting a gas law.

5. **Failing to identify initial and final conditions.**

   Write down the known values before substituting.

6. **Incorrect rearrangement of equations.**

   Rearrange the equation carefully to make the unknown quantity the subject.

7. **Confusing the two Gay-Lussac laws.**

   Remember that pressure law concerns pressure and temperature at constant volume, while combining volumes concerns reacting gas-volume ratios.

8. **Confusing rate and time in diffusion.**

   For a fixed distance, a faster-diffusing gas takes less time.

9. **Ignoring the reaction equation coefficients.**

   For gas-volume calculations, use the balanced equation.

10. **Assuming all gases behave ideally under all conditions.**

    Real gases deviate from ideal behaviour, particularly at high pressure and low temperature.

20.0 SUMMARY OF IMPORTANT FORMULAE

Concept	Formula
Pressure	$P = \frac{F}{A}$
Celsius to kelvin	$T = t + 273$
Boyle's law	$P_1V_1 = P_2V_2$
Charles' law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Combined gas law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Gay-Lussac's pressure law	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Graham's law	$\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$
Avogadro's law	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Ideal gas equation	$PV = nRT$

21.0 REVISION QUESTIONS

SECTION A: MULTIPLE-CHOICE QUESTIONS

Choose the correct answer from the alternatives provided.

1. Which statement best describes a gas?

A. It has a fixed shape and fixed volume.
B. It has a fixed volume but no fixed shape.
C. It has neither a fixed shape nor a fixed volume.
D. It has a fixed shape but no fixed volume.

2. Gas pressure is caused by:

A. Attraction between gas particles.
B. Collisions of gas particles with container walls.
C. The colour of gas particles.
D. The mass of the container only.

3. Which unit is used to measure pressure in the SI system?

A. Newton
B. Pascal
C. Joule
D. Mole

4. A temperature of 27°C is equivalent to:

A. 200 K
B. 273 K

C. 300 K
D. 327 K

5. Boyle's law relates:

A. Pressure and temperature.
B. Volume and temperature.
C. Pressure and volume.
D. Volume and number of moles.

6. According to Boyle's law, when the pressure of a fixed mass of gas increases at constant temperature, its volume:

A. Increases.
B. Decreases.
C. Remains constant.
D. Becomes zero.

7. Which mathematical expression represents Boyle's law?

A. $P_1V_1 = P_2V_2$
B. $P_1T_1 = P_2T_2$
C. $V_1T_1 = V_2T_2$
D.

8. A gas occupies 100 cm^3 at 200 kPa. What volume will it occupy at 100 kPa at constant temperature?

A. 25 cm^3
B. 50 cm^3
C. 100 cm^3
D. 200 cm^3

9. Which law explains the expansion of a gas when heated at constant pressure?

A. Boyle's law
B. Charles' law
C. Graham's law
D. Avogadro's law

10. Charles' law states that gas volume is directly proportional to:

A. Pressure.
B. Absolute temperature.
C. Density.
D. Molar mass.

11. Which temperature scale must be used in Charles' law calculations?

A. Celsius
B. Fahrenheit
C. Kelvin
D. Centigrade only

12. A gas occupies 200 cm^3 at 300 K. What volume will it occupy at 600 K at constant pressure?

A. 100 cm^3
B. 200 cm^3

C. 300 cm^3
D. 400 cm^3

13. The theoretical temperature at which an ideal gas would have zero volume by Charles' law extrapolation is:

A. 0°C
B. −100°C
C. −273°C
D. 100°C

14. The combined gas law relates:

A. Pressure, volume and temperature.
B. Pressure and mass only.
C. Volume and density only.
D. Temperature and molar mass only.

15. Which equation represents the combined gas law?

A. $P_1V_1 = P_2V_2$
B. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
C. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
D. $PV = n$

16. Gay-Lussac's pressure law states that pressure is directly proportional to absolute temperature when:

A. Volume is constant.
B. Pressure is constant.
C. Temperature is constant.
D. Mass is changing.

17. A gas is heated in a sealed rigid container. Its pressure increases because:

A. The particles stop moving.
B. The particles move faster and collide more energetically with the walls.
C. The number of particles decreases.
D. The container volume increases.

18. Diffusion is the movement of particles from:

A. Low concentration to high concentration only.
B. High concentration to low concentration.
C. High pressure to low temperature only.
D. Low temperature to high pressure only.

19. Which gas diffuses fastest under the same conditions?

A. Oxygen
B. Carbon dioxide
C. Hydrogen
D. Chlorine

20. Graham's law relates the rate of diffusion to:

A. The square root of molar mass.
B. The square root of volume.

C. Pressure only.
D. Temperature only.

21. Which gas diffuses faster under the same conditions?

A. NH_3
B. HCl
C. CO_2
D. SO_2

22. The molar mass of hydrogen gas is:

A. 1 g/mol
B. 2 g/mol
C. 16 g/mol
D. 32 g/mol

23. Avogadro's law states that equal volumes of gases at the same temperature and pressure contain equal numbers of:

A. Atoms only.
B. Molecules.
C. Grams.
D. Neutrons.

24. Which expression represents Avogadro's law?

A. $V \propto n$
B.
C. $P \propto T$
D. $r \propto M$

25. The conventional molar volume of an ideal gas at 0°C and 1 atmosphere is approximately:

A. 2.24 dm^3
B. 11.2 dm^3
C. 22.4 dm^3
D. 44.8 dm^3

26. Gay-Lussac's law of combining volumes states that reacting gas volumes are in:

A. Random ratios.
B. Simple whole-number ratios.
C. Decimal ratios only.
D. Equal ratios in all reactions.

27. What is the volume ratio of hydrogen to oxygen in the formation of water vapour?

$$2H_2 + O_2 \rightarrow 2H_2O$$

A. 1:1
B. 1:2
C. 2:1
D. 2:3

28. What volume of oxygen reacts with 60 cm^3 of hydrogen to form water vapour?

A. 15 cm^3
B. 30 cm^3

C. 60 cm^3
D. 120 cm^3

29. Which of the following represents the ideal gas equation?

A. $PV = nRT$
B.
C.
D. $PT = VR$

30. Which condition favours ideal gas behaviour?

A. High pressure and low temperature.
B. Low pressure and high temperature.
C. High pressure and high temperature.
D. Low temperature and high density.

SECTION B: STRUCTURED QUESTIONS

1. Kinetic theory of gases
   a) Define a gas.
   b) State six assumptions of the kinetic theory of gases.
   c) Explain how gas pressure develops inside a closed container.
   d) Explain why gases are easily compressed.
   e) State four properties of gases.
   f) Explain the effect of increasing temperature on the average kinetic energy of gas particles.

2. Gas pressure
   a) Define pressure.
   b) State the SI unit of pressure.
   c) Write the mathematical expression for pressure.
   d) A force of 200 N acts on an area of 0.5 m^2. Calculate the pressure exerted.
   e) Convert 150 kPa into pascals.
   f) Convert 2 atmospheres into kilopascals.

3. Boyle's law
   a) State Boyle's law.
   b) Explain Boyle's law using the kinetic theory of gases.
   c) Describe an experiment to verify Boyle's law.
   d) State three precautions taken during the experiment.
   e) Sketch a graph of pressure against volume.
   f) Explain why the graph obtained is a rectangular hyperbola.

4. Boyle's law calculations

A gas occupies 500 cm^3 at a pressure of 100 kPa.

a) Calculate its volume when pressure increases to 250 kPa at constant temperature.

b) Calculate the pressure when the original gas volume is reduced to 200 cm^3.

c) Explain what happens to the volume when pressure is reduced by half.

5. Charles' law

a) State Charles' law.

b) Explain Charles' law using the kinetic theory of gases.

c) Describe an experiment to verify Charles' law.

d) State four precautions taken during the experiment.

e) Explain the importance of using kelvin rather than degrees Celsius in gas law calculations.

f) Sketch a graph of volume against absolute temperature.

6. Charles' law calculations

A gas occupies 250 cm^3 at 27°C.

a) Convert 27°C into kelvin.

b) Calculate the volume of the gas at 127°C if pressure remains constant.

c) Explain why the volume increases when the temperature increases.

7. Combined gas law

a) State the combined gas law.

b) Derive the combined gas law from Boyle's law and Charles' law.

c) State three conditions necessary when applying the combined gas law.

d) A gas occupies 400 cm^3 at 100 kPa and 300 K. Calculate its volume at 150 kPa and 450 K.

8. Gay-Lussac's pressure law

a) State Gay-Lussac's pressure law.

b) Explain the law using the kinetic theory of gases.

c) Describe an experiment that can be used to demonstrate the law.

d) State three applications of the law.

e) A gas has a pressure of 120 kPa at 300 K. Calculate its pressure at 450 K if volume remains constant.

9. Graham's law of diffusion

a) Define diffusion.

b) State Graham's law of diffusion.

c) Explain why lighter gases diffuse faster than heavier gases at the same temperature.

d) State five factors affecting the rate of diffusion of gases.

e) Explain why a white ring forms closer to the hydrogen chloride end when ammonia and hydrogen chloride gases diffuse towards each other.

10. Graham's law calculations
    a) Calculate the ratio of the rates of diffusion of hydrogen and oxygen.
    b) Calculate the ratio of the rates of diffusion of ammonia and hydrogen chloride.
    c) A gas with molar mass 4 g/mol diffuses twice as fast as another gas. Determine the molar mass of the second gas.
    d) Explain the relationship between diffusion rate and molar mass.

11. Avogadro's law
    a) State Avogadro's law.
    b) Explain why equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules.
    c) Define molar volume.
    d) State the conventional molar volume of an ideal gas at 0°C and 1 atmosphere.
    e) A gas occupies 4 dm^3 when it contains 0.2 moles. Calculate the volume occupied by 0.5 moles at the same temperature and pressure.

12. Gay-Lussac's law of combining volumes
    a) State Gay-Lussac's law of combining volumes.
    b) Explain the relationship between the law and balanced chemical equations.
    c) State the volume ratio of nitrogen, hydrogen and ammonia in the Haber process.
    d) Calculate the volume of hydrogen required to react with 10 dm^3 of nitrogen.
    e) Calculate the volume of ammonia formed.

13. Gas reactions

Hydrogen reacts with oxygen to produce water vapour according to the equation:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

    a) State the volume ratio of the reacting gases.
    b) Calculate the volume of oxygen needed to react with 100 cm^3 of hydrogen.
    c) Calculate the volume of water vapour produced.
    d) Explain the conditions under which the volume ratios are valid.

14. Applications of gas laws
    a) Explain how Boyle's law applies to breathing.
    b) Explain how Charles' law applies to hot-air balloons.
    c) Explain how Gay-Lussac's pressure law applies to vehicle tyres.
    d) Explain why aerosol containers should not be exposed to excessive heat.
    e) State four applications of gas laws in Kenyan industries.

15. Integrated gas law calculations

A fixed mass of gas occupies 600 cm^3 at a pressure of 100 kPa and temperature of 300 K.

a) Calculate its volume when pressure changes to 150 kPa and temperature changes to 450 K.

b) Calculate the final pressure if the volume is maintained at 600 cm^3 and temperature increases to 450 K.

c) State the gas law used in each calculation.

16. Diffusion in everyday life

a) Explain the role of diffusion in human respiration.

b) Explain how diffusion contributes to the spreading of perfume in a room.

c) State three differences between diffusion in gases and diffusion in liquids.

d) Explain why diffusion generally occurs faster at higher temperatures.

e) State four applications of diffusion in agriculture and medicine.

17. Ideal gas equation

a) Define an ideal gas.

b) Derive the ideal gas equation from the gas laws.

c) State the meaning of each symbol in the equation:

$$PV = nRT$$

d) Calculate the volume occupied by 0.5 moles of an ideal gas at 300 K and 100 kPa.

e) State two conditions under which real gases behave approximately ideally.

18. Practical investigations

a) Describe an experiment to investigate the relationship between pressure and volume of a fixed mass of gas.

b) Describe an experiment to investigate the relationship between volume and temperature.

c) State the apparatus required in each experiment.

d) State four precautions necessary when carrying out gas law experiments.

e) Explain the importance of recording accurate experimental data.

19. Industrial applications and safety

a) Explain five applications of gas laws in industry.

b) State four dangers associated with pressurised gas cylinders.

c) Explain why gas cylinders should be stored away from excessive heat.

d) State five safety precautions when handling gases in a chemistry laboratory.

e) Explain how gas leaks can affect human health and the environment.

20. Comprehensive comparison

a) Differentiate between Boyle's law and Charles' law.

b) Distinguish between Charles' law and Gay-Lussac's pressure law.

c) Differentiate between Gay-Lussac's pressure law and Gay-Lussac's law of combining volumes.

d) Compare Avogadro's law and Graham's law.

e) Explain how the combined gas law brings together Boyle's law and Charles' law.

Revision notes

Gas Laws

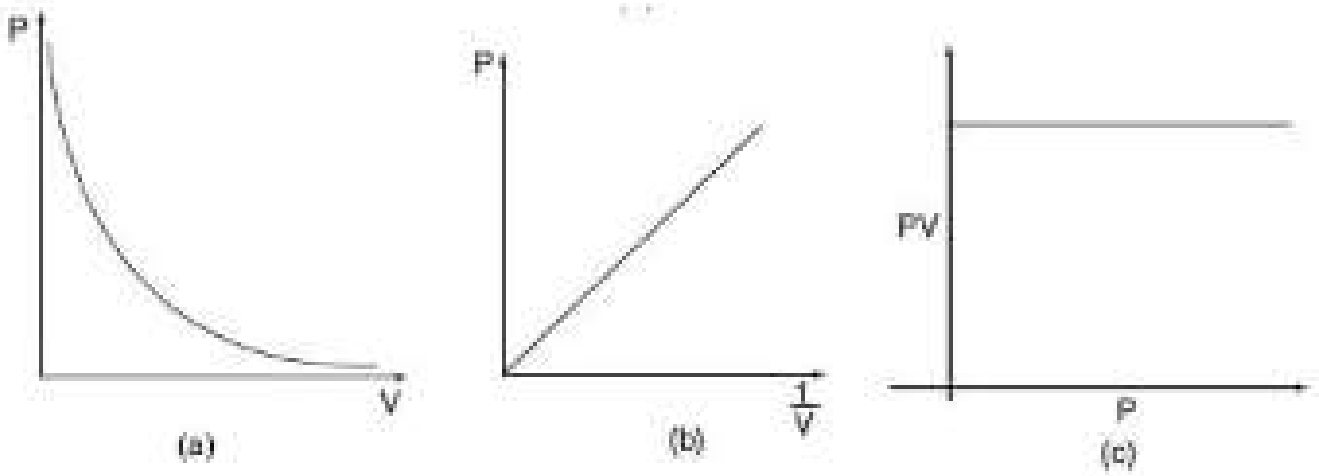
Boyle's Law

Boyle's law states that **the pressure of a fixed mass of a gas is inversely proportional to its volume, provided the temperature is kept constant.**

Stated in symbols;

P α , or P = k ×

So, PV = constant



The graph of P against V is a smooth curve, as shown in (a), while that of P against is a straight line passing through the origin. That of PV against P is a straight line parallel to the xaxis.

Since **PV = constant;**

$P_1V_1 = P_2V_2$ = constant, for any given mass of a gas.

Using Kinetic Theory to Explain Boyle's Law

1. If the volume of a vessel containing a fixed mass of gas is halved, the number of molecules per unit volume will be doubled. The number of collisions per unit time, and therefore the rate of change of momentum, will also be doubled. Consequently, halving the volume of the gas doubles the pressure, which is the import of the Boyle's law.
2. In a demonstration using a syringe, as volume of gas is reduced, there is increase in number of molecular collisions resulting to increase in pressure

Problem:

A sample of chlorine gas occupies a volume of 946mL at a pressure of 726 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL.

Given:

V_1 = 946 ml

P_1 = 726 mmHg

p_2 = ?

v_2 = 154 mL

$P_1V_1 = P_2V_2$

$P_1V_1/V_2 = P_2$

(760 mmHg)(946 mL)/154mL = P_2

4460 mmHg = P_2

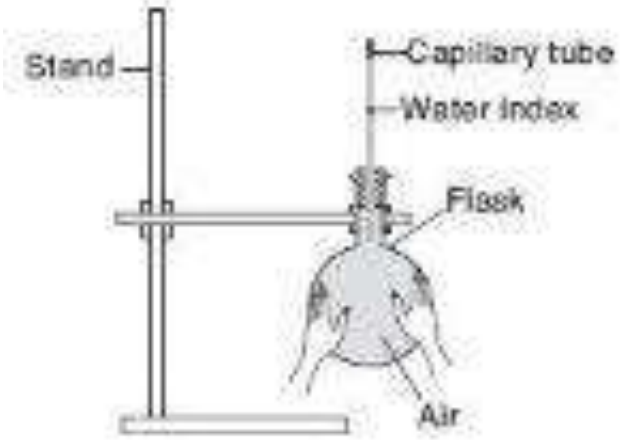
Charles' Law

- Charles law relates the volume of a gas with its absolute (or Kelvin) temperature.

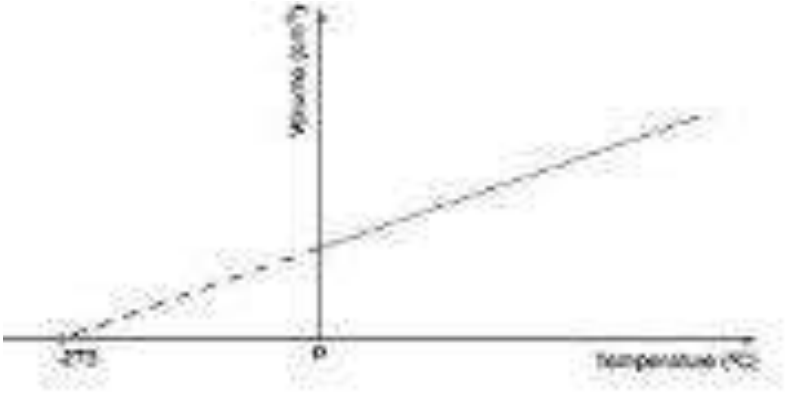
- Charles' law states that **the volume of a fixed mass of gas is directly proportional to its absolute temperature if the pressure is kept constant.**

In symbols, Charles' law can be stated as follows;
V α T or V = kT, where k is a constant.
Hence,
$V_1/T_1 = V_2/T_2$

Experiment to verify Charles law



- As the temperature rises, the height h (volume) also increases. A plot of volume against temperature is represented alongside.
- The graph is a straight line, indicating proportional changes in volume and temperature. However, it does not pass through the origin.
- If the graph is extrapolated, it cuts the temperature axis at about –273°C. At this temperature, the volume of the gas is assumed to be zero.



- This temperature, –273°C, at which the volume, pressure of the gas and kinetic energy of the particles are assumed to be zero is ideally the lowest temperature a gas can attain and therefore called **absolute zero.**

Using Kinetic Theory to explain Charles' Law

- When a gas is heated, the **kinetic energy and the velocity of the molecules increases.** As the temperature rises, the molecules **move faster.**
- If the volume of the container were constant, the pressure resulting from the collisions of the molecules with the walls would **increase due to greater rate of change of momentum per unit time.**
- But since Charles' law requires that the pressure be constant, then **the volume must increase** accordingly so that although the molecules are moving faster, the **number** of collisions at the walls of the container per unit time is **reduced**, since the distance between the walls is increased by increasing the volume.

Relation Between Celsius and Absolute scale

- The zero Kelvin (0 K) corresponds to –273°C while 0°C corresponds to 273 K.
- It follows that to change from Celsius to Kelvin, we add 273 to the Celsius temperature, i.e.;

θ°C = T = (θ + 273) K

Example 1

The temperature of a gas is -42°C. What is this temperature on the Kelvin scale?

Solution

Temperature T = (- 42 + 273) K

= 231 K

Example

Problem: A sample of carbon monoxide gas occupies 3.20 L at 125°C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

V_1 = 3.2L

T_1 = 125°C → K Conversion Required → 125°C + 273 = 398K

T_2 = ?

V_2= 1.54L

$V_1/T_1 = V_2/T_2$

T_2 = (1.54L)(398k)/3.2L

T_2 = 192K

Combined Gas Law

- A general gas law relating the changes in pressure, volume and the absolute temperature can be derived from the three gas laws.
- Consider a fixed mass of gas which is being changed from state A to state B through an intermediate state C, as shown below. From A to C, the gas is heated at constant pressure

P1. By Charles' law;

$V_1/T_1 = V_2/T_2$

From C to B, the gas pressure is changed from P_1 to P_2 at constant temperature T_2.

By Boyle's law, $P_1Vc = P_2V_2$

$V_2 = P_1Vc/P_2$

But $Vc = V_1T_2/T_1$

Therefore: $V_2 = P_1V_1T_2/P_2T_1$

Re-arrange: $P_1V_1/T_1 = P_2V_2/T_2$

- This is referred to as the **ideal gas equation**

Assumptions of Gas laws

- When explaining the gas laws using the kinetic theory, both **the size of molecules** and the **inter-molecular forces** are assumed to be negligible.

A real gas is defined as a gas that does not obey gas laws at all standard pressure and temperature conditions.

- Real gases have velocity, volume and mass. A real gas is a gas occupies space and the molecules have interactions.
- Real gases have molecules with definite volumes and therefore the idea of zero volume or zero pressure is not real. Real gases get liquified before zero volume is reached.

An ideal gas is defined as a gas that obeys gas laws at all conditions of pressure and temperature.

- Ideal gases have velocity and mass. They do not have volume. When compared to the total volume of the gas the volume occupied by the gas is negligible. It does not condense and does not have triple point
- An ideal gas is a theoretical gas composed of many randomly moving particles that are not subject to interparticle interactions.

QUESTIONS

(1). A mass of 1200 cm^3 of oxygen at 27°C and a pressure 1.2 atmosphere is compressed until its volume is 600 cm^3 and its pressure is 3.0 atmosphere. Calculate the temperature of the gas after compression in °C. (2 marks)

(2). A gas occupies a volume of 400cm^3 at 500k and 1 atmospheric pressure. What will be the temperature of the gas when the volume and pressure of the gas is 100cm^3 and 0.5 atmospheric pressure respectively? (2 marks)

(3). 200$cm^3$of gas "p" at s.t.p was cooled and the volume contracted to 160cm^3. Calculate the new temperature of the gas in °C if the pressure is kept constant. (2 marks)

(4). A gas occupies a volume of 400cm^3 at 500k and 1 atmosphere pressure. What will be the temperature of the gas when the volume and pressure of the gas is 100cm^3 and 0.5 atmospheres respectively. (2 marks)

(5). A certain mass of gas occupies 0.15dm^3 at 293K and 98,648.5 Pa. Calculate its volume at 101325 Pa and 273K (2 marks)

Graham's Law

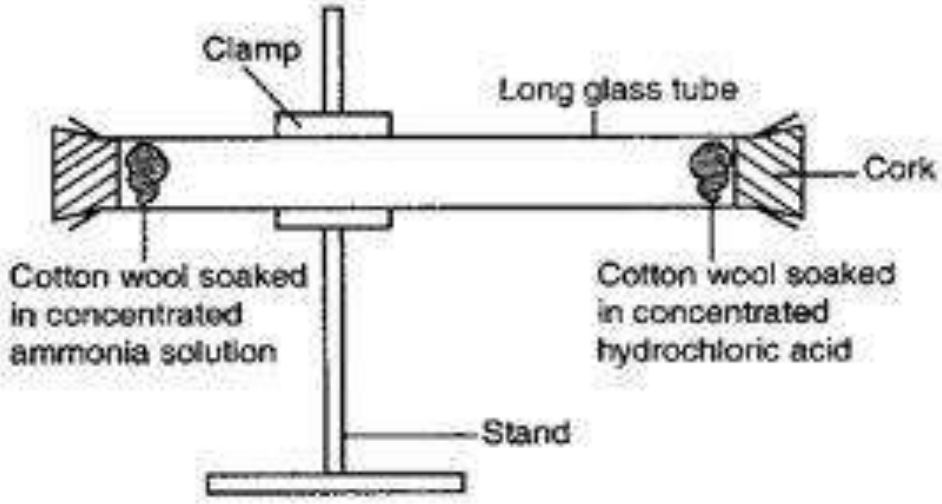
DIFFUSION

- The continuous random motion of gases differ from gas to the other.
- The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion.**
- The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas**.
- The density of a gas depends on its relative molecular mass. i.e. **The** higher **the density the** higher **the relative atomic mass and thus the** lower **the rate of diffusion.**

Examples

1. Carbon (IV)oxide(CO_2) has a molar mass of 44g. Nitrogen(N_2) has a molar mass of 28g. (N_2)is thus lighter or less dense than Carbon (IV)oxide(CO_2). N_2 diffuses faster than CO_2.
2. Ammonia(NH_3) has a molar mass of 17g. Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus about **twice** lighter or less dense than Ammonia(NH_3). Ammonia(NH_3) diffuses twice faster than N_2.
3. Ammonia(NH_3) has a molar mass of 17g. Hydrogen chloride gas has a molar mass of 36.5g. Both gases on contact react to form **white fumes** of ammonium chloride.

- When a cotton wool dipped in ammonia and another cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower



The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**. Grahams law states that:

- **"The rate of diffusion of a gas is inversely proportional to the square root of its density, at the same temperature and pressure"**

$$\text{Rate (R)} \alpha = \frac{1}{\sqrt{\text{density (d)}}} \qquad \text{Rate (R)} = \frac{\text{Constant}}{\sqrt{\text{density}}}$$

$$\text{Therefore } R_A\sqrt{d_A} = R_B\sqrt{d_B}$$

$$\frac{R_A}{R_B} = \sqrt{\frac{d_B}{d_A}}$$

Since density is directly proportional to molecular mass; The law can also be expressed as

$$\text{Rate} = \frac{\text{Constant}}{\sqrt{\text{molecular mass}}}$$

$\frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$ where M_A and M_B are relative molecular masses of the gases A and B respectively.

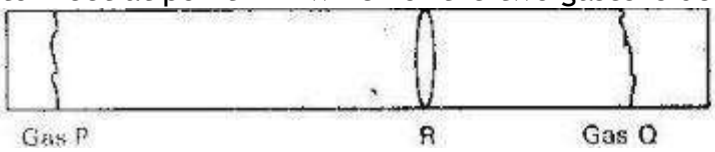
Since Rate is inversely proportional to time $\left(\text{Rate } \alpha \frac{1}{\text{Time}}\right)$

$\frac{R_A}{R_B} = \frac{T_B}{T_A}$ Where T_A and T_B are the times of diffusion of gases A and B respectively.

QUESTIONS

(1). Explain why the volume of a gas increases when its temperature is increased at a constant pressure. (1mk)

(2). Cotton wool pads were socked with concentrated solutions of gas "p" and gas "Q" the pads were then placed of the opposite ends of a long horizontal glass tube at the same time. The tube was then immediately corked at both ends as shown the diagram below. After sometimes the gases were observed to meet at point "R" which of the two gases is dense? Explain your answer



(3). A given volume of Ozone (O_3) diffused from a certain apparatus in 96 seconds. Calculate the time taken by equal volume of carbon (IV) oxide (CO_2) to diffuse under the same condition(O= 16) (C=12) (2mks)

(4). A few crystals of potassium manganate VII were carefully placed in a beaker at one spot. The beaker was left undisturbed for two hours. State and explain the observation that was made. (2mks)

(5). $60cm^3$ of oxygen gas diffused through a porous partition in 50 seconds. How long would it take $60cm^3$ of sulphur (IV) oxide gas to diffuse through the same conditions? (S= 32.0) (O=16.0). (3mks)

(6).

a. State the Graham's law of diffusion (1 mark) b. The molar masses of gases w and x are 16.0 and 44.0 respectively. If the rate of diffusion of w through a porous material is $12cm^3S^{-1}$, calculate the rate of diffusion of x through the same material. (2marks)

(7). Calculate the R.F.M of gas "A" given that the time taken for equal volumes of oxygen and gas "A" to diffuse through a hole is 20 seconds and 24 seconds respectively (O =16.0) (2mks)

(8). A certain volume of CO_2 gas takes 200 seconds to diffuse through porous plug. How long would it take the same volume of HCL to diffuse under the same condition? (3mks)

(9). Calculate the relative formula mass of gas **A** given that the time taken for equal volumes of oxygen to diffuse through the use same hole is 20 seconds and 24 seconds respectively (O = 16.0)

TOPICAL REVISION QUESTIONS

1. A sample of unknown compound gas X is shown by analysis to contain Sulphur and Oxygen. The gas requires 28.3 seconds to diffuse through a small aperture into a vacuum. An identical number of oxygen molecules pass through the same aperture in 20 seconds. Determine the molecular mass of gas X (O= 16, S= 32)
2.
   a. State Graham's Law of diffusion
   b. Gas **V** takes 10 seconds to diffuse through a distance of one fifth of a meter. Another gas W takes the same time to diffuse through a distance of 10 cm. if the relative molecular mass of gas **V** is 16.0; calculate the molecular mass of **W**
3.
   a. State Charles' Law
   b. The volume of a sample of nitrogen gas at a temperature of 291K and 1.0×10^5 Pascals was 3.5×10^{-2} m^3. Calculate the temperature at which the volume of the gas would be 2.8×10^{-2} m^3 at 1.0×10^5 pascals.
4. 60 cm^3 of oxygen gas diffused through a porous partition in 50 seconds. How long would it take 60 cm^3 of sulphur (IV) oxide gas to diffuse through the same partition under the sane conditions? (S = 32.0, O = 16.0)
5.
   a. State Graham's law of diffusion
   b. 30cm^3 of hydrogen chloride gas diffuses through a porous pot in 20 seconds. How long would it take 42cm^3 of sulphur(IV) oxide gas to diffuse through the same pot under the same conditions (H =1 Cl = 35.5 S = 32 O =16)
6.
   a. State **Boyles law**
   b. Sketch a graph that represents Charles' law
   c. A gas occupied a volume of 250 cm^3 at -23° C and 1 atmosphere. Determine its volume at 127° C when pressure is kept constant.
7. A factory produces Calcium Oxide from Calcium Carbonate as shown in the equation below:-

   $$CaCO_{3\,(s)} \xrightarrow{heat} CaO_{\,(s)} + CO_{2\,(g)}$$

   What volume of Carbon (IV) Oxide would be produced from 1000kg of Calcium Carbonate at s.t.p (Ca = 40, C = 12, O = 16, Molar gas volume at s.t.p = 22.4dm^3)
8. A fixed mass of gas occupies 200 cm^3 at a temperature of 23oC and pressure of 740mmHg. Calculate the volume of the gas at -25oC and 780mmHg pressure
9. Gas **K** diffuses through a porous material at a rate of 12cm^3 s^{-1} where as S diffuses through the same material at a rate of 7.5cm^3s^{-1}. Given that the molar mass of **K** is 16, calculate the molar mass of **S**
10. State Gay Lussac's law
11.
    a. What is the relationship between the rate of diffusion of a gas and its molecular mass?
    b. A sample of Carbon (IV) Oxide takes 200 seconds to diffuse across a porous plug. How long will it take the same amount of Carbon (II) Oxide to diffuse through the same plug?(C=12, O=16)
12. Below are structures of particles. Use it to answer questions that follow. In each case only electrons in the outermost energy level are shown

    key
    P = Proton
    N = Neutron
    X = Electron
    a. Identify the particle which is an anion
    b. Choose a pair of isotopes. Give a reason

13. The figure below shows two gases P and Q diffusing from two opposite ends 18 seconds after the experiment
    a. Which of the gases has a lighter density?
    b. Given that the molecular mass of gas Q is 17, calculate the molecular mass of P
14. Identify the particles that facilitate the electric conductivity of the following substances
    a. Sodium metal
    b. Sodium Chloride solution
    c. Molten Lead Bromide
15. Gas B takes 110 seconds to diffuse through a porous pot, how long will it take for the same amount of ammonia to diffuse under the same conditions of temperature and pressure? (RMM of B = 34 RMM of ammonia = 17)
16. A gas occupies 5dm^3 at a temperature of -27^{o}C and 1 atmosphere pressure. Calculate the volume occupied by the gas at a pressure of 2 atmospheres and a temperature of 127^{o}C
17. A fixed mass of gas occupies 200 cm^3 at a temperature of 23^{0}c and a pressure of 740 mm Hg. Calculate the volume of the gas at -25^{0}c and 790 mm Hg pressure.
18.
    a. State the Graham's law
    b. 100cm^3 of Carbon (IV) oxide gas diffused through a porous partition in 30seconds. How long would it take 150cm3 of Nitrogen (IV) oxide to diffuse through the same partition under the same conditions? (C = 12.0, N = 14.0, O = 16.0)

Answers

1.

X: $t_1 = 28.3$sec $RMM = ?$
Q_2: $t_2 = 20.0$sec $RMM = 32$

$T \propto \sqrt{MM}$ ✓

$$\frac{T_1}{T_2} = \sqrt{\frac{X}{32}}$$

$$\left(\frac{T_1}{T_2}\right)^2 = \frac{X}{32}$$ ✓

$$\left(\frac{28.3}{T_2}\right)^2 = \frac{X}{32}$$ ✓

$$X = \frac{28.3^2 \times 32}{400}$$ ✓

$X = 64$ ✓

2.

a. The rate of diffusion of a gas is inversely proportional to the square root of its density under the same conditions of temperature and pressure

b.

Rate of gas $V = {}^1/_5 \times \frac{100cm}{10sec}$
$= 2cm/sec$ ✓ ½
Rate of $W = \frac{10cm}{10sec}$
$= 1cm/sec$ ✓ ½

$$\frac{RV}{RW} = \sqrt{\frac{MW}{MV}} \quad = \frac{2}{1} = \sqrt{\frac{MW}{16}}$$

$$\frac{2^2}{1} = \frac{MW}{16}$$

$$\frac{4}{1} = \frac{MW}{16}$$

$$MW = \frac{4}{1} \times 16$$

$MW = 64$

3.

a. The volume of a fixed mass of a gas is directly proportional to its absolute temperature at constant Pressure

b.

Apply combined gas law; $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$V_1 = 3.5 \times 10^{-2}\ m^3$ $V_2 = 2.8 \times 10^{-2} m^3$
$P_1 = 1.0 \times 10^5 Pa$ $P_2 = 1.0 \times 105 Pa$ ✓ ½
$T_1 = 291K$ $T_2 = ?$

$T_2 = \frac{P_2V_2T_1}{P_1V_1}$

$T_2 = \frac{1.0 \times 10^5 Pa \times 2.8 \times 10^{-2} m^3 \times 291K}{1.0 \times 10^5 Pa \times 3.5 \times 10^{-2} m^3}$

$T_2 = 232.8k$ ✓

4.

$\frac{TsO_2}{TO_2} = \frac{R.M.N.SO_2}{R.M.MO_2}$ ✓½

$SO_2 = 32 + (16 \times 2) = 64$ ✓½

$O_2 = (16 \times 2) = 32$ ✓½

$\frac{TsO_2}{50} = \sqrt{\frac{64}{32}}$ ✓½ $= 70.75$ ✓½

5.

a. The rate of diffusion of a fixed mass of a gas is inversely proportional to the square root of it density at constant temperature and pressure

b.

b) $RHCl = \frac{30\ cm^3}{20\ se} = 1.5\ cm^3/sec$

$\frac{RHCL}{RSO_2} = \frac{\sqrt{MSO_2}}{\sqrt{MHCL}}$

$\frac{(1.5)^2}{RSO_2} = \frac{\sqrt{64}}{\sqrt{36.5}}$

$(RSO_2)^2 = \frac{2.25 \times 36.5}{64}$

$RSO_2 = \sqrt{\frac{2.25 \times 36.5}{64}}$

$= 1.133\ cm^3/Sec$

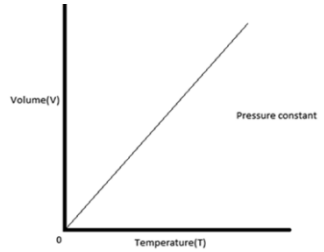
$\frac{1.133\ cm^3}{42\ cm^3} = \frac{1\ sec}{42 \times 1}{1.133}$

$= 37\ sec$

6.

a. Boyles' law For a fixed mass of a gas, volume is inversely promotional to pressure at constant temperature

b.



c) $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ √½ $V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2}$ √½

$\frac{250 \times 273 - 23}{273 + 127}$ √½

$= 156.5cm^3$

7.

a) RFM of $CaCO_3 = 40 + 12 + 48$

$= 100kg.$ √½

∴ 100 kg of $CaCO_3 \equiv 22.4dm^3$ of $CO_2(g)$

1000 kg " " ⟶ ?

$= \frac{22.4 \times 1000}{100}$ √1 $= 224\ dm^3$ √½

8.

$T_1 = 23+273 = 296$ $T_2 = -25+273 = 248$

$V_1 = 200cm^3$ $V_2 = ?$

$P_1 = 740mmHg$ $P_2 = 780mmHg$

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$\frac{740x200}{296}$ √1 $= \frac{780x?}{248}$ √1

$\therefore x = \frac{740 \times 200 \times 248}{296 \times 780}$

$= 158.974\ cm^3$ √1 (penalize ½ mark for units)

9.

$\frac{Rk}{Rs} = \sqrt{\frac{Ms}{Mk}}$

$\therefore \frac{12}{7.2} = \sqrt{\frac{x}{16}}$ √½

$X = \frac{12^2}{7.2^2} \times 16$ √½

$= 44.464$ √

10. When gases combine they do so in volume which bear a simple ratio to one another and to the product if gaseous under standard temperature and pressure

11.

a. Rate of diffusion is whereby proportional to molecular mass of a gas. √1

b.

$\frac{TCO_2}{TCO} = \frac{\sqrt{MCO_2}}{\sqrt{MCO}}$ √½

$\Rightarrow \frac{200}{T} = \sqrt{\frac{44}{28}} = \sqrt{\frac{44}{28}}\ 11$ √½

$\Rightarrow \left(\frac{200}{T}\right)^2 = \frac{11}{7}$

$\Rightarrow \frac{T}{200} = \frac{\sqrt{7}}{\sqrt{11}}$

$\Rightarrow T = 200.0.79772$ √½ $= 159.5$ Seconds. √½

12.

a. Y √1

b. Z and W √1 have same atomic number but different mass number. √1

13.

a. Gas P

b.

$$\frac{RO}{RP} = \sqrt{\frac{RMMP}{RMMQ}}$$

$$\frac{18}{54} = \sqrt{\frac{x}{17}}$$

$$\frac{1^2}{3^2} = \left(\sqrt{\frac{x}{17}}\right)^2$$

$$\frac{1}{9} = \frac{x}{17}$$

$$9x = 17$$

$$x = {}^{17}/_{9}$$

$$x = 1.88$$

14.

a. Delocalized electrons
b. Mobile ions
c. Mobile ions

15.

$$\frac{TNH_3}{TB} \quad \frac{MNH_3}{MB} \surd ½$$

$$\frac{TNH_3}{TB} = \frac{17}{34}$$

$$\frac{TNH_3}{110} = \frac{17}{34} \surd ½$$

$$TNH3 = 110 X \frac{17}{34} \surd ½ \quad = 77.78 \text{ seconds } \surd ½$$

16.

$$\frac{P_1V_1}{T1} = \frac{P_2V_2}{T_2}$$

$$\frac{1x5}{246} = \frac{2xV_2}{400}$$

$$V2 = \frac{400 x1 x 5}{2 x246}$$

$$= 4.065 \text{ dm}^3$$

17.

a.

$V_1 = 200 \text{ cm}^3$ $V_2 = ?$

$T_1 = 296 K$ $T_2 = 284K$

$P_1 = 740 \text{ mmHg}$ $P_2 = 780 \text{ mm Hg}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{740 \text{ mm Hg} \times 200\text{cm}^3 \times 248K}{296K \times 780 \text{ mm Hg}}$$

$$= 158.97 \text{ cm}^3$$

b. 601 √1

18.

a. Grahams law states
Under the same conditions of pressure and temperature, the rate of diffusion of a gas is inversely proportional to the square root of its density

b.

$$\frac{\text{Time } CO_2}{\text{Time } NO_2} = \frac{\sqrt{M_rCO_2}}{M_rNO_2}$$

Where 100cm^3 of CO_2 takes 30 seconds

$\therefore$ 150cm3 of CO_2 takes ${}^{30}/_{100} \times 150$

$= 45 \text{ seconds} \surd$

$$\frac{45^2}{TNO_2} = 0.975$$

$$\frac{45}{TNO_2} = \frac{\sqrt{44}}{46} \quad TNO_2 = \frac{45}{0.978}$$

$TNO_2 = 46 \text{ sec}$

OR

$$\frac{RCO_2}{RNO_2} = \frac{\sqrt{M_rNO_2}}{M_rCO_2}$$

$$\text{But } RCO_2 = \frac{100\text{cm}^3}{30 \text{ s}} = 3.33 \text{ cm}^3 \text{ per sec}$$

$$\frac{3.33}{RNO_2} = \frac{\sqrt{46}}{44}$$

$$= 1.0225$$

$$RNO_2 = \frac{3.33}{1.0225}$$

$= 3.26 \text{ cm}^3 \text{ per second}$

$$\text{Time for No} = \frac{150 \text{ cm}^3}{3.26\text{cm sec}^{-1}} = 46 \text{ secs}$$

STRAND 2.0: PHYSICAL CHEMISTRY

SUB-STRAND 2.3: REACTION RATES AND CHEMICAL EQUILIBRIUM

INTRODUCTION

Chemical reactions take place at different speeds. Some reactions occur almost instantly, while others take hours, months or even years to show noticeable changes.

For example:

- Burning of magnesium ribbon takes place rapidly.
- Reaction between dilute hydrochloric acid and sodium carbonate produces carbon dioxide quickly.
- Rusting of iron takes place slowly.
- Fermentation of milk takes several hours or days.
- Formation of petroleum and some minerals takes millions of years.

Chemists study the speed at which chemical reactions occur because reaction rates are important in controlling industrial processes, improving product yields, reducing production costs and ensuring safety.

Some chemical reactions can proceed in both forward and reverse directions. Such reactions are called reversible reactions. When the forward and reverse reactions occur at equal rates, the system reaches a state known as dynamic chemical equilibrium.

Chemical equilibrium is important in the manufacture of fertilisers, nitric acid and sulphuric acid, among other industrial chemicals.

PART A: REACTION RATES

4.0 MEANING OF A CHEMICAL REACTION

A chemical reaction is a process in which one or more substances called reactants are converted into new substances called products through the breaking and formation of chemical bonds.

During a chemical reaction:

- Reactant particles interact with one another.
- Existing chemical bonds may break.
- New chemical bonds form.
- Atoms are rearranged.
- New substances with different properties are produced.
- Energy may be released or absorbed.

Example: Reaction between hydrogen and oxygen

Hydrogen reacts with oxygen to form water.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Reactants:

- Hydrogen gas.
- Oxygen gas.

Product:

- Water.

Example: Reaction between magnesium and oxygen

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Magnesium burns in oxygen with a bright white flame to form magnesium oxide.

5.0 MEANING OF RATE OF CHEMICAL REACTION

The rate of a chemical reaction is the change in the amount or concentration of a reactant used up or product formed per unit time.

In simple terms, reaction rate describes how fast or slowly a chemical reaction takes place.

A fast reaction produces noticeable changes within a short time, while a slow reaction takes a longer time.

5.1 Examples of fast chemical reactions

1. Burning of magnesium ribbon in air.
2. Reaction between dilute hydrochloric acid and sodium carbonate.
3. Reaction between silver nitrate and sodium chloride solutions.
4. Combustion of petrol in a vehicle engine.
5. Neutralisation between a strong acid and a strong alkali.

5.2 Examples of slow chemical reactions

1. Rusting of iron.
2. Fermentation of milk.
3. Decomposition of organic matter.
4. Weathering of limestone.
5. Ripening of fruits through chemical and biochemical changes.

5.3 Importance of studying reaction rates

Studying reaction rates helps chemists to:

1. Control the speed of industrial chemical processes.
2. Increase the production of useful substances.
3. Reduce the time required for manufacturing.
4. Minimise production costs.
5. Improve the efficiency of chemical reactions.
6. Prevent dangerous reactions from occurring too rapidly.

7. Determine suitable storage conditions for chemicals.
8. Understand biological processes such as digestion and respiration.
9. Improve food preservation methods.
10. Develop suitable catalysts for industrial processes.

6.0 COLLISION THEORY

6.1 Definition of collision theory

Collision theory is a scientific explanation which states that chemical reactions occur when reactant particles collide with one another with sufficient energy and in the correct orientation.

According to collision theory, not every collision between reactant particles results in a chemical reaction.

Only collisions that have sufficient energy and suitable orientation lead to the formation of products. These are called effective collisions.

6.2 Main principles of collision theory

Collision theory is based on the following principles:

1. Reactant particles are in continuous motion.

Atoms, ions and molecules move continuously. Their movement causes them to collide with one another.

2. Chemical reactions occur through collisions.

Reactant particles must come into contact before chemical bonds can break and new bonds can form.

3. Not all collisions are effective.

Some particles collide but separate without reacting because they lack sufficient energy or have unsuitable orientation.

4. Particles must possess sufficient energy.

The colliding particles must have energy equal to or greater than the activation energy for the reaction to occur.

5. Particles must have suitable orientation.

Reactant particles must approach one another in a way that allows the necessary bonds to break and new bonds to form.

6. Effective collisions produce products.

When particles collide with sufficient energy and correct orientation, a reaction can take place.

6.3 Effective and ineffective collisions

Effective collision

An effective collision is a collision between reactant particles that has sufficient energy and the correct orientation to result in a chemical reaction.

Ineffective collision

An ineffective collision is a collision that does not produce a chemical reaction because the particles lack sufficient energy or have unsuitable orientation.

6.4 Factors determining whether a collision is effective

1. Energy of the colliding particles.
2. Orientation of the colliding particles.
3. Nature of the reacting substances.
4. Temperature of the reacting system.
5. Presence or absence of a catalyst.

6.5 Relationship between collision frequency and reaction rate

Collision frequency refers to the number of collisions between reactant particles per unit time.

Generally:

- More frequent collisions increase the chances of effective collisions.
- Fewer collisions reduce the chances of effective collisions.
- A greater number of effective collisions per unit time produces a faster reaction.

Example: When the concentration of hydrochloric acid is increased, there are more acid particles in a given volume. This increases the frequency of collisions with the surface of magnesium, resulting in a faster reaction.

7.0 ACTIVATION ENERGY

7.1 Definition of activation energy

Activation energy is the minimum energy that reacting particles must possess during a collision for a chemical reaction to occur.

It is represented by the symbol:

$$E_a$$

Its common unit is kilojoules per mole (kJ mol^{-1}).

7.2 Explanation of activation energy

For a chemical reaction to occur:

1. Reactant particles must collide.
2. They must possess sufficient kinetic energy.
3. Existing bonds may need to be weakened or broken.
4. An unstable arrangement of atoms called the activated complex or transition state may form.
5. New bonds form to produce the products.

The minimum energy needed to reach the activated complex is the activation energy.

7.3 Meaning of activated complex

An activated complex is a temporary, unstable arrangement of atoms formed during the conversion of reactants into products.

It exists at the highest-energy point along the reaction pathway.

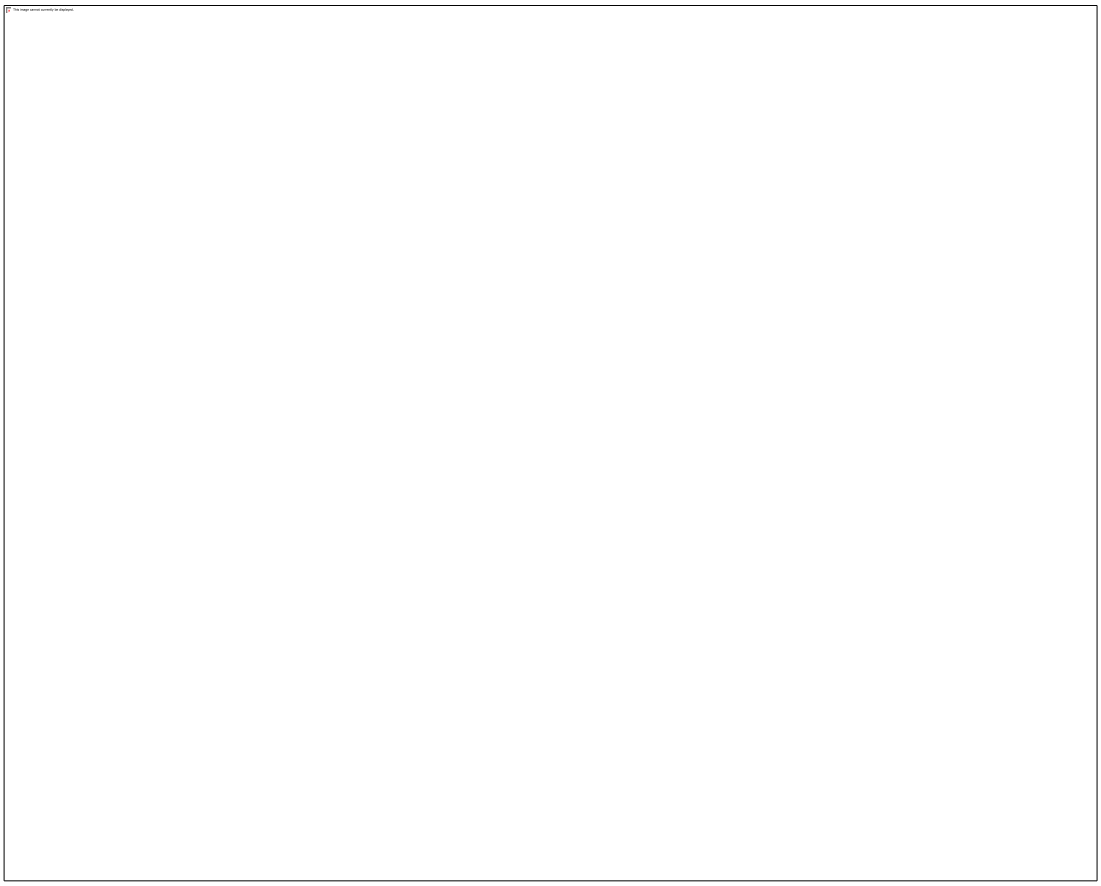
7.4 Effect of activation energy on reaction rate

- A reaction with a high activation energy generally proceeds slowly under the same conditions.
- A reaction with a low activation energy generally proceeds faster under the same conditions.
- A catalyst provides an alternative reaction pathway with lower activation energy.
- Lower activation energy allows a greater proportion of particles to react successfully at a given temperature.

7.5 Energy profile diagram

An energy profile diagram shows how the energy of a reacting system changes as a reaction proceeds.

Energy profile of an exothermic reaction



The products have lower energy than the reactants. The reaction releases energy.

Explanation:

- The vertical axis represents energy.
- The horizontal axis represents the progress of the reaction.
- The reactants begin at a higher energy level than the products.
- The peak represents the activated complex.
- The energy difference between the reactants and the peak is the activation energy.
- The energy difference between the reactants and products is the energy change of the reaction.

7.6 Energy profile of an endothermic reaction

In an endothermic reaction, the products have higher energy than the reactants.

Energy profile of an endothermic reaction



The products have higher energy than the reactants. The reaction absorbs energy.

7.7 Difference between exothermic and endothermic reactions

Exothermic reaction	Endothermic reaction
Releases energy to the surroundings.	Absorbs energy from the surroundings.
Products have lower energy than reactants.	Products have higher energy than reactants.
Enthalpy change is negative.	Enthalpy change is positive.
Surroundings may become warmer.	Surroundings may become cooler.
Example: combustion of methane.	Example: thermal decomposition of calcium carbonate.

Key equations:

For an exothermic reaction:

$$\Delta H < 0$$

For an endothermic reaction:

$$\Delta H > 0$$

8.0 FACTORS AFFECTING THE RATE OF CHEMICAL REACTIONS

The **rate of a chemical reaction** refers to how fast reactants are converted into products. It is influenced by several factors that affect the frequency of collisions between reacting particles and the number of collisions that result in a chemical reaction.

The main factors affecting the rate of a chemical reaction are:

1. Concentration of reactants.
2. Temperature.
3. Surface area of solid reactants.
4. Pressure of gaseous reactants.
5. Presence of a catalyst.
6. Light intensity in photochemical reactions.
7. Nature of the reacting substances.

8.1 Effect of concentration on the rate of reaction

Definition of concentration

Concentration is the amount of solute dissolved in a given volume of solution.

It is commonly expressed in moles per cubic decimetre (mol dm^{-3}).

$$\text{Concentration} = \frac{\text{Moles of solute}}{\text{Volume of solution in dm}^3}$$

How concentration affects the rate of reaction

Increasing the concentration of a reactant generally increases the rate of reaction.

This is because:

8. A concentrated solution contains more reactant particles per unit volume.
9. The particles are closer together.
10. The particles collide more frequently.
11. The number of effective collisions per unit time increases.
12. The reaction takes place faster.

When the concentration decreases:

- There are fewer reactant particles per unit volume.
- Collisions occur less frequently.
- The number of effective collisions decreases.
- The reaction proceeds more slowly.

Example: Reaction between magnesium and hydrochloric acid

Magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen gas.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

When magnesium reacts with more concentrated hydrochloric acid:

- Hydrogen gas is produced at a faster rate.
- Magnesium dissolves more quickly.
- The reaction takes a shorter time to finish.

When dilute hydrochloric acid is used:

- Hydrogen gas is produced more slowly.
- Magnesium dissolves at a slower rate.
- The reaction takes a longer time to finish.

Experimental investigation: Effect of concentration on reaction rate

Aim: To investigate how the concentration of an acid affects the rate of reaction.

Apparatus and materials:

- Magnesium ribbon.
- Hydrochloric acid of different concentrations.
- Conical flask.
- Measuring cylinder.
- Stopwatch.
- Gas syringe.
- Delivery tube.
- Retort stand.

Procedure:

1. Measure a fixed volume of dilute hydrochloric acid and pour it into a conical flask.
2. Add a measured length or mass of magnesium ribbon to the acid.
3. Immediately fit a stopper connected to a gas syringe.
4. Start the stopwatch.
5. Record the volume of hydrogen gas produced at regular time intervals.
6. Repeat the experiment using hydrochloric acid of different concentrations.
7. Keep the temperature, mass of magnesium, surface area of magnesium and volume of acid constant where appropriate.
8. Compare the volume of hydrogen gas produced per unit time.

Expected observations:

The more concentrated hydrochloric acid produces hydrogen gas at a faster rate than the less concentrated acid.

Conclusion:

Increasing the concentration of a reactant increases the rate of reaction when other conditions are kept constant.

Illustration: A concentrated solution contains more reactant particles in the same volume than a dilute solution.

8.2 Effect of temperature on the rate of reaction

Definition of temperature

Temperature is a measure of the degree of hotness or coldness of a substance.

Increasing the temperature of a reacting mixture generally increases the rate of a chemical reaction.

Explanation using collision theory

When the temperature of a reacting mixture is increased:

1. The reactant particles gain kinetic energy.

2. The particles move faster.
3. The particles collide more frequently.
4. A greater proportion of particles acquire energy equal to or greater than the activation energy.
5. The number of effective collisions per unit time increases.
6. The reaction proceeds faster.

When the temperature decreases:

- The particles have less average kinetic energy.
- Fewer particles have sufficient energy to react.
- The number of effective collisions decreases.
- The reaction proceeds more slowly.

Important point: An increase in temperature does not only increase the frequency of collisions. It also increases the proportion of particles with enough energy to overcome the activation energy barrier. This is a major reason why reaction rates increase with temperature.

Experimental investigation: Effect of temperature on reaction rate

Aim: To investigate the effect of temperature on the rate of reaction.

Materials and apparatus:

- Sodium thiosulphate solution.
- Dilute hydrochloric acid.
- Conical flask.
- Measuring cylinder.
- Thermometer.
- Water bath.
- Stopwatch.
- White paper marked with a black cross.

Procedure:

7. Measure a fixed volume of sodium thiosulphate solution into a conical flask.
8. Place the flask in a water bath set at a selected temperature.
9. Allow the solution to reach the required temperature.
10. Place the flask over a white sheet of paper marked with a black cross.
11. Add a fixed volume of dilute hydrochloric acid to the flask.
12. Start the stopwatch immediately.
13. Observe the cross through the reaction mixture.
14. Record the time taken for the cross to disappear as sulphur forms and makes the mixture cloudy.
15. Repeat the experiment at different temperatures while keeping the volumes and concentrations of the reactants constant.

Chemical equation:

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

Observations:

- At higher temperatures, the cross disappears more quickly.
- At lower temperatures, the cross takes longer to disappear.

Conclusion:

Increasing the temperature increases the rate of reaction.

Safety precaution: Sulphur dioxide is an irritating and harmful gas. The experiment should be carried out in a well-ventilated laboratory, preferably in a fume cupboard, using small quantities of chemicals.

8.3 Effect of surface area on the rate of reaction

Definition of surface area

Surface area is the total exposed area of a solid substance that is available for contact with other reacting substances.

The surface area of a solid reactant affects the rate at which it reacts.

Explanation

When a solid reactant is broken into smaller pieces or converted into powder:

1. Its total exposed surface area increases.
2. More particles of the solid become available for reaction.
3. More collisions take place between the solid and the other reactant.
4. The number of effective collisions per unit time increases.
5. The reaction proceeds faster.

A large lump of a solid has a smaller exposed surface area than the same mass of the solid in powdered form.

Example: Reaction of calcium carbonate with hydrochloric acid

Calcium carbonate reacts with hydrochloric acid to produce calcium chloride, carbon dioxide and water.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

When powdered calcium carbonate reacts with hydrochloric acid:

- The reaction takes place faster.
- Carbon dioxide is produced more rapidly.

When large marble chips react with the same acid:

- The reaction takes place more slowly.
- Carbon dioxide is produced at a lower rate.

Illustration: Powdered calcium carbonate exposes a larger total surface area to the reacting solution than an equal mass of large marble chips.

Experimental investigation: Effect of surface area on reaction rate

Aim: To compare the rates of reaction of calcium carbonate in large-chip and powdered forms.

Procedure:

1. Measure equal masses of large marble chips and powdered calcium carbonate.
2. Put equal volumes of hydrochloric acid of the same concentration into two separate conical flasks.
3. Add the marble chips to the first flask.
4. Add the powdered calcium carbonate to the second flask.
5. Collect the carbon dioxide produced using gas syringes.
6. Record the volume of gas produced at regular time intervals.
7. Compare the rates of gas production in the two experiments.

Expected observation:

Powdered calcium carbonate produces carbon dioxide faster than large marble chips.

Conclusion:

Increasing the surface area of a solid reactant increases the rate of reaction, provided the other conditions remain constant.

8.4 Effect of pressure on the rate of reaction

Pressure mainly affects reactions involving gaseous reactants.

8.4 Effect of pressure on the rate of reaction

Pressure mainly affects reactions involving gaseous reactants.

Increasing the pressure of reacting gases at constant temperature generally increases the rate of reaction.

Explanation

1. Increasing pressure compresses the gas into a smaller volume.
2. The number of gas particles per unit volume increases.
3. The gas particles become closer together.
4. The frequency of collisions increases.
5. The number of effective collisions per unit time increases.
6. The reaction proceeds faster.

Example: Formation of ammonia

Nitrogen reacts with hydrogen to form ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Increasing the pressure of nitrogen and hydrogen increases the frequency of collisions between the reacting gas particles. This generally increases the rate at which ammonia is formed.

Applications of pressure in reaction rates

1. Industrial manufacture of ammonia.
2. Industrial gas-phase chemical reactions.
3. Manufacture of methanol from synthesis gas.
4. Certain combustion processes involving gaseous fuels.
5. Industrial catalytic reactions involving gases.

Important: Pressure has a significant effect on reactions involving gases. It does not generally have the same direct effect on reactions involving only solids and liquids because these substances are much less compressible.

8.5 Effect of catalysts on the rate of reaction

Definition of a catalyst

A catalyst is a substance that changes the rate of a chemical reaction without being permanently consumed in the overall chemical reaction.

Most catalysts used in industrial processes increase the rate of reaction.

How catalysts work

1. A catalyst provides an alternative reaction pathway.
2. The alternative pathway has a lower activation energy.

3. A greater proportion of reacting particles can overcome the energy barrier.
4. The number of effective collisions increases.
5. The reaction proceeds faster.

A catalyst does not:

- Change the overall chemical equation.
- Change the overall enthalpy change of the reaction.
- Alter the equilibrium constant at a given temperature.
- Change the equilibrium position of a reversible reaction.
- Get permanently used up in the overall reaction.

Examples of catalysts

Catalyst	Reaction or application
Iron	Manufacture of ammonia in the Haber process
Vanadium(V) oxide, V_2O_5	Manufacture of sulphur trioxide in the Contact process
Platinum-rhodium gauze	Oxidation of ammonia in the Ostwald process
Manganese(IV) oxide, MnO_2	Decomposition of hydrogen peroxide
Nickel	Hydrogenation of vegetable oils

Example: Decomposition of hydrogen peroxide

Hydrogen peroxide decomposes to form water and oxygen.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

Manganese(IV) oxide can be used as a catalyst. It increases the rate at which oxygen is produced without being permanently consumed in the overall reaction.

Experimental investigation: Effect of a catalyst

Aim: To investigate the effect of manganese(IV) oxide on the rate of decomposition of hydrogen peroxide.

Materials and apparatus:

- Hydrogen peroxide solution.
- Manganese(IV) oxide.
- Conical flask.
- Delivery tube.
- Gas syringe.
- Stopwatch.

Procedure:

1. Place a measured volume of hydrogen peroxide solution in a conical flask.
2. Connect the flask to a gas syringe.
3. Record the volume of oxygen produced at regular time intervals.
4. Repeat the experiment using the same volume and concentration of hydrogen peroxide.
5. Add a small measured quantity of manganese(IV) oxide.
6. Record the volume of oxygen produced at the same time intervals.
7. Compare the rates of oxygen production.

Observation:

The mixture containing manganese(IV) oxide produces oxygen more rapidly.

Conclusion:

Manganese(IV) oxide increases the rate of decomposition of hydrogen peroxide by acting as a catalyst.

8.6 Effect of light on the rate of reaction

Light can increase the rate of certain chemical reactions. Such reactions are known as **photochemical reactions**.

Definition of a photochemical reaction

A photochemical reaction is a chemical reaction initiated or influenced by light energy.

Explanation

1. Reactant molecules absorb light energy.
2. The absorbed energy may excite the molecules.
3. The excited molecules become more reactive.
4. The number of particles able to undergo reaction increases.
5. The reaction rate may increase.

Examples of photochemical reactions

A. Decomposition of silver chloride

Silver chloride decomposes in the presence of light.

$$2AgCl(s) \xrightarrow{\text{light}} 2Ag(s) + Cl_2(g)$$

Observation: White silver chloride gradually turns grey because metallic silver is formed.

B. Decomposition of silver bromide

$$2AgBr(s) \xrightarrow{\text{light}} 2Ag(s) + Br_2(g)$$

Silver bromide is sensitive to light and has been used in traditional photographic materials.

C. Photosynthesis

Plants use light energy to drive the chemical reactions involved in photosynthesis.

A simplified overall equation is:

$$6CO_2 + 6H_2O \xrightarrow{\text{light}} C_6H_{12}O_6 + 6O_2$$

Applications of light in chemical reactions

1. Photography.
2. Solar-powered chemical processes.
3. Photochemical manufacture of certain chemicals.
4. Photosynthesis in plants.
5. Light-activated polymerisation and curing processes.

8.7 Effect of the nature of reactants

Different substances react at different rates because of differences in their chemical properties and bonding.

Factors that influence the rate include:

- Type of chemical bonds.
- Strength of chemical bonds.
- Reactivity of the elements or compounds.
- Physical state of the reactants.
- Whether the reaction involves ions, molecules or atoms.

Examples

1. Ionic reactions in aqueous solutions are often rapid because oppositely charged ions can react when they meet.
2. Some reactions involving strong covalent bonds are slower because considerable energy is needed to break or reorganise the bonds.
3. Sodium reacts rapidly with water, while copper does not react readily with cold water.

8.8 Summary of factors affecting reaction rates

Factor	Effect on reaction rate	Explanation using collision theory
Increased concentration	Usually increases the rate	More particles are present per unit volume, resulting in more frequent collisions.
Increased temperature	Increases the rate	More particles have energy equal to or greater than the activation energy.
Increased surface area	Increases the rate	More exposed particles are available for collisions.
Increased gas pressure	Usually increases the rate	Gas particles are closer together and collide more frequently.
Addition of a catalyst	Increases the rate of a catalysed reaction	An alternative reaction pathway has a lower activation energy.
Increased suitable light intensity	May increase the rate	More reactant molecules absorb photons and become activated.
Nature of reactants	The rate varies	Differences in bonding and chemical reactivity affect the reaction pathway.

9.0 MEASURING THE RATE OF A CHEMICAL REACTION

9.1 Meaning of measuring reaction rate

Measuring the rate of a chemical reaction involves determining how quickly reactants are used up or products are formed during the reaction.

The rate can be measured by monitoring changes in:

- Concentration.
- Mass.
- Volume of gas.
- Amount of reactant.
- Amount of product.
- Other suitable physical properties.

9.2 General expression for reaction rate

The average rate of reaction can be expressed as:

$$\text{Average rate} = \frac{\text{Change in quantity}}{\text{Time taken}}$$

When concentration is measured:

$$\boxed{\text{Rate} = \frac{\text{Change in concentration}}{\text{Time}}}$$

For a reactant being used up:

$$\text{Rate} = -\frac{\Delta[\text{Reactant}]}{\Delta t}$$

For a product being formed:

$$\text{Rate} = \frac{\Delta[\text{Product}]}{\Delta t}$$

The negative sign for a reactant indicates that its concentration decreases with time.

A common unit of reaction rate is:

$$mol\ dm^{-3}\ s^{-1}$$

Other units may be used depending on the quantity being measured.

9.3 Methods of determining reaction rates

A. Measuring the volume of gas produced

This method is suitable for reactions that produce a gas.

Example:

Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Procedure:

1. Place hydrochloric acid in a conical flask.
2. Add a measured mass of magnesium.
3. Connect the flask to a gas syringe.
4. Collect the hydrogen gas produced.
5. Record the volume of gas at regular time intervals.
6. Calculate the rate of gas production.

$$\text{Average rate of gas production} = \frac{\text{Volume of gas produced}}{\text{Time}}$$

The units may be $cm^3\ s^{-1}$ or $dm^3\ s^{-1}$.

B. Measuring loss in mass

This method is suitable when a gaseous product escapes from the reaction vessel.

Example:

Calcium carbonate reacts with hydrochloric acid to produce calcium chloride, carbon dioxide and water.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Procedure:

1. Place a conical flask containing hydrochloric acid on a balance.
2. Add calcium carbonate to the flask.
3. Record the initial mass.
4. Record the mass at regular time intervals.
5. Carbon dioxide escapes from the flask.
6. The mass of the apparatus and its contents decreases.
7. Calculate the loss in mass per unit time.

$$\text{Rate of mass loss} = \frac{\text{Loss in mass}}{\text{Time}}$$

C. Measuring changes in concentration

The concentration of a reactant or product can be measured at different times using suitable analytical methods.

Examples of methods include:

- Titration.
- Colorimetry.
- Conductivity measurements.
- pH measurements, where appropriate.

10.0 GRAPHS OF REACTION RATES

10.1 Introduction

Chemical reactions occur at different rates. Some reactions take place very quickly, while others take a long time to complete.

A reaction-rate graph is a graphical representation showing how the amount or concentration of a reactant or product changes with time.

Reaction-rate graphs are important because they help learners to:

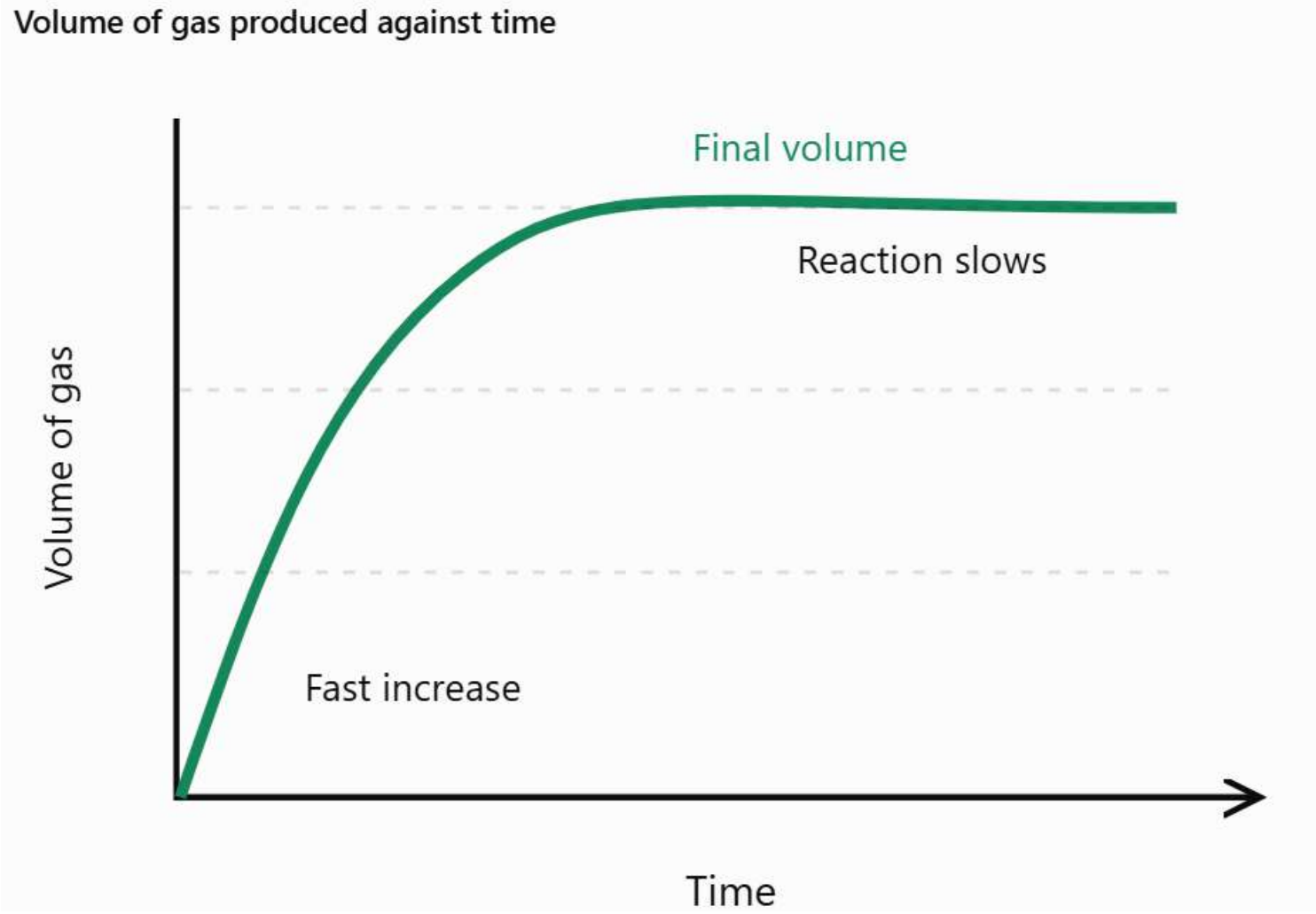
1. Understand how reaction rates change as reactions proceed.
2. Compare the rates of different chemical reactions.
3. Determine the rate of a reaction at a particular time.
4. Identify when a reaction has stopped producing additional products.
5. Investigate the effects of changing reaction conditions.
6. Interpret experimental results accurately.
7. Understand the relationship between reactant concentration and reaction rate.

The two common types of reaction-rate graphs are:

1. Graphs of product formed against time.
2. Graphs of reactant concentration against time.

10.2 Graph of Product Formed Against Time

When a chemical reaction takes place, the amount of product formed generally increases with time.



For example, magnesium reacts with dilute hydrochloric acid to produce magnesium chloride and hydrogen gas.

Chemical equation:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The volume of hydrogen gas produced can be measured at regular time intervals using a suitable gas collection apparatus.

Changes observed during the reaction

1. At the beginning of the reaction, the concentration of reactants is relatively high.
2. The frequency of effective collisions is initially high.
3. Hydrogen gas is produced rapidly.
4. As the reaction proceeds, reactants are gradually used up.
5. The frequency of effective collisions decreases.
6. The rate of product formation gradually decreases.
7. Eventually, the graph becomes horizontal when no further gas is being produced.

Illustration: Volume of gas produced against time

Explanation of the graph

- The curve rises steeply at the beginning because the reaction is initially fast.
- The curve gradually becomes less steep as the reaction slows down.
- The horizontal section indicates that the volume of gas is no longer increasing.
- The final volume represents the total amount of gas produced under the experimental conditions.

Important: A horizontal graph indicates that the measured quantity is no longer changing. In an irreversible reaction, this commonly means that the reaction has stopped. In a reversible reaction, it may indicate that dynamic equilibrium has been established.

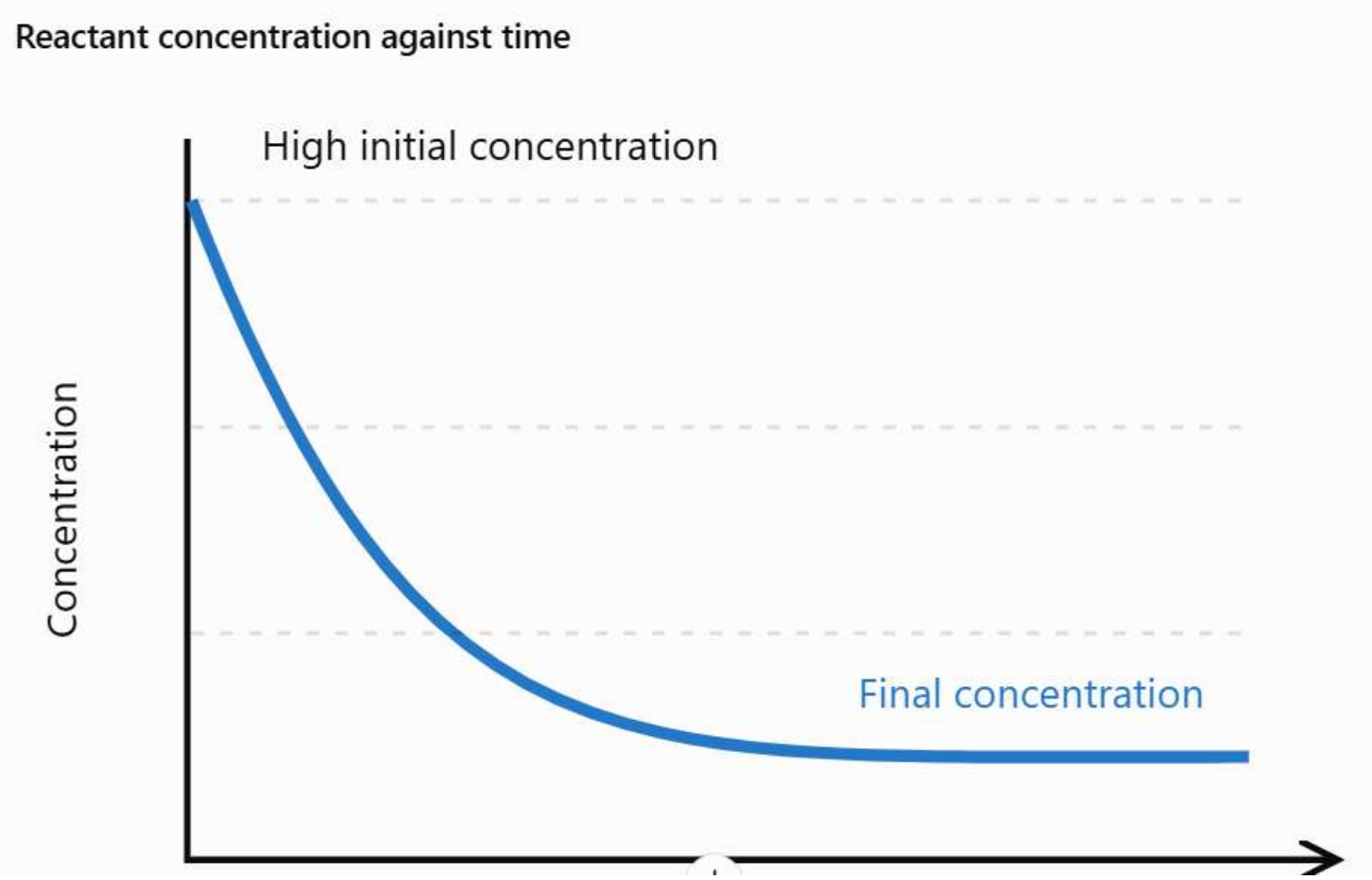
10.3 Graph of Reactant Concentration Against Time

During a chemical reaction, reactants are consumed to form products. Therefore, the concentration of a reactant generally decreases as the reaction proceeds.

Changes observed

1. At the beginning, the concentration of the reactant is relatively high.
2. The reaction rate is usually high because many reactant particles are present per unit volume.
3. As the reaction proceeds, reactant particles are gradually used up.
4. The concentration of the reactant decreases.
5. The frequency of effective collisions generally decreases.
6. The rate of decrease in concentration becomes slower.
7. The graph eventually becomes horizontal when the concentration stops changing.

Illustration: Reactant concentration against time



Explanation of the graph

- The curve falls steeply at the beginning because the reactant is being consumed rapidly.
- The curve becomes less steep as the reaction slows down.
- The horizontal section indicates that the concentration is no longer changing.
- In a reversible reaction, the horizontal section may indicate that dynamic equilibrium has been established.

10.4 Interpreting the Gradient of a Reaction-Rate Graph

The gradient of a graph is the change in the vertical-axis quantity divided by the change in the horizontal-axis quantity.

For a graph of volume of gas against time:

$$\text{Gradient} = \frac{\text{Change in gas volume}}{\text{Change in time}}$$

Therefore:

$$\boxed{\text{Rate} = \frac{\Delta V}{\Delta t}}$$

Where:

- ΔV = change in the volume of gas.
- Δt = time taken.

The gradient indicates the rate of change of the measured quantity.

Interpretation of gradients

1. **Steep gradient:** Indicates a high rate of reaction or a rapid change in the measured quantity.
2. **Gentle gradient:** Indicates a low rate of reaction or a slow change in the measured quantity.
3. **Horizontal gradient:** Indicates that the measured quantity is not changing.
4. **Steeper initial gradient:** Indicates a higher initial reaction rate.
5. **Changing gradient:** Indicates that the reaction rate is changing as the reaction proceeds.

Worked Example 1

A chemical reaction produces 60 cm^3 of gas in 20 seconds. Calculate the average rate of gas production.

Step 1: Identify the given information.

- Volume of gas produced = 60 cm^3.
- Time taken = 20 seconds.

Step 2: Write the formula.

$$\text{Average rate} = \frac{\text{Change in volume}}{\text{Time taken}}$$

Step 3: Substitute the values.

$$\text{Average rate} = \frac{60}{20}$$

Step 4: Calculate.

Answer: The average rate of gas production is 3 cm^3 per second.

Worked Example 2

A reaction produces 100 cm^3 of gas in 25 seconds. Calculate the average rate of gas production.

Step 1: Write the formula.

$$\text{Average rate} = \frac{\text{Change in volume}}{\text{Time}}$$

Step 2: Substitute the values.

$$\text{Average rate} = \frac{100}{25}$$

Step 3: Calculate.

Answer: The average rate of gas production is 4 cm^3 per second.

10.5 Comparing Reaction Rates Using Graphs

Reaction-rate graphs can be used to compare the rates of different chemical reactions.

When comparing two reaction-rate graphs:

1. The reaction with the steeper initial gradient has a higher initial reaction rate.
2. The reaction with the gentler initial gradient has a lower initial reaction rate.
3. The reaction that reaches its final value sooner completes the measured reaction sooner.
4. Two reactions may produce the same final amount of product but at different rates.
5. A catalyst may make a reaction reach its final amount sooner without changing the theoretical final amount of product.

Worked Example

Two experiments involve the reaction of equal amounts of magnesium with hydrochloric acid.

- Experiment A uses dilute hydrochloric acid.
- Experiment B uses more concentrated hydrochloric acid.

Assuming all other conditions are kept constant, Experiment B generally has a steeper initial gradient.

Explanation:

The more concentrated hydrochloric acid contains more acid particles per unit volume. This increases the frequency of collisions between reacting particles and the frequency of effective collisions. Consequently, the reaction proceeds faster.

10.6 Importance of Reaction-Rate Graphs

Reaction-rate graphs are important because they help learners to:

1. Determine the rate of a chemical reaction.
2. Compare the rates of different reactions.
3. Understand how reaction rates change with time.
4. Identify when product formation stops.
5. Investigate the effects of concentration and temperature.
6. Explain the effectiveness of catalysts.
7. Interpret experimental results scientifically.
8. Determine average reaction rates from experimental data.
9. Estimate instantaneous reaction rates using the gradient of a tangent.
10. Understand the relationship between reactant consumption and product formation.

11.0 CATALYSTS AND THEIR EFFECT ON ENERGY PROFILE DIAGRAMS

11.1 Meaning of Catalysis

Catalysis is the process by which the rate of a chemical reaction is changed by the presence of a catalyst.

A **catalyst** is a substance that changes the rate of a chemical reaction without being permanently consumed in the overall reaction.

Most catalysts studied at this level increase the rate of reaction.

Catalysts are widely used in laboratories, industries and living organisms.

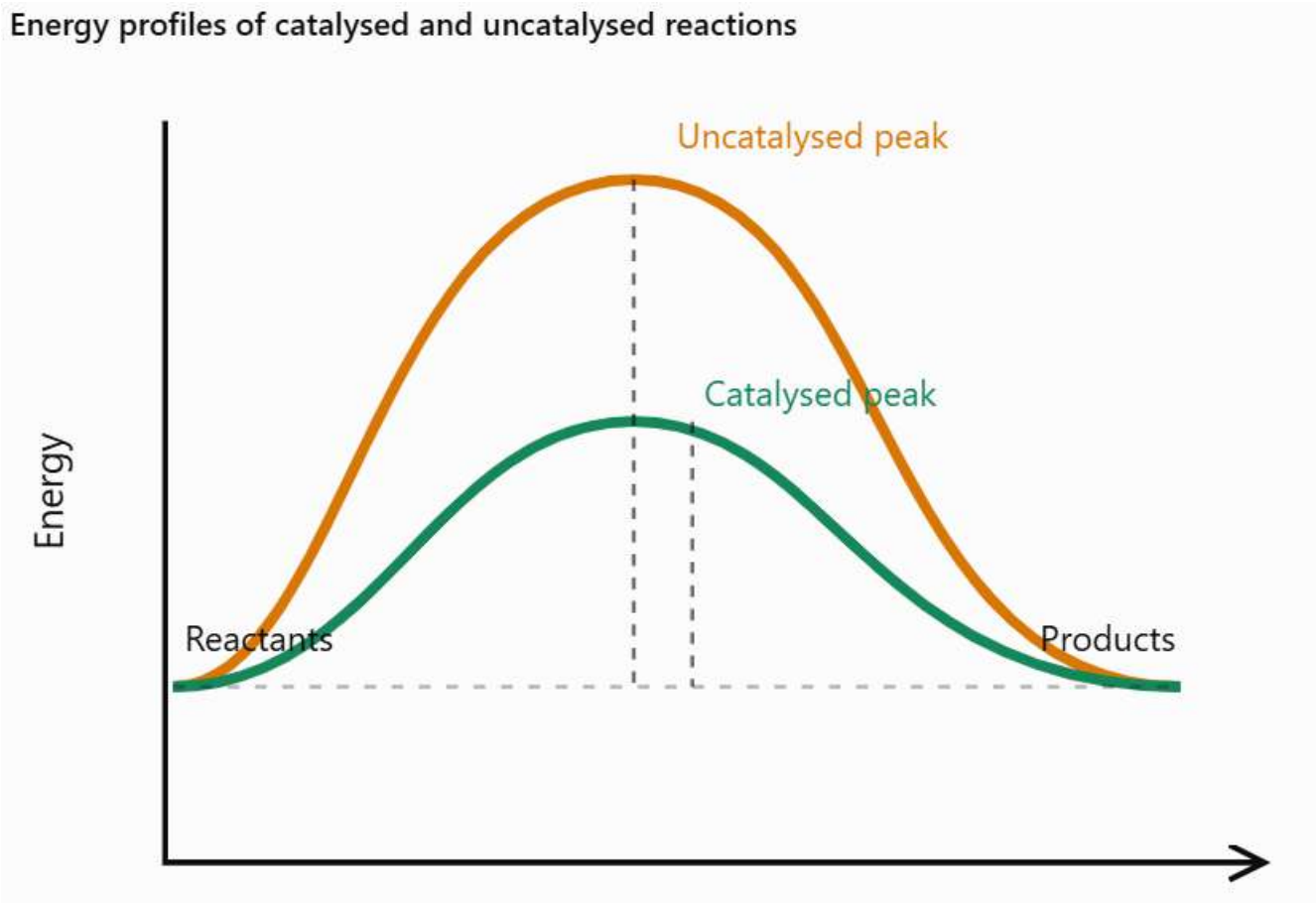
11.2 How a Catalyst Affects Activation Energy

A catalyst increases the rate of a reaction by providing an alternative reaction pathway that requires less activation energy.

A catalyst:

1. Provides an alternative reaction pathway.
2. Lowers the activation energy required for that pathway.
3. Allows a greater proportion of reacting particles to possess sufficient energy to react.
4. Increases the number of successful reactions per unit time.
5. Increases the reaction rate.
6. Is regenerated during the overall reaction and is not permanently consumed.

Important: A catalyst does not increase the kinetic energy of reacting particles. It provides a reaction pathway with a lower activation energy.



11.3 Meaning of Activation Energy

Activation energy is the minimum energy that reacting particles must possess for a successful reaction to occur through a particular reaction pathway.

It is represented by the symbol:

$$E_a$$

A chemical reaction can occur when reacting particles collide with:

- Sufficient energy.
- Appropriate orientation.

Particles that collide with energy below the activation energy generally do not react.

A catalyst increases the reaction rate by providing an alternative pathway with a lower activation energy.

11.4 Energy Profile Diagrams With and Without a Catalyst

An energy profile diagram shows the changes in energy as a chemical reaction proceeds.

The diagram has two main axes:

- **Vertical axis:** Energy.
- **Horizontal axis:** Progress of reaction or reaction pathway.

An energy profile diagram can be used to compare catalysed and uncatalysed reactions.

Illustration: Energy profiles of catalysed and uncatalysed reactions

Explanation of the diagram

1. The uncatalysed pathway has a higher energy peak.
2. The catalysed pathway has a lower energy peak.
3. The difference between the energy level of the reactants and the peak represents activation energy.
4. The catalyst provides an alternative pathway with a lower activation energy.
5. The energy levels of the reactants and products remain unchanged.
6. The overall energy change of the reaction remains unchanged.

Important: A catalyst does not change the overall energy change of a reaction. It only changes the reaction pathway and lowers the activation energy.

11.5 Types of Catalysts

Catalysts can be classified according to their physical state in relation to the reactants and according to their nature.

A. Homogeneous Catalysts

A homogeneous catalyst is a catalyst that exists in the same physical phase as the reactants.

For example:

- An acid catalyst in a reaction involving aqueous reactants.
- Sulphuric acid used as a catalyst in certain reactions involving liquid reactants.

Characteristics of homogeneous catalysts

1. The catalyst and reactants exist in the same physical phase.

2. They mix uniformly in the reaction mixture.
3. The catalyst participates in the reaction mechanism.
4. The catalyst is regenerated during the overall reaction.
5. The catalyst may form temporary intermediate compounds during the reaction.

B. Heterogeneous Catalysts

A heterogeneous catalyst is a catalyst that exists in a different physical phase from the reactants.

Examples:

1. Solid iron used in the Haber process with gaseous nitrogen and hydrogen.
2. Solid vanadium(V) oxide used in the Contact process with gaseous reactants.
3. Solid nickel used in the hydrogenation of vegetable oils.

Characteristics of heterogeneous catalysts

1. The catalyst and reactants exist in different physical phases.
2. The reaction often takes place on the surface of the solid catalyst.
3. Reactant particles become attached to the catalyst surface through adsorption.
4. Chemical bonds in the reactants may be weakened, making reaction easier.
5. Products are formed and released from the catalyst surface.
6. The catalyst can be used repeatedly if it remains active.

C. Enzyme Catalysts

Enzymes are biological catalysts produced by living organisms.

They increase the rate of biochemical reactions in living cells.

Examples of enzyme catalysts

Enzyme	Function
Amylase	Helps break down starch into simpler sugars.
Catalase	Speeds up the decomposition of hydrogen peroxide into water and oxygen.
Proteases	Help break down proteins into smaller peptides and amino acids.
Lipase	Helps break down fats into fatty acids and glycer-ol.

Importance of enzymes

1. They control metabolic reactions.
2. They help in the digestion of food.
3. They support the functioning of living organisms.
4. They enable biochemical reactions to occur rapidly under suitable biological conditions.
5. They are important in food processing and some industrial processes.

11.6 Importance of Catalysts

Catalysts are important in everyday life and industrial processes because they:

1. Increase the rate of chemical reactions.
2. Lower the activation energy required for a reaction pathway.
3. Improve industrial production efficiency.
4. Reduce energy consumption in some industrial processes.

5. Help control chemical reactions in living organisms.
6. Enable reactions to occur under more suitable conditions.
7. Are important in pollution-control systems.
8. Help in the manufacture of fuels, fertilisers and other chemicals.
9. Reduce the time required for some industrial processes.
10. Support the development of more efficient chemical manufacturing methods.

11.7 Industrial Applications of Catalysts

Industrial process	Catalyst	Product or purpose
Haber process	Iron	Manufacture of ammonia.
Contact process	Vanadium(V) oxide	Oxidation of sulphur dioxide to sulphur trioxide.
Hydrogenation of vegetable oils	Nickel	Conversion of unsaturated oils into more saturated products.
Catalytic converters	Platinum, palladium and rhodium	Reduction of harmful vehicle exhaust emissions.
Ostwald process	Platinum-rhodium gauze	Oxidation of ammonia during nitric acid manufacture.

Explanation of selected applications

1. Haber process

Iron is used as a catalyst to increase the rate of reaction between nitrogen and hydrogen during ammonia manufacture.

2. Contact process

Vanadium(V) oxide increases the rate of oxidation of sulphur dioxide to sulphur trioxide.

3. Hydrogenation of vegetable oils

Nickel is used to speed up the addition of hydrogen to unsaturated vegetable oils.

4. Catalytic converters

Platinum, palladium and rhodium are used in vehicle exhaust systems to help convert harmful gases into less harmful substances.

5. Ostwald process

Platinum-rhodium gauze is used to catalyse the oxidation of ammonia during nitric acid manufacture.

11.8 Effect of Catalysts on Chemical Equilibrium

A catalyst does not increase the equilibrium yield of a reversible reaction at a fixed temperature.

Instead, it:

11. Increases the rate of the forward reaction.
12. Increases the rate of the reverse reaction.
13. Helps the system reach equilibrium more quickly.
14. Does not change the equilibrium position.
15. Does not change the equilibrium constant at a fixed temperature.

Key point: A catalyst affects how quickly equilibrium is reached, not the final equilibrium composition.

PART B: REVERSIBLE REACTIONS AND CHEMICAL EQUILIBRIUM

12.0 REVERSIBLE AND IRREVERSIBLE REACTIONS

Chemical reactions can be classified according to whether the products can react to reform the original reactants.

The two main categories are:

1. Reversible reactions.
2. Irreversible reactions.

12.1 Reversible Reactions

Definition

A reversible reaction is a chemical reaction in which the products can react under suitable conditions to form the original reactants.

In a reversible reaction:

1. Reactants form products in the forward reaction.
2. Products can react to form the original reactants in the reverse reaction.
3. Both reactions can occur in the same reaction system.
4. The forward and reverse reactions may eventually occur at equal rates in a closed system.
5. Dynamic equilibrium may be established when the conditions are suitable.

A reversible reaction is represented by a double arrow:

$$\rightleftharpoons$$

General representation

$$A + B \rightleftharpoons C + D$$

Where:

- A and B are reactants.
- C and D are products.
- The forward reaction converts A and B into C and D.
- The reverse reaction converts C and D into A and B.

Examples of reversible reactions

1. Formation of ammonia

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Nitrogen reacts with hydrogen to form ammonia. Ammonia can also decompose to form nitrogen and hydrogen under suitable conditions.

2. Formation of hydrogen iodide

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

Hydrogen reacts with iodine to form hydrogen iodide. Hydrogen iodide can also decompose to form hydrogen and iodine.

3. Dehydration and hydration of copper(II) sulphate

$$CuSO_4 \cdot 5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(g)$$

Hydrated copper(II) sulphate loses water of crystallisation when heated to form anhydrous copper(II) sulphate. The anhydrous salt can regain water to form the hydrated salt.

4. Sublimation and deposition of iodine

$$I_2(s) \rightleftharpoons I_2(g)$$

Solid iodine changes directly into iodine vapour when heated. On cooling, the vapour changes back into solid iodine.

This is a reversible physical change involving sublimation and deposition, rather than a chemical reaction.

12.2 Irreversible Reactions

Definition

An irreversible reaction is a chemical reaction that proceeds predominantly in one direction under the conditions of the reaction, with the products not readily reforming the original reactants.

It is represented using a single forward arrow:



Examples of irreversible reactions

1. Combustion of methane

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Methane burns in oxygen to form carbon dioxide and water.

2. Reaction between silver nitrate and sodium chloride

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

A white precipitate of silver chloride is formed.

3. Neutralisation between hydrochloric acid and sodium hydroxide

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

An acid reacts with a base to form salt and water.

4. Burning of magnesium

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Magnesium burns in oxygen to form magnesium oxide.

5. Decomposition of potassium chlorate on heating

$$2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$$

Potassium chlorate decomposes on heating to form potassium chloride and oxygen.

12.3 Differences Between Reversible and Irreversible Reactions

Reversible reactions	Irreversible reactions
Can proceed in forward and reverse directions.	Proceed predominantly in one direction under the

Reversible reactions	Irreversible reactions
	reaction conditions.
Products can reform reactants.	Products do not readily reform the original reactants.
Represented by a double arrow.	Represented by a single arrow.
Can establish dynamic equilibrium in a closed system.	Usually proceed towards completion under the conditions used.
Example: Haber process.	Example: Burning of magnesium.

Important: The distinction depends on reaction conditions. Some reactions that appear irreversible in an ordinary laboratory can be reversed under specially controlled conditions.

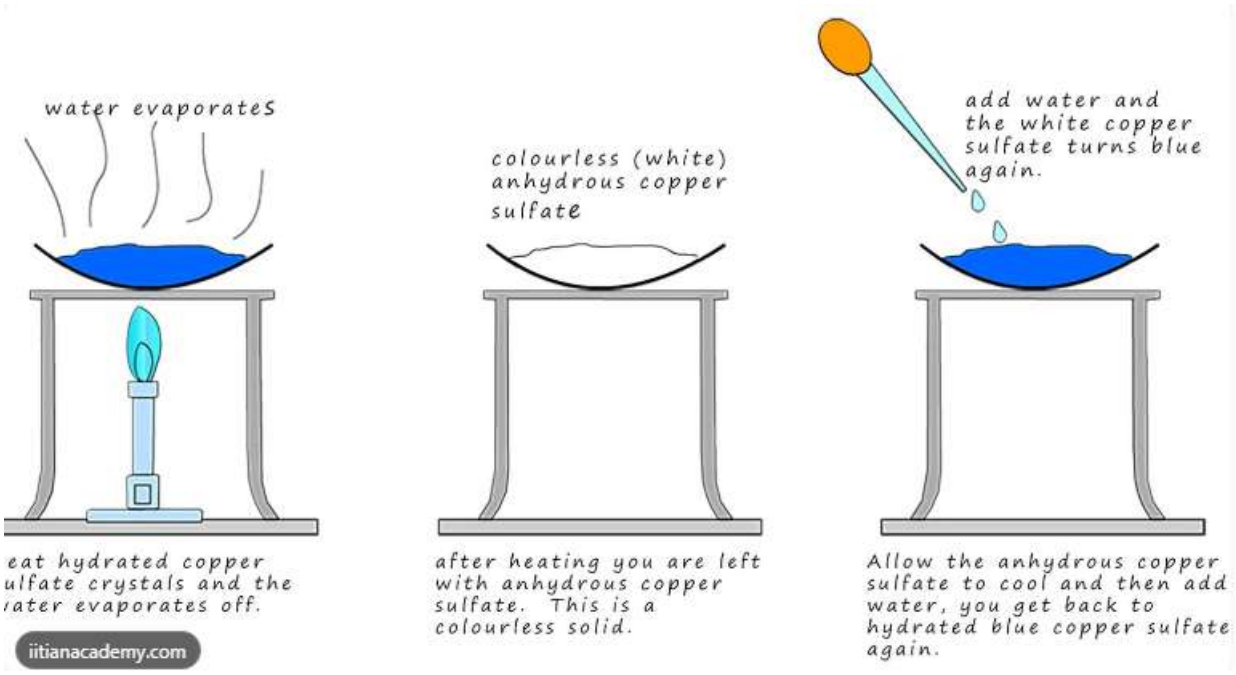
13.0 EXPERIMENTS TO DEMONSTRATE REVERSIBLE CHANGES

Reversible changes can be demonstrated through practical activities in the chemistry laboratory.

The experiments below illustrate changes that can occur in both directions under suitable conditions.

13.1 Heating Hydrated Copper(II) Sulphate

Aim



To demonstrate a reversible change involving hydrated copper(II) sulphate.

Background Information

Hydrated copper(II) sulphate contains water of crystallisation.

Its chemical formula is:

$$CuSO_4 \cdot 5H_2O$$

It is blue in colour.

When heated, it loses its water of crystallisation and forms white anhydrous copper(II) sulphate.

When water is added to the white solid, the blue hydrated compound is formed again.

Materials and Apparatus

- Hydrated copper(II) sulphate crystals.
- Dry test tube.
- Test-tube holder.
- Bunsen burner or suitable heating source.
- Dropper.
- Water.
- Protective goggles.

Procedure

1. Place a small quantity of blue hydrated copper(II) sulphate crystals in a dry test tube.
2. Hold the test tube using a suitable test-tube holder.
3. Heat the crystals gently.
4. Observe the colour change.
5. Observe any water droplets forming on the cooler part of the test tube.
6. Remove the heat source and allow the test tube to cool.
7. Add a few drops of water to the white solid.
8. Observe the colour change.

Observations

1. Blue hydrated copper(II) sulphate crystals turn white on heating.
2. Water vapour is released.
3. Water droplets may condense on the cooler part of the test tube.
4. Addition of water to the white solid restores the blue colour.

Chemical Equation

$$\boxed{CuSO_4 \cdot 5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(g)}$$

Explanation

Forward change: Heating

1. Heating provides energy to the hydrated copper(II) sulphate.
2. The hydrated salt loses its water of crystallisation.
3. White anhydrous copper(II) sulphate is formed.
4. Water vapour is released.

Reverse change: Adding water

5. Water is added to the anhydrous copper(II) sulphate.
6. The anhydrous salt combines with water.
7. Blue hydrated copper(II) sulphate is formed again.

Conclusion

The change is reversible because hydrated copper(II) sulphate can be converted into anhydrous copper(II) sulphate by heating, while the anhydrous salt can be converted back by adding water.

Colour Changes Summarised

Blue hydrated copper(II) sulphate

$$CuSO_4 \cdot 5H_2O$$

↓ Heating

White anhydrous copper(II) sulphate

$CuSO_4$

↓ Adding water

Blue hydrated copper(II) sulphate

Safety Precautions

1. Wear protective goggles.
2. Use only a small quantity of copper(II) sulphate.
3. Avoid inhaling copper(II) sulphate dust.
4. Do not point the test tube towards anyone.
5. Handle hot glassware using a suitable holder.
6. Allow the test tube to cool before adding water.
7. Dispose of copper-containing waste according to laboratory instructions.

13.2 Heating Zinc Oxide

Aim

To observe the reversible colour change of zinc oxide when heated and allowed to cool.





Background Information

Zinc oxide is a white solid at room temperature.

When strongly heated, it changes to yellow. On cooling, it returns to its original white colour.

This makes zinc oxide useful for demonstrating a reversible temperature-dependent colour change.

Materials and Apparatus

- Zinc oxide powder.
- Heat-resistant container or suitable crucible.
- Spatula.
- Bunsen burner or suitable heating source.
- Tongs.
- Protective goggles.

Procedure

1. Place a small quantity of zinc oxide in a suitable heat-resistant container.
2. Heat it strongly using an appropriate heating source.
3. Observe the colour of the solid while hot.
4. Remove the heat source.
5. Allow the zinc oxide to cool.
6. Observe the colour again.

Observations

1. Zinc oxide is white at room temperature.
2. On strong heating, it becomes yellow.
3. On cooling, it returns to white.

Explanation

The yellow colour observed at high temperature is associated with temperature-dependent changes in the solid's electronic structure and defect population.

When zinc oxide cools, these changes are reversed and the white appearance returns.

Important Scientific Clarification

This is a reversible temperature-dependent colour change in a solid.

It is not a straightforward chemical reaction in which zinc oxide decomposes into zinc and oxygen.

The colour change can be represented as:

$$ \text{White zinc oxide} \xrightleftharpoons[\text{cooling}]{\text{heating}} \text{Yellow zinc oxide} $$

This representation shows the reversible colour change, not a balanced chemical reaction.

Conclusion

Zinc oxide changes from white to yellow when heated and returns to white when cooled. The change is reversible and depends on temperature.

Safety Precautions

1. Wear protective goggles.
2. Avoid inhaling zinc oxide dust.
3. Use a suitable heat-resistant container.
4. Handle hot apparatus using tongs.
5. Do not touch the heated solid.
6. Allow the apparatus to cool before handling.

REVISION CHECKPOINT QUESTIONS

A. SHORT-ANSWER QUESTIONS

1. Define the term reaction rate.

2. Explain what is meant by the gradient of a reaction-rate graph.

3. State three features of a graph showing product formed against time.

4. Explain why the rate of product formation usually decreases as a reaction proceeds.

5. State what a horizontal section of a product-against-time graph indicates.

6. Explain how reactant concentration changes as a reaction proceeds.

7. Distinguish between a steep gradient and a gentle gradient.

8. Define catalysis.

9. What is a catalyst?

10. Define activation energy.

11. Explain how a catalyst increases the rate of a chemical reaction.

12. Distinguish between homogeneous and heterogeneous catalysts.

13. Give three examples of enzyme catalysts and state their functions.

14. State four industrial applications of catalysts.

15. Explain why a catalyst does not change the equilibrium yield of a reversible reaction at a fixed temperature.

16. Define a reversible reaction.

17. Define an irreversible reaction.

18. State the symbols used to represent reversible and irreversible reactions.

19. Give three examples of reversible reactions.

20. Give three examples of irreversible reactions.

21. Describe the reversible change involving hydrated copper(II) sulphate.

22. State the colour change observed when zinc oxide is heated and cooled.

23. Explain why the colour change of zinc oxide should not be described as a simple decomposition reaction.

B. APPLICATION QUESTIONS

Question 1: Interpreting Reaction-Rate Graphs

A learner plotted a graph of the volume of hydrogen gas produced against time during the reaction between magnesium and dilute hydrochloric acid.

a) Describe the shape of the graph expected.

b) Explain why the graph becomes less steep as time increases.

c) What does a horizontal section of the graph indicate?

Question 2: Comparing Reaction Rates

Two experiments are carried out using equal masses of magnesium. In Experiment A, dilute hydrochloric acid is used, while in Experiment B, more concentrated hydrochloric acid is used.

d) Which experiment is expected to have a faster initial reaction rate?

e) Explain your answer using collision theory.

f) State one factor that must be kept constant for a fair comparison.

Question 3: Reversible Changes in Copper(II) Sulphate

A learner heats blue hydrated copper(II) sulphate crystals and obtains a white solid.

g) Name the white solid formed.

h) Write the chemical equation for the change.

i) Explain how the original blue crystals can be obtained again.

Question 4: Reversible Colour Change in Zinc Oxide

A learner heats zinc oxide and observes that it changes from white to yellow.

j) State the colour observed when the solid is cooled.

k) Explain the cause of the colour change.

l) Why is the observation not evidence that zinc oxide has decomposed into zinc and oxygen?

15.0 KEY POINTS TO REMEMBER

1. Reaction-rate graphs show how reactant or product quantities change with time.
2. A steep gradient represents a high rate of change.
3. A gentle gradient represents a low rate of change.
4. A horizontal section shows that the measured quantity is no longer changing.
5. The gradient of a product-against-time graph can be used to determine the rate of product formation.
6. Catalysts provide alternative reaction pathways with lower activation energy.
7. Catalysts increase reaction rates without being permanently consumed in the overall reaction.
8. Catalysts do not change the equilibrium position or equilibrium constant at a fixed temperature.
9. Homogeneous catalysts exist in the same physical phase as the reactants.
10. Heterogeneous catalysts exist in a different physical phase from the reactants.
11. Enzymes are biological catalysts that speed up biochemical reactions.
12. Reversible reactions can proceed in both forward and reverse directions.
13. Irreversible reactions proceed predominantly in one direction under the conditions used.
14. Hydrated copper(II) sulphate changes from blue to white on heating.
15. Adding water to anhydrous copper(II) sulphate restores the blue hydrated salt.
16. Zinc oxide changes from white to yellow when heated and returns to white on cooling.
17. The colour change of zinc oxide is a reversible temperature-dependent change, not a simple chemical decomposition.

13.3 Heating iodine

Iodine is a solid that sublimes when heated.

Definition of sublimation

Sublimation is the direct change of a substance from the solid state to the gaseous state without passing through the liquid state.

Procedure:

1. Place a few iodine crystals in a suitable dry container.
2. Heat gently using an approved demonstration setup.
3. Observe the purple vapour produced.
4. Allow the vapour to cool on a cold surface.
5. Observe the formation of solid iodine crystals.

Observations:

- Grey-black iodine crystals produce purple vapour when heated.
- On cooling, the purple vapour forms solid iodine crystals.

Representation:

$$I_2(s) \rightleftharpoons I_2(g)$$

This is a reversible physical change involving sublimation and deposition.

Applications and significance:

1. Demonstrates that some physical changes can be reversed.
2. Helps learners understand the effect of temperature on changes of state.
3. Provides an example of dynamic equilibrium between a solid and its vapour in a closed container.

Safety: Iodine vapour is harmful and irritating. Use very small quantities in a fume cupboard under teacher supervision.

13.4 Heating candle wax

Candle wax undergoes reversible physical changes when heated and cooled.

Procedure:

1. Place a small piece of candle wax in a suitable heat-resistant container.
2. Heat it gently.
3. Observe the wax melting.
4. Remove the heat.
5. Allow the molten wax to cool.
6. Observe the wax solidifying.

Observations:

- Solid wax changes into liquid wax when heated.
- Liquid wax changes back into solid wax when cooled.

Representation:

$$\text{Solid wax} \rightleftharpoons \text{Liquid wax}$$

Important: Melting and solidification of wax are physical changes, not chemical reactions. They demonstrate reversible changes of state, not a chemical equilibrium involving different chemical substances.

14.0 MEANING OF CHEMICAL EQUILIBRIUM

14.1 Definition of chemical equilibrium

Chemical equilibrium is the state in a reversible chemical reaction in which the rate of the forward reaction is equal to the rate of the reverse reaction, resulting in no overall change in the concentrations of reactants and products.

Chemical equilibrium is established in a closed system under constant conditions.

14.2 Explanation of chemical equilibrium

Consider the reversible reaction:

$$A + B \rightleftharpoons C + D$$

When the reaction begins:

1. Reactants A and B collide and form products C and D.
2. The forward reaction initially proceeds rapidly because the concentration of reactants is high.
3. As products accumulate, the reverse reaction begins and increases in rate.
4. As reactants are used up, the forward reaction gradually slows down.
5. The reverse reaction continues as long as products are present.
6. Eventually, the forward and reverse reaction rates become equal.
7. The system reaches dynamic equilibrium.

14.3 Meaning of dynamic equilibrium

Dynamic equilibrium is a condition in which forward and reverse reactions continue to occur at equal rates.

It is described as *dynamic* because both reactions continue even though there is no observable overall change in the concentrations of reactants and products.

14.4 Characteristics of chemical equilibrium

1. It is established only in a reversible reaction.
2. It is established in a closed system.
3. The forward reaction continues to occur.
4. The reverse reaction continues to occur.
5. The rates of the forward and reverse reactions are equal.
6. The concentrations of reactants and products remain constant.
7. The concentrations of reactants and products are not necessarily equal.
8. The macroscopic properties of the system remain constant.
9. Equilibrium can be reached from either direction under suitable conditions.
10. Changes in temperature, pressure or concentration can disturb the equilibrium position.

14.5 Difference between static and dynamic equilibrium

Static equilibrium	Dynamic equilibrium
No movement or change occurs.	Forward and reverse processes continue.
No opposing processes are taking place.	Opposing processes occur at equal rates.
Example: A book resting on a table.	Example: A reversible chemical reaction at equilibrium.

14.6 Conditions necessary for chemical equilibrium

1. The reaction must be reversible.
2. The system must be closed so that reactants and products do not escape.
3. Temperature should remain constant for a given equilibrium state.
4. Other relevant conditions, such as pressure, should be controlled.
5. Sufficient time must be allowed for equilibrium to be established.

14.7 Concentration changes during the establishment of equilibrium

Consider the reaction:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

Initially:

- Hydrogen and iodine concentrations are relatively high.
- The forward reaction rate is high.
- The reverse reaction rate is initially low if little or no hydrogen iodide is present.

As the reaction proceeds:

- Hydrogen and iodine concentrations decrease.
- Hydrogen iodide concentration increases.
- The forward reaction rate decreases.
- The reverse reaction rate increases.

At equilibrium:

- The forward reaction rate equals the reverse reaction rate.
- Concentrations remain constant.
- Both reactions continue to occur.

Concentration changes during the establishment of equilibrium: The concentration of reactants decreases while the concentration of products increases. Eventually, both become constant.

14.8 Reaction rates during equilibrium

At equilibrium, the forward and reverse reaction rates are equal and remain constant.

Important: At equilibrium, the concentrations of reactants and products are constant, but they are not necessarily equal.

15.0 POSITION OF CHEMICAL EQUILIBRIUM

The position of equilibrium describes the relative amounts of reactants and products present in a system at equilibrium.

15.1 Equilibrium favouring products

An equilibrium is said to favour products when relatively more products than reactants are present at equilibrium.

For example:

$$A + B \rightleftharpoons C + D$$

If relatively more C and D are present at equilibrium, the equilibrium lies towards the product side.

15.2 Equilibrium favouring reactants

An equilibrium favours reactants when relatively more reactants than products are present at equilibrium.

15.3 Importance of equilibrium position

Understanding equilibrium position helps chemists to:

1. Predict the amount of product formed.
2. Select suitable reaction conditions.
3. Improve industrial product yields.
4. Reduce waste in chemical manufacturing.
5. Improve the efficiency of chemical production.
6. Determine suitable temperature and pressure conditions.

16.0 LE CHATELIER'S PRINCIPLE

16.1 Definition of Le Chatelier's principle

Le Chatelier's principle states that:

When a system at equilibrium is subjected to a change in concentration, pressure or temperature, the equilibrium shifts in the direction that opposes the imposed change.

The principle helps chemists predict how reversible reactions respond to changes in their conditions.

16.2 Meaning of equilibrium shift

An equilibrium shift is a change in the relative amounts of reactants and products as a result of changing the conditions of a system at equilibrium.

An equilibrium may shift:

- To the right, towards products.
- To the left, towards reactants.

Important: A shift does not mean that the reaction stops in one direction. Both forward and reverse reactions continue, but their rates become unequal temporarily until a new equilibrium is established.

17.0 FACTORS AFFECTING THE POSITION OF CHEMICAL EQUILIBRIUM

The main factors affecting equilibrium position are:

1. Concentration.
2. Pressure.
3. Temperature.

A catalyst affects the time taken to reach equilibrium but does not change the equilibrium position.

17.1 Effect of concentration on equilibrium position

Changing the concentration of a reactant or product can shift the position of equilibrium.

According to Le Chatelier's principle:

- Increasing the concentration of a reactant shifts equilibrium towards the products.
- Decreasing the concentration of a reactant shifts equilibrium towards the reactants.
- Increasing the concentration of a product shifts equilibrium towards the reactants.
- Removing a product shifts equilibrium towards the products.

Example: Haber process

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

A. Increasing nitrogen concentration

When more nitrogen is added:

1. The concentration of nitrogen increases.
2. The equilibrium is disturbed.
3. The system responds by using up some of the added nitrogen.
4. Equilibrium shifts to the right.
5. More ammonia is formed.

B. Increasing hydrogen concentration

When more hydrogen is added:

- Equilibrium shifts to the right.
- More ammonia is formed.

C. Removing ammonia

When ammonia is removed:

- The product concentration decreases.
- Equilibrium shifts to the right.
- More ammonia is formed to oppose the removal.

D. Increasing ammonia concentration

When ammonia is added:

- Equilibrium shifts to the left.
- More nitrogen and hydrogen are formed.

Applications of concentration changes

1. Adding reactants to increase product formation.
2. Removing products to improve product yield.
3. Recycling unreacted reactants in industrial plants.
4. Controlling product concentrations in chemical manufacturing.

17.2 Effect of pressure on equilibrium position

Pressure affects equilibria involving gases when the total number of gaseous moles differs between the reactant and product sides.

According to Le Chatelier's principle:

- Increasing pressure favours the side with fewer moles of gas.
- Decreasing pressure favours the side with more moles of gas.
- If both sides have equal numbers of gaseous moles, changing pressure has no significant effect on the equilibrium position under ideal conditions.

Example 1: Haber process

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Count the gaseous moles.

Reactant side:

$$1 + 3 = 4 \text{ moles}$$

Product side:

$$2 \text{ moles}$$

Increasing pressure

1. Pressure increases.
2. The system responds by favouring the side with fewer gaseous moles.
3. The product side has fewer gaseous moles.
4. Equilibrium shifts to the right.
5. Ammonia yield increases.

Decreasing pressure

1. Pressure decreases.
2. The system favours the side with more gaseous moles.
3. The reactant side has more gaseous moles.
4. Equilibrium shifts to the left.

5. Ammonia yield decreases.

Example 2: Formation of hydrogen iodide

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

Reactant side:

$$1 + 1 = 2 \text{ moles}$$

Product side:

$$2 \text{ moles}$$

Both sides contain equal numbers of gaseous moles. Therefore, changing pressure does not significantly shift the equilibrium position under ideal gas assumptions.

Example 3: Formation of sulphur trioxide

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Reactant side:

$$2 + 1 = 3 \text{ moles}$$

Product side:

$$2 \text{ moles}$$

Increasing pressure shifts equilibrium to the right because the product side has fewer gaseous moles.

Summary of pressure effects

Change in pressure	Equilibrium response
Increase pressure	Shifts towards fewer gaseous moles
Decrease pressure	Shifts towards more gaseous moles
Equal gaseous moles on both sides	No significant shift

Important: Pure solids and liquids are not included when counting gaseous moles for predicting pressure effects.

17.3 Effect of temperature on equilibrium position

Temperature affects equilibrium position depending on whether the forward reaction is exothermic or endothermic.

A. Exothermic forward reaction

An exothermic reaction releases heat. Heat can be treated as a product in the forward reaction.

General representation:

$$A + B \rightleftharpoons C + D + \text{Heat}$$

When temperature increases:

- Additional heat is supplied.
- The system opposes the increase by favouring the reverse, endothermic reaction.
- Equilibrium shifts to the left.

- Product yield decreases.

When temperature decreases:

- Heat is removed.
- The system favours the forward, exothermic reaction.
- Equilibrium shifts to the right.
- Product yield increases.

B. Endothermic forward reaction

An endothermic reaction absorbs heat. Heat can be treated as a reactant.

General representation:

$$A + B + \text{Heat} \rightleftharpoons C + D$$

When temperature increases:

When temperature increases:

- Additional heat is supplied.
- The system favours the forward, endothermic reaction.
- Equilibrium shifts to the right.
- Product yield increases.

When temperature decreases:

- Heat is removed.
- The system favours the reverse, exothermic reaction.
- Equilibrium shifts to the left.
- Product yield decreases.

Summary of temperature effects

Type of forward reaction	Increase in temperature	Decrease in temperature
Exothermic	Shifts to the left	Shifts to the right
Endothermic	Shifts to the right	Shifts to the left

Example 1: Haber process

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The forward reaction is exothermic.

$$\Delta H = -92\,kJ\,mol^{-1}$$

Therefore:

- Increasing temperature shifts equilibrium to the left.
- Decreasing temperature shifts equilibrium to the right.

A lower temperature favours ammonia formation, although the reaction becomes slower. Industrial production therefore uses a compromise temperature.

Example 2: Thermal decomposition of calcium carbonate

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

The forward reaction is endothermic.

Therefore:

- Increasing temperature favours the forward reaction.
- More calcium oxide and carbon dioxide are produced.
- Decreasing temperature favours the reverse reaction.

17.4 Effect of a catalyst on equilibrium

A catalyst increases the rates of both the forward and reverse reactions.

It does not change:

- The equilibrium position.
- The equilibrium constant at a fixed temperature.
- The theoretical equilibrium yield.

However, it reduces the time needed to reach equilibrium.

Example: Haber process

Iron is used as a catalyst in the manufacture of ammonia. It enables the reaction to reach equilibrium faster, but it does not shift the equilibrium towards ammonia.

17.5 Effect of changing the volume of a gaseous equilibrium system

Changing the volume of a gaseous system changes its pressure.

- Decreasing the volume increases pressure and favours the side with fewer gaseous moles.
- Increasing the volume decreases pressure and favours the side with more gaseous moles.

This is another way of applying the pressure principle.

17.6 Effect of inert gases

The effect of adding an inert gas depends on the conditions.

- At constant volume, adding an inert gas does not change the partial pressures of the reacting gases and does not shift the equilibrium.
- At constant pressure, adding an inert gas increases the volume and may shift the equilibrium towards the side with more gaseous moles.

For introductory equilibrium questions, learners should first identify whether the volume or pressure is being kept constant.

18.0 EQUILIBRIUM CONSTANT

18.1 Meaning of equilibrium constant

The equilibrium constant is a numerical value that describes the relationship between the concentrations or partial pressures of products and reactants at equilibrium at a particular temperature.

For a general reaction:

$$aA + bB \rightleftharpoons cC + dD$$

The concentration equilibrium constant is:

$$\boxed{K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

Where:

- K_c is the equilibrium constant in terms of concentration.
- Square brackets represent equilibrium concentrations.
- a, b, c, d are the coefficients in the balanced chemical equation.

Pure solids and pure liquids are omitted from the equilibrium expression.

18.2 Meaning of the magnitude of the equilibrium constant

- A large K_c indicates that products are favoured at equilibrium.
- A small K_c indicates that reactants are favoured at equilibrium.
- A value of K_c around 1 indicates that neither side is overwhelmingly favoured.

18.3 Important properties of the equilibrium constant

1. Its value depends on temperature.
2. It is not changed by adding a catalyst.
3. It is not changed by changing the concentration of reactants or products.
4. It is not changed by changing pressure at constant temperature.
5. It describes the equilibrium composition of a reaction system.
6. It is calculated using the balanced chemical equation.

18.4 Examples of writing equilibrium expressions

Example 1

For the reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The equilibrium expression is:

$$\boxed{K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}}$$

Example 2

For the reaction:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

The equilibrium expression is:

$$\boxed{K_c = \frac{[HI]^2}{[H_2][I_2]}}$$

Example 3

For the reaction:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

The solids are omitted from the equilibrium expression.

$$\boxed{K_c = [CO_2]}$$

These expressions apply to the specified equilibrium reactions at a fixed temperature.

PART C: INDUSTRIAL APPLICATIONS OF CHEMICAL EQUILIBRIUM

19.0 IMPORTANCE OF CHEMICAL EQUILIBRIUM IN INDUSTRY

Chemical equilibrium is important in industries where reversible reactions are used to manufacture useful products.

Industrial chemists apply Le Chatelier's principle to select suitable:

- Temperatures.
- Pressures.
- Reactant concentrations.
- Catalysts.
- Gas recycling methods.
- Product removal methods.

The main aim is to obtain a high product yield at a reasonable production rate and cost.

Product yield refers to the amount of product obtained from a chemical reaction compared with the amount theoretically expected.

The main industrial processes studied are:

1. Haber process.
2. Ostwald process.
3. Contact process.

20.0 HABER PROCESS: INDUSTRIAL MANUFACTURE OF AMMONIA

20.1 Meaning of the Haber process

The Haber process is an industrial method used to manufacture ammonia by reacting nitrogen gas with hydrogen gas under controlled conditions of temperature and pressure in the presence of an iron catalyst.

Ammonia is an important industrial chemical used mainly in the manufacture of fertilisers and other nitrogen-containing compounds.

20.2 Importance of ammonia

Ammonia is used in:

1. Manufacture of nitrogenous fertilisers.
2. Production of ammonium nitrate.
3. Production of ammonium sulphate.
4. Manufacture of urea fertiliser.
5. Manufacture of nitric acid.
6. Manufacture of cleaning products.
7. Refrigeration systems in certain industrial applications.
8. Production of other nitrogen-containing chemicals.

20.3 Raw materials used in the Haber process

The main raw materials are:

A. Nitrogen

Nitrogen is obtained from air. Air contains approximately 78% nitrogen by volume.

B. Hydrogen

Hydrogen is commonly obtained industrially from natural gas through steam reforming, followed by processes that convert carbon monoxide into carbon dioxide and produce additional hydrogen.

Hydrogen can also be produced by other methods, including electrolysis of water.

20.4 Chemical equation for the Haber process

$$\boxed{N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)}$$

The forward reaction is exothermic.

$$\Delta H = -92\, kJ\, mol^{-1}$$

This means that energy is released when ammonia is formed from nitrogen and hydrogen.

20.5 Conditions used in the Haber process

The industrial process uses approximately the following conditions:

Condition	Typical operating value
Temperature	About 400–450°C
Pressure	About 150–250 atmospheres
Catalyst	Finely divided iron
Reactant ratio	1 volume nitrogen : 3 volumes hydrogen
Reaction type	Reversible and exothermic

Actual operating conditions vary between industrial plants.

20.6 Why high pressure is used

Consider the equation:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Number of gaseous moles on the reactant side:

$$1 + 3 = 4 \text{ moles}$$

Number of gaseous moles on the product side:

$$2 \text{ moles}$$

According to Le Chatelier's principle:

- Increasing pressure favours the side with fewer gaseous moles.
- The product side has fewer gaseous moles.
- Equilibrium shifts to the right.
- More ammonia is produced.

Therefore, high pressure increases the equilibrium yield of ammonia.

Why is extremely high pressure not used?

1. Very high pressure requires expensive equipment.
2. Stronger reaction vessels are needed.
3. Compression consumes energy.
4. Maintenance and safety costs increase.

5. The additional improvement in yield may not justify the cost.

Industrial plants therefore use a pressure that provides a practical balance between yield, production rate, cost and safety.

20.7 Why a moderate temperature is used

The forward reaction is exothermic.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3 + \text{Heat}$$

According to Le Chatelier's principle:

- Lower temperature favours ammonia formation.
- Higher temperature favours the reverse reaction.

However, very low temperatures make the reaction too slow.

A temperature of approximately 400–450°C is therefore used as a compromise between:

- A favourable equilibrium yield.
- A sufficiently fast reaction rate.
- Economical industrial production.

20.8 Why an iron catalyst is used

Finely divided iron is used as a catalyst.

Its importance includes:

1. Providing an alternative reaction pathway.
2. Lowering the activation energy.
3. Increasing the rate at which equilibrium is reached.
4. Improving the efficiency of the industrial process.

The catalyst does not change the equilibrium position or the equilibrium yield at a fixed temperature.

20.9 Why nitrogen and hydrogen are mixed in a 1:3 ratio

The balanced equation is:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The mole ratio is:

$$1:3:2$$

Therefore:

- One mole of nitrogen reacts with three moles of hydrogen.
- Two moles of ammonia are formed.

The gases are supplied in approximately the stoichiometric ratio of 1:3.

20.10 Steps involved in the Haber process

1. Nitrogen is separated from air.
2. Hydrogen is obtained from an appropriate industrial source.
3. The nitrogen and hydrogen are purified to remove substances that may poison the catalyst.
4. The gases are mixed in the correct ratio.
5. The mixture is compressed to the required pressure.
6. The compressed gases are heated to the operating temperature.

7. The gases pass over the iron catalyst.
8. Ammonia is formed.
9. The reaction mixture is cooled.
10. Ammonia is condensed and separated.
11. Unreacted nitrogen and hydrogen are recycled into the reactor.

Flow diagram: Haber process

Nitrogen from air

Hydrogen from industrial production



Purification and mixing

Nitrogen : Hydrogen = 1 : 3



Compression and heating

High pressure and approximately 400–450°C



Catalytic reactor

Iron catalyst

$N_2 + 3H_2 \rightleftharpoons 2NH_3$



Cooling and condensation



Ammonia collected

Liquid ammonia is separated and stored safely.

Unreacted nitrogen and hydrogen are recycled to the reactor.

20.11 Why ammonia is removed from the reaction mixture

Ammonia is cooled and condensed after leaving the reactor.

This is important because:

1. Ammonia is separated from unreacted nitrogen and hydrogen.
2. Removing ammonia reduces its concentration in the reaction mixture.
3. Equilibrium is encouraged to shift towards the product side when the gases are returned to the reactor.
4. More ammonia can be produced.
5. The unreacted gases can be reused, reducing waste.

20.12 Importance of the Haber process in Kenya

The Haber process is important to Kenyan agriculture and industry because ammonia is used to manufacture nitrogen-containing fertilisers.

These fertilisers:

- Supply nitrogen to crops.
- Promote vegetative growth.
- Improve agricultural productivity when used appropriately.
- Support food production.
- Contribute to the supply of fertilisers used by farmers.

Examples of nitrogenous fertilisers include:

- Urea.
- Ammonium nitrate.
- Ammonium sulphate.

CBE application: Learners can investigate how nitrogen fertilisers support crop production in Kenyan farming communities and discuss the environmental effects of excessive fertiliser use.

21.0 OSTWALD PROCESS: INDUSTRIAL MANUFACTURE OF NITRIC ACID

21.1 Meaning of the Ostwald process

The Ostwald process is an industrial method used to manufacture nitric acid by the catalytic oxidation of ammonia.

Nitric acid is an important industrial chemical used in the manufacture of fertilisers, explosives and other chemicals.

21.2 Importance of nitric acid

Nitric acid is used in:

1. Manufacture of ammonium nitrate fertiliser.
2. Production of other nitrogen-containing fertilisers.
3. Manufacture of explosives.
4. Production of dyes.
5. Manufacture of certain plastics and chemicals.
6. Metal treatment and cleaning processes.
7. Chemical laboratory applications.

21.3 Raw materials used

The main raw materials are:

1. Ammonia.
2. Oxygen from air.
3. Water.

Ammonia used in the Ostwald process is commonly obtained from the Haber process.

21.4 Stages of the Ostwald process

The industrial manufacture of nitric acid involves three major chemical stages.

Stage 1: Catalytic oxidation of ammonia

Ammonia is mixed with air and passed over a platinum-rhodium catalyst at a high temperature.

$$\boxed{4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)}$$

Conditions:

- Temperature: approximately 800–950°C.
- Catalyst: platinum-rhodium gauze.
- Oxygen supplied in excess.

The reaction is highly exothermic.

Explanation:

1. Ammonia reacts with oxygen.
2. Nitric oxide is produced.
3. Water vapour is also formed.
4. The catalyst increases the reaction rate.

Stage 2: Oxidation of nitric oxide

Nitric oxide reacts with additional oxygen to form nitrogen dioxide.

$$\boxed{2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)}$$

This reaction is exothermic.

Nitrogen dioxide is a brown gas.

Stage 3: Absorption of nitrogen dioxide in water

Nitrogen dioxide is absorbed in water in an absorption tower to produce nitric acid.

An overall representation is:

$$\boxed{3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)}$$

The nitric oxide produced can be oxidised again to nitrogen dioxide and recycled.

21.5 Flow diagram of the Ostwald process

Ammonia

From ammonia production

Air

Source of oxygen



Catalytic oxidation

Platinum-rhodium gauze, high temperature

$NH_3 \rightarrow NO$



Oxidation of nitric oxide

$2NO + O_2 \rightarrow 2NO_2$



Absorption tower

Nitrogen dioxide reacts with water



Nitric acid

HNO_3

Nitric oxide formed during absorption can be oxidised again and recycled.

21.6 Role of equilibrium in the Ostwald process

Equilibrium is important in the oxidation of ammonia and nitric oxide.

For the oxidation of nitric oxide:

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

Reactant side:

$$2 + 1 = 3 \text{ gaseous moles}$$

Product side:

2 gaseous moles

Therefore:

- Increasing pressure favours nitrogen dioxide formation.
- Decreasing pressure favours the reactant side.

Since the reaction is exothermic:

- Lower temperatures favour nitrogen dioxide formation.
- Higher temperatures favour the reverse reaction.

Industrial operation must also consider reaction rate, heat removal, catalyst performance and the practical design of the plant.

21.7 Importance of the Ostwald process in Kenya

Nitric acid is important in the manufacture of fertilisers, especially ammonium nitrate.

It also has applications in:

- Chemical manufacturing.
- Metal processing.
- Laboratory analysis.
- Production of industrial chemicals.

Environmental consideration: Nitrogen oxides can contribute to air pollution and acid rain. Industrial plants must therefore use suitable emission-control systems.

22.0 CONTACT PROCESS: INDUSTRIAL MANUFACTURE OF SULPHURIC ACID

22.1 Meaning of the Contact process

The Contact process is an industrial method used to manufacture sulphuric acid by oxidising sulphur dioxide to sulphur trioxide and then converting the sulphur trioxide into sulphuric acid.

Sulphuric acid is one of the most important industrial chemicals.

22.2 Importance of sulphuric acid

Sulphuric acid is used in:

1. Manufacture of phosphate fertilisers.
2. Manufacture of detergents.
3. Production of chemicals.
4. Petroleum refining.
5. Manufacture of lead-acid batteries.
6. Metal processing.
7. Manufacture of pigments and dyes.
8. Chemical laboratory processes.

22.3 Raw materials used in the Contact process

The main raw materials include:

1. Sulphur or sulphur-containing materials.
2. Oxygen from air.
3. Water.

Sulphur dioxide can be obtained by burning sulphur or roasting suitable sulphide ores.

22.4 Stages of the Contact process

Stage 1: Production of sulphur dioxide

Sulphur is burned in air to produce sulphur dioxide.

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

The reaction is exothermic.

Stage 2: Oxidation of sulphur dioxide to sulphur trioxide

Sulphur dioxide reacts with oxygen to form sulphur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Conditions:

- Temperature: approximately 400–450°C.
- Catalyst: vanadium(V) oxide, V_2O_5.
- Pressure: usually close to atmospheric pressure or slightly above, depending on plant design.
- Oxygen is supplied in sufficient quantity.

This is the most important equilibrium stage in the Contact process.

Stage 3: Absorption of sulphur trioxide

Sulphur trioxide is absorbed in concentrated sulphuric acid to form oleum.

$$SO_3(g) + H_2SO_4(l) \rightarrow H_2S_2O_7(l)$$

Oleum is then carefully diluted with water to produce sulphuric acid.

$$H_2S_2O_7(l) + H_2O(l) \rightarrow 2H_2SO_4(l)$$

Important: Sulphur trioxide is not normally absorbed directly into water on an industrial scale because the reaction is highly exothermic and produces a fine mist of sulphuric acid that is difficult to condense and collect.

22.5 Flow diagram of the Contact process

Sulphur



Burning in air

$$S + O_2 \rightarrow SO_2$$



Purification and drying of gases



Catalytic converter

Approximately 400–450°C, vanadium(V) oxide catalyst

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$



Absorption in concentrated sulphuric acid

Oleum formation and controlled dilution



Sulphuric acid

$$H_2SO_4$$

22.6 Explanation of equilibrium in the Contact process

The important reversible reaction is:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

The forward reaction is exothermic.

$$\Delta H < 0$$

Effect of temperature

- Lower temperature favours the formation of sulphur trioxide.
- Higher temperature favours the reverse reaction.
- A temperature of approximately 400–450°C is used as a compromise between yield and reaction rate.

Effect of pressure

Reactant side:

$$2 + 1 = 3 \text{ gaseous moles}$$

Product side:

$$2 \text{ gaseous moles}$$

Therefore:

- Increasing pressure favours sulphur trioxide formation.
- Lower pressure favours the reactant side.

In practice, only moderate pressure is used because the improvement in yield may not justify the extra cost of operating at high pressure.

Effect of catalyst

Vanadium(V) oxide:

- Lowers activation energy.
- Increases the reaction rate.
- Allows equilibrium to be reached faster.
- Does not change the equilibrium position.

22.7 Why gases are purified before entering the catalytic converter

The gases are purified because impurities may:

1. Poison or damage the catalyst.
2. Reduce catalyst efficiency.
3. Cause corrosion of equipment.
4. Reduce the effectiveness of the production process.

The gases are also dried to remove moisture that could interfere with the process.

22.8 Importance of the Contact process in Kenya

Sulphuric acid is important in:

- Fertiliser production.
- Chemical manufacturing.
- Industrial processing.
- Battery manufacture.
- Metal treatment.

Phosphate fertiliser manufacture is particularly important to agriculture because phosphorus is an essential plant nutrient.

22.9 Environmental considerations in the Contact process

1. Sulphur dioxide emissions should be controlled because the gas causes respiratory irritation and contributes to acid rain.
2. Industrial gases should be treated before release.
3. Catalysts should be managed properly.
4. Acid spills should be contained and handled safely.
5. Industrial wastewater should be treated before discharge.

23.0 COMPARISON OF THE THREE INDUSTRIAL PROCESSES

Feature	Haber process	Ostwald process	Contact process
Main product	Ammonia	Nitric acid	Sulphuric acid
Main raw materials	Nitrogen and hydrogen	Ammonia and air	Sulphur or sulphide sources and air
Important reversible reaction	$N_2 + 3H_2 \rightleftharpoons 2NH_3$	$4NH_3 + 5O_2 \rightleftharpoons 4NO + 6H_2O$	$2SO_2 + O_2 \rightleftharpoons 2SO_3$
Catalyst	Iron	Platinum-rhodium gauze	Vanadium(V) oxide
Typical temperature	400–450°C	800–950°C for ammonia oxidation	400–450°C
Pressure	High, approximately 150–250 atm	Depends on plant design	Near atmospheric or slightly above

Feature	Haber process	Ostwald process	Contact process
Important use	Fertiliser manufacture	Fertiliser and chemical manufacture	Fertiliser and chemical manufacture

24.0 HOW INDUSTRIAL CHEMISTS OPTIMISE EQUILIBRIUM

Industrial chemists must balance product yield, reaction speed, cost, energy use and safety.

Factor	Benefit	Possible disadvantage
Lower temperature for an exothermic reaction	Favours product formation at equilibrium	Reaction may become too slow
Higher pressure where fewer gas moles are formed	Increases product yield	Requires expensive, strong equipment
Catalyst	Increases reaction rate	May be costly or require replacement
Increased reactant concentration	Can shift equilibrium towards products	May increase raw material costs
Continuous product removal	Can favour further product formation	Requires additional separation equipment
Recycling unreacted gases	Improves raw material utilisation	Requires extra equipment and energy

Main lesson: The best industrial conditions are not necessarily those that give the highest theoretical equilibrium yield. They are conditions that provide a practical combination of high yield, adequate speed, safety and profitability.

PART D: COMPETENCY-BASED LEARNING ACTIVITIES AND APPLICATIONS

25.0 SUGGESTED LEARNING EXPERIENCES

The following activities help learners develop practical skills, scientific reasoning and the ability to apply knowledge of reaction rates and chemical equilibrium.

Activity 1: Investigating collision theory

Task: Demonstrate how reactant particles collide during chemical reactions.

Learners should:

1. Discuss the kinetic theory of matter.
2. Use balls or particle models to represent reacting particles.
3. Demonstrate collisions between moving particles.
4. Identify effective and ineffective collisions.
5. Explain why some collisions result in chemical reactions.
6. Relate the activity to activation energy.

Expected competency: Critical thinking and problem-solving.

Activity 2: Investigating factors affecting reaction rates

Divide learners into groups and assign each group one factor to investigate.

Group	Investigation
Group 1	Effect of concentration

Group	Investigation
Group 2	Effect of temperature
Group 3	Effect of surface area
Group 4	Effect of a catalyst
Group 5	Effect of light
Group 6	Effect of pressure using a suitable gas reaction or simulation

Each group should:

1. State the aim of the investigation.
2. Identify the apparatus and materials required.
3. Formulate a prediction.
4. Conduct the investigation safely.
5. Record observations.
6. Draw suitable graphs.
7. Explain the results using collision theory.
8. Present findings to the class.

Activity 3: Drawing reaction-rate graphs

Learners should:

1. Obtain experimental data.
2. Plot a graph of volume of gas against time.
3. Plot concentration against time where data are available.
4. Identify steep and gentle gradients.
5. Determine average reaction rates.
6. Compare the rates of different reactions.
7. Explain why the reaction rate changes as time progresses.

Activity 4: Investigating reversible reactions

Learners should:

1. Observe the heating of hydrated copper(II) sulphate.
2. Record the colour change.
3. Add water to the anhydrous product.
4. Observe the reverse change.
5. Write the chemical equation.
6. Explain why the reaction is reversible.

Activity 5: Demonstrating chemical equilibrium

Learners should:

1. Use a suitable reversible reaction or computer simulation.
2. Observe how reactant and product concentrations change.
3. Discuss the forward and reverse reactions.
4. Explain why the concentrations become constant.
5. Distinguish between dynamic and static equilibrium.

Activity 6: Investigating Le Chatelier's principle

Learners should:

1. Discuss a reversible reaction.
2. Predict the effect of changing concentration.
3. Predict the effect of changing pressure.
4. Predict the effect of changing temperature.
5. Use simulations to observe equilibrium shifts.
6. Compare their predictions with observations.

Activity 7: Industrial applications of equilibrium

Learners should research the following industrial processes:

- Haber process.
- Ostwald process.
- Contact process.

They should identify:

1. Raw materials.
2. Main chemical equations.
3. Operating conditions.
4. Catalysts.
5. Industrial products.
6. Economic importance.
7. Environmental challenges.

Learners should present their findings through charts, posters, oral presentations or digital slides.

PART E: GENERAL SUMMARY AND REVISION

29.0 SUMMARY OF IMPORTANT CONCEPTS

A. Reaction rates

- Reaction rate measures how fast reactants are converted into products.
- Collision theory explains how chemical reactions occur.
- Effective collisions require sufficient energy and correct orientation.
- Activation energy is the minimum energy required for an effective reaction pathway.
- Increasing temperature generally increases reaction rate.
- Increasing concentration generally increases reaction rate.
- Increasing the surface area of a solid reactant increases reaction rate.
- Increasing gas pressure generally increases the rate of reactions involving gases.
- Catalysts provide alternative pathways with lower activation energy.
- Light can initiate or increase the rate of photochemical reactions.

B. Measuring reaction rates

Reaction rates can be measured by:

- Measuring the volume of gas produced.
- Measuring loss in mass.
- Measuring changes in concentration.
- Monitoring the formation of a precipitate.
- Measuring changes in suitable physical properties.

C. Reversible reactions and equilibrium

- Reversible reactions proceed in both forward and reverse directions.
- Chemical equilibrium occurs when the forward and reverse reaction rates are equal.
- Equilibrium is dynamic.
- Concentrations remain constant at equilibrium.
- Reactant and product concentrations are not necessarily equal.
- Equilibrium is established in a closed system.

D. Le Chatelier's principle

- Increasing reactant concentration generally shifts equilibrium towards products.
- Increasing product concentration generally shifts equilibrium towards reactants.
- Increasing pressure favours the side with fewer gaseous moles.
- Decreasing pressure favours the side with more gaseous moles.
- Increasing temperature favours the endothermic direction.
- Decreasing temperature favours the exothermic direction.
- A catalyst does not change equilibrium position or equilibrium yield.

E. Industrial applications

- The Haber process manufactures ammonia.
- The Ostwald process manufactures nitric acid.
- The Contact process manufactures sulphuric acid.
- Industrial chemists select conditions that balance yield, speed, safety and cost.

30.0 KEY FORMULAE AND EQUATIONS

Concept	Formula or equation
Average reaction rate	$\frac{\text{Change in quantity}}{\text{Time}}$
Rate of disappearance of reactant	$-\frac{\Delta[\text{Reactant}]}{\Delta t}$
Rate of formation of product	$\frac{\Delta[\text{Product}]}{\Delta t}$
Relative rate in cross-disappearance experiment	$\frac{1}{t}$
General reversible reaction	$A + B \rightleftharpoons C + D$
Haber process	$N_2 + 3H_2 \rightleftharpoons 2NH_3$
Ostwald oxidation of ammonia	$4NH_3 + 5O_2 \rightleftharpoons 4NO + 6H_2O$
Oxidation of nitric oxide	$2NO + O_2 \rightleftharpoons 2NO_2$
Contact process	$2SO_2 + O_2 \rightleftharpoons 2SO_3$
Equilibrium constant	$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

31.0 IMPORTANT CHEMISTRY DEFINITIONS

Term	Definition
Chemical reaction	A process in which reactants are converted into products.
Reaction rate	Change in the amount or concentration of a reactant or product per unit time.

Term	Definition
Collision theory	An explanation that chemical reactions occur through effective collisions between reacting particles.
Effective collision	A collision with sufficient energy and suitable orientation that results in a reaction.
Activation energy	Minimum energy required for particles to react through a particular reaction pathway.
Activated complex	A temporary, unstable arrangement of atoms formed during a reaction.
Catalyst	A substance that changes the rate of a reaction without being permanently consumed in the overall reaction.
Photochemical reaction	A chemical reaction initiated or influenced by light.
Reversible reaction	A reaction that can proceed in forward and reverse directions.
Irreversible reaction	A reaction that proceeds predominantly in one direction under the conditions used.
Chemical equilibrium	A state in which the forward and reverse reaction rates are equal.
Dynamic equilibrium	Equilibrium in which forward and reverse processes continue at equal rates.
Equilibrium position	The relative amounts of reactants and products present at equilibrium.
Le Chatelier's principle	A principle describing how an equilibrium system responds to an imposed change.
Equilibrium constant	A numerical value describing the equilibrium relationship between products and reactants at a specified temperature.
Exothermic reaction	A reaction that releases heat to the surroundings.
Endothermic reaction	A reaction that absorbs heat from the surroundings.
Industrial catalyst	A catalyst used to increase the rate of an industrial chemical process.
Product yield	The amount of product obtained compared with the theoretical amount expected.

REVISION QUESTIONS

SUB-STRAND 2.3: REACTION RATES AND CHEMICAL EQUILIBRIUM

SECTION A: MULTIPLE-CHOICE QUESTIONS

Instructions: Choose the correct answer from the alternatives provided.

1. The rate of a chemical reaction refers to:

A. The total mass of reactants used.

B. The change in amount or concentration of a substance per unit time.

C. The temperature at which a reaction occurs.

D. The total energy stored in a substance.

2. According to collision theory, a chemical reaction occurs when particles:

A. Move slowly towards one another.

B. Collide with any amount of energy.

C. Collide with sufficient energy and correct orientation.

D. Have equal masses.

3. Which of the following describes activation energy?

A. Energy released during a reaction.

B. Minimum energy required for a reaction to occur through a particular pathway.

C. Total energy of the products.

D. Energy absorbed by the surroundings only.

4. Which of the following collisions is effective?

A. A collision with insufficient energy.

B. A collision with unsuitable orientation.

C. A collision with sufficient energy and correct orientation.

D. A collision in which particles do not come into contact.

5. Increasing the concentration of a reactant generally increases reaction rate because:

A. Particles become larger.

B. The number of particles per unit volume increases.

C. The activation energy increases.

D. The reaction becomes irreversible.

6. Which factor increases reaction rate by increasing the kinetic energy of particles?

A. Decreasing pressure.

B. Decreasing surface area.

C. Increasing temperature.

D. Removing a catalyst.

7. Powdered calcium carbonate reacts faster with hydrochloric acid than large marble chips because powdered calcium carbonate has:

A. Lower concentration.

B. Greater surface area.

C. Lower temperature.

D. Higher activation energy.

8. Which of the following substances is commonly used as a catalyst in the decomposition of hydrogen peroxide?

A. Sodium chloride.

B. Manganese(IV) oxide.

C. Calcium carbonate.

D. Sodium hydroxide.

9. A catalyst increases the rate of a reaction by:

A. Increasing the enthalpy change.

B. Increasing the equilibrium constant.

C. Providing an alternative pathway with lower activation energy.

D. Increasing the mass of the reactants.

10. Which of the following is a photochemical reaction?

A. Rusting of iron.

B. Decomposition of silver chloride in sunlight.

C. Neutralisation of an acid.

D. Dissolving sodium chloride in water.

11. Which of the following methods can be used to measure the rate of a reaction that produces a gas?

A. Measuring the colour of the container.

B. Measuring the volume of gas produced per unit time.

C. Measuring the mass of the thermometer.

D. Measuring the length of the delivery tube.

Question 12. The unit of reaction rate when concentration is measured in mol dm^{-3} and time in seconds is:

A. mol dm^{-3} s^{-1}.

B. mol dm^{-3}.

C. mol s^{-1}.

D. dm^3 s^{-1}.

13. A graph of volume of gas produced against time becomes horizontal when:

A. The reaction rate is increasing.

B. The reaction has produced no gas.

C. No further gas is being produced.

D. The temperature becomes zero.

14. Which of the following is a reversible reaction?

A. Burning of magnesium.

B. Formation of ammonia in the Haber process.

C. Complete combustion of methane.

D. Precipitation of silver chloride.

Question 15. The symbol used to represent a reversible reaction is:

A. $\rightarrow$

B. $\leftarrow$

C. $\rightleftharpoons$

D. $=$

16. Chemical equilibrium is established when:

A. Reactants are completely used up.

B. Products are completely converted into reactants.

C. Forward and reverse reaction rates are equal.

D. Both reactions stop completely.

17. At dynamic equilibrium:

A. The forward reaction stops.

B. The reverse reaction stops.

C. Both reactions continue at equal rates.

D. All reactants have disappeared.

18. Which condition is necessary for establishing chemical equilibrium?

A. An open system.

B. A closed system.

C. Continuous removal of all products.

D. Continuous addition of reactants.

19. Le Chatelier's principle explains:

A. The effect of pressure, concentration and temperature on equilibrium.

B. The atomic structure of metals.

C. The preparation of insoluble salts only.

D. The movement of electrons in wires.

20. Increasing the concentration of a reactant in a reversible reaction generally shifts equilibrium:

A. Towards the reactants.

B. Towards the products.

C. In neither direction.

D. Towards the catalyst.

21. Increasing pressure in a gaseous equilibrium favours the side with:

A. More gaseous moles.

B. Fewer gaseous moles.

C. More solid particles.

D. More liquid particles.

22. For an exothermic forward reaction, increasing temperature shifts equilibrium:

A. To the right.

B. To the left.

C. Towards the catalyst.

D. Without any change in reaction direction.

23. A catalyst added to a reversible reaction at equilibrium:

A. Increases the equilibrium yield.

B. Shifts equilibrium towards products.

C. Makes equilibrium occur faster without changing its position.

D. Stops the reverse reaction.

24. Which catalyst is used in the Haber process?

A. Nickel.

B. Iron.

C. Platinum.

D. Vanadium(V) oxide.

25. The main product of the Haber process is:

A. Nitric acid.

B. Sulphuric acid.

C. Ammonia.

D. Methanol.

26. The main catalyst used in the Contact process is:

A. Iron.

B. Nickel.

C. Vanadium(V) oxide.

D. Manganese(IV) oxide.

27. The main product of the Ostwald process is:

A. Ammonia.

B. Nitric acid.

C. Sulphuric acid.

D. Hydrochloric acid.

28. The catalyst used in the oxidation of ammonia during the Ostwald process is:

A. Platinum-rhodium gauze.

B. Iron.

C. Copper.

D. Zinc oxide.

29. Which of the following is a major use of ammonia?

A. Manufacture of nitrogenous fertilisers.

B. Manufacture of table salt only.

C. Production of pure water only.

D. Manufacture of glass only.

30. Why is unreacted nitrogen and hydrogen recycled in the Haber process?

A. To increase wastage.

B. To reduce the amount of ammonia formed.

C. To improve raw material utilisation.

D. To reduce the pressure to zero.

SECTION B: STRUCTURED QUESTIONS

31: INTRODUCTION TO REACTION RATES

a) Define the term rate of chemical reaction.

b) Give five examples of chemical reactions that occur at different rates in everyday life.

c) State five reasons why the study of reaction rates is important in industry.

d) Explain the difference between a fast reaction and a slow reaction.

e) State four factors that affect the rate of chemical reactions.

32: COLLISION THEORY

a) Define collision theory.

b) State six principles of collision theory.

c) Distinguish between effective and ineffective collisions.

d) State three conditions necessary for an effective collision.

e) Explain why not all collisions between reactant particles lead to a chemical reaction.

f) Describe how increasing temperature affects the frequency of effective collisions.

33: ACTIVATION ENERGY

a) Define activation energy.

b) Explain the meaning of an activated complex.

c) State the importance of activation energy in chemical reactions.

d) Explain why a reaction with a lower activation energy may proceed faster than one with a higher activation energy at the same temperature.

e) Draw and label an energy profile diagram for an exothermic reaction.

f) Draw and label an energy profile diagram for an endothermic reaction.

34: FACTORS AFFECTING REACTION RATES

a) Explain how concentration affects the rate of a chemical reaction.

b) Explain how temperature affects reaction rate using collision theory.

c) Explain how surface area affects the rate of reaction between calcium carbonate and hydrochloric acid.

d) Explain the effect of pressure on the rate of reactions involving gaseous reactants.

e) Explain how catalysts increase reaction rates.

f) Explain how light affects the rate of photochemical reactions.

g) State five examples of catalysts and their applications.

35: EXPERIMENTAL INVESTIGATION OF REACTION RATES

a) Describe an experiment to investigate the effect of concentration on the rate of reaction between magnesium and hydrochloric acid.

b) State the apparatus and materials required.

c) Identify four variables that should be controlled.

d) State three observations that may be made.

e) Explain why the volume of hydrogen produced should be recorded at regular time intervals.

f) State three precautions necessary during the experiment.

36: TEMPERATURE AND SURFACE AREA

a) Describe an experiment to investigate the effect of temperature on the reaction between sodium thiosulphate and hydrochloric acid.

b) Write the balanced chemical equation for the reaction.

c) Explain why the cross disappears faster at higher temperatures.

d) Describe an experiment to investigate the effect of surface area on the reaction between calcium carbonate and hydrochloric acid.

e) State the expected observations.

f) Explain the results using collision theory.

37: CATALYSTS

a) Define a catalyst.

b) State five characteristics of catalysts.

c) Explain how a catalyst lowers the activation energy of a reaction.

d) Differentiate between homogeneous and heterogeneous catalysts.

e) Give four examples of catalysts and their industrial or laboratory applications.

f) Explain why a catalyst does not change the equilibrium position of a reversible reaction.

38: MEASUREMENT OF REACTION RATES

a) State four methods used to measure the rate of chemical reactions.

b) Explain how the volume of gas produced can be used to determine reaction rate.

c) Describe how loss in mass can be used to monitor a reaction.

d) Explain how a precipitate can be used to monitor the rate of reaction.

e) A reaction produces 180 cm^3 of gas in 60 seconds. Calculate its average rate of gas production.

f) The concentration of a reactant decreases from 0.90 mol dm^{-3} to 0.30 mol dm^{-3} in 40 seconds. Calculate its average rate of disappearance.

39: GRAPHS OF REACTION RATES

a) Draw a graph showing the volume of gas produced against time during a chemical reaction.

b) Explain why the gradient is steep at the beginning of the reaction.

c) Explain why the gradient gradually decreases.

d) State what a horizontal section of the graph indicates.

e) Explain how the gradient of a graph can be used to compare reaction rates.

f) Distinguish between average rate and instantaneous rate of reaction.

40: REVERSIBLE AND IRREVERSIBLE REACTIONS

a) Define a reversible reaction.

b) Define an irreversible reaction.

c) State five differences between reversible and irreversible reactions.

d) Write the chemical equation for the formation of ammonia.

e) Give four examples of irreversible reactions.

f) Explain why reversible reactions are important in industrial chemical processes.

41: EXPERIMENTS ON REVERSIBLE CHANGES

a) Describe an experiment to demonstrate the reversible reaction involving hydrated copper(II) sulphate.

b) State the colour of hydrated copper(II) sulphate.

c) State the colour of anhydrous copper(II) sulphate.

d) Write the chemical equation for the reversible change.

e) Describe the effect of heating and cooling zinc oxide.

f) Explain the reversible changes that occur when iodine is heated and then cooled.

42: CHEMICAL EQUILIBRIUM

a) Define chemical equilibrium.

b) Explain the meaning of dynamic equilibrium.

c) State six characteristics of chemical equilibrium.

d) State four conditions necessary for establishing chemical equilibrium.

e) Explain why the concentrations of reactants and products remain constant at equilibrium.

f) Distinguish between static and dynamic equilibrium.

43: LE CHATELIER'S PRINCIPLE

a) State Le Chatelier's principle.

b) Explain the meaning of equilibrium shift.

c) State how an equilibrium system responds when the concentration of a reactant is increased.

d) Explain the effect of removing a product from an equilibrium system.

e) State how an increase in pressure affects gaseous equilibrium systems.

f) Explain the effect of increasing temperature on an exothermic reversible reaction.

g) Explain the effect of decreasing temperature on an endothermic reversible reaction.

44: EFFECT OF CONCENTRATION ON EQUILIBRIUM

Consider the reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

a) State the effect of increasing the concentration of nitrogen.

b) State the effect of increasing the concentration of hydrogen.

c) Explain the effect of increasing the concentration of ammonia.

d) Explain why removing ammonia from the reaction mixture increases ammonia production.

e) State two industrial applications of changing reactant concentration to improve product yield.

45: EFFECT OF PRESSURE ON EQUILIBRIUM

Consider the reaction:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

a) Calculate the total number of gaseous moles on the reactant side.

b) Calculate the total number of gaseous moles on the product side.

c) Explain the effect of increasing pressure on the equilibrium position.

d) Explain the effect of decreasing pressure.

e) State why pure solids and liquids are not included when counting gaseous moles.

f) Give two other examples of reversible reactions where pressure affects equilibrium position.

46: EFFECT OF TEMPERATURE ON EQUILIBRIUM

a) Differentiate between exothermic and endothermic reactions.

b) Explain the effect of increasing temperature on an exothermic forward reaction.

c) Explain the effect of decreasing temperature on an exothermic forward reaction.

d) Explain the effect of increasing temperature on an endothermic forward reaction.

e) State why industrial processes may use moderate temperatures instead of very low temperatures.

f) Explain how temperature affects the equilibrium yield and reaction rate differently.

47: EQUILIBRIUM CONSTANT

a) Define equilibrium constant.

b) Write the expression for K_c for the reaction:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

c) Write the expression for K_c for:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

f) Explain the meaning of a large equilibrium constant.

g) Explain the meaning of a small equilibrium constant.

h) State three factors that do not change the equilibrium constant at a fixed temperature.

48: HABER PROCESS

a) Define the Haber process.

b) State the raw materials used in the Haber process.

c) Write the balanced chemical equation for ammonia production.

d) State the typical industrial temperature and pressure used.

e) Name the catalyst used.

f) Explain why high pressure favours ammonia production.

g) Explain why a moderate temperature is used.

h) Explain why unreacted nitrogen and hydrogen are recycled.

i) State five uses of ammonia.

49: OSTWALD PROCESS

a) Define the Ostwald process.

b) State the raw materials used in the manufacture of nitric acid.

c) Write the chemical equation for the catalytic oxidation of ammonia.

d) Name the catalyst used.

e) Write the equation for the oxidation of nitric oxide to nitrogen dioxide.

f) Write the equation for the absorption of nitrogen dioxide in water.

g) State four uses of nitric acid.

h) Explain the importance of equilibrium in the Ostwald process.

50: CONTACT PROCESS

a) Define the Contact process.

b) State the raw materials used.

c) Write the equation for the production of sulphur dioxide.

d) Write the reversible equation for the oxidation of sulphur dioxide.

e) Name the catalyst used.

f) State the typical temperature used in the catalytic converter.

g) Explain the effect of pressure on the equilibrium position.

h) Explain why moderate temperature is used.

i) Explain why sulphur trioxide is absorbed in concentrated sulphuric acid rather than directly in water.

j) State five uses of sulphuric acid.

51: COMPARISON OF INDUSTRIAL PROCESSES

a) Differentiate between the Haber process and the Ostwald process.

b) State the main product of each of the three industrial processes studied.

c) Compare the catalysts used in the Haber, Ostwald and Contact processes.

d) State two similarities between the Haber process and the Contact process.

e) Explain the importance of equilibrium in industrial manufacturing.

f) State four factors that industrial chemists consider when selecting operating conditions.

52: INDUSTRIAL APPLICATIONS AND ENVIRONMENTAL CONSERVATION

a) Explain five ways in which chemical equilibrium is applied in industries.

b) State four environmental problems associated with industrial chemical processes.

c) Explain how sulphur dioxide emissions can affect the environment.

d) State four measures used to control industrial air pollution.

e) Explain the importance of catalysts in reducing industrial energy consumption.

f) Discuss the importance of nitrogenous fertilisers in Kenyan agriculture.

53: INTEGRATED APPLICATION

A chemical company manufactures ammonia using nitrogen and hydrogen.

The reaction is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

a) Identify the reactants and product.

b) State the mole ratio of nitrogen to hydrogen.

c) Calculate the total number of gaseous moles on the reactant side.

d) Calculate the total number of gaseous moles on the product side.

e) Explain the effect of increasing pressure on the equilibrium position.

f) Explain the effect of increasing temperature.

g) State the catalyst used.

h) Explain the role of the catalyst.

i) Explain why ammonia is removed from the reaction mixture.

j) State four uses of ammonia in Kenya.

54: EXPERIMENTAL DATA INTERPRETATION

The following data were obtained during an experiment investigating the reaction between magnesium and hydrochloric acid.

Time (seconds)	Volume of hydrogen (cm^3)
0	0
10	24
20	42
30	54
40	60
50	60

a) Plot a graph of volume of hydrogen against time.

b) Calculate the average rate of gas production during the first 20 seconds.

c) Calculate the average rate of gas production between 20 and 40 seconds.

d) State the time when gas production stops increasing.

e) Explain why the rate decreases with time.

f) State two ways of increasing the rate of the reaction.

55: COMPREHENSIVE APPLICATION OF THE SUB-STRAND

a) Explain collision theory.

b) Describe the factors affecting the rate of chemical reactions.

c) Explain how activation energy influences reaction rate.

d) Describe how reaction rates can be measured experimentally.

e) Differentiate between reversible and irreversible reactions.

f) Explain the meaning of dynamic equilibrium.

g) State and explain Le Chatelier's principle.

h) Discuss the effects of concentration, pressure and temperature on equilibrium position.

i) Explain the industrial applications of equilibrium in the Haber, Ostwald and Contact processes.

j) Discuss the importance of responsible handling of chemicals in industrial and laboratory settings.

Revision summary notes Reaction Rates

Introduction

- The rate of a chemical reaction is the time taken for a given mass/amount of products to be formed.
- The rate of a chemical reaction is also the time taken for a given mass/amount of reactant to be consumed /used up.
- Some reactions are too slow to be determined. e.g rusting ,decomposition of hydrogen peroxide and weathering.
- Some reactions are too fast and instantaneous e.g. neutralization of acid and bases/alkalis in aqueous solution and double decomposition/precipitation.
- Other reactions are explosive and very risky to carry out safely e.g. reaction of potassium with water and sodium with dilute acids.

The collision theory

- The collision theory is an application of the Kinetic theory of matter which assumes matter is made up of small/tiny/minute particles like ions atoms and molecules.

The collision theory proposes that;

- For a reaction to occur, reacting particles must collide.
- Not all collisions between reacting particles are successful in a reaction. Collisions that initiate a chemical reaction are called **successful / fruitful/ effective collisions**
- The speed at which particles collide is called collision **frequency.**
- The higher the collision frequency the higher the chances of successful / fruitful/ effective collisions to form products.
- The higher the chances of successful collisions, the faster the reaction. The average distance between solid particles from one another is too big for them to meet and collide successfully.
- Dissolving substances in a solvent ,make the solvent a medium for the reaction to take place.
- The solute particle distance is reduced as the particle ions are free to move in the solvent medium.
- Successful collisions take place if the particles colliding have the required energy and right orientation which increases their vibration and intensity of successful / fruitful/ effective collisions to form products.

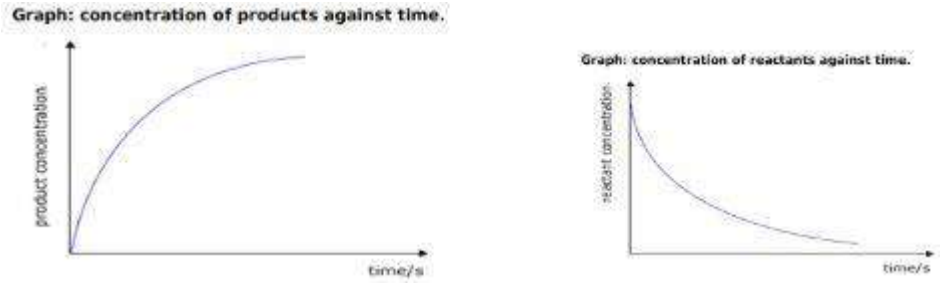
Measurements of reaction rates

Reaction rates are measured in terms of;

- Ø How much product appears in a given time.
- Ø How much reactant disappears in a given time

Rates curves are then plotted

These are plots of concentration or properties, which are functions of concentration against time



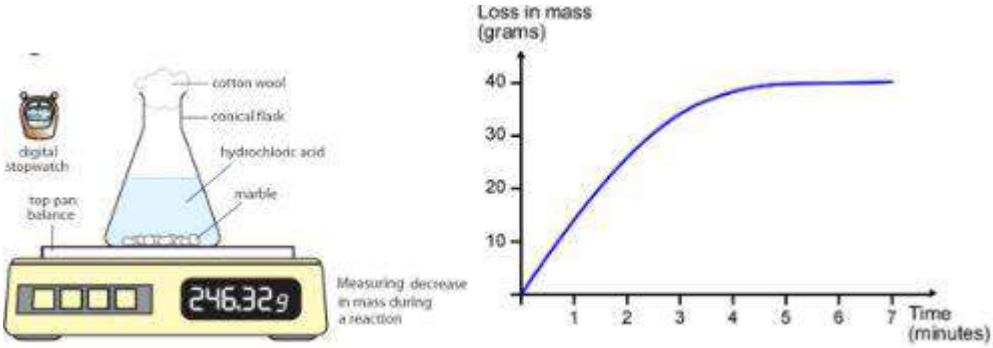
Apparatus used to determine rates of reaction

Examples:

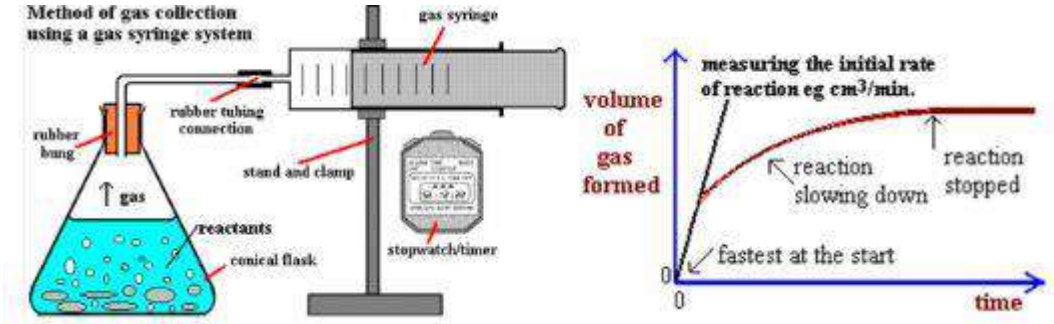
Reaction between dilute HCl and marble chips.

This can be by:

Measuring the decrease in mass with time

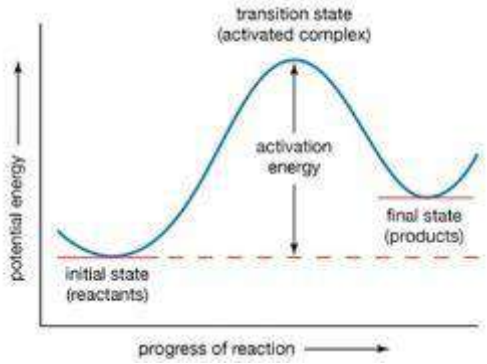


Measuring the volume of the gas produced with time and then plotting a graph



The Activation Energy(Ea) theory

- The Enthalpy of activation (ΔHa) /Activation Energy(Ea) is the minimum amount of energy which the reactants must overcome before they react.
- Activation Energy(Ea) is usually required /needed in bond breaking of the reacting particles.
- Bond breaking is an endothermic process that require an energy input.
- The higher the bond energy the slower the reaction to start of.
- Activation energy does not influence whether a reaction is exothermic or endothermic.



Factors determining the value of activation energy:

- Strength of the bonds in the particles of the reactants;
- Whether the reaction is exothermic or exothermic;
- Presence of catalysts

<u>Factors affecting the rate of reaction</u>

Factors that increase the rate of a reaction does so by:

- Increasing number of effective collisions;
- Lowering the activation energy; since a reaction with high activation energy is slow at ordinary conditions;

These factors are:

- Concentration of reactants
- Temperature
- Surface area (size of the particles
- Catalyst Light
- Pressure

(a). Concentration of reactants.

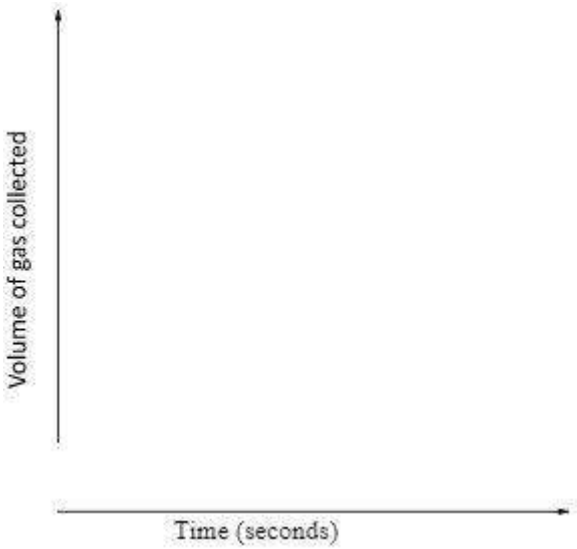
- Increase in concentration of one of the reactants increases the rate of the reaction;

Reason:

- Higher concentration results in a greater probability of collision hence a higher rate of reaction

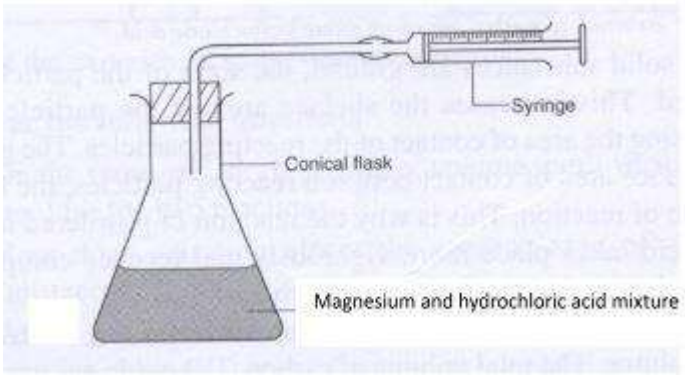
Example: reaction between Mg s) and 2M and 4M hydrochloric acid.

Graphical representation



- The maximum volume of a gas collected in both experiments is the same because the number of moles of HCl is the same
- Curve for 4M HCl is steeper indicating a faster rate because the acid is more concentration
- The 4M HCl contains more H^+ ions per unit volume than 2M HCl. This reaction takes a shorter time to go to completion
- The start of the flat part of the curve indicates the end of the reaction;

Diagram:



Note:

Rate of reaction can also be verified by measuring the rate of disappearance of reactants per unit time.

- Reactants with 4M HCl will disappear faster;
- Curve for a faster reaction rate is always to the left of the reference curve;

Worked example:

In an experiment to determine the effect of concentration on the rate of a reaction, various concentrations of sodium thiosulphate were reacted with equal volume of hydrochloric acid and the time taken for the precipitate to obscure a cross on paper put under the reaction beaker was determined.

The results are shown in the table below.

Volume of $S_2O_3^{2-}$ (cm^3)	Volume of H_2O, cm^3	Volume of HCl, cm^3	Time for cross to disappear, (s)
50	0	10	20
40	10	10	25
30	20	10	34
20	30	10	52
10	40	10	70

(a). Plot a graph of time (vertical axis) against volume of aqueous sodium thiosulphate. (3marks)

(b). From the graph estimate the time for the cross to disappear when 10 cm^3 of hydrochloric acid is added to a mixture of 35cm^3 of aqueous thiosulphate and 15cm^3 of water. (2 marks)

(c). What is the effect on the rate of the reaction of adding more water to the aqueous sodium thiosulphate? Explain your answer. (2 marks)

(d). On the same graph, plot the curves you would obtain if the experiment were repeated at:

(i). At 45°C; (1mark)

(ii). Using 10 cm^3 of less concentrated hydrochloric acid. (1mark)

Factors affecting rate of reaction; Temperature

Effect of temperature.

- An increase in temperature increases the rate of a chemical reaction.

Reasons:

- Increase in temperature increases the kinetic energy of the reacting particles;
- The particles therefore move faster leading to more frequent and more effective collisions.
- The rate of reaction therefore increases;

Example: reaction between sodium thiosulphate and dilute HCl(aq)

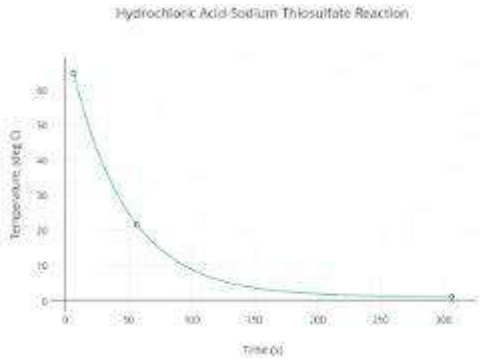
Stoichiometric equation:

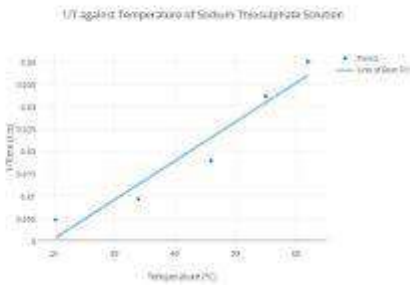
$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + S(s) + SO_2(g)$$

Ionic equation:

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l);$$

- The time taken to precipitate the same amount of sulphur by the same volume of acid at different temperatures is measured.
- Concentration of thiosulphate solution and the acid are kept constant.





The rate is proportionate to 1/ t;

Hence;

- The higher the temperature the shorter the time for completion of the reaction and the higher the rate of the reaction;

Alternatively:

You can plot a graph of 1/ t against temperature;

QUESTION

Reaction of sodium thiosulphate with dilute Hydrochloric acid

Procedure:

(i).

- Measure 20cm^3 of 0.05M sodium thiosulphate into a 50cm^3 glass beaker. Place the beaker on a white piece of filter paper with ink mark 'X' on it. Determine and record its temperature as room temperature in table 2 below. Measure 20cm^{3}of 0.1M Hydrochloric acid solution using a 50cm^3 measuring cylinder.
- Put the acid into the beaker containing sodium thiosulphate.
- Immediately start off the stop watch/clock.
- Determine the time taken for the ink mark 'X' to become invisible /obscured when viewed from above.
- Measure another 20cm^3 separate portion of the thiosulphate into a beaker, Heat the solution to 30°c.
- Add the acid into the beaker and repeat the procedure above. Complete table 2 below using different temperatures of the thiosulphate.

Sample results: Table 2

Temperature of $Na_2S_2O_3$	Room temperature	30	40	50	60
Time taken for mark X to be obscured /invisible (seconds)	50.0	40.0	20.0	15.0	10.0
Reciprocal of time($^1/_t$)	0.02	0.025	0.05	0.0667	0.1

Plot a graph of temperature(x-axis) against 1/t

(ii). Sodium thiosulphate solution reacts with dilute hydrochloric acid according to the following equation.

$S_2O_3^{2-}$ (aq) + 2H+ (aq) → H_2O (l) + SO_2 + S(s)

In an experiment to study how the rate of reaction varies with concentration, $10cm^3$ of 0.4M sodium thiosulphate was mixed with $10cm^3$. Of 2M hydrochloric acid in a flask. The flask was placed in a white paper marked with across X. The time taken for the cross X become invisible when viewed from above was noted and recorded in the table below. The experiment was repeated three times as the temperature using the volumes in

the table and the results recorded as shown in the table below.

Experiment	Volume of 0.4M thiosulphate (cm3)	Volume of water (cm3)	Volume of 2M HCl(cm3)	Time (Sec)
1	10	0	10	16
2	7.5	2.5	10	23
3	5.0	5.0	10	32
4	2.5	7.5	10	72

a.

(i) On the grid below, plot a graph of the volume of thiosulphate (Vertical axis) against time taken for the cross (X) to become invisible)

(ii) From the graph determine how long it would take for the cross to become invisible if the experiment was done.

- Using $6cm^3$ of the 0.4M thiosulphate
- Using $6cm^3$of 0.2M thiosulphate solution

b.

Using values for experiment I. Calculate;

(i) Moles of thiosulphate used

(ii) Moles of hydrochloric acid used

(iii) Explain which of the two reactants in experiment I controlled the rate of the reaction? Explain

c. Give two precautions which should be taken in experiment I controlled the rate of the reaction? Explain

Factors Affecting Rate of Reaction; Surface Area

Effect of particle size (surface area).

- A decrease in particle size (surface area) increases the rate of a chemical reaction; and an increase in particle size decreases the rate of a chemical reaction.

Reason:

- A smaller particle size means a larger total surface area and this offers a large surface on which the reacting particles can collide;

- This means more collisions, hence more chances of effective collisions leading to a higher rate of reaction;

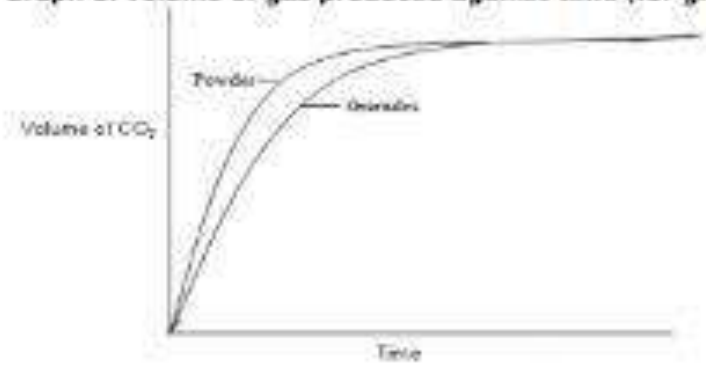
Example: Reaction between dilute HCl and marble chips

The larger $CaCO_3$ granules undergo a slower rate of reaction than the finely powdered $CaCO_3$

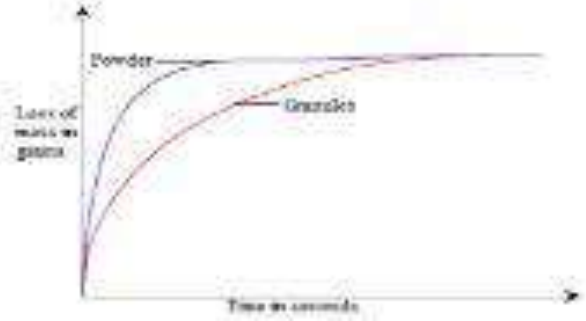
Equation:

$CaCO_3(S) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$

Graph 1: volume of gas produced against time (for granules and powder)



Graph 2: loss in mass of calcium carbonate against time



QUESTIONS

(1). Exactly 3.0g of powdered carbonate of metal M of formula MCO_3 were mixed with excess dilute hydrochloric acid. The mass of the reaction vessel and its contents were recorded at various times. From these readings, the total loss in mass of the reaction vessel and its contents was calculated and recorded as shown in the table below. The experiment was carried out at room temperature.

Time (secs)	0	30	60	90	120	150	180	210
Total loss in mass (g)	0	0.08	0.37	0.90	1.19	1.28	1.32	1.32

a.

(i). Plot a graph of total loss in mass against time. (1mark)

(ii). Determine the total loss in mass after 100seconds (2marks)

b.

(i). Write an equation for the reaction that occurs. (1mark)

(ii). Calculate the number of moles of:

- Carbon (IV) oxide gas produced. (2 marks)
- MCO_3 used. (2 marks)

(iii). Calculate the relative atomic mass of metal M. (3 marks)

c.

On the same axis sketch a curve that you are likely to obtain if the experiment was repeated:

(i). 9. at 50°C. Explain. (2 marks)

(ii). Using MCO_3 granules instead of the powder. Explain.(2marks)

Factors Affecting Rate of Reaction; Catalyst

<u>Effect of catalysts</u>

- Catalysts are substances which alter the rate of a chemical reaction without being consumed, e.g. finely divided iron in the harber process;
- They alter the rate of a chemical reaction but remain unchanged at the end of the reaction;
- They can therefore be reused at the end of the reaction;

Modes of action of catalysts:

- Provision of a surface over which particles can react.

Note:

Due to this solid catalysts are more effective when used in powder form;

- Formation of short-lived intermediate compounds with products which then break up to give the products and the catalyst;
- By adsorption;

Note:

- Generally, a catalyst increases the rate of a reaction by providing a different pathway of lower activation energy.
- This means more collisions can overcome this energy barrier and result in a reaction.
- It lowers the activation energy of the reaction;
- Catalysts are also reaction-specific;

Some common catalysts:

Catalyst	Reaction catalyzed.
Manganese (IV) oxide	Dissociation of hydrogen peroxide to give oxygen gas and water;
Vanadium (V) oxide or platinum	Contact process for the conversion of sulphur (IV) oxide to sulphur (VI) oxide;
Platinum	Manufacture of nitric acid in the Oswald's process;
Finely divided iron or vanadium	Haber process in the manufacture of ammonia;
Nickel	Hydrogenation of oils to make fats; Hydrogenation of unsaturated hydrocarbons
Copper (II) oxide	Oxidation of ethanol to produce ethanal;
Titanium (IV) chloride	Ziegler method of polymerizing alkenes

QUESTION

(1). Hydrogen peroxide decomposes slowly to water and oxygen under normal conditions. When a little manganese (IV) oxide is added to the solution the rate of decomposition is enhanced. The results of an experiment on the decomposition of hydrogen peroxide in presence of manganese (IV) oxide are shown below. Study the results and answer the questions that follow.

Volume of oxygen (cm^3)	0	19	27	33	36	38	39	40	40
Time (seconds)	0	30	60	90	120	150	180	210	240

a. Plot a graph of volume of oxygen produced (vertical axis) against time. (3 marks)

b.What is the rate of reaction:

(i). During the first 30 seconds. (2 marks)

(ii). Between 30 seconds and 60 seconds. (2 marks)

c. Explain why:

(i). The slope of the curve is steeper at the start of the reaction. (2 marks)

(ii). The curve goes flat at the end of the reaction. (2 marks)

(d)

(i). Suggest the role of manganese (IV) oxide in the decomposition of hydrogen peroxide. (1mark)

(ii). Comment on change in mass of manganese (IV) oxide at the end of the reaction. Explain.(2 marks)

Factors Affecting Rate of Reaction; Light

Effect of light on the rate of reaction

- The effect of heating and illuminating substances is the same;
- In both cases the constituent particles absorb radiant energy leading to an increase in the number of particles with activation energy resulting in increased rate of reaction;
- Light energizes the particles involved in a reaction;
- This increases the chances of effective collisions per unit time thus increasing the rate of reaction.
- Light of higher frequencies give higher reaction rates e.g. UV Light;
- Examples of reactions affected by light are those involving halogens:

Examples of reactions affected by light.

- Reaction between $Cl_2(g)$ and $H_2(g)$ does not take place in the dark but is explosive in bright light;



- For the reaction between methane and bromine; decolourisation of bromine only occurs in presence of light;

$CH_4(g) + Br_2(l) \xrightarrow{Light} CH_3Br(g) + HBr(g)$

Experiment: Effect of light on the decomposition of silver bromide.

Procedure:

- About $20cm^3$ of 0.1M potassium bromide is put in a glass beaker.
- $5cm^3$ of 0.05M silver nitrate solution is added.
- The resulting pale yellow precipitate is divided into three portions in 3 separate test tubes.
- One of the test tubes in immediately placed in a dark cupboard; the second on a bench and the third is placed in a direct source of light e.g. sunlight.

Observations.

- Formation of a pale yellow precipitate of silver bromide when silver nitrate reacts with potassium bromide.

Equation:

$KBr\ (aq) + AgNO_3(aq) \rightarrow AgBr(s) + KNO_3(aq)$

(Pale yellow)

- **Test tube in light**: precipitate changes colour from pale yellow to grey.
- **Test tube on the bench**: slight change in colour from pale yellow to slight grey.
- **Test tube in dark cupboard**: no observable (noticeable) colour change in precipitate.

Explanation:

- Light decomposes silver bromide to metallic silver (hence the grey colour) and bromine.

Equation:

$2AgBr(aq) \xrightarrow{Light} 2Ag(s) + Br_2(g)$

(grey)

- No observable change in test tube placed in darkness due to lack of light;
- The degree of decomposition and hence change depends on the light intensity falling on the test tubes;

- The rate of decomposition of silver bromide increases with increase in light intensity.

Conclusion:

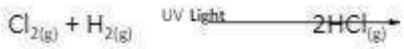
- Light affects the rate of some chemical reactions by energizing the particles involved in a reaction hence increasing the chances of effective collisions per unit time thus increasing the rate of reaction.

Application of effect of light on reaction rate.

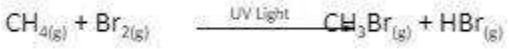
- Processing of black and white photographic films is done in dark to prevent the decomposition of the silver bromide that is usually used to coat photographic plates;

Further examples of reactions affected by light:

(i).



(ii).



(iii).



(f). Pressure

- Increase in pressure increases the rate of reaction involving the formation of small volume of product because of increase in concentration and slight increase in temperature
- In a given volume the higher the number of molecules of a given gas in a container, the greater the pressure;
- Increase in pressure causes the same effect as increase in concentration;
- Thus the rate of reactions involving gases can be increased by increasing the pressure of the gases

Note:

- From gas laws, decreases in volume results in increase in pressure. Thus, reactions accompanied by decrease in volume move faster to completion.

<u>Reversible reactions</u>

- Are reactions which can be made to go to either direction (forward or backward) by changing conditions such as temperatures, pressure etc.
- In such cases none of the reactants is completely used up and so the reaction does not go into completion.

Note:

Reversible reactions are of two types:

- Reversible physical changes e.g. heating ice, iodine etc.
- Reversible chemical changes e.g. heating hydrated copper (II) sulphate, Haber process, decomposition of limestone, heating blue cobalt chloride, ammonium chloride etc.

Examples:

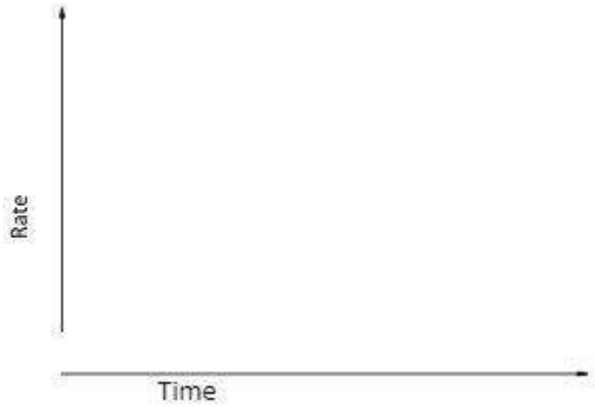
- Reaction of nitrogen and hydrogen to give ammonia in the **Haber process**.

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$

- The reversible sign means the reaction can reach a **state of equilibrium** if left undisturbed.

Explanations of reversible reactions:

Consider the curves below:



Point (a): forward reaction.

- At the start of the reaction, the rate of the forward reaction is faster.
- The concentration of the reactants is greatest at the beginning.
- The rate of forward reaction and the concentration of reactants decreases with time as the reaction proceeds.

Point (b): backward reaction.

- At zero time the rate of the backward reaction is zero.
- The concentration of the products is lowest at this time.
- The rate of the backward reaction and concentration of products increases as the forward reaction proceeds.

Point (c): equilibrium.

- It comes to a time (t) when the rate of forward reaction is equal to the rate of backward reaction.
- This balance is called equilibrium.

Equilibrium:

- Is the state at which the rate of forward reaction is equal to the rate of backward reaction.

Types of equilibrium:

Static equilibrium:

- Refers to a situation when two opposing forces balance each other and whatever was happening before comes to a standstill.

Dynamic equilibrium:

- Refers to a situation when two opposing processes, the forward and reverse, continue taking place but at the same rate.
- The equilibrium is said to be **dynamic** and this **state of balance** can be reached from either direction.

Examples:

$3Fe\ (s) + 4H_2O\ (g) \leftrightarrows Fe_3O_4\ (s) + 4H_2\ (g)$

The Equilibrium state here is established when:

- Either steam is passed over heated iron in a closed container; or
- When hydrogen is passed over heated iron oxide;

Characteristics of equilibrium systems.

(i) The concentrations of reactants and products do not change after the equilibrium has been reached unless the system is disturbed.

Note:

- Observable properties of reactants or products such as colour, mass, volume, PH, temperature etc can be used to detect whether a system is in equilibrium or not.

Examples:

Property	How it can be used to detect attainment of equilibrium
Colour	The colour intensity remains constant at equilibrium
Volume.	Total volume of solution remains constant at equilibrium;
Precipitate	The height of the precipitate remains constant;

(ii). The equilibrium can be reached from either direction.

(iii). All the reactants and products are present in the system.

Factors affecting the position of equilibrium

Such factors act as a strain on the state of equilibrium and the system reacts in a way to oppose the change.

The factors are:

- Temperature
- Pressure
- Concentration
- Catalysts.

(a). Temperature

Endothermic reactions

- Increase in temperature favours the endothermic reactions;

Example:

N_2O_4 (g) ⇆ $2NO_2$ (g) ; ΔH = + ve

Pale yellow ***dark brown***

- Increase in temperature shifts the equilibrium to the right; since the reaction is endothermic and hence more yield of nitrogen (IV) oxide; hence dark brown fumes will be observed;
- Decrease in temperature shifts the equilibrium to the left; since the reaction is endothermic favoured by low temperatures; hence formation of N_2O_4; and the pale yellow colour is observed;

Exothermic reactions:

- Decrease in temperature favours exothermic reactions.

Example:

$2SO_2(g) + O_2(g) ⇆ 2SO_3(g)$; ΔH = -ve

Thus:

- Decrease in temperature causes more yield of sulphur (VI) oxide; because the equilibrium shifts to the right; since the reaction is exothermic favoured by low temperatures;
- Increase in temperature causes SO_3 to decompose decreasing its yields. Increase in temperature shifts the equilibrium to the left; because the reaction is exothermic which is favoured by low temperatures;

QUESTIONS

(1). Given the equation's

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$, $\Delta H = -92KJMol^{-}$

What happens to the position of equilibrium when Temperature is raised

(2). Consider the following equilibrium reaction;

$2SO + O_2 \rightleftharpoons 2SO_2 \quad \Delta H = -Ve$

What will be the effect of increasing the temperature on the yield of sulphur (VI) oxide?

(3). NO_2 and N_2O_4 gases exist in equilibrium at 20°C as shown in the equation below;

$2NO_2(g) \rightleftharpoons N_2O_4(g)$ ΔH= -ve

Brown Pale yellow

State and explain the observation that would be made when;

(4). A syringe containing the mixture at 20°C is heated to 40°C

(5). Consider the reversible reaction below;

$H_2 + l_2(g) \rightleftharpoons 2Hl(g)$ $\Delta H = -53KJ$

Explain how the following factor would affect the concentration of Hydrogen iodide.

a. Increase in temperature.

Factors affecting Equilibrium; Pressure

Pressure

- Increase in pressure favours the side with fewer numbers of gaseous molecules since this is the side where pressure is reduced (low).

Example:

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$

- Increase in pressure favours the production of ammonia; by shifting the equilibrium position to the right; because the volume of gaseous reactants is higher than the volume of gaseous products.
- Decrease in pressure leads to the production of less ammonia (ammonia decomposes); by shifting the equilibrium position to the left; because the volume of gaseous reactants is higher than the volume of gaseous products.

$N_2O_4(g) \leftrightarrows 2NO_2(g)$

Pale yellow dark brown

Increase in pressure:

- Shifts the equilibrium to the left; leading to formation of more N_2O_4 hence mixture turns pale yellow; because the volume of gaseous products (NO_2) is higher than the volume of gaseous reactants (N_2O_4).

Decrease in temperature

- Shifts the equilibrium position to the right; leading to formation of more NO_2 hence colour turns dark brown; because the volume of gaseous reactants (N_2O_4) is lower than the volume of gaseous products (NO_2);

Note:

- Change in pressure has no effect on the equilibrium mixture where gaseous molecules on the two sides are equal

Example:

$N_2(g) + O_2(g) \leftrightarrows 2NO(g)$

- Increase or decrease in pressure has no effect on position of equilibrium; because the volume of gaseous reactants is equal to volume of gaseous products;

Note:

- Generally reactions involving only solids and, or liquids are **not affected by pressure change because they are not compressible;**

Summary: effects of pressure on equilibrium:

Reaction	Effects of pressure change on position of equilibrium	
	Increase	Decrease
$N_2O_{4(g)}$ $2NO_{2(g)}$	Equilibrium shifts to the left; more N_2O_4 is formed;	Equilibrium shifts to the right; more NO_2 is formed;
$N_{2(g)} + 3H_{2(g)}$ $2NH_{3(g)}$	- More NH_3 is formed; equilibrium shifts to the right; since volume of gaseous reactants is higher than volume of gaseous products	- More N_2 and H_2 are formed; equilibrium shifts to the left; since volume of gaseous reactants is higher than volume of gaseous products
$2SO_{2(g)} + O_{2(g)}$ $2SO_{3(g)}$	- More SO_3 is formed; equilibrium shifts to the right (forward reaction is favoured); since volume of gaseous reactants is higher than volume of gaseous products	- Less SO_3 is formed (more SO_2 and O_2 are formed); equilibrium shifts to the left (backward reaction is favoured); since volume of gaseous reactants is higher than volume of gaseous products
$4NH_{3(g)} + 5O_{2(g)}$ $4NO_{(g)} + 6H_2O_{(g)}$	- More NH_3 and O_2 is formed; equilibrium shifts to the left; since volume of gaseous reactants is lower than volume of gaseous products	- Less NH_3 and O_2 is formed (more NO and H_2O are formed); equilibrium shifts to the right; since volume of gaseous reactants is lower than volume of gaseous products
$H_{2(g)} + Cl_{2(g)}$ $2HCl_{(g)}$	- No effect on the equilibrium; because volume of gaseous reactants and volume of gaseous products is the same;	- No effect on the equilibrium; because volume of gaseous reactants and volume of gaseous products is the same;

QUESTIONS

(1). Given the equation's

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$, $\Delta H = -92KJMol^{-1}$

What happens to the position of equilibrium when The pressure of they system is reduced

(2). NO_2 and N_2O_4 gases exist in equilibrium at 20°C as shown in the equation below;

$2NO_2(g) \leftrightarrows N_2O_4(g)$ ΔH= -ve

Brown Pale yellow

State and explain the observation that would be made when; The gaseous mixture in a syringe compressed

(3). Consider the reversible reaction below

$H_2(g) + l_2(g) \leftrightarrows 2Hl(g)$ ΔH = -53KJ

Explain how each of the following factors would affect the concentration of Hydrogen iodide.

a. Increase in pressure.

(4). Nitrogen reacts with hydrogen according to the equation below:-

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$ ΔH = -92KJ

How would the yield of ammonia be affected by increase in :-

a. Pressure

Factors Affecting Equilibrium; Concentration

Concentration

- Consider the equilibrium in bromine water system:

$Br_2(aq) + H_2O(l) \leftrightarrows OBr^-(aq) + Br^-(aq) + 2H^+(aq)$

Yellow orange ***Colourless***

Addition of sodium hydroxide to the equilibrium:

- When NaOH(aq) is added to the equilibrium, the concentration of H^+ decrease and the rate of forward reaction are favoured (equilibrium shifts to the right); the reaction of bromine with water increases and there is loss of colour of bromine water.

Explanation:

- Addition of sodium hydroxide provides hydroxyl ions into the equilibrium.
- The hydroxyl ions react with the hydrogen ions (on the right of equilibrium) to form water.

Equation:

$OH^-(aq) + H^+(aq) \leftrightarrows H_2O(l)$

- This process removes hydrogen ions from the equilibrium mixture.
- This shifts the equilibrium to the right hence formation of more products;
- This leads to a change in colour from yellow orange to colourless;

Addition of hydrochloric acid:

- Addition of HCl(aq) is added concentration of H^+ increases; more bromine is formed and the orange-yellow colour of bromine water becomes more intense;

Explanation:

- Addition of hydrochloric acid introduces more hydrogen ions into the equilibrium;
- The hydrogen ions react with the **colourless bromide** and **hypobromite** ions to form **yellow-orange** aqueous bromine;
- This shifts the equilibrium to the left hence the increase in the intensity of the yellow-orange colour of bromine water;

Further examples:

(i). Given the equilibrium:

$SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

Removal of sulphur (VI) oxide:

- Reducing the concentration of SO_3 by removing it causes more sulphur (IV) oxide to be converted to sulphur (IV) oxide.

Addition of either oxygen or sulphur (IV) oxide.

- Addition of either sulphur (IV) oxide or oxygen to the equilibrium shifts the equilibrium to the right; due to increase of concentration of products hence more yield of sulphur (VI oxide;

Addition of pyrogallic acid:

- Addition of **pyrogallic acid** into the equilibrium shifts the equilibrium to the left; since pyrogallic acid dissolves oxygen gas reducing concentration of reactants on the left hence less yield of sulphur (VI) oxide;

(ii). **Given the equilibrium:**

$2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow 2Cr_2O_7^{2-}(aq) + H_2O(l)$

Yellow ***Orange***

Addition of sodium hydroxide to the equilibrium:

- When NaOH(aq) is added to the equilibrium, the concentration of H^+ decrease and the rate of backward reaction are favoured (equilibrium shifts to the left); the reaction of dichromate ions with water to form chromate ions and hydrogen ions increases and there is change in colour to yellow.

Explanation:

- Addition of sodium hydroxide provides hydroxyl ions into the equilibrium.
- The hydroxyl ions react with the hydrogen ions (on the left of equilibrium) to form water.

Equation:

$OH^-(aq) + H^+(aq) \leftrightarrows H_2O(l)$

- This process removes hydrogen ions from the equilibrium mixture.
- This shifts the equilibrium to the left hence formation of more reactants (chromate and hydrogen ions);
- This leads to a change in colour from orange to yellow;

Addition of hydrochloric acid:

- Addition of HCl(aq) is added concentration of H^+ increases; more dichromate solution is formed and the orange colour of dichromate ions become more intense;

Explanation:

- Addition of hydrochloric acid introduces more hydrogen ions into the equilibrium;
- The hydrogen ions react with the **yellow chromate** solution to form **orange** dichromate solution (and water);
- This shifts the equilibrium to the right hence the increase in the intensity of the orange colour of dichromate solution;

Generally:

- A change in concentration disturbs the already established equilibrium by making the reaction rate in one direction faster
- The reaction then proceeds predominantly in that direction until equilibrium is re established

QUESTIONS

(1).

a. What conditions is necessary for an equilibrium to be established? (1 mk)

b. When calcium carbonate is heated, the equilibrium shown below is established

$CaCO_3(s) \leftrightarrows CaO(s) + CO_2\ (g)$

How would be the position of the equilibrium be affected if a small amount of dilute potassium hydroxide is added to the equilibrium mixture? Explain (2 mks)

(2). In a closed system, aqueous iron (III) chloride reacts with hydrogen sulphide gas as shown in the equation below

$2FeCL_3(aq) + H_2S(g) \leftrightarrows 2FeCl_2(aq) + 2HCl(aq) + S\ (s)$

State and explain the observation that would be made if dilute hydrochloric acid is added to the system at equilibrium (2mks)

Industrial applications of chemical equilibrium.

- The ability to change the position of an equilibrium by varying the conditions has been important in industrial processes as industrialists aim at obtaining maximum products at minimum cost and shortest time possible.
- Conditions required to obtain **greatest yield** of products at **minimum costs** and **shortest time** possible are called **optimum conditions**.
- Such optimum conditions are obtained by continuous removal of products hence reduction in its concentration or varying external factors like temperature and pressure.

Summary:

Optimum conditions for common industrial processes.

Condition	Optimum condition for:	
	➢ **Haber process** $N_{2(g)} + 3H_{2(g)} \quad 2NH_{3(g)}$; ΔH=-92KJMol-	➢ **The contact process** $2SO_{3(g)} + O_{2(g)} \quad 2SO_{3(g)}$; ΔH = -197Kjmol-
Temperature	- The reaction is **exothermic** hence favoured by low temperatures; which shift the equilibrium to the right hence more yield of ammonia gas; - However rate at which the $NH_{3(g)}$ will be produced would be too slow	- Forward reaction is exothermic thus increase in temperature favours backward reaction; shifts the equilibrium to the left hence less yield of $SO_{2(g)}$; - Low temperature will favour forward reaction thus shifts the

	and thus uneconomical; an optimum temperature of about 450°C is thus normally used;	equilibrium to the right; hence more production of $SO_{2(g)}$; However rate of yield will be too slow and hence optimum temperature for maximum yield are set at about 450°C;
Pressure	Increase in pressure favours the forward reaction; shifting the equilibrium to the right hence more yield of ammonia; because the volume of gaseous reactants is higher than the volume of gaseous products; - However the cost of producing and maintaining high pressures in a system is very high; - Thus for maximum yield the optimum pressure is 200 atmospheres;	- An increase in pressure will shift the equilibrium to the left leading to low yield of sulphur (IV) oxide; since the volume of gaseous reactants is lower than the volume of gaseous products; - Thus the optimum pressure used is atmospheric pressure, which gives a percentage conversion of $SO_{2(g)}$ to $SO_{3(g)}$ of about 96%.
Concentration	- Removal of ammonia gas shifts the equilibrium to the right hence more yield of ammonia; since this removal lowers concentration of $NH_{3(g)}$ - thus for maximum yield ammonia is removed as soon as it is produced; so that more $H_{2(g)}$ and $N_{2(g)}$ can continue reacting;	- Removal of sulphur (VI) oxide gas shifts the equilibrium to the right hence more yield of $SO_{3(g)}$; since this removal lowers its concentration; - thus for maximum yield $SO_{3(g)}$ is removed as soon as it is produced; so that more $SO_{2(g)}$ and $O_{2(g)}$ can continue reacting;
Catalyst	- Main catalyst used is finely divided iron; platinum is a better catalyst but is very expensive and easily poisoned by impurities, hence may increase cost of production;	- A catalyst of vanadium (V) oxide is used to increase the rate of reaction;

QUESTIONS

(1). Consider the equilibrium reaction below

$2SO_4\ (g) + O_2(g) \leftrightarrows 2SO_3\ (g)\ \Delta H = -ve$

Which of the following will increase the yield of sulphur (vi) oxide

- Addition of catalyst
- Increase in pressure
- Increase in temperature
- Doubling the volume of the system (1mk)

(2).

a. State the Le chatelier's principle (1mk)

b. Carbon (II) oxide gas reacts with steam according to the reaction;

$CO(g) + H_2O(g) \leftrightarrows H_2(g) + C0_2(g)$

What would be the effect of increasing the pressure of the system at equilibrium? Explain (2mks)

c. When the reaction in (ii) above was carried out at lower temperature, the yields of hydrogen and carbon (IV) oxide increased. What is the sign of ΔH for the reaction? Explain (2mks)

(3). Methanol is manufactured from carbon (IV) oxide and hydrogen gas according to the equation.

$CO_2(g) + 3H_2 \leftrightarrows CH_3OH(g) + H_2O(g)$

The reaction is carried out in the presence of a chromium catalyst at 700K and 30kpa. Under these conditions an equilibrium is reached when 2% of the carbon (IV) oxide is converted to methanol

a. How does the rate of the forward reaction compare with that of the reverse reaction when 2% of the carbon (IV) oxide is converted to methanol? (1mk)

b. Explain how each of the following would affect the yield of methanol

- Reduction in pressure (2 mks)
- Using a more efficient catalyst (2mks)

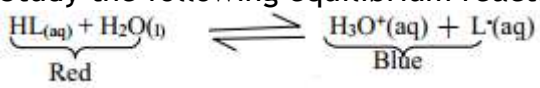
c. If the reaction is carried out at 500K and 30 kpa, the percentage of carbon (IV) oxide converted to methanol is higher than 2%

(i). What is the sign of Δ H for the reaction? Give a reason (2mks)

(ii). Explain why in practice the reaction is carried out at 700J but NOT at 500K (1 mk)

TOPICAL REVISION QUESTIONS

1. Study the following equilibrium reaction and answer the questions that follow:-

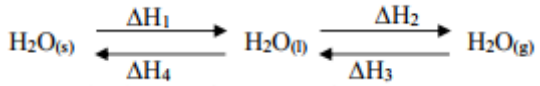


Given that in an acid solution, $H3O^{+}_{(aq)}$ act in place of hydrogen ions, H^{+}, according to the equation.

$H_3O^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow 2H_2O(l)$

Explain what would be observed when potassium hydroxide solution is added to the above equilibrium mixture

2. The scheme below shows the energy changes that take place between ice, water and steam.



Study it and answer the questions that follow:

   a. What name is given to the energy change ΔH_4?

   b. What is the sign ΔH_3, give a reason

3. The table below gives bond energies for three covalent bonds

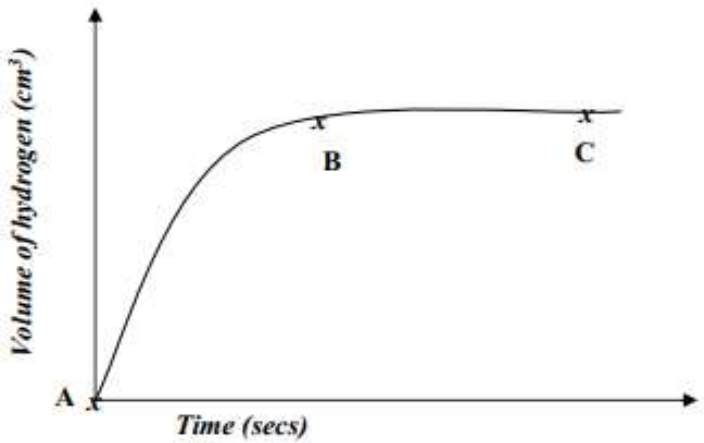
Bond	Bond energy (KJmol-1)
H -H	435
Cl - Cl	240
H - Cl	430

4.

   a. Calculate the energy change for the following reaction:
   $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

   b. Sketch an energy level diagram for the reaction in (a) above

5. The sketch below was obtained when 2g of magnesium was reacted with excess of 2M hydrochloric acid. The volume of hydrogen evolved was then plotted against time as shown below:



   a. On the same axis plot the graph that would be obtained if 1M hydrochloric acid was used instead of 2M hydrochloric acid. Explain

   b. Explain the significance of the flat portion **BC** of the curve

6. In a closed system an equilibrium exists between Nitrogen (IV) Oxide and dinitrogen tetraoxide as shown in the equation below:

$N_2O_{4\,(g)}$ ⇌ $2NO_{2\,(g)}$; $\Delta H = +27.5KJ$

Pale yellow — Reddish brown

a. State and explain the observation when a glass syringe containing the equilibrium mixture is immersed in ice-cold water

b. If the piston of the syringe is pushed, state the effect on the position of the equilibrium

7. The table below gives the volumes of the gas produced when different volumes of 2M hydrochloric acid were reacted with 1.0g of a lump of an alloy of Magnessium and copper at room temperature

Volume of 2M hydrochloric acid (cm^3)	Volume of gas (cm^3)
0	0
10	240
20	480
30	600
40	600
50	600

a. Write an equation for the reaction that occurred

b. On the grid provided below, plot a graph of the volume of the gas produced (vertical axis) against the volume of acid added (Note that before the reaction comes to completion, the volume of the gas produced is directly proportional to the volume of the acid added)

c. From the graph, determine:

i. The volume of the gas produced if 13.0cm^3 of 2M hydrochloric acid had been used

ii. The volume of 2M hydrochloric acid required for the reaction to go to completion

d. State and explain the effect on the rate of production of the gas if:

i. 1.0g of the lump of the alloy were replaced by 1.0g powder of the alloy

ii. The reaction was carried out at 35^oC.

8. In a series of experiments in which magnesium ribbon of uniform width reacted with 2.0M Hydrochloric acid, the rates of evolution of hydrogen gas were found to be as follows:-

Length of ribbon (cm)	1.0	2.0	3.0	4.0	5.0	6.0	7.0
Rate of Evolution of hydrogen (cm^3/min)	1.1	1.8	2.7	3.6	4.6	5.46	

a.

i. Draw a graph of rate of evolution of hydrogen gas against length of ribbon

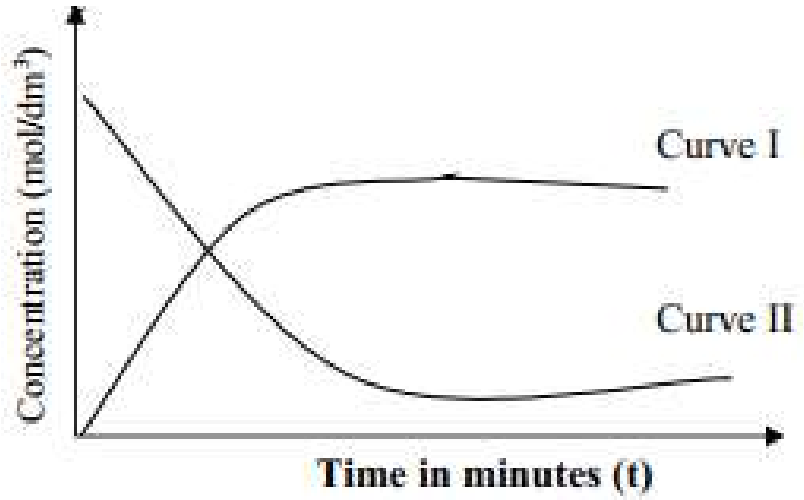
ii. What conclusion can you make from this graph? *

iii. Determine the rate of evolution of hydrogen gas from a piece of magnesium ribbon 12cm long under the same conditions *

iv. With dotted line, sketch on the same axis the graph that would be obtained if all the ribbons were ground into powder *

b. The curves below represent the changes in concentration of substances E and F with time in reaction

$E_{(g)} \rightleftharpoons F_{(g)}$



i. Which curve represents the change in the concentration of substance F? Give a reason

ii. Give one reason for the shapes of the curves after two minutes

9. A typical electrolysis cell uses a current of 40,000 amperes. Calculate the mass (in Kg of aluminium produced in one hour). (Al = 27) (Faraday = 96500 Coloumbs)

10. The table below shows the volumes of nitrogen (IV) oxide gas produced when different volumes of 1M nitric acid were each reacted with 0.635g of copper at room temperature.

a. Give a reason why hydrochloric acid can not be used instead of nitric acid

b. Explain how the rate of the reaction between copper and nitric acid would be affected if the temperature of the reaction mixture was raised

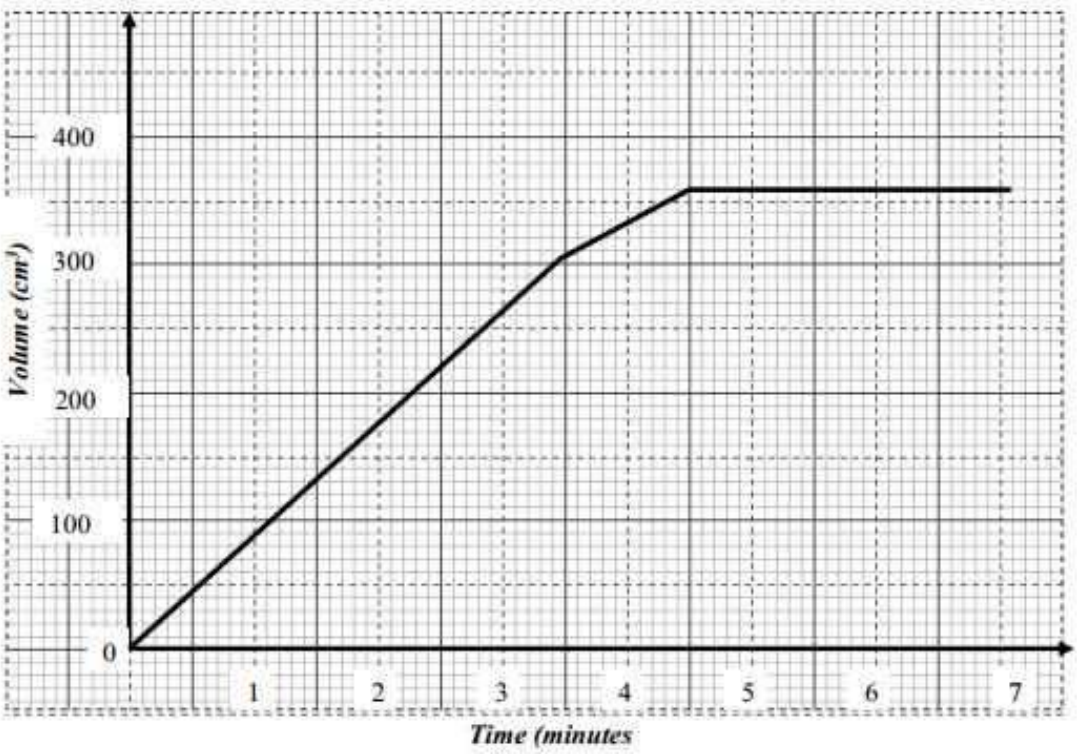
c. On the grid provided below, plot a graph of the volume of the gas produced (vertical axis) against volume of acid

d. Using the graph, determine the volume of:

i. Nitrogen (IV) oxide produced when 30cm^3 of 1M nitric acid were reacted with 0.635g of copper

ii. 1M nitric acid which would react completely with 0.635g of copper

11. The graph below represents the volume of gas collected against time when dilute sulphuric acid is reacted with Zinc granules:-



a. Determine the rate of reaction between the 1st and 3rd minute

b. When did the reaction stop?

c. Give a possible reason for the reaction to stop

12. The equation below represents a reaction that takes place in an industrial process
$4NH_{3(g)} + 5O_{2(g)} \rightarrow 6H_2O_{(g)} + 4NO_{(g)}$

a. Name the catalyst used

b. What are the other conditions for the reaction?

c. Why are the products cooled before being oxidised?

13. Nitrogen reacts with hydrogen according to the equation below:-
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ $\Delta H = -92KJ$

a. How would the yield of ammonia be affected by increase in :-

i. Pressure

ii. temperature

b. The ammonia produced is isolated form the other gases from time to time. How does this affect the equilibrium?

14. Explain how you would obtain an insoluble salt XSO_4 when you are provided with the following:-

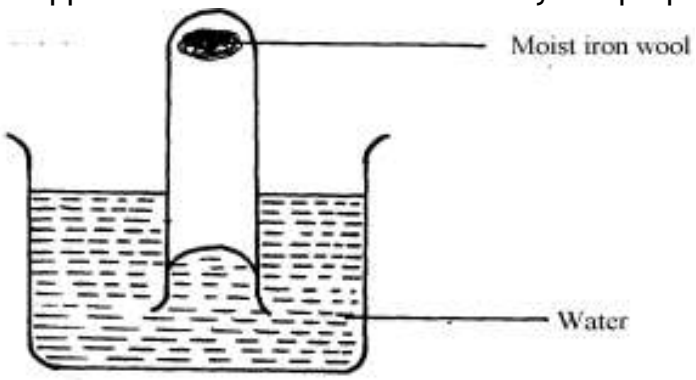
a. Distilled water

b. Solid YSO_4 which is soluble in water

c. Solid salt $X(NO_3)_2$

15. Metal R and S were used to form a cell. The following half equations show the standard electrode potentials of the half cells. (R and S are not actual symbols of the element)
$R^{2+}_{(aq)} + 2e^- \rightleftharpoons R_{(s)}$ $E^\theta = -2.04V$
$S^{2+}_{(aq)} + 2e- \rightleftharpoons S_{(s)}$ $E^\theta = -0.47V$
Write the full equation for the cell and calculate the e.m.f

16. The apparatus below were used to study the properties of air



a. State two observations made at the end of the experiment

b. Give one simple method that can be used to prevent rusting

17. Equal volumes of 1M monobasic acids K and L were each reacted with excess zinc granules. The table below shows the volumes of the gas produced after two minutes

Acid	Volume of gas (cm^3)
K	40
L	100

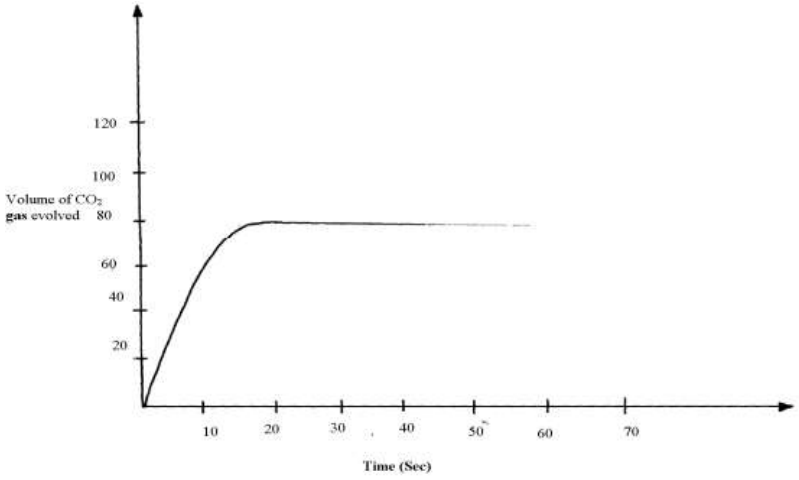
a. Explain the difference in the volumes of the gas produced

b. How can the production of the gas be increased?

18. The following is a thermochemical equation for the reaction between hydrogen and oxygen
$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H = -287kJmol\text{-}1$
Calculate the bond energy between the elements in water given that:
$O = O = +496kJmol^{-1}$ $H\text{-}H = +435kJmol^{-1}$

19. $AgClO_{2(s)} \rightleftharpoons Ag_{(s)} + \frac{1}{2}Cl_{2(g)} + O_{2(g)}$ $\Delta H = 0.00KJ/mol$
What is the effect on the position of equilibrium of the above system if temperature is decreased? Give a reason

20. Sodium carbonate was reacted with dilute sulphuric (VI) acid at 25°C. The volume of carbon (IV) Oxide gas liberated was recorded at 10seconds interval. Below is a graph of the volume of carbon (IV) Oxide gas evolved against time.

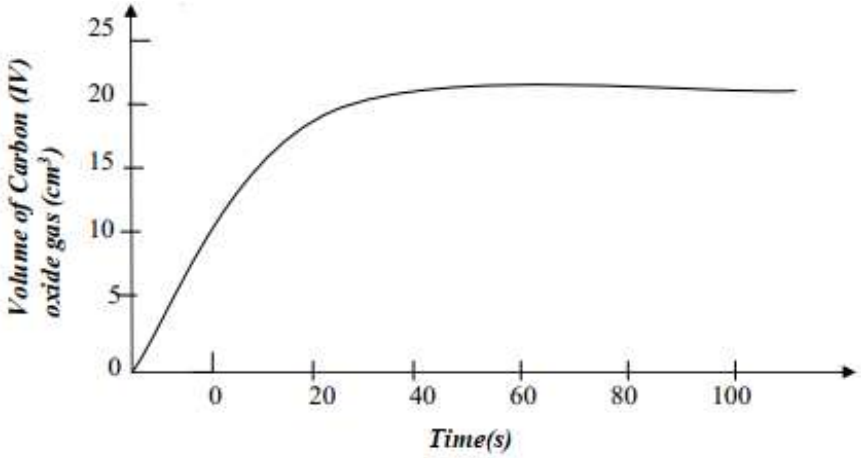


On the same axes, sketch the curve labelled V that would be obtained if Barium carbonate was used instead of sodium carbonate. (All conditions remain constant)

21.

a. What is meant by activation energy?

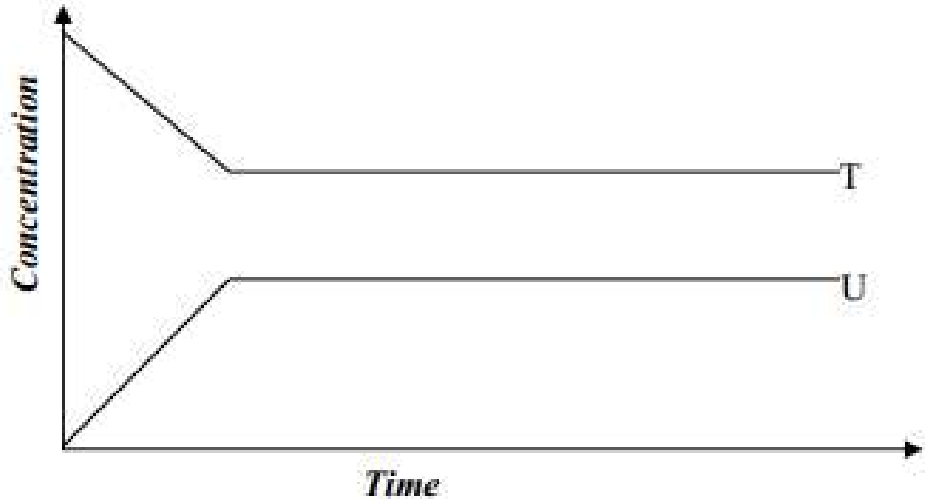
b. A certain mass of unground compound X1 reacted with excess dilute hydrochloric acid. The volume of carbon (IV) oxide gas liberated was measured after every 20 seconds. The results were presented as shown in the graph below:-



i. On the same axis, sketch the curve that would be obtained if the experiment was repeated using ground compound X1

ii. Explain the shape of your curve in (b) (i) above

22. The sketch below shows the rate at which substance T is converted into U. Study it and answer the questions that follows:-



When the equilibrium has been established the two curves become horizontal after sometime. Explain the effect of the amount of the two reactants and products

23. Elements A, B, C, and D are not actual symbols, have atomic numbers 19, 9, 12 and 10 respectively.

a. Which two elements represent non-metals

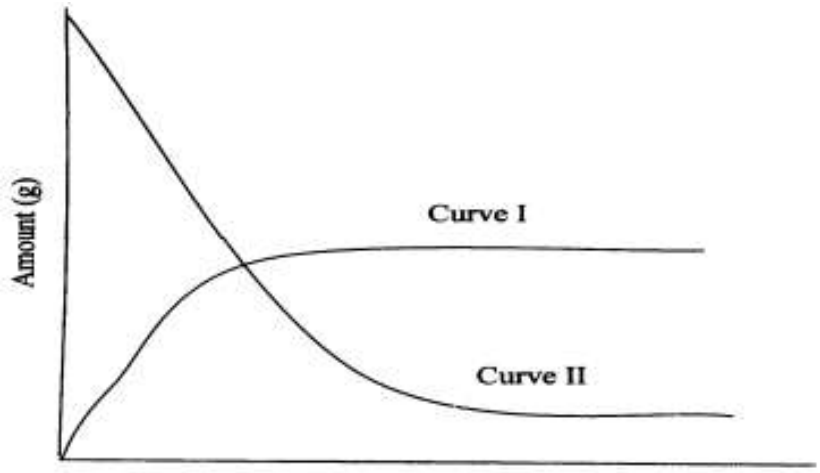
b. Write the formula of the compound formed between elements B and C and identity the bond present in the compound

24. An equilibrium is established between nitrogen tetra -oxide and nitrogen (IV) oxide as shown below: State and explain what happens when temperature is increased

$N_2O_{4(l)}$ ⇌ $2NO_{2(g)}$
Pale yellow Red-brown fumes

25. The graph below shows the amount of calcium carbonate and calcium chloride varying with time in the reactions:

$CaCO_{3\,(s)} + 2HCl_{\,(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{\,(g)} + CO_{2\,(g)}$



a. Which curve shows the amount of calcium chloride varying with time? (lmk)

b. Explain why the two curves become horizontal after a given period of time. (lmk)

c. Sketch on the graph how curve II would appear if the experiment was repeated using a more dilute hydrochloric acid solution (lmk)

26. An equilibrium is established between CrO4 and H+ ions as shown below:
$\underset{\text{(Yellow)}}{2CrO_{4(aq)}} + 2H^{+}_{(aq)}) \rightleftharpoons \underset{\text{(Orange)}}{Cr_2O_{7(aq)}} + H_2O_{(l)}$

State the effect on the equilibrium when;

a. Pressure is increased

b. Oxygen gas is added

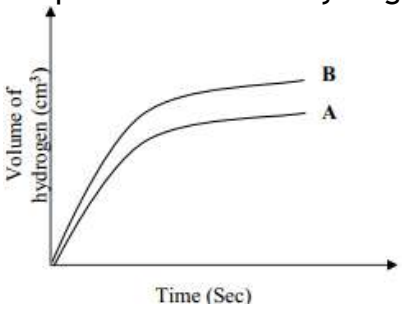
27. State and explain and explain the observation made when aqueous sodium hydroxide is added to the equilibrium mixture

28. Two experiments were carried out as follows and the volume of hydrogen gas evolved measured at intervals of 10seconds for 100seconds.

a. 8cm of magnesium ribbon was added to 1M hydrochloric acid

b. 8cm of magnesium ribbon was added to 0.5M hydrochloric acid

Graphs of volume of hydrogen evolved against time were plotted



c. Which of the graph was obtained for reaction (i) above? Explain

d. Explain the general shape of the graph

29. Bromine dissolves in water forming a brown solution, according to the dynamic equation below.
$Br_{2(aq)} + H_2O_{(l)} \rightleftharpoons 2H^{+}_{(aq)} + Br^{-}_{(aq)} + OBr^{-}_{(aq)}$
State and explain the observation that could be made if a solution of sodium hydroxide is added to the system

Answers

1. colour changes from red to blue
   H_3O^+ ions and $L^-_{(aq)}$ ions which form red solution.

2.

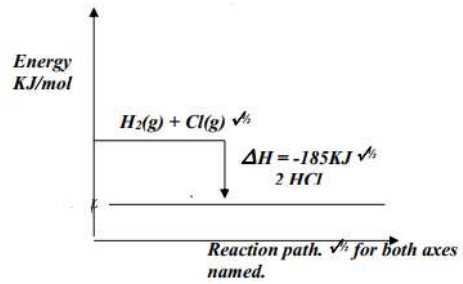
   a. ΔH_4 – latent heat of fusion

   b. ΔH_3 - is negative particles lose hat/process is exothermic/heat is given out (any)

3.

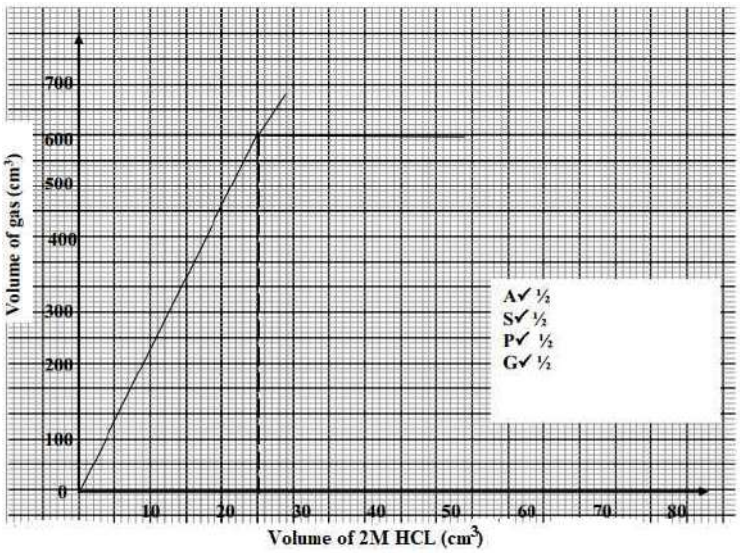
   a. $H\text{-}H_{(g)} + Cl - Cl_{(g)} \rightarrow 2H - Cl$
   Bonds broken bonds formed
   |H-H = 435 KJ 2H – Cl = 430 x 2
   |Cl – Cl = 240 KJ. = 860 KJ.
   Total = 675 KJ.
   ΔH_R = 860 + 675 ✓½
   = -185KJ ✓½

   b.



4.

   a.



   - Graph should be less steeper showing lower reaction rate since HCl is less concentrated. ✓½

   b. Graph flattens out at BC showing that all the magnesium has been used up, hence, no reaction is taking place ✓½ and there is therefore no evolution of hydrogen gas. ✓½
   The volume of the gas, therefore, remains constant. ✓½

5.

a. Pale yellow liquid is observed. ✓1 Backward reaction is favoured since ✓½ it is exothermic. Dinitrogen tetra oxide liquefies✓½ at very low temperature to pale yellow liquid.

b. Pressure increase, and favours backward reaction ✓½ which is at lower pressure; hence equilibrium shifts to the right. ✓½

6.

a. $Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$ ✓1

b.

c.

i. Showing on the graph. ✓½ Answer ✓½

ii. Showing on the graph. ✓½ Answer ✓½

d.

i. The rate of reaction increases. ✓1
The surface area of particles has been increased ✓½ thus increasing the area ✓½ of contact of the reacting particles.

ii. The rate of reaction increases. ✓1
Increase in temperature results in crease in the kinetic energy of the particles. This makes the particles move faster and collide more frequently with sufficient energy to cause more effective collision per given time. ✓1

7.

a.

i. Drawn on graph paper A = ½ mk
S = ½ mk
P = 1 mk
C = 1 mk

ii. Rate of evolution of hydrogen gas increases with increase in length of magnesium ribbon.

iii. Read from the student's graph. 1 mk – showing on graph
1 mk – for answer.

iv. Shown on the graph paper.

b.

i.

i. Curve I Reason: F increases as E decreases.

ii. Equilibrium is achieved.

8. Q = 40000 x 60 x 60 = 144000000c

Mass of Al = $\frac{144000000 \times 27}{3 \times 96500}$

= 13.43kg

9.

a. Hydrochloric acid is a weaker oxidizing agent which cannot oxidize copper to form Nitrogen (VI) Oxide gas

b. It increases ✓ 1 mark
Molecules/ particles acquire the necessary activation energy// Kinetic energy.
This increases the frequency of collisions hence the rate of reaction ✓ ½ mark

c. Graph
- Scale 1 mark with axis well labeled
- Plotting + all points correct 1 mark
- 5 correct points ½ mark
- Less than 5 points 0 mark
- Correct smooth curve 1 mark

TOTAL 3 marks

d.

i. $360cm^3$ Read correct value from graph + .05

ii. $40cm^3$ = Value from graph + .05
Read where it levels off

10.

a. $\frac{260 - 85}{2} = \frac{175}{2} = 87.5cm^3/mn$;

b. 4½ min;

c. Zinc was used up/H_2SO_4 used up;

11.

a. Platinum / Platinum Rhodium

b. Pressure -9atm (½ mk)
Temp - 700°C - 900°C (½ mk)

c. Reaction is exothermic

12.

a.

i. Will increase;

ii. Decrease;

b.

13.

a. Dissolve solid∫ ½ YSO_4 to obtain∫ ½ YSO_4 in solution,

b. Dissolve ∫ ½ $X(NO_3)_2$ in water to obtain ∫ ½ X $(NO_3)_2$ solution.

c. Mix the two above solutions

d. Filter to obtain XSO_4 solid residue, rinse with water and dry by heating∫ ½ under asbestos pad.

14. $R_{(s)} + S^{2+}_{(aq)} \rightarrow R^{2+}_{(aq)} + S_{(s)}$ ∫1 (1 mk)
E.m.f = 0.47 - (-2.04)
= -0.47 + 2.04
= 1.57V ∫1 (1 mk)

15.

a. – Water level rises. √1 (1 mk)
Grey Iron wool changes to brown. √1 (1 mk) 3 (any one of these scores)

b. – Oiling and greasing. √1 (1 mk)
- Painting.

16.

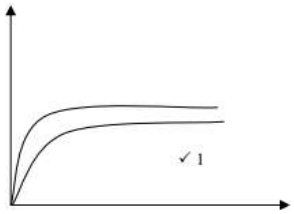
a. L is more ionized √1 than K hence reacts faster √1 producing higher volume of a gas. Or L is a stronger acid therefore ionized faster than K a weaker acid

b. Increasing the temeprature√1 using zinc powder/increasing the concentration of acid.

17. Energy of reaction = Bond breakage + Bond formation. √1
Bond formation = Energy of reaction – Bond Breakage
= -287 – 931 √1 3
= - 1218 K Joules per mole. √1 (3 mks)

18. - No effect on the position of the equilibrium
- RXM is neither endothermic nor exothermic hence not affected by changes in temperature enthalpy is zero.

19.

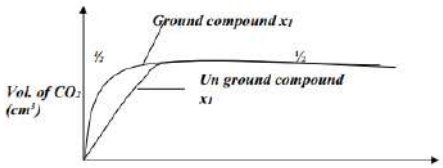


20.

a. the minimum energy required by the reaction particles to cause a successful collision to form product

b.

i.



NB. Sketch curve should be to the left and both flatten not at the same final volume

ii. curve is steep to the left due to the size of particles of ground compound x is reduced,√ ½ increasing surface area√ ½ of the particles thus increasing area of contact of

21. At equilibrium there will be very little of T that has reacted.√1At equilibrium there will be a lot of T and very little V produced hence equilibrium lies to the left or forms the reactants√1

22. - CB_2
- Ionic bond

23. - Intensity of red-brown fumes increases.
- High temperature vapourizes liquid nitrogen tetra-oxide to form nitrogen (IV) oxide that is red-brown.

24.

a. Curve 1

b. After sometime, the rate of formation of $CaCl_2$ or rate of depletion of $CaCO_3$ become to low that cant be evaluated

25.

a. Equilibrium shifts o the left, more CO_2 formed (Increase in pressure favors reaction producing fewer molecules)

b. Equilibrium shifts to the left, more $CO_{2(g)}$ formed

26. The solution turns yellow. Equilibrium shifts to the left when NaOH is added, the OH^- ions react with H^+ ions forcing more of $cr_2O_7^{2-}$ and H_2O to react forming more H^+ and crO_4^{2-} ions the reaction particles causing higher rate of reaction and twice shorter time✓ ½

27.

28.

a. B ; The acid had higher concentration (½mk)

b. The rate of reaction is initially high (½mk) because of high concentration of the reactant but decreases (½mk) steadily as the concentration also decreases.

29. Yellow/brown colour of bromine water ((½mk) fades or becomes colourless because sodium hydroxide solution provides OH^- ions which reacts with H^+ ions to form water (½mk) shifts the equilibrium to the right

STRAND 3.0: ORGANIC CHEMISTRY

SUB-STRAND 3.1: HYDROCARBONS

INTRODUCTION TO ORGANIC CHEMISTRY

Meaning of organic chemistry

Organic chemistry is the branch of chemistry that deals with the study of carbon compounds, particularly hydrocarbons and their derivatives.

- Most organic compounds contain carbon and hydrogen.
- Some also contain oxygen, nitrogen, sulphur, phosphorus and halogens.
- Organic compounds occur naturally in living organisms and can also be manufactured industrially.
- Carbon is central to organic chemistry because it forms stable covalent bonds with itself and many other elements.

Examples of organic compounds

1. Methane – CH_4
2. Ethane – C_2H_6
3. Ethanol – C_2H_5OH
4. Glucose – $C_6H_{12}O_6$
5. Ethanoic acid – CH_3COOH
6. Propane – C_3H_8

Why carbon forms many compounds

Carbon forms a very large number of compounds because of the following properties:

1. Tetravalency

- Carbon has four electrons in its outermost energy level.
- It requires four additional electrons to achieve a stable electronic arrangement.
- It therefore forms four covalent bonds with other atoms.
- These bonds may be single, double or triple bonds.

2. Catenation

- Catenation is the ability of an element to form bonds with atoms of the same element.
- Carbon atoms bond strongly with other carbon atoms.
- This enables carbon to form long chains, branched chains and rings.
- For example, propane contains a chain of three carbon atoms.

3. Formation of multiple bonds

- Carbon can form single, double and triple covalent bonds.
- A single bond involves one shared pair of electrons.
- A double bond involves two shared pairs of electrons.
- A triple bond involves three shared pairs of electrons.

4. Formation of stable covalent bonds

- Carbon forms strong covalent bonds with hydrogen, oxygen, nitrogen and other elements.
- These bonds make many organic compounds stable under ordinary conditions.

5. Isomerism

- Carbon compounds can have the same molecular formula but different structural arrangements.
- Such compounds are called isomers.
- This increases the number of possible organic compounds.

Sources of organic compounds

Organic compounds occur in nature and are also produced through industrial processes.

A. Natural sources

1. Plants

Plants produce many organic compounds through biological processes.

Examples include:

- Glucose and other carbohydrates produced during photosynthesis.
- Vegetable oils found in seeds such as sunflower, sesame and groundnuts.
- Cellulose found in plant cell walls.
- Natural rubber obtained from latex.
- Essential oils found in aromatic plants.

2. Animals

Animals contain organic compounds that form important components of their bodies.

Examples:

- Proteins in muscles, skin and other tissues.
- Fats stored in body tissues.
- Glucose in blood.
- Glycogen stored in the liver and muscles.
- Hormones and enzymes.

3. Crude oil (petroleum)

- Crude oil is a naturally occurring mixture of hydrocarbons found beneath the Earth's surface.
- It forms from the remains of ancient marine organisms that were buried under sediments and transformed over millions of years under heat and pressure.
- It is extracted through oil wells.
- It is refined to obtain useful fractions such as petrol, kerosene, diesel and lubricating oils.

4. Natural gas

- Natural gas is a naturally occurring mixture of gaseous hydrocarbons.
- Its major component is methane.
- It may also contain ethane, propane and butane.
- It occurs in underground reservoirs, sometimes together with crude oil.

5. Coal

- Coal is a carbon-rich fossil fuel formed from ancient plant remains.
- It contains a complex mixture of carbon-based substances.
- Heating coal in the absence of air produces useful substances such as coal gas, coke and coal tar.
- Coal tar can be used as a source of various organic chemicals.

B. Industrial sources

Organic compounds are manufactured in industries through chemical reactions and processing.

Examples include:

- Plastics manufactured from petroleum-derived monomers.
- Synthetic fibres such as nylon and polyester.
- Synthetic rubber used in tyres.
- Detergents and cleaning agents.
- Paints, adhesives and solvents.
- Pharmaceuticals and cosmetics.
- Ethanol produced through fermentation or industrial processes.

Kenyan context: Petroleum products used in transport, cooking and manufacturing are important in Kenya. Natural organic materials such as sugarcane, maize, tea, coffee, vegetable oils and animal products also provide sources of organic compounds.

HYDROCARBONS

Meaning of hydrocarbons

A **hydrocarbon** is an organic compound that contains only carbon and hydrogen atoms.

Hydrocarbons are the simplest organic compounds.

Examples:

Hydrocarbon	Molecular formula
Methane	CH_4
Ethane	C_2H_6
Propane	C_3H_8
Butane	C_4H_{10}
Ethene	C_2H_4
Propene	C_3H_6
Ethyne	C_2H_2

Classification of hydrocarbons

Hydrocarbons can be classified according to the nature of the bonds between carbon atoms.

The two main classes are:

1. Saturated hydrocarbons.
2. Unsaturated hydrocarbons.

Saturated hydrocarbons

Saturated hydrocarbons are hydrocarbons in which carbon atoms are joined together by single covalent bonds only.

They are mainly represented by the homologous series known as alkanes.

Examples:

- Methane – CH_4

- Ethane – C_2H_6
- Propane – C_3H_8
- Butane – C_4H_{10}
- Pentane – C_5H_{12}

Characteristics of saturated hydrocarbons

3. They contain carbon-carbon single bonds only.
4. They have the maximum possible number of hydrogen atoms for their carbon skeletons.
5. They are generally less reactive than alkenes and alkynes.
6. They undergo substitution reactions under suitable conditions.
7. They undergo combustion in oxygen.
8. They are commonly found in crude oil and natural gas.

Example: Ethane

Molecular formula: C_2H_6

Structural formula:

```
   H   H
   |   |
H - C - C - H
   |   |
   H   H
```

Each carbon atom forms four covalent bonds.

Unsaturated hydrocarbons

Unsaturated hydrocarbons are hydrocarbons that contain at least one carbon-carbon double bond or triple bond.

They are classified into:

- Alkenes – contain at least one carbon-carbon double bond.
- Alkynes – contain at least one carbon-carbon triple bond.

A. Alkenes

Alkenes are unsaturated hydrocarbons containing at least one carbon-carbon double bond.

Examples:

- Ethene – C_2H_4
- Propene – C_3H_6
- Butene – C_4H_8

Structural formula of ethene:

```
   H   H
    \ /
     C = C
    / \
   H   H
```

Characteristics of alkenes

1. They contain at least one carbon-carbon double bond.
2. They are unsaturated.

3. They generally undergo addition reactions.
4. They decolourise bromine water under appropriate test conditions.
5. They can undergo polymerisation.
6. They burn in oxygen to form carbon dioxide and water when combustion is complete.

B. Alkynes

Alkynes are unsaturated hydrocarbons containing at least one carbon-carbon triple bond.

Examples:

- Ethyne - C_2H_2
- Propyne - C_3H_4
- Butyne - C_4H_6

Structural formula of ethyne:

H - C ≡ C - H

Characteristics of alkynes

7. They contain at least one carbon-carbon triple bond.
8. They are unsaturated.
9. They undergo addition reactions.
10. They can decolourise bromine water under suitable conditions.
11. They undergo combustion in oxygen.
12. Ethyne burns in oxygen to produce a very hot flame used in oxy-acetylene welding.

3.5 Differences between saturated and unsaturated hydrocarbons

Saturated hydrocarbons	Unsaturated hydrocarbons
Contain carbon-carbon single bonds only	Contain at least one double or triple bond
Include alkanes	Include alkenes and alkynes
Generally less reactive	Generally more reactive in addition reactions
Do not normally decolourise bromine water without suitable reaction conditions	Decolourise bromine water under suitable conditions
Undergo substitution reactions	Commonly undergo addition reactions
Have fewer possibilities for addition reactions	Can undergo addition and polymerisation reactions

4.0 HOMOLOGOUS SERIES OF HYDROCARBONS

4.1 Meaning of a homologous series

A **homologous series** is a family of organic compounds that have:

- The same general molecular formula.
- Similar chemical properties.
- The same functional group or characteristic structural feature.
- Successive members differing by a CH_2 unit.
- Gradual changes in physical properties as molecular size increases.

4.2 Characteristics of a homologous series

13. Members have the same general formula.
14. Members have similar chemical properties.

15. Consecutive members differ by CH_2.
16. Members can be prepared using similar general methods.
17. Physical properties change gradually with increasing molecular size.
18. Members can be represented using a common structural pattern.

4.3 Main homologous series of hydrocarbons

Series	Characteristic bond	General formula
Alkanes	C–C	C_nH_{2n+2}
Alkenes	C=C	C_nH_{2n}
Alkynes	C≡C	C_nH_{2n-2}

These general formulae apply to open-chain hydrocarbons containing one characteristic unsaturated bond in the case of alkenes and alkynes.

4.4 The alkane homologous series

Alkanes are saturated hydrocarbons containing only carbon-carbon single bonds.

Their general formula is:

$$\boxed{C_nH_{2n+2}}$$

Where:

- C represents carbon.
- H represents hydrogen.
- n represents the number of carbon atoms.

First ten members of the alkane series

Number of carbon atoms	Name	Molecular formula
1	Methane	CH_4
2	Ethane	C_2H_6
3	Propane	C_3H_8
4	Butane	C_4H_{10}
5	Pentane	C_5H_{12}
6	Hexane	C_6H_{14}
7	Heptane	C_7H_{16}
8	Octane	C_8H_{18}
9	Nonane	C_9H_{20}
10	Decane	$C_{10}H_{22}$

Worked example 1: Determining the molecular formula of an alkane

Determine the molecular formula of an alkane containing 6 carbon atoms.

General formula:

$$C_nH_{2n+2}$$

Given:

$$n = 6$$

Substitute:

$$H = 2(6) + 2$$

$$H = 12 + 2 = 14$$

Therefore:

$$\boxed{C_6H_{14}}$$

The hydrocarbon is hexane.

Worked example 2

Determine the number of hydrogen atoms in an alkane containing 9 carbon atoms.

$$H = 2n + 2$$

$$H = 2(9) + 2 = 20$$

Therefore, the molecular formula is:

$$\boxed{C_9H_{20}}$$

4.5 The alkene homologous series

Alkenes are unsaturated hydrocarbons containing one carbon-carbon double bond in their open-chain structures.

Their general formula is:

$$\boxed{C_nH_{2n}}$$

The value of n must be at least 2 for an open-chain alkene.

First ten members of the simple alkene series

Number of carbon atoms	Name	Molecular formula
2	Ethene	C_2H_4
3	Propene	C_3H_6
4	Butene	C_4H_8
5	Pentene	C_5H_{10}
6	Hexene	C_6H_{12}
7	Heptene	C_7H_{14}
8	Octene	C_8H_{16}
9	Nonene	C_9H_{18}
10	Decene	$C_{10}H_{20}$

Worked example 3

Determine the molecular formula of an alkene containing 7 carbon atoms.

General formula:

$$C_nH_{2n}$$

$$n = 7$$

$$H = 2(7) = 14$$

Molecular formula:

$$\boxed{C_7H_{14}}$$

4.6 The alkyne homologous series

Alkynes are unsaturated hydrocarbons containing one carbon-carbon triple bond in their open-chain structures.

Their general formula is:

$$\boxed{C_nH_{2n-2}}$$

The value of n must be at least 2.

First ten members of the simple alkyne series

Number of carbon atoms	Name	Molecular formula
2	Ethyne	C_2H_2
3	Propyne	C_3H_4
4	Butyne	C_4H_6
5	Pentyne	C_5H_8
6	Hexyne	C_6H_{10}
7	Heptyne	C_7H_{12}
8	Octyne	C_8H_{14}
9	Nonyne	C_9H_{16}
10	Decyne	$C_{10}H_{18}$

Worked example 4

Determine the molecular formula of an alkyne containing 5 carbon atoms.

General formula:

$$C_nH_{2n-2}$$

$$H = 2(5) - 2$$

$$H = 10 - 2 = 8$$

Therefore:

$$\boxed{C_5H_8}$$

4.7 Comparing the three homologous series

Property	Alkanes	Alkenes	Alkynes
Type	Saturated	Unsaturated	Unsaturated
Characteristic bond	Single	Double	Triple
General formula	C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}

Property	Alkanes	Alkenes	Alkynes
First member	Methane	Ethene	Ethyne
First molecular formula	CH_4	C_2H_4	C_2H_2
Main reaction type	Substitution	Addition	Addition

5.0 NOMENCLATURE OF HYDROCARBONS

5.1 Meaning of nomenclature

Nomenclature is the systematic method of naming chemical compounds according to internationally accepted rules.

Organic compounds are named using the rules developed by the International Union of Pure and Applied Chemistry (IUPAC).

IUPAC naming helps chemists to identify compounds accurately regardless of the country or language in which they are studied.

5.2 Naming hydrocarbons according to the number of carbon atoms

The first part of the name of a simple hydrocarbon indicates the number of carbon atoms in its longest continuous carbon chain.

Number of carbon atoms	Prefix
1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-
9	Non-
10	Dec-

5.3 Naming according to the type of bond

The suffix indicates the type of carbon-carbon bond.

Type of hydrocarbon	Suffix	Example
Alkane	-ane	Ethane
Alkene	-ene	Ethene
Alkyne	-yne	Ethyne

Examples:

- CH_4 - methane.

- C_2H_6 – ethane.
- C_2H_4 – ethene.
- C_2H_2 – ethyne.
- C_3H_8 – propane.
- C_3H_6 – propene.
- C_3H_4 – propyne.

5.4 Rules for naming straight-chain hydrocarbons

Rule 1: Identify the longest continuous carbon chain.

- Count the number of carbon atoms in the longest continuous chain.
- Select the appropriate prefix.
- The longest chain forms the parent name.

Rule 2: Identify the type of carbon-carbon bond.

- Single bond: use the suffix -ane.
- Double bond: use the suffix -ene.
- Triple bond: use the suffix -yne.

Rule 3: Number the carbon chain where necessary.

- For alkenes and alkynes, number the chain from the end nearest the multiple bond.
- The position of the multiple bond should be given the lowest possible number.

Rule 4: Identify and number branches.

- Identify any side chains attached to the parent chain.
- Give each branch its appropriate number.
- Common branches include methyl and ethyl groups.

Rule 5: Write the complete name.

- Use numbers to show the positions of branches and multiple bonds.
- Separate numbers from words using hyphens.
- Use commas to separate numbers.

5.5 Naming simple alkanes

Alkanes contain only single bonds.

Examples:

Molecular for-mula	Structural arrange-ment	IUPAC name
CH_4	One carbon	Methane
C_2H_6	Two-carbon chain	Ethane
C_3H_8	Three-carbon chain	Propane
C_4H_{10}	Four-carbon chain	Butane
C_5H_{12}	Five-carbon chain	Pentane

5.6 Naming alkenes

The position of the double bond is indicated by a number.

For example:

But-1-ene

Structural formula:

$$CH_2 = CH - CH_2 - CH_3$$

- The longest chain contains four carbon atoms.
- The prefix is but-.
- The double bond begins at carbon 1.
- The suffix is -ene.

Name: **but-1-ene**.

But-2-ene

$$CH_3 - CH = CH - CH_3$$

- The longest chain contains four carbon atoms.
- The double bond begins at carbon 2.
- The name is but-2-ene.

5.7 Naming alkynes

The position of the triple bond is indicated by a number.

But-1-yne

$$HC \equiv C - CH_2 - CH_3$$

Name: **but-1-yne**.

But-2-yne

$$CH_3 - C \equiv C - CH_3$$

Name: **but-2-yne**.

5.8 Naming branched-chain hydrocarbons

A branched-chain hydrocarbon has one or more carbon-containing side chains attached to the main carbon chain.

Common branches include:

Branch	Formula
Methyl	$-CH_3$
Ethyl	$-CH_2CH_3$
Propyl	$-CH_2CH_2CH_3$

Example 1: 2-methylpropane

Structural formula:

```
    CH3
     |
CH3 — CH — CH3
```

Explanation:

19. The longest continuous chain contains three carbon atoms.
20. The parent hydrocarbon is propane.
21. There is a methyl branch attached to carbon 2.
22. The name is 2-methylpropane.

Example 2: 2-methylbutane

```
     CH3
      |
CH3 — CH — CH2 — CH3
```

Explanation:

- The longest chain contains four carbon atoms.
- The parent name is butane.
- A methyl branch is attached to carbon 2.
- The correct name is 2-methylbutane.

Important naming principles

- Select the longest continuous carbon chain.
- Number the chain to give substituents the lowest possible locants.
- If there is a multiple bond, it takes priority in numbering over ordinary alkyl branches.
- Identical substituents are indicated using prefixes such as di-, tri- and tetra-.
- List different substituents alphabetically, ignoring multiplicative prefixes such as di- and tri-.

6.0 ISOMERISM IN HYDROCARBONS

6.1 Meaning of isomerism

Isomerism is the phenomenon in which two or more compounds have the same molecular formula but different arrangements of atoms.

The compounds are called **isomers**.

Isomers have the same number and types of atoms but differ in the way the atoms are arranged.

6.2 Importance of studying isomerism

23. It explains why compounds with the same molecular formula may have different properties.
24. It helps in identifying and naming organic compounds.
25. It explains the large number of organic compounds.
26. It is useful in petroleum refining and the manufacture of fuels.
27. It is important in the production of medicines, plastics and other industrial chemicals.

6.3 Types of isomerism studied in hydrocarbons

The main types considered at this level include:

28. Structural isomerism.
29. Positional isomerism.

6.4 Structural isomerism

Structural isomerism occurs when compounds have the same molecular formula but different arrangements of the carbon skeleton or connectivity of atoms.

For example, butane and 2-methylpropane both have the molecular formula:

$$C_4H_{10}$$

However, their carbon arrangements differ.

Isomer 1: Butane

$$CH_3 - CH_2 - CH_2 - CH_3$$

Isomer 2: 2-methylpropane

```
     CH3
      |
CH3 — CH — CH3
```

Both compounds have four carbon atoms and ten hydrogen atoms, but their structures differ.

6.5 Structural isomers of pentane

Pentane has the molecular formula:

$$C_5H_{12}$$

It has three structural isomers.

1. Pentane

$$CH_3 - CH_2 - CH_2 - CH_2 - CH_3$$

2. 2-methylbutane

$$CH_3 - CH(CH_3) - CH_2 - CH_3$$

3. 2,2-dimethylpropane

```
      CH3
       |
CH3 — C — CH3
       |
      CH3
```

These compounds have the same molecular formula but different carbon skeletons.

6.6 Positional isomerism

Positional isomerism occurs when compounds have the same molecular formula and the same carbon skeleton, but the position of a multiple bond or substituent differs.

For example, butene has the molecular formula:

$$C_4H_8$$

It has two positional isomers.

1. But-1-ene

$$CH_2 = CH - CH_2 - CH_3$$

2. But-2-ene

$$CH_3 - CH = CH - CH_3$$

The two compounds have the same molecular formula and carbon skeleton, but the double bond is in a different position.

6.7 Positional isomerism in alkynes

Butyne has the molecular formula:

$$C_4H_6$$

It has two positional isomers.

But-1-yne

$$HC \equiv C - CH_2 - CH_3$$

But-2-yne

$$CH_3 - C \equiv C - CH_3$$

6.8 Differences between structural and positional isomerism

Structural isomerism	Positional isomerism
Carbon skeleton or connectivity differs	Position of a multiple bond or substituent differs
May involve straight and branched chains	Usually involves the same carbon skeleton
Example: butane and 2-methylpropane	Example: but-1-ene and but-2-ene
Same molecular formula	Same molecular formula

6.9 Practical activity: Modelling isomers

Materials

- Molecular modelling kit.
- Coloured balls representing carbon and hydrogen.
- Connecting rods representing covalent bonds.

Procedure

30. Construct a model of butane.
31. Rearrange the carbon atoms to construct 2-methylpropane.
32. Compare the two structures.
33. Construct but-1-ene and but-2-ene.
34. Identify the differences in the positions of their double bonds.

Expected observations

- Butane and 2-methylpropane have different carbon skeletons.
- But-1-ene and but-2-ene have the same carbon skeleton but different double-bond positions.

Conclusion: Isomers have the same molecular formula but differ in the arrangement of their atoms.

7.0 FRACTIONAL DISTILLATION OF CRUDE OIL

7.1 Meaning of crude oil

Crude oil is a naturally occurring dark, oily liquid consisting mainly of a mixture of hydrocarbons.

It is a fossil fuel formed from the remains of ancient organisms over millions of years.

Crude oil is not a pure substance. It contains many different hydrocarbons with different molecular sizes and boiling points.

7.2 Composition of crude oil

Crude oil contains:

- Alkanes.
- Cycloalkanes.
- Aromatic hydrocarbons.
- Small amounts of sulphur-containing compounds.
- Nitrogen-containing compounds.
- Oxygen-containing compounds.
- Traces of metals and other impurities.

The exact composition varies according to the oil field from which the crude oil is obtained.

7.3 Meaning of fractional distillation

Fractional distillation is the process of separating a mixture of liquids into different fractions based on differences in their boiling points.

Crude oil is separated into useful fractions by fractional distillation in an oil refinery.

7.4 Principle of fractional distillation

Fractional distillation is based on the fact that different hydrocarbons have different boiling points.

- Hydrocarbons with low boiling points vaporise easily and rise higher in the fractionating column.
- Hydrocarbons with high boiling points condense lower down the column.
- The fractions are collected at different levels according to their boiling ranges.

7.5 Apparatus used in fractional distillation

The main equipment includes:

35. Furnace or heating unit.
36. Fractionating column.
37. Trays or packing material.
38. Condensers.
39. Pipes for collecting different fractions.
40. Storage tanks.

7.6 Process of fractional distillation of crude oil

41. Crude oil is first cleaned to remove water, salts and other impurities.
42. It is heated in a furnace to a high temperature.
43. Most of the hydrocarbons vaporise.
44. The hot vapour mixture enters the fractionating column near its lower section.
45. The column is hot at the bottom and cooler towards the top.
46. Vapours rise through the column.
47. As the vapours rise, they cool.
48. Different hydrocarbons condense at different levels depending on their boiling points.
49. The condensed liquids are collected through outlets at different levels.
50. Some heavy residue remains at the bottom of the column.

7.7 Main fractions of crude oil and their uses

Fraction	Approximate carbon range	Main uses

Fraction	Approximate carbon range	Main uses
Refinery gases	C1–C4	Cooking and heating fuels; chemical manufacture
Petrol (gasoline)	C5–C10	Fuel for petrol engines
Naphtha	Approximately C5–C12	Petrochemical feedstock and chemical manufacture
Kerosene	Approximately C10–C16	Jet fuel, heating and lighting
Diesel oil	Approximately C14–C20	Fuel for diesel engines
Lubricating oil	Higher hydrocarbons	Lubricating machinery and engines
Fuel oil	Heavy hydrocarbons	Industrial heating and marine fuel
Bitumen/residue	Very large molecules	Road surfacing and roofing

Note: Carbon-number ranges and boiling ranges overlap and vary with the source of crude oil and refinery operating conditions. Fractions are mixtures, not pure compounds.

7.8 Properties of fractions obtained from crude oil

As the size of hydrocarbon molecules increases:

51. Boiling points generally increase.
52. Viscosity generally increases.
53. Volatility generally decreases.
54. Flammability generally decreases.
55. The fractions become less easily ignited.
56. Density generally increases, although this is not a universal rule for every pair of fractions.

7.9 Uses of crude oil fractions in Kenya

1. Refinery gases

- Used as fuel gases.
- Liquefied petroleum gas (LPG) is used in domestic cooking.
- Used in some industrial heating processes.

2. Petrol

- Used in cars, motorcycles and other petrol-powered engines.
- Used in small generators and machinery.

3. Kerosene

- Used in some domestic stoves and lamps.
- Used as aviation turbine fuel after appropriate refining and specification.

4. Diesel

- Used in buses, lorries and tractors.
- Used in generators and some industrial equipment.

5. Lubricating oils

- Used to reduce friction between moving machine parts.
- Used in motor vehicle engines and industrial machines.

6. Bitumen

- Used in road construction.
- Used in roofing and waterproofing materials.

7.10 Cracking of petroleum fractions

Cracking is the process of breaking large hydrocarbon molecules into smaller, more useful molecules through heating, sometimes in the presence of a catalyst.

It is important because heavy fractions are often less useful as fuels than lighter fractions.

For example:

$$C_{10}H_{22} \rightarrow C_5H_{12} + C_5H_{10}$$

Decane produces pentane and pentene in this illustrative cracking equation.

Types of cracking

57. **Thermal cracking**
    - Uses high temperatures and often high pressure.
    - Breaks large hydrocarbon molecules into smaller molecules.
58. **Catalytic cracking**
    - Uses a catalyst, commonly a zeolite.
    - Allows cracking to occur under conditions different from thermal cracking.
    - Often produces branched and smaller hydrocarbons suitable for petrol production.

Importance of cracking

- Increases the supply of petrol-range hydrocarbons.
- Produces alkenes used in polymer manufacture.
- Makes better use of heavy petroleum fractions.
- Provides raw materials for the petrochemical industry.

8.0 BONDING IN HYDROCARBONS

8.1 Covalent bonding in hydrocarbons

Hydrocarbons contain covalent bonds between carbon atoms and hydrogen atoms.

A **covalent bond** is a chemical bond formed when atoms share one or more pairs of electrons.

Carbon has four valence electrons and normally forms four covalent bonds in simple hydrocarbons.

Hydrogen has one valence electron and forms one covalent bond.

8.2 Types of covalent bonds

A. Single covalent bond

A single covalent bond involves the sharing of one pair of electrons between two atoms.

It is represented by a single line:

$$C - C$$

Example: Ethane.

$$CH_3 - CH_3$$

B. Double covalent bond

A double covalent bond involves the sharing of two pairs of electrons between two atoms.

It is represented by:

$$C = C$$

Example: Ethene.

$$CH_2 = CH_2$$

C. Triple covalent bond

A triple covalent bond involves the sharing of three pairs of electrons between two atoms.

It is represented by:

$$C \equiv C$$

Example: Ethyne.

$$HC \equiv CH$$

8.3 Differences between single, double and triple bonds

Single bond	Double bond	Triple bond
One shared electron pair	Two shared electron pairs	Three shared electron pairs
Represented by -	Represented by =	Represented by ≡
Found in alkanes	Found in alkenes	Found in alkynes
Generally longest of the three	Shorter than a single bond	Generally shortest of the three

8.4 Shapes of simple hydrocarbon molecules

- Methane has a tetrahedral shape.
- Ethene has a trigonal planar arrangement around each carbon atom.
- Ethyne has a linear arrangement around the carbon atoms.

These shapes result from the arrangement of electron pairs around the carbon atoms.

9.0 METHODS OF PREPARING HYDROCARBONS

Hydrocarbons can be prepared in the laboratory and obtained industrially.

At this level, learners study laboratory preparation of methane, ethene and ethyne.

9.1 Laboratory preparation of methane

Methane is the simplest alkane and has the molecular formula:

$$CH_4$$

It can be prepared in the laboratory by heating a mixture of sodium ethanoate and soda lime.

Soda lime is a mixture of sodium hydroxide and calcium oxide.

Chemical equation

$$CH_3COONa + NaOH \xrightarrow{\Delta} CH_4 + Na_2CO_3$$

Sodium ethanoate reacts with sodium hydroxide to produce methane and sodium carbonate.

Apparatus and materials

- Hard glass test tube.
- Delivery tube.
- Retort stand and clamp.
- Bunsen burner or suitable heating source.
- Gas jar.
- Sodium ethanoate.
- Soda lime.
- Appropriate safety equipment.

Procedure

59. Mix sodium ethanoate with soda lime.
60. Place the mixture in a hard glass test tube.
61. Fit the test tube with a delivery tube.
62. Heat the mixture carefully.
63. Allow the gas produced to pass through the delivery tube.
64. Collect the gas using an appropriate gas-collection method.

Observations

- A colourless gas is produced.
- The gas is methane.

Test for methane

- Methane burns in air with a blue flame when sufficient oxygen is available.
- It does not decolourise bromine water under ordinary test conditions in the absence of ultraviolet light.

Safety precautions

- Carry out the preparation in a well-ventilated laboratory.
- Keep the apparatus away from uncontrolled flames while collecting the gas.
- Check that the apparatus is properly fitted.
- Do not inhale the gas.
- Ensure the delivery tube is not blocked.
- Follow the teacher's instructions when heating glass apparatus.

9.2 Laboratory preparation of ethene

Ethene is the simplest alkene.

Molecular formula:

$$C_2H_4$$

Ethene can be prepared by dehydrating ethanol using concentrated sulphuric acid or by passing ethanol vapour over heated aluminium oxide.

Chemical equation

$$CH_3CH_2OH \xrightarrow[170°C]{\text{conc. } H_2SO_4} CH_2 = CH_2 + H_2O$$

Ethanol loses a molecule of water to form ethene.

Apparatus and materials

- Round-bottomed flask or suitable hard glass apparatus.
- Delivery tube.
- Heating source.
- Ethanol.
- Concentrated sulphuric acid.
- Gas jar.
- Retort stand and clamp.

Procedure

65. Place ethanol in the reaction flask.
66. Add concentrated sulphuric acid carefully under teacher supervision.
67. Heat the mixture under controlled conditions.
68. Ethene gas is produced and passed through the delivery tube.
69. Collect the gas using an appropriate method.

Observations

- A colourless gas is produced.
- The gas is ethene.

Tests for ethene

70. Pass the gas through bromine water.
71. The orange-brown colour of bromine water disappears.
72. This indicates the presence of unsaturation.

Safety precautions

- Concentrated sulphuric acid is corrosive.
- Ethanol is flammable.
- Heating must be controlled.
- The preparation should be carried out by a trained teacher or under close supervision.
- Avoid inhaling the gas.

9.3 Laboratory preparation of ethyne

Ethyne, also known as acetylene, is the simplest alkyne.

Molecular formula:

$$C_2H_2$$

It is prepared by reacting calcium carbide with water.

Chemical equation

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

Calcium carbide reacts with water to produce ethyne and calcium hydroxide.

Apparatus and materials

- Suitable reaction flask.
- Delivery tube.

- Gas jar.
- Calcium carbide.
- Water.
- Retort stand and clamp.

Procedure

73. Place a small quantity of calcium carbide in a suitable reaction flask.
74. Add water carefully.
75. Allow the gas produced to pass through the delivery tube.
76. Collect the gas using an appropriate method.
77. Carry out suitable tests under teacher supervision.

Observations

- Effervescence occurs.
- A colourless gas is produced.
- Calcium hydroxide is formed in the reaction vessel.

Tests for ethyne

- Ethyne decolourises bromine water under suitable conditions.
- It burns in air with a luminous, often sooty flame because of its relatively high carbon content.

Safety precautions

- Calcium carbide reacts with water to produce a flammable gas.
- Keep ethyne away from naked flames during preparation and collection.
- Use small quantities.
- Avoid inhalation.
- Commercial calcium carbide may contain impurities that produce unpleasant or harmful gases.
- The experiment must be performed in a well-ventilated laboratory under teacher supervision.

9.4 Summary of laboratory preparation methods

Hydrocarbon	Reactants	Main product
Methane	Sodium ethanoate and soda lime	CH_4
Ethene	Ethanol	C_2H_4
Ethyne	Calcium carbide and water	C_2H_2

10.0 PHYSICAL PROPERTIES OF HYDROCARBONS

Physical properties are characteristics that can be observed or measured without changing the chemical composition of a substance.

The physical properties of hydrocarbons include:

78. Colour.
79. Odour.
80. Density.
81. Melting point.
82. Boiling point.

83. Solubility.

10.1 Colour

- Most pure simple hydrocarbons are colourless.
- Methane, ethane, propane and butane are colourless gases.
- Many liquid hydrocarbons such as hexane are also colourless.
- Crude oil is usually dark brown or black because it contains a complex mixture of substances and impurities.

10.2 Odour

- Many simple pure hydrocarbons have little or no distinctive smell.
- Some petroleum products have characteristic odours.
- Commercial LPG is normally given an added odorant to make gas leaks easier to detect.
- The smell of commercial LPG is therefore not the natural smell of pure propane or butane.

10.3 Density

Density is the mass of a substance per unit volume.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

- Most hydrocarbons are less dense than water.
- Many liquid hydrocarbons form a separate layer on water.
- Hydrocarbon vapours have different densities depending on their molecular masses.

Example

A liquid hydrocarbon has a mass of 60 g and a volume of 100 cm^3. Calculate its density.

$$\text{Density} = \frac{60}{100}$$

10.4 Melting point

Melting point is the temperature at which a solid changes into a liquid at a specified pressure.

- Small hydrocarbons generally have low melting points.
- Melting points generally increase with molecular size, although the pattern is not perfectly regular.
- The shape and symmetry of molecules influence their melting points.

10.5 Boiling point

Boiling point is the temperature at which a liquid changes into a gas throughout the liquid at a specified pressure.

- Small hydrocarbons generally have low boiling points.
- Boiling points generally increase as the number of carbon atoms increases.
- Larger molecules have stronger intermolecular attractions.
- More energy is therefore required to separate the molecules.

Example

Methane has a much lower boiling point than decane because methane molecules are much smaller and have weaker intermolecular attractions.

10.6 Solubility

- Hydrocarbons are generally insoluble or only very slightly soluble in water.
- They are soluble in many non-polar organic solvents.
- This is because hydrocarbons are non-polar substances while water is polar.
- Hydrocarbons do not normally mix uniformly with water.

Practical example

When petrol or kerosene is accidentally spilled on water, it tends to form a separate layer rather than dissolving.

10.7 Trends in physical properties

As the carbon chain length increases:

Property	General trend
Boiling point	Increases
Viscosity	Increases
Volatility	Decreases
Flammability	Generally decreases
Physical state	Changes from gases to liquids and eventually solids

Explanation of the trends

- Larger hydrocarbon molecules have more electrons.
- Their electron clouds are more easily distorted.
- This increases London dispersion forces between molecules.
- More energy is required to separate larger molecules.
- Therefore, boiling points generally increase with molecular size.

11.0 CHEMICAL PROPERTIES OF ALKANES

Alkanes are relatively unreactive compared with alkenes and alkynes because they contain strong carbon-carbon and carbon-hydrogen single bonds.

Their main chemical reactions include:

84. Combustion.
85. Substitution reactions.
86. Cracking.

11.1 Combustion of alkanes

Combustion is the chemical reaction in which a substance reacts with oxygen, releasing energy.

Alkanes are commonly used as fuels because they release heat energy when burned.

There are two main types of combustion:

- Complete combustion.
- Incomplete combustion.

A. Complete combustion

Complete combustion occurs when a hydrocarbon burns in sufficient oxygen.

The products are carbon dioxide and water.

General equation

$$\text{Hydrocarbon} + O_2 \rightarrow CO_2 + H_2O$$

Example 1: Complete combustion of methane

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Example 2: Complete combustion of propane

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Observations

- A blue flame may be observed when combustion is efficient.
- Heat and light are released.
- Carbon dioxide and water are formed.

Importance

- Provides energy for cooking.
- Provides energy for heating.
- Powers engines and generators.
- Supports industrial heating processes.

B. Incomplete combustion

Incomplete combustion occurs when a hydrocarbon burns in insufficient oxygen.

Possible products include:

- Carbon monoxide.
- Carbon (soot).
- Water.

Example 1: Formation of carbon monoxide

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O$$

Example 2: Formation of carbon

$$CH_4 + O_2 \rightarrow C + 2H_2O$$

Dangers of incomplete combustion

87. Carbon monoxide is poisonous.
88. Soot contributes to air pollution.
89. Incomplete combustion wastes fuel.
90. Soot can cause respiratory problems.
91. Carbon monoxide reduces the ability of blood to transport oxygen.

Safety measures

- Ensure proper ventilation when using fuel-burning appliances.
- Maintain cooking stoves and heaters.
- Avoid using charcoal burners in enclosed rooms.
- Install appropriate carbon monoxide detectors where available.
- Ensure chimneys and exhaust systems are functioning properly.

11.2 Substitution reactions of alkanes

Substitution reaction is a chemical reaction in which an atom or group of atoms in a molecule is replaced by another atom or group.

Alkanes undergo substitution reactions with halogens such as chlorine in the presence of ultraviolet light or sunlight.

A. Substitution reaction of methane with chlorine

When methane reacts with chlorine in ultraviolet light, a hydrogen atom is replaced by a chlorine atom.

Chemical equation

$$CH_4 + Cl_2 \overset{UV}{\rightarrow} CH_3Cl + HCl$$

Products:

- Chloromethane.
- Hydrogen chloride.

Further substitution can occur:

$$CH_3Cl + Cl_2 \overset{UV}{\rightarrow} CH_2Cl_2 + HCl$$

$$CH_2Cl_2 + Cl_2 \overset{UV}{\rightarrow} CHCl_3 + HCl$$

$$CHCl_3 + Cl_2 \overset{UV}{\rightarrow} CCl_4 + HCl$$

These products are formed when successive hydrogen atoms are replaced by chlorine atoms.

B. Substitution reaction of ethane with chlorine

Ethane reacts with chlorine in ultraviolet light to form chloroethane and hydrogen chloride.

$$C_2H_6 + Cl_2 \overset{UV}{\rightarrow} C_2H_5Cl + HCl$$

Importance of substitution reactions

- They are used in the manufacture of halogenated organic compounds.
- Some halogenated compounds are used as industrial intermediates.
- They demonstrate the chemical behaviour of saturated hydrocarbons.

12.0 CHEMICAL PROPERTIES OF ALKENES

Alkenes are more reactive than alkanes because they contain a carbon-carbon double bond.

The double bond contains a sigma bond and a pi bond. The pi bond is relatively more easily broken during addition reactions.

The main reactions of alkenes include:

92. Combustion.
93. Addition reactions.
94. Polymerisation.
95. Oxidation.

12.1 Combustion of alkenes

Alkenes burn in oxygen to form carbon dioxide and water during complete combustion.

Example: Combustion of ethene

$$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$$

Observations

- Heat and light are released.
- A flame is produced.
- Carbon dioxide and water are formed during complete combustion.

Incomplete combustion may produce carbon monoxide or soot.

12.2 Addition reactions of alkenes

Addition reaction is a reaction in which atoms or groups of atoms are added across a carbon-carbon double bond to form a more saturated compound.

The double bond is converted into a single bond as new atoms or groups attach to the carbon atoms.

A. Addition of hydrogen (hydrogenation)

Hydrogen reacts with an alkene in the presence of a suitable catalyst to form an alkane.

Example: Ethene

$$CH_2 = CH_2 + H_2 \xrightarrow{Ni} CH_3 - CH_3$$

Ethene reacts with hydrogen to form ethane.

Conditions

- Nickel catalyst.
- Appropriate heating and reaction conditions.

Uses

- Hydrogenation is used in the manufacture of some processed fats and oils.
- It is also used in industrial chemical manufacture.

B. Addition of bromine

Alkenes react with bromine to form dibromoalkanes.

Example: Ethene

$$CH_2 = CH_2 + Br_2 \rightarrow CH_2Br - CH_2Br$$

The product is 1,2-dibromoethane.

Observation

- Orange-brown bromine water loses its colour.
- This is used as a test for unsaturation.

C. Addition of hydrogen halides

Hydrogen halides such as hydrogen chloride and hydrogen bromide add across the double bond of an alkene.

Example: Ethene and hydrogen chloride

$$CH_2 = CH_2 + HCl \rightarrow CH_3CH_2Cl$$

The product is chloroethane.

D. Addition of water (hydration)

Water can add across the double bond of an alkene under suitable conditions to form an alcohol.

Example: Ethene

$$CH_2 = CH_2 + H_2O \xrightarrow{\text{acid catalyst}} CH_3CH_2OH$$

The product is ethanol.

12.3 Polymerisation of alkenes

Polymerisation is the chemical process in which many small molecules called monomers join together to form a large molecule called a polymer.

Alkenes undergo addition polymerisation because their double bonds can open and form new carbon-carbon single bonds.

Example: Polymerisation of ethene

$$n(CH_2 = CH_2) \rightarrow [-CH_2 - CH_2 -]_n$$

The product is poly(ethene), commonly called polythene.

Uses of poly(ethene)

- Packaging bags.
- Water pipes.
- Plastic containers.
- Insulation of electrical cables.
- Packaging films.
- Household products.

Other examples of alkene polymerisation

Monomer	Polymer	Common uses
Ethene	Poly(ethene)	Bags, containers and pipes
Propene	Poly(propene)	Ropes, containers and fibres
Chloroethene (vinyl chloride)	Poly(chloroethene) (PVC)	Pipes, cable insulation and floor-ing

12.4 Oxidation of alkenes

Alkenes undergo oxidation reactions under suitable conditions.

A common school laboratory test involves acidified or dilute potassium manganate(VII), depending on the intended reaction conditions.

Test using cold, dilute alkaline potassium manganate(VII)

- Add the alkene to the potassium manganate(VII) solution.
- The purple colour disappears under suitable conditions.
- A diol may be formed.

For example, ethene can be oxidised to ethane-1,2-diol under suitable mild conditions.

$$CH_2 = CH_2 + [O] + H_2O \rightarrow HOCH_2CH_2OH$$

Observation

The purple potassium manganate(VII) solution is decolourised, often with formation of brown manganese(IV) oxide depending on the conditions.

Importance

- Oxidation reactions are used in chemical manufacturing.
- They provide methods for converting alkenes into useful oxygen-containing organic compounds.
- They help identify unsaturation in organic compounds.

13.0 CHEMICAL PROPERTIES OF ALKYNES

Alkynes are unsaturated hydrocarbons containing a carbon-carbon triple bond.

Their triple bond allows them to undergo addition reactions.

13.1 Combustion of alkynes

Alkynes burn in oxygen to produce carbon dioxide and water during complete combustion.

Example: Complete combustion of ethyne

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

Ethyne burns in air with a luminous and often sooty flame.

When ethyne burns in pure oxygen, it produces an extremely hot flame.

13.2 Addition reactions of alkynes

Alkynes can undergo addition reactions because of the presence of a carbon-carbon triple bond.

A. Addition of hydrogen

Hydrogen adds to ethyne in the presence of a suitable catalyst.

First stage:

$$C_2H_2 + H_2 \rightarrow C_2H_4$$

Ethyne forms ethene.

Second stage:

$$C_2H_4 + H_2 \rightarrow C_2H_6$$

Ethene forms ethane.

Overall:

$$C_2H_2 + 2H_2 \rightarrow C_2H_6$$

B. Addition of bromine

Ethyne reacts with bromine.

First addition:

$$HC \equiv CH + Br_2 \rightarrow CHBr = CHBr$$

Further addition:

$$CHBr = CHBr + Br_2 \rightarrow CHBr_2 - CHBr_2$$

Observation

Bromine water is decolourised under suitable conditions.

C. Addition of hydrogen chloride

Hydrogen chloride adds across the triple bond of ethyne.

$$HC \equiv CH + HCl \rightarrow CH_2 = CHCl$$

The product is chloroethene, which is used to manufacture PVC.

14.0 TESTS FOR SATURATION AND UNSATURATION

14.1 Meaning of a test for unsaturation

A test for unsaturation is a chemical test used to determine whether a hydrocarbon contains carbon-carbon double or triple bonds.

The common laboratory tests include:

96. Bromine water test.
97. Potassium manganate(VII) test.

14.2 Bromine water test

Bromine water is an orange-brown solution of bromine in water.

Procedure

98. Place a small quantity of bromine water in a test tube.
99. Add a small quantity of the hydrocarbon.
100. Shake gently.
101. Observe any colour change.

Expected observations

Substance	Observation under ordinary test conditions
Alkane	Orange-brown colour remains
Alkene	Orange-brown colour disappears
Alkyne	Orange-brown colour disappears

Explanation

- Alkenes and alkynes contain multiple carbon-carbon bonds.
- Bromine adds across these bonds.
- The coloured bromine is consumed during the reaction.
- Alkanes do not normally react with bromine water in the absence of ultraviolet light or other suitable reaction conditions.

Example: Ethene

$$CH_2 = CH_2 + Br_2 \rightarrow CH_2Br - CH_2Br$$

14.3 Potassium manganate(VII) test

Potassium manganate(VII), also called potassium permanganate, is a purple chemical reagent.

Procedure

102. Place dilute potassium manganate(VII) solution in a test tube.
103. Add the hydrocarbon sample.
104. Shake gently.
105. Observe the colour change.

Observation

- Unsaturated hydrocarbons may decolourise the purple solution under suitable conditions.
- A brown precipitate of manganese(IV) oxide may form in neutral or alkaline conditions.
- Saturated hydrocarbons do not normally produce this reaction under the same mild conditions.

Conclusion

Decolourisation of the reagent under suitable test conditions indicates the presence of unsaturation.

14.4 Comparison of the two tests

Test	Reagent	Positive result for unsaturation
Bromine water test	Orange-brown bromine water	Decolourisation
Potassium manganate(VII) test	Purple potassium manganate(VII)	Decolourisation, sometimes with brown precipitate

Laboratory safety

- Bromine is harmful and corrosive.
- Potassium manganate(VII) is an oxidising agent.
- Use small quantities.
- Wear protective goggles, gloves and laboratory coats.
- Avoid skin contact and inhalation.
- Dispose of chemicals according to laboratory instructions.

15.0 USES OF HYDROCARBONS

Hydrocarbons are important in domestic, transport, agricultural and industrial activities.

15.1 Uses as fuels

Hydrocarbons release energy during combustion.

Examples:

- Methane is used as a fuel in natural gas systems.
- Propane and butane are components of LPG used for cooking.
- Petrol is used in petrol engines.
- Diesel is used in diesel engines.
- Kerosene is used in some heating and lighting applications and as aviation fuel after suitable re-fining.

15.2 Manufacture of plastics

Alkenes are used to manufacture polymers.

Examples:

- Ethene produces poly(ethene).
- Propene produces poly(propene).
- Chloroethene produces PVC.

15.3 Manufacture of synthetic fibres

Hydrocarbon-derived chemicals are used to manufacture synthetic fibres.

Examples include:

- Nylon.
- Polyester.
- Acrylic fibres.

These materials are used in clothing, ropes, carpets and other products.

15.4 Manufacture of synthetic rubber

Hydrocarbon-based monomers are used in manufacturing synthetic rubber.

Synthetic rubber is used in:

- Vehicle tyres.
- Conveyor belts.
- Footwear.
- Seals and gaskets.
- Industrial hoses.

15.5 Manufacture of solvents

Some hydrocarbons are used as organic solvents in industrial applications.

Examples include:

- Hexane.
- Heptane.
- Other petroleum-derived solvents.

They are used in extraction, cleaning and chemical manufacturing.

15.6 Manufacture of lubricants

Higher petroleum fractions are processed into lubricating oils and greases.

They are used to:

- Reduce friction.
- Reduce wear.
- Remove heat from moving machine parts.
- Protect metallic surfaces from damage.

15.7 Manufacture of chemicals

Hydrocarbons are used as raw materials in the petrochemical industry.

They are used to produce:

- Alcohols.
- Organic acids.

- Detergents.
- Paints.
- Adhesives.
- Fertiliser-related chemical feedstocks.
- Pharmaceutical intermediates.

15.8 Other uses

- Bitumen is used in road construction.
- Paraffin wax is used in candles and protective coatings.
- Petroleum-derived chemicals are used in packaging materials.
- Some hydrocarbons are used as refrigerants, subject to environmental and safety regulations.

16.0 ENVIRONMENTAL EFFECTS OF HYDROCARBONS

Although hydrocarbons are useful, their extraction, transportation, processing and use can cause environmental problems.

16.1 Air pollution

Burning hydrocarbons can release:

- Carbon dioxide.
- Carbon monoxide.
- Nitrogen oxides from high-temperature combustion.
- Unburnt hydrocarbons.
- Soot and particulate matter.

Effects

106. Respiratory illnesses.
107. Reduced air quality.
108. Formation of photochemical smog.
109. Contribution to global warming.
110. Damage to plants and ecosystems.

16.2 Global warming

- Complete combustion of hydrocarbons releases carbon dioxide.
- Carbon dioxide is a greenhouse gas.
- Increased atmospheric greenhouse gas concentrations contribute to global warming and climate change.

16.3 Oil spills

Oil spills may occur during extraction, storage and transportation.

Effects

- Contamination of water bodies.
- Death of aquatic organisms.
- Damage to aquatic habitats.
- Contamination of shorelines.
- Disruption of fishing and other economic activities.

16.4 Soil pollution

- Leakage of petroleum products can contaminate soil.
- Soil organisms may be harmed.
- Plant growth may be affected.
- Contaminated land may become unsuitable for some agricultural activities.

16.5 Plastic pollution

Many plastics are manufactured from hydrocarbon-derived raw materials.

Improper disposal of plastics can cause:

- Blockage of drainage systems.
- Harm to livestock and wildlife.
- Pollution of rivers, lakes and oceans.
- Formation of microplastics.
- Long-term environmental contamination.

16.6 Measures for reducing environmental pollution

111. Use fuels efficiently.
112. Maintain vehicles and machines to reduce incomplete combustion.
113. Use cleaner energy sources where practical.
114. Prevent leakage during fuel transportation and storage.
115. Recycle suitable plastic materials.
116. Avoid burning plastics and other waste materials.
117. Dispose of used oil through approved collection and treatment systems.
118. Enforce environmental regulations.
119. Restore areas damaged by oil spills.
120. Promote public awareness on responsible fuel and plastic use.

16.7 Safe disposal of organic waste

Organic waste should be handled according to its chemical nature.

Examples of safe practices

- Separate biodegradable organic waste from hazardous chemical waste.
- Compost suitable biodegradable materials.
- Recycle suitable plastics and packaging.
- Store waste solvents in properly labelled containers.
- Do not pour petroleum products into drains, rivers or soil.
- Dispose of laboratory chemicals through approved procedures.
- Use designated collection centres for used engine oil and other hazardous waste.

17.0 SUMMARY OF IMPORTANT CONCEPTS

Concept	Main point
Organic chemistry	Study of carbon compounds
Hydrocarbon	Compound containing carbon and hydrogen only
Saturated hydrocarbon	Contains carbon-carbon single bonds only

Concept	Main point
Unsaturated hydrocarbon	Contains at least one carbon-carbon double or triple bond
Alkane	Saturated hydrocarbon
Alkene	Hydrocarbon containing a carbon-carbon double bond
Alkyne	Hydrocarbon containing a carbon-carbon triple bond
Homologous series	Family of compounds with similar chemical properties and a common general formula
Isomerism	Same molecular formula but different arrangement of atoms
Fractional distillation	Separation of crude oil into fractions by boiling-point differences
Cracking	Breaking large hydrocarbon molecules into smaller molecules
Combustion	Reaction with oxygen that releases energy
Substitution	Replacement of an atom or group in a molecule
Addition	Addition of atoms or groups across a multiple bond
Polymerisation	Joining many monomers to form a polymer
Oxidation	Reaction involving an increase in oxidation state or addition of oxygen/removal of hydrogen in the appropriate context

17.1 Essential general formulae

Homologous series	General formula
Alkanes	C_nH_{2n+2}
Alkenes	C_nH_{2n}
Alkynes	C_nH_{2n-2}

17.2 Important chemical equations

Preparation of methane

$$CH_3COONa + NaOH \xrightarrow{\Delta} CH_4 + Na_2CO_3$$

Preparation of ethene

$$C_2H_5OH \xrightarrow{\text{conc. } H_2SO_4} C_2H_4 + H_2O$$

Preparation of ethyne

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

Complete combustion of methane

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Substitution of methane

$$CH_4 + Cl_2 \xrightarrow{UV} CH_3Cl + HCl$$

Hydrogenation of ethene

$$C_2H_4 + H_2 \xrightarrow{Ni} C_2H_6$$

Addition of bromine to ethene

$$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$$

Polymerisation of ethene

$$n(CH_2 = CH_2) \rightarrow [-CH_2 - CH_2 -]_n$$

18.0 REVISION QUESTIONS

Grade 11 Chemistry – Strand 3.0: Organic Chemistry, Sub-strand 3.1: Hydrocarbons

SECTION A: MULTIPLE-CHOICE QUESTIONS

Choose the correct answer from the alternatives provided.

1. Which of the following statements best defines an organic compound?

A. A compound containing only oxygen and hydrogen.

B. A compound based mainly on carbon.

C. A compound containing only metals.

D. A compound containing only carbon and oxygen.

2. Which of the following is a hydrocarbon?

A. Ethanol.

B. Ethanoic acid.

C. Propane.

D. Glucose.

3. The ability of carbon atoms to bond with one another to form long chains is known as:

A. Tetravalency.

B. Catenation.

C. Sublimation.

D. Neutralisation.

4. Which of the following is a natural source of hydrocarbons?

A. Crude oil.

B. Sodium chloride.

C. Limestone.

D. Water.

5. The main component of natural gas is:

A. Ethene.

B. Ethyne.

C. Methane.

D. Butene.

6. Which of the following is a saturated hydrocarbon?

A. Ethene.

B. Ethyne.

C. Propane.

D. Propene.

7. The general formula of open-chain alkanes is:

A. C_nH_{2n}.

B. C_nH_{2n+2}.

C. C_nH_{2n-2}.

D. C_nH_n.

8. Which of the following is the molecular formula of hexane?

A. C_6H_{10}.

B. C_6H_{12}.

C. C_6H_{14}.

D. C_6H_{16}.

9. An open-chain alkene containing five carbon atoms has the molecular formula:

A. C_5H_{12}.

B. C_5H_{10}.

C. C_5H_8.

D. C_5H_6.

10. Which of the following is the general formula of open-chain alkynes containing one triple bond?

A. C_nH_{2n+2}.

B. C_nH_{2n}.

C. C_nH_{2n-2}.

D. C_nH_{n+2}.

11. The first member of the alkene homologous series is:

A. Methene.

B. Ethene.

C. Propene.

D. Butene.

12. Which suffix is used in naming alkynes?

A. -ane. B. -ene.

C. -yne. D. -ol.

13. The correct IUPAC name of $CH_2 = CH - CH_2 - CH_3$ is:

A. Butane.

B. But-1-ene.

C. But-2-ene.

D. But-1-yne.

14. The compound $CH_3 - C \equiv C - CH_3$ is named:

A. But-1-yne.

B. But-2-yne.

C. But-1-ene.

D. But-2-ene.

15. Which of the following pairs are structural isomers?

A. Ethane and ethene.

B. Butane and 2-methylpropane.

C. Methane and ethane.

D. Ethene and ethyne.

16. But-1-ene and but-2-ene are examples of:

A. Structural chain isomerism.

B. Positional isomerism.

C. Optical isomerism.

D. Functional group isomerism.

17. Fractional distillation of crude oil is based mainly on differences in:

A. Colour.

B. Density.

C. Boiling points.

D. Solubility in water.

18. Which fraction of crude oil is commonly used as fuel in diesel engines?

A. Refinery gases.

B. Petrol.

C. Diesel oil.

D. Bitumen.

19. The process of breaking large hydrocarbon molecules into smaller molecules is called:

A. Polymerisation.

B. Cracking.

C. Neutralisation.

D. Hydration.

20. Which type of bond is found between carbon atoms in ethene?

A. Single bond. B. Double bond.

C. Triple bond. D. Ionic bond.

21. Methane can be prepared in the laboratory by heating sodium ethanoate with:

A. Water.

B. Soda lime.

C. Hydrochloric acid.

D. Sodium chloride.

22. Ethyne is prepared by reacting calcium carbide with:

A. Ethanol.

B. Oxygen.

C. Water.

D. Hydrogen.

23. The main product formed when ethanol undergoes dehydration is:

A. Ethane.

B. Ethene.

C. Ethyne.

D. Methane.

24. Which of the following is a physical property of hydrocarbons?

A. Combustion.

B. Substitution.

C. Boiling point.

D. Polymerisation.

25. Most hydrocarbons are insoluble in water because they are:

A. Ionic.

B. Non-polar.

C. Metallic.

D. Strong acids.

26. Complete combustion of methane produces:

A. Carbon and hydrogen.

B. Carbon dioxide and water.

C. Carbon monoxide and hydrogen.

D. Methanol and oxygen.

27. Which poisonous gas may be produced during incomplete combustion of hydrocarbons?

A. Oxygen. B. Nitrogen.

C. Carbon monoxide. D. Hydrogen.

28. Methane reacts with chlorine in ultraviolet light through a:

A. Substitution reaction.

B. Neutralisation reaction.

C. Polymerisation reaction.

D. Precipitation reaction.

29. Which reagent is commonly used to test for unsaturation?

A. Sodium chloride solution.

B. Bromine water.

C. Limewater.

D. Distilled water.

30. When ethene is passed through bromine water, the expected observation is:

A. Formation of a blue solution.

B. Formation of a white precipitate.

C. Decolourisation of bromine water.

D. No visible change under suitable test conditions.

31. The addition of hydrogen to ethene produces:

A. Ethyne.

B. Ethane.

C. Methane.

D. Propane.

32. The polymer formed from ethene is:

A. Poly(ethene).

B. Poly(propene).

C. PVC.

D. Nylon.

33. Which hydrocarbon is commonly used in oxy-acetylene welding?

A. Methane. B. Ethane.

C. Ethene. D. Ethyne.

34. Which of the following is an environmental effect of burning fossil fuels?

A. Reduction in atmospheric carbon dioxide.

B. Increased greenhouse gas emissions.

C. Complete prevention of air pollution.

D. Increased soil fertility in all areas.

35. Which of the following is an appropriate way of disposing of used engine oil?

A. Pouring it into a river.

B. Pouring it into a drainage system.

C. Taking it to an approved collection or treatment facility.

D. Burning it in an open space.

SECTION B: STRUCTURED QUESTIONS

1. Introduction to organic chemistry

a) Define organic chemistry. (2 marks)

b) Explain four reasons why carbon forms a large number of compounds. (4 marks)

c) State five natural sources of organic compounds. (5 marks)

d) Define a hydrocarbon. (2 marks)

e) Give four examples of hydrocarbons and their molecular formulae. (4 marks)

2. Classification of hydrocarbons

a) Define the term saturated hydrocarbon. (2 marks)

b) Define the term unsaturated hydrocarbon. (2 marks)

c) Distinguish between alkanes, alkenes and alkynes. (6 marks)

d) State three differences between saturated and unsaturated hydrocarbons. (3 marks)

3. Homologous series

a) Define a homologous series. (2 marks)

b) State four characteristics of a homologous series. (4 marks)

c) Write the general formulae for:

i. Alkanes. (1 mark)

ii. Alkenes. (1 mark)

iii. Alkynes. (1 mark)

d) Determine the molecular formulae of the following:

i. An alkane containing 8 carbon atoms. (2 marks)

ii. An alkene containing 6 carbon atoms. (2 marks)

iii. An alkyne containing 7 carbon atoms. (2 marks)

4. Nomenclature of hydrocarbons

a) What is meant by IUPAC nomenclature? (2 marks)

b) Write the prefixes used for hydrocarbons containing:

i. 1 carbon atom.

ii. 3 carbon atoms.

iii. 5 carbon atoms.

iv. 8 carbon atoms. (4 marks)

d) Name the following compounds:

i. $CH_3CH_2CH_3$. (1 mark)

ii. $CH_2 = CHCH_3$. (1 mark)

iii. $CH_3C \equiv CH$. (1 mark)

iv. $CH_2 = CHCH_2CH_3$. (1 mark)

v. $CH_3CH = CHCH_3$. (1 mark)

d) Draw the structural formulae of:

i. Butane.

ii. But-1-ene.

iii. But-2-yne. (6 marks)

5. Isomerism

a) Define isomerism. (2 marks)

b) Explain structural isomerism. (2 marks)

c) Explain positional isomerism. (2 marks)

d) Draw the structural formulae of the two structural isomers of C_4H_{10}. (4 marks)

e) Draw the positional isomers of butene. (4 marks)

f) State two importance of studying isomerism. (2 marks)

6. Fractional distillation of crude oil

a) Define crude oil. (2 marks)

b) Explain the principle of fractional distillation. (3 marks)

c) Describe the process of fractional distillation of crude oil. (6 marks)

d) Name five fractions obtained from crude oil and give one use of each. (10 marks)

e) Explain why different fractions condense at different levels in the fractionating column. (3 marks)

7. Cracking of hydrocarbons

a) Define cracking. (2 marks)

b) Distinguish between thermal cracking and catalytic cracking. (4 marks)

c) Explain three reasons why cracking is important in petroleum refining. (6 marks)

d) Write a balanced equation showing one possible cracking reaction of decane. (2 marks)

8. Bonding in hydrocarbons

a) Define a covalent bond. (2 marks)

b) Explain how a carbon-carbon single bond is formed. (2 marks)

c) Distinguish between single, double and triple covalent bonds. (6 marks)

d) Draw the structural formulae of:

i. Methane.

ii. Ethene.

iii. Ethyne. (6 marks)

9. Preparation of methane

a) State the laboratory method used to prepare methane. (2 marks)

b) Name the two main reactants used. (2 marks)

c) Write the balanced chemical equation for the preparation of methane. (2 marks)

d) Describe the laboratory procedure for preparing methane. (5 marks)

e) State three precautions taken during the experiment. (3 marks)

10. Preparation of ethene

a) Name the compound used to prepare ethene in the laboratory. (1 mark)

b) Name the reagent used to dehydrate ethanol in the common laboratory method. (1 mark)

c) Write the balanced chemical equation for the preparation of ethene. (2 marks)

d) Describe how ethene can be prepared in the laboratory. (5 marks)

e) State two physical properties of ethene. (2 marks)

f) State two precautions during the experiment. (2 marks)

11. Preparation of ethyne

a) State the reactants used in the laboratory preparation of ethyne. (2 marks)

b) Write the balanced chemical equation for the reaction. (2 marks)

c) Describe the observations made during the experiment. (3 marks)

d) State three safety precautions during the preparation. (3 marks)

12. Physical properties of hydrocarbons

a) State six physical properties used to describe hydrocarbons. (6 marks)

b) Explain why hydrocarbons are generally insoluble in water. (3 marks)

c) Describe the trend in boiling points as the size of alkane molecules increases. (3 marks)

d) Explain why larger alkane molecules generally have higher boiling points than smaller alkane molecules. (4 marks)

13. Combustion of alkanes

a) Define combustion. (2 marks)

b) Distinguish between complete and incomplete combustion. (4 marks)

c) Write balanced chemical equations for:

i. Complete combustion of methane. (2 marks)

ii. Complete combustion of propane. (2 marks)

iii. Incomplete combustion of methane to form carbon monoxide. (2 marks)

d) State three dangers associated with incomplete combustion. (3 marks)

14. Substitution reactions

a) Define a substitution reaction. (2 marks)

b) State the conditions necessary for methane to react with chlorine. (2 marks)

c) Write a balanced equation for the reaction between methane and chlorine. (2 marks)

d) Write an equation for the reaction between ethane and chlorine. (2 marks)

e) Explain why the reaction between methane and chlorine is classified as substitution. (3 marks)

15. Addition reactions of alkenes

a) Define an addition reaction. (2 marks)

b) Explain why alkenes undergo addition reactions more readily than alkanes. (3 marks)

c) Write balanced equations for the reaction of ethene with:

i. Hydrogen. (2 marks)

ii. Bromine. (2 marks)

iii. Hydrogen chloride. (2 marks)

iv. Water. (2 marks)

d) State one industrial use of hydrogenation. (2 marks)

16. Polymerisation

a) Define polymerisation. (2 marks)

b) Distinguish between a monomer and a polymer. (2 marks)

c) Write the equation for the polymerisation of ethene. (3 marks)

d) Name the polymer formed from ethene. (1 mark)

e) State four uses of poly(ethene). (4 marks)

f) Explain two environmental problems associated with improper disposal of plastics. (4 marks)

17. Oxidation of alkenes

a) Define oxidation in terms of organic chemistry. (2 marks)

b) Name a reagent that can be used to test for unsaturation through oxidation. (2 marks)

c) State the observation when an alkene reacts with cold, dilute alkaline potassium manganate(VII). (2 marks)

d) Name the type of organic product that may be formed when ethene undergoes mild oxidation. (2 marks)

18. Tests for unsaturation

a) Name two chemical reagents used to test for unsaturation. (2 marks)

b) Describe how bromine water is used to distinguish between an alkane and an alkene. (4 marks)

c) State the expected observation when bromine water is added to:

i. Hexane. (1 mark)

ii. Hexene. (1 mark)

d) Explain why alkenes decolourise bromine water. (3 marks)

19. Uses of hydrocarbons

a) State six uses of hydrocarbons in daily life. (6 marks)

b) Explain the importance of hydrocarbons in the transport industry. (3 marks)

c) State four uses of petroleum-derived materials in Kenya. (4 marks)

d) Explain why hydrocarbons are important raw materials in the petrochemical industry. (3 marks)

20. Environmental conservation

a) State four environmental problems associated with the use of hydrocarbons. (4 marks)

b) Explain three effects of oil spills on the environment. (6 marks)

c) State five measures that can be used to reduce pollution caused by petroleum products and plastics. (5 marks)

d) Explain why used engine oil should not be poured into rivers or drainage systems. (3 marks)

19.0 PRACTICAL AND APPLICATION-BASED QUESTIONS

21. Investigation of unsaturation

A learner was provided with three colourless liquids labelled A, B and C. The learner added bromine water to each liquid and recorded the following observations.

Liquid	Observation
A	Orange-brown colour remained
B	Orange-brown colour disappeared
C	Orange-brown colour disappeared

a) Which liquid or liquids are likely to contain unsaturated hydrocarbons? (2 marks)

b) Which liquid is likely to be a saturated hydrocarbon? (1 mark)

c) Explain the observations. (4 marks)

d) State one precaution when carrying out the experiment. (1 mark)

22. Application of cracking

A petroleum refinery has a large supply of heavy hydrocarbons but experiences high demand for petrol and petrochemical raw materials.

a) Identify the process that can be used to convert the heavy hydrocarbons into smaller molecules. (1 mark)

b) Explain how the process works. (3 marks)

c) State two useful products that may be obtained. (2 marks)

d) Explain why the process is important to the petroleum industry. (3 marks)

23. Hydrocarbons in domestic energy use

A household in Kenya uses LPG for cooking.

a) Name two hydrocarbons that are common components of LPG. (2 marks)

b) Explain why LPG is suitable for domestic cooking. (3 marks)

c) State three safety precautions when using LPG. (3 marks)

d) Explain why an odorant is added to commercial LPG. (2 marks)

SUMMARY NOTES ON HYDROCARBONS

Organic chemistry

- Is the branch of chemistry is the study of the structure, properties, composition, reactions, and preparation of carbon-containing compounds excluding oxides of carbon, hydrogen carbonates and carbonates.
- Carbon is **tetravalent.** It is able to form stable covalent bonds with itself and many non-metals like hydrogen, nitrogen, Oxygen and halogens to form a variety of compounds.

.Hydrocarbons

- Are compounds of hydrogen and carbon only; and are the simplest organic compounds.

Main groups of hydrocarbons

Are classified on the basis of the type of bonds found within the carbon atoms.

- **Alkanes:** Are hydrocarbons in which carbon atoms are linked by single covalent bonds.
- **Alkenes:** Carbon atoms are held by at least one double bond.
- **Alkynes:** Have at least one triple bond between any two carbon atoms.

Saturated hydrocarbons

- They are compounds in which each carbon atom has only single covalent bonds, throughout the structure.

Unsaturated hydrocarbons

- Are hydrocarbons which contain at least one double or triple bond, between any two adjacent carbon atoms.

Examples: All alkenes and Alkynes.

Homologous series

- Refers to a group of organic compounds that have the same general formula, whose consecutive members differ by a similar unit, and usually have similar chemical properties.

Characteristics of a Homologous series.

1. Can be represented by a general formula;
2. Have similar chemical properties.
3. Have similar structures and names.
4. They show a steady gradation of physical properties
5. Can usually be prepared by similar methods.

Alkanes

- Alkanes are hydrocarbons that are composed of **carbon and hydrogen** atoms with **single** covalent bonds.
- Alkanes are saturated hydrocarbons; they are called saturated because they contain the maximum amount of hydrogen atoms possible.
- They are the simplest hydrocarbons with the general formula**; CnH2n + 2** where n = number of carbon atoms in the molecule.
- Since carbon is **tetravalent** ,each atom of carbon in the alkane MUST always be bonded using **four covalent bond** /four shared pairs of electrons.
- Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkane MUST always be bonded using one covalent bond/one shared pair of electrons
- Their name end with a surfix **- ane** after the prefix
- The alkanes have the prefix **meth_,eth_,prop_** and **but_** to represent 1,2,3 and 4 carbons in the compound. All other use the **numeral** prefix **pent_,Hex_,hept_**
- The **general formula/molecular formular** of a compound is the number of each atoms of elements making the compound
- The **structural formula** is the **arrangement / bonding** of atoms of each element making the compound The formula is written without showing the bonds and each carbon is written separately, with the atoms or groups that are attached to it

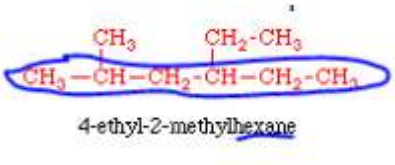
Skeletal notation:

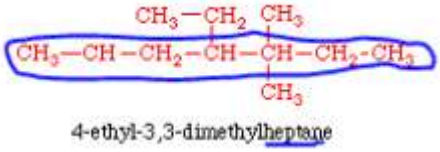
- In this notation we do not show the carbon atoms at all. We show only the carbon skeleton as zigzag lines. The skeleton has no carbon or hydrogen atoms written in it. We assume that carbon atoms are at each end and a carbon atom where two lines meet.

Alkanes Nomenclature and Naming

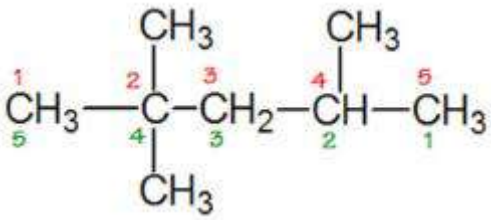
Rules of naming

Identify the longest continuous carbon chain to get/determine the parent alkane





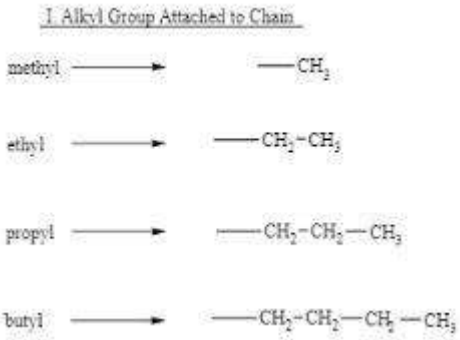
Number the longest chain from the end of the chain that is near the branches so as the branch get the lowest number possible.



Determine the position, number and type of branches.

Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane.

Name them fluoro-,chloro-,bromo-,iodo- if they are halogens



R–X

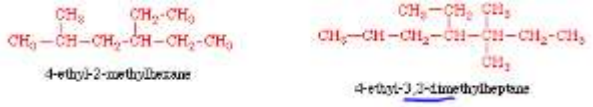
X = F, Cl, Br, I

Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of branches attached to the parent alkane.

The branches should be named in an alphabetic order

E.g ethyl comes before methyl

And bromo comes before chloro which comes before ethyl



QUESTIONS.

(1). Draw the structure of:

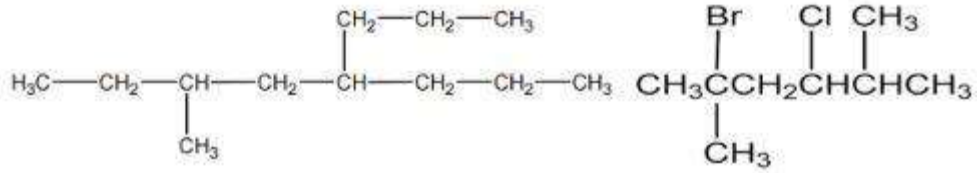
a. 2-methylpentane

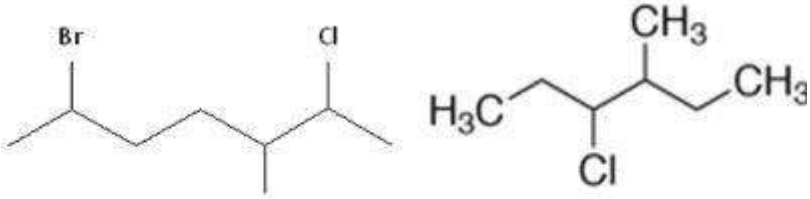
b. 2,2-dimethylpentane

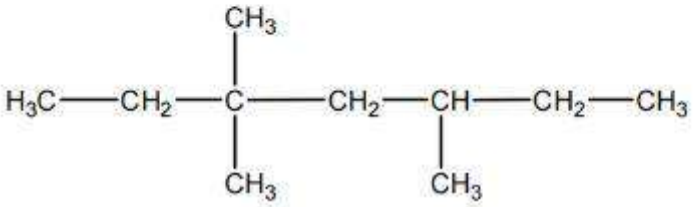
c. 2,2,3-trimethylbutane

d. 1,1,1-trichloro-2,2-dimethyl-butane

(2). Name the following structures







Isomers of Alkanes

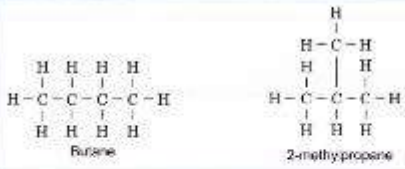
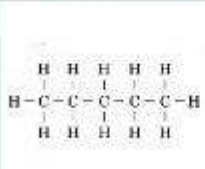
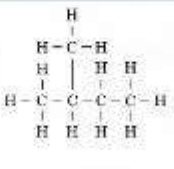
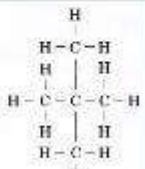
Isomerism

1. Is a situation whereby two or more compounds have similar molecular formulae but different structural formula.
2. Such compounds are called **isomers**, i.e compounds with the same molecular formula but different structural formula.

Isomers have different physical and chemical properties.

The 1st three alkanes do not form isomers. Isomers are named by using the IUPAC (International Union of Pure and Applied Chemistry) system of nomenclature /naming.

Table *Names and structural formulae of isomers of alkanes*

n	Molecular formula, C_nH_{2n+2}	Structural formula	Number of isomers
4	C_4H_{10}	 Butane 2-methylpropane	2
5	C_5H_{12}	 Pentane  2-methylbutane  2,2-dimethylpropane	3

QUESTIONS

(1). Draw and name all the isomers of;

a. Hexane

b. Octane

c. Nonane

d. Decane

Natural Sources of Alkanes

Occurrence of Alkanes

1. There are 3 known natural sources:

Natural gas: this consists of mainly of methane;

Crude oil:

1. Consists of a mixture of many Alkanes
2. It can be separated into its components by **fractional distillation**.

Reason:

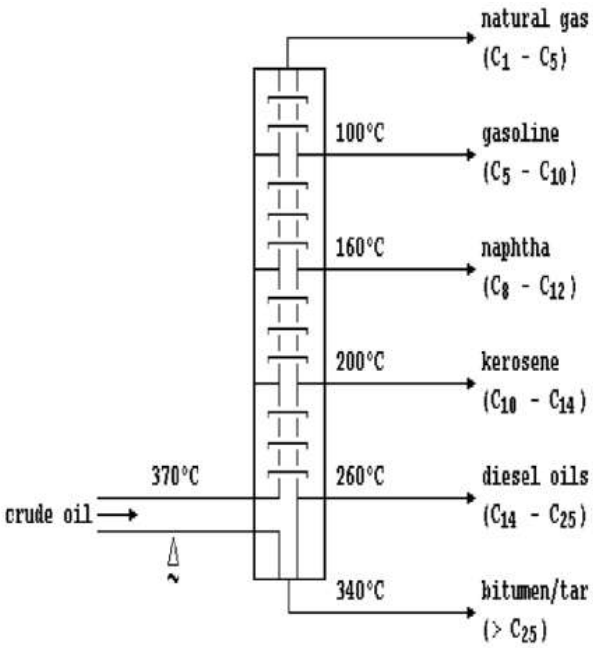
1. The different components have different boiling points

Biogas:

This contains about 60-75% of methane gas/marshy gas.

Crude oil is a naturally occurring viscous black oil. The raw material is made up of complex organic hydrocarbon structures and is formed from living matter that existed many millions of years ago

1. The liquid crude oil is heated within a chamber, to about 350 – 400 °C, until most of it evaporates and forms a gas. Any liquid is drained off as the crude oil gas enters the fractionating column. The bottom of the fractionating column has the highest temperature and the top of the column has the lowest temperature.
2. Long hydrocarbons condense first. Long chain alkanes have stronger intermolecular forces of attraction between their molecules and therefore require large force to separate them and hence a higher boiling point.
3. Short hydrocarbons condense last. The small chain alkanes have weaker intermolecular forces of attraction and hence can be broken down by a small force hence the lower boiling point.



Appearance

1. Intensity of the colour increases with increase in boiling point.
2. Boiling point increases with increasing number of carbon atoms.
3. The higher the number of carbon atoms, the higher the number of covalent bonds.
4. Thus the first fraction to be distilled (lab gas) is colorless while the last distillates (between) is dark black in colour

Viscosity

1. Increases with increasing boiling point;
2. The fractions with low boiling points are less viscous while the fraction with the highest boiling point is semi-solid;

Inflammability:

1. Decreases with increasing boiling points.
2. The gaseous fractions, with least boiling points readily catches fire // burn, while the semi-solid fractions with very high boiling points are almost non-combustible.

Uses of the various fractions of crude oil.

No. f carbon atom per molecule	Fractions	Uses
1-4	Gases	Laboratory gases and gas cookers
5-12	Petrol	Fuel in petrol engines
9-16	Kerosene (paraffin)	Fuel for jet engines (aeroplanes) and domestic uses
15-18	Light diesel oils	Fuel for heavy diesel engines e.g. for ships
18-25	Diesel oils	Fuel for diesel engines
20-70	Lubricating oils	Used for smooth running of engine parts
>70	Bitumen	Road tarmacking

Cracking sound of alkanes

To increase the supply of short chain alkanes long chain alkanes are broken up in the process called **cracking of alkanes**

There are two ways of cracking of alkanes

1. Thermal cracking
2. Catalytic cracking [Zeolite]

During cracking short chain alkanes, alkenes and hydrogen gas are produced

In thermal cracking a refined crude oil fraction containing long chain alkanes is heated to a temperature of **700 degrees.** The result is a mixture of short chain alkanes and alkenes.

In catalytic cracking a **catalyst** is used so that the long chain alkane can be broken at a lower temperature of 450 degrees. The catalyst used is powdered **silica, alumina or zeolite**

QUESTIONS

(1). Propane can be changed into methane and ethane as shown in the equation below.

$$CH_3CH_2CH_3(g) \xrightarrow{\text{high temperature}} CH_4(g) + C_2H_4(g)$$

a. Name the process undergone by propane.

(2). The molecular formula of a hydrocarbon is C_6H_{14}. The hydrocarbon can be converted into two other hydrocarbons as shown by the equation below.

$C_6H_{14} \rightarrow C_2H_6 + x$

a. Name and draw the possible structural formula of x.

(3). Crude oil is a source of many compounds that contain carbon and hydrogen only

a. Name the process used to separate the components of crude oil

b. On what two physical properties of the above components does the separation depend (2mks)

(4). The table below gives information about the major constituents of crude oil. Study it and answer the questions that follow.

Constituent	Boiling point (°C)
Gases	Below 40
Petrol	40 - 175
Kerosene	175 - 250
Diesel oil	250 - 350
Lubricating oil	350 - 400
Bitumen	Above 400

a. Which one of the constituents of crude oil has molecules with the highest number of carbon atoms? Explain

b. Name the process you would use to separate a mixture of petrol and diesel and explain how the separation takes place

c. Explain why the constituent of crude oil have a range of boiling point.

(5). Below is a list of major components of crude oil and their boiling points.

Component	**Boiling point (^{0}C)**
Bitumen	Above 400
Lubricating oil	350 - 400
Petrol	40 - 175
Gases	Below 40

(i) What is the name of the process by which the constituents of crude oil can be separated?

(ii) Give one use of the gases component.

(iii) Give the order by which the components are obtained from the mixture, starting with the first.

Laboratory Preparation of Alkanes

Preparation of Alkanes

1. Alkanes, like any other Homologous series have similar chemical properties.
2. In a school laboratory, alkanes may be prepared from the reaction of a sodium alkanoate with solid sodium hydroxide /soda lime.
3. Generally any alkane can be represented form the reaction represented by the following equation:

$$C_nH_{2n+1}COONa + NaOH_{(aq)} \rightarrow C_nH_{2n+2} + Na_2CO_{3(aq)};$$

Thus;

1. Methane can be prepared form sodium ethanoate (CH3COONa)
2. Ethane can be prepared form sodium propanoate (CH3CH2COONa)
3. Propane can be prepared form sodium Butanoate (CH3CH2CH2COONa)

Procedure

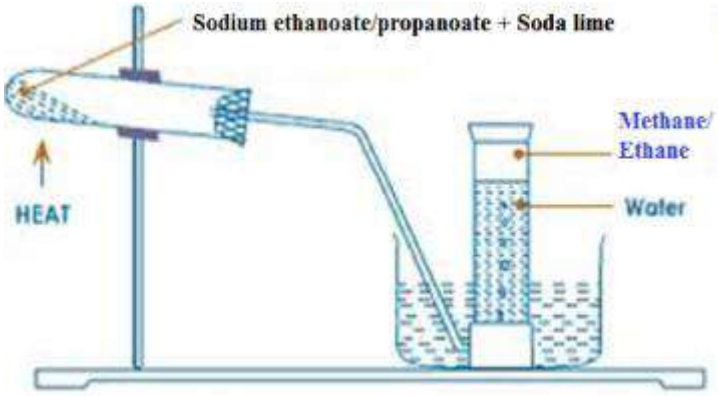
1. About 5g of sodium ethanoate and an equal mass of soda lime is put in a hard glass test tube, upon mixing them thoroughly in a mortar.
2. The mixture is heated thoroughly in the test-tube.

Observation

1. A colorless gas collects over water

Reasons:

1. Methane does not react with and is insoluble in water.



Equation

$CH_3COONa + NaOH_{(s)} \rightarrow CH_{4(g)} + Na_2CO_{3(aq)}$

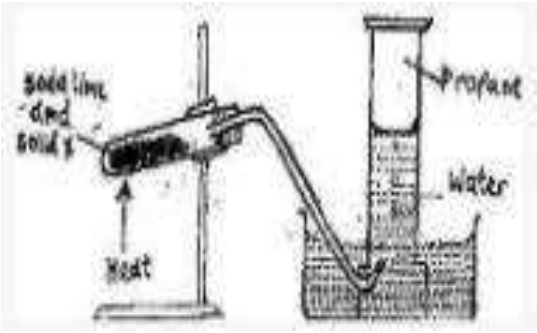
Sodium ethanoate ***sodalime*** ***Methane*** ***Sodium carbonate***

1. Methane is a non-poisonous, colorless and odourless gas. It is insoluble in water but soluble in organic solvents such as ethanol.
2. The melting and boiling points of alkanes increase with increase in number of carbon atoms. The longer the chain, the stronger the intermolecular forces of attraction (Van der Waals forces).
3. The first four alkanes are gases, the next six are liquids, while the rest are solids.
4. The density of straight chain alkanes increase with increasing carbon chain as the intermolecular forces increases.
5. This reduces the volume occupied by a given mass of the compound.

Alkane	General formula	Melting point(K)	Boiling point(K)	Density gcm^{-3}	State at room(298K) temperature and pressure atmosphere (101300Pa)
Methane	CH4	90	112	0.424	gas
Ethane	CH_3CH_3	91	184	0.546	Gas
Propane	$CH_3CH_2CH_3$	105	231	0.501	gas
Butane	$CH_3(CH_2)_2CH_3$	138	275	0.579	gas
Pentane	$CH_3(CH_2)_3CH_3$	143	309	0.626	liquid
Hexane	$CH_3(CH_2)_4CH_3$	178	342	0.657	liquid
Heptane	$CH_3(CH_2)_5CH_3$	182	372	0.684	liquid
Octane	$CH_3(CH_2)_6CH_3$	216	399	0.703	liquid
Nonane	$CH_3(CH_2)_7CH_3$	219	424	0.708	liquid
Octane	$CH_3(CH_2)_8CH_3$	243	447	0.730	liquid

QUESTIONS;

(1). The figure below represents the set up that was used to crack an alkane.

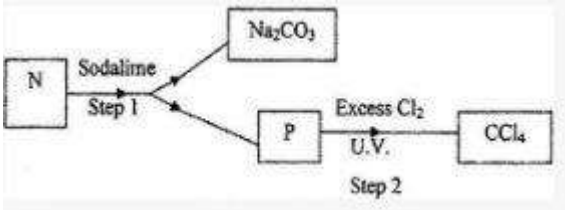


a. Write the formular of solid x

b. Explain why calcium oxide is added to the mixture above.

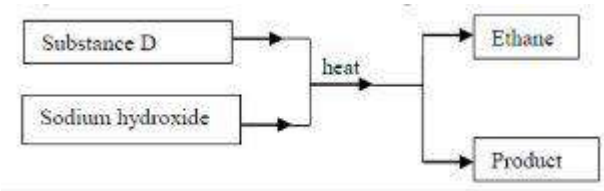
c. Write a balanced equation for the reaction that occurs in the flask.

(2). Study the flow chart below and answer the questions that follow.



a. Identify N and P

(3). Study the flow chart below and answer the question that follow.



Name the substance D

Chemical Properties and Uses of Alkanes

Chemical properties

Burning

- It is flammable and burns in excess air // oxygen with a pale blue non-luminous flame to give carbon (IV) oxide ad water vapour.

Equation:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Note:

- In a limited supply of air, the flame is **luminous**.

Reason:

- This is due to incomplete combustion of the methane.
- A mixture of methane and air **explodes** violently when ignited if the volume ratio is approximately 1:10 and this is often the cause of fatal explosions in coal mines.

Reaction with Bromine water and acidified potassium permanganate

- When methane is bubbled through bromine water the red brown colour of bromine persists; and when bubbled through acidified potassium manganate (VII) solution; the purple colour of the solution remains.
- Thus it has no effect on either bromine water or acidified potassium permanganate

.

Reason:

- It is a saturated hydrocarbon.

Substitution reactions

- A substitution reaction is one in which one atom replaces another atom in a molecule.

The substitution of Bromine in methane.

Procedure:

- A sample of Methane (CH_4) is placed in a boiling tube and to it is added some bromine gas.
- The tube is stoppered, and the mixture shaken, then allowed to stand and exposed to **ultra-violet lamp.**

Observations

- The red colour of Bromine begins to fade, and the pungent smell of hydrogen bromide (HBr) gas is detectable when the stopper is removed.
- A moist blue litmus paper also turns **red** on dipping into the resultant mixture.

Equation $CH_4(g) + Br_2(g) \rightarrow CH_3Br(g) + HBr(g)$

Explanation

- For a chemical reaction to occur, bonds must be broken.
- The light energy (V.V. light) splits the Bromine molecule into free atoms, which are very reactive species.
- Similarly the energy breaks the weaker carbon – hydrogen bonds, and not the stronger carbon – carbon bonds
- The free bromine atoms can then substitute (replace one of the hydrogen atoms of methane, resulting unto bromomethane and hydrogen bromide gas.

- This process can be repeated until all hydrogen atoms in CH_4 are replaced.
- The substitution reactions can also occur with chlorine, forming chloromethane dichloromethane, trichloromethane (chloroform) and tetrachloromethane (carbon tetrachloride) respectively.

USES OF GASES

(i)

- Most alkanes are used as fuel e.g.
- Methane is used as biogas in homes.
- Butane is used as the Laboratory gas.

(ii)

- On cracking, alkanes are a major source of Hydrogen for the manufacture of ammonia/Haber process.

(ii)

- Methane is used in manufacture of Carbon black which is a component in printers ink.

(iv)

- Methane is used in manufacture of useful industrial chemicals like methanol, methanol, and chloromethane.

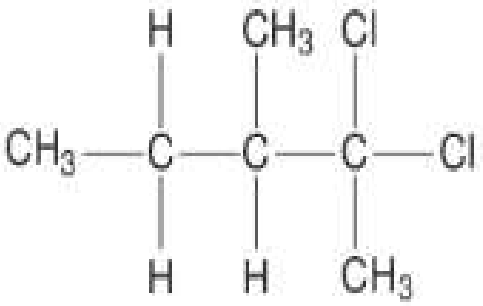
QUESTIONS

(1). When the alkane, hexane, is heated to high temperature, one of the products is ethene.

a. Write the equation for the reaction. (1 mark)

b. Name the process described in (b). (1 mark)

c. A different chlorinated compound shown below. name this compound and state its empirical formula.



(2). State the uses of alkanes

(3). Propane and chlorine react as shown below

$CH_3CH_2CH_3 + Cl_2 \rightarrow CH_3CH_2CH_2CL + HCl$

a. Name the type of reaction that takes place.

b. State the conditions under which this reaction takes place.

(4). Name and draw the structure of the compound formed when methane react with excess chlorine in presence of U.V light.

Alkenes

1. Are hydrocarbons with at least one carbon-carbon double bond, and have the general formula CnH2n.
2. They thus form a homologous series – with the simplest member being ethane.

In organic chemistry, an alkene is an unsaturated hydrocarbon that contains at least one carbon–carbon double bond

The double bond is the **functional** group. The carbon atoms are linked by at least one **double** bond to each other and single bonds to hydrogen atoms.

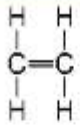
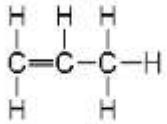
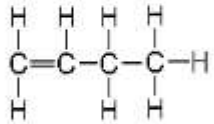
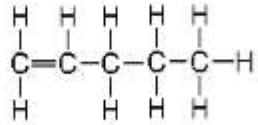
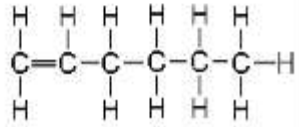
Their name end with a suffix – ene after the prefix

The alkenes have the prefix **eth_,prop_** and **but_** to represent 1,2,3 and 4 carbons in the compound. All other use the numeral prefix **pent_,Hex_,hept_**

The **general formula/molecular formular** of a compound is the number of each atoms of elements making the compound

The **structural formula** is the **arrangement / bonding** of atoms of each element making the compound The formula is written without showing the bonds and each carbon is written separately, with the atoms or groups that are attached to it.

Skeletal notation: In this notation we do not show the carbon atoms at all. We show only the carbon skeleton as zigzag lines. The skeleton has no carbon or hydrogen atoms written in it. We assume that carbon atoms are at each end and a carbon atom where two lines meet.

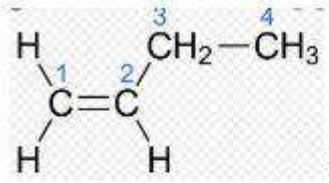
ALKENE	CARBON NUMBER	FORMULA	STRUCTURE
ETHENE	2	C_2H_4	
PROPENE	3	C_3H_6	
BUTENE	4	C_4H_8	
PENTENE	5	C_5H_{10}	
HEXENE	6	C_6H_{12}	

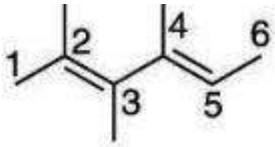
Rules of Naming

RULES

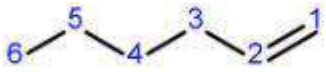
1. The longest chain is the one containing the double bond
2. The parent name should include the position of the double bond. And if there are 2 or more double bonds their positions and the words di, tri should be indicated before the suffix –ene
3. When numbering the double bond should be given the list number possible

If there are more than one substituent groups name them alphabetically (refer to naming in alkanes

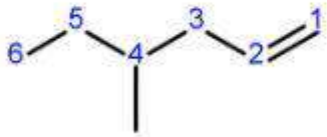




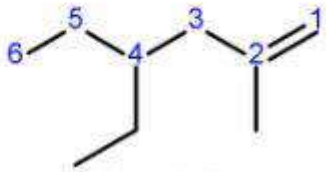
2,3,4-trimethyl-2,4-hexadiene



hex-1-ene



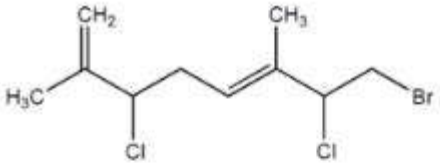
4-methylhex-1-ene

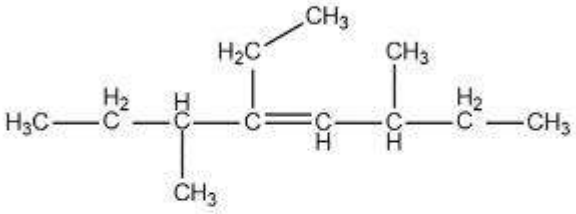


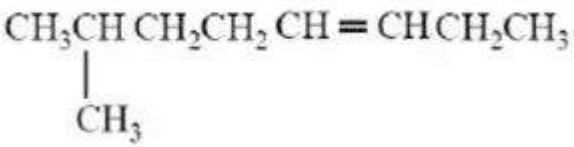
4-ethyl-2-methylhex-1-ene

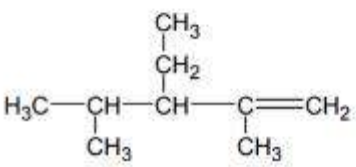
QUESTIONS

(1). Name the following alkenes.









(2). Draw the following compounds

a. 2,3-dimethylpent-2-ene

b. 2,2-dibromo-3-methylhep-3-ene

c. 2-methyl but-2-ene

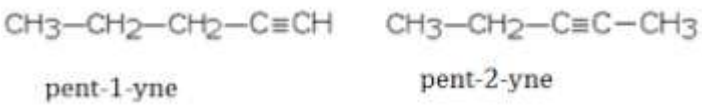
Isomerism in Alkenes

Just like alkanes, alkynes exhibit isomerism. However since they have a triple bond in their structure, they can exhibit two types of isomerism.

(i) Positional isomerism

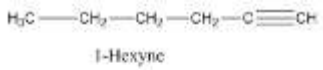
(ii) Structural isomerism.

Positional isomers for alkynes are constitutional isomers that have the same carbon skeleton and the same triple bond but differ from each other in the location of the triple bond in the carbon chain

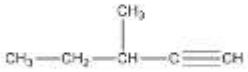


pent-1-yne pent-2-yne

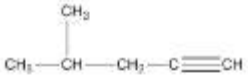
Structural isomers for alkynes are isomers that have the same triple bond in the same position but differ from each other in the connectivity of the carbon skeleton.



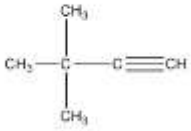
1-Hexyne



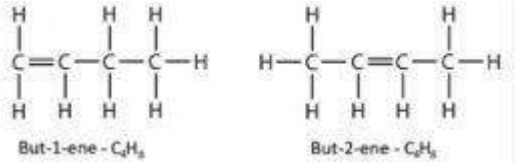
3-Methylpentyne



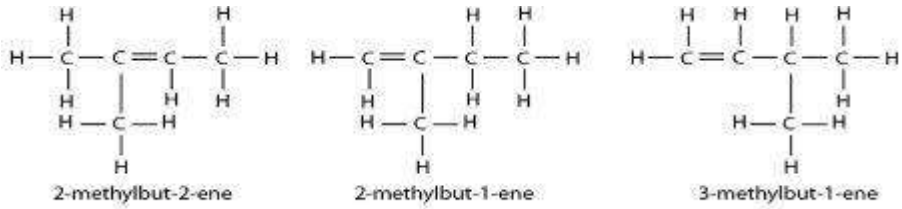
4-Methylpentyne



3,3-Dimethyl-but-1-yne



But-1-ene - C_4H_8 But-2-ene - C_4H_8



2-methylbut-2-ene 2-methylbut-1-ene 3-methylbut-1-ene

Question

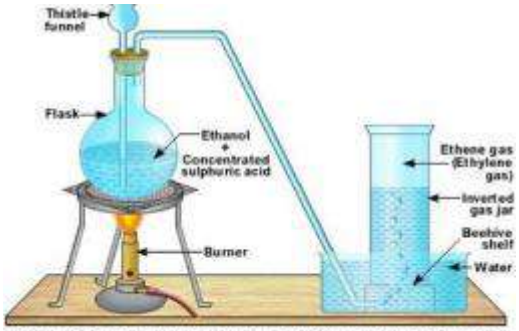
1. What is meant by isomerism?
2. Draw and name two isomers of Butyne

Laboratory Preparation of Alkenes

Preparation of Ethene

Procedure

A mixture of ethanol and concentrated sulfuric acid in the ratio 1:2 respectively are heated in a flask to a temp. of 160^{0}C – 180^{0}C.



Labelled diagram of the apparatus for the laboratory preparation of ethene gas.

Observation

1. A colourless gas results; and is collected over water.

Reason:

1. It is insoluble, unreactive and lighter than water.

Equation

Explanation

At 160^{0}C – 180^{0}C the conc. H_2SO_4 dehydrates the ethanol, removing a water molecule form it and the remaining C and H atoms rearrange and combine to form Ethene which is collected as colourless gas.

$$CH_3\text{-}CH_2\text{-}OH \xrightarrow{\text{conc } H_2SO_4} CH_2{=}CH_2 + H_2O$$

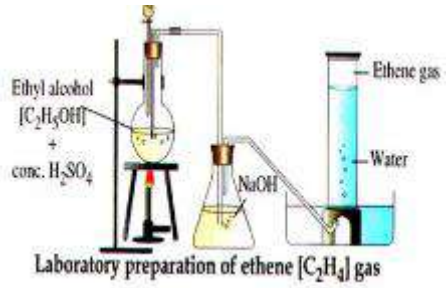
Note:

At temperature below 140^{0}C, a different compound called ether is predominantly formed.

Ethene can also be prepared by passing hot aluminum oxide over ethanol. The later of which acts as a catalyst i.e

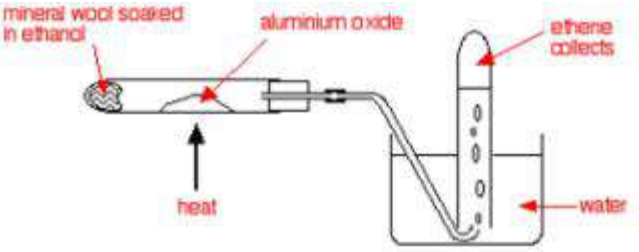
1. The first reaction encountered in this experiment was that for the production of ethene by the dehydration of the alcohol, ethanol, over a heated catalyst, in this case concentrated sulphuric acid at 180^{0}C.
2. The acid catalysts normally used are either concentrated sulphuric acid or concentrated phosphoric(V) acid, H_3PO_4.
3. Concentrated sulphuric acid produces inconvenient and rather messy results, however. Not only is it an acid, but it is also a strong oxidising agent and so it oxidises some of the alcohol to carbon dioxide and at the same time is reduced itself to sulphur dioxide.

4. Both of these gases have to be removed from the alkene before it can be used industrially. These gases which are produced are passed through sodium hydroxide solution to remove the carbon dioxide and sulphur dioxide produced from the side reactions.
5. Sand is used to prevent bumping which may cause cracking of the flask



Laboratory preparation of ethene [C_2H_4] gas

aluminium oxide can also be used.



$$C_2H_5OH \xrightarrow[-H_2O]{Al_2O_3,\ 623K} CH_2{=}CH_2$$

ethanol → ethene

PHYSICAL PROPERTIES

Alkane	General formula	Melting point (°C)	Boiling point (°C)	State at room(298K) temperature and pressure atmosphere (101300Pa)
Ethene	CH_2CH_2	-169	-104	gas
Propene	$CH_3\ CHCH_2$	-145	-47	gas
Butene	$CH_3CH_2\ CHCH_2$	-141	-26	gas
Pent-1-ene	$CH_3(CH_2\ CHCH_2$	-138	30	liquid
Hex-1-ene	$CH_3(CH_2)\ CHCH_2$	-98	64	liquid

•Like alkanes, alkenes are colourless gases, solids and liquids that are not poisonous.

• They are slightly soluble in water.

•The solubility in water decrease as the carbon chain/ molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene.

•Melting/boiling point increase as carbon chain increase.

•This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.

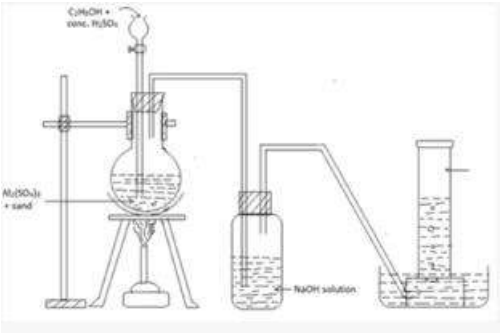
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The 1st four straight chain alkenes (ethene,propene,but-1-ene and pent-1-ene)are gases at room temperature/ pressure

QUESTIONS

(1). Study the diagram below and answer the questions that follow;



a. Name gas Q

b. Write an equation for the formation of gas Q

c. What is the role of the sand in the set up?

d. What name is given to the type of reaction taking place in the set up above?

e. The heating in this experiment should be maintained at 160 degrees - 180 degrees. Explain

Chemical Properties and Uses of Alkenes

Reactions of ethene/chemical properties

(a) Burning/combustion

Just like an alkenes and alkanes, ethene burn in air, producing carbon dioxide and large quantities of heat.

Equation;

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$

Caution: Mixtures of air and ethene can be explosive and must be handled very carefully.

b) Additional reactions:

Is a reaction in which are molecule adds to another to form a single product occur in alkenes due to presence of a double bond.

(1).With oxidizing agents

i) Reaction with acidified potassium permanganate.

Procedure:

Ethene is bubbled into a test tube containing acidified potassium.

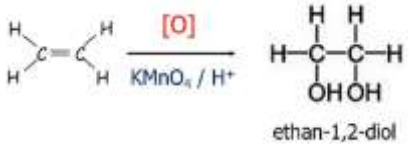
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The purple colour of the solution disappears.

Explanation:

1. Ethene reduces the potassium permanganate.
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1. The net effect of the above reaction is the addition of two –OH groups to the double bond forming ethan-1, 2-dio(ethylene glycol).
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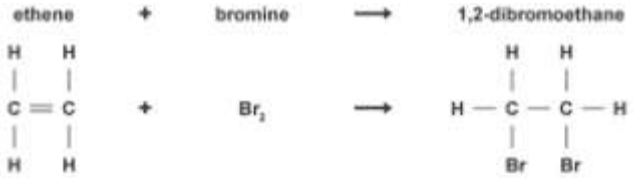
The chlorine (greenish yellow) also gets decolourised due to the addition of its atoms on the double bond.

Note:

Alkenes react with and decolourise halogens and potassium permanganate by additional reaction at room temperature and pressure.

The reaction site is the double bond and hence/all alkenes will react in a similar manner.

Example; Butene and Bromine



(iii) Reaction with Bromine water

Bromine is dissolved in water and reacted with ethene.

Equation:

Further examples of additional reactions

(3). Addition of hydrogen halides

Hydrogen halides reacts with alkene to form a halogen alkane.

The double bond in the alkene break and form a single bond



Ethene Bromoethane

(i). With hydrobromic acid; HBr (aq)

With sulphuric acid

(4)Addition of Ethene with sulphuric acid

Note:

When ethyl hydrogen sulphate is hydrolysed, ethanol is formed.

In this reaction, water is added to ethyl hydrogen sulphate and the mixture warmed.

(5). Ethene with Hydrogen i.e. Hydrogenation.

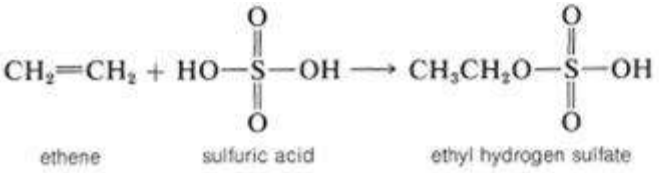
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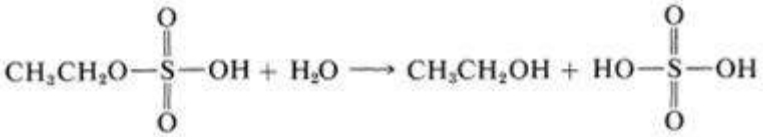
Conditions;

1. moderate temperature and pressure
2. Nickel catalyst/palladure catalyst

Equation:



ethene sulfuric acid ethyl hydrogen sulfate



ethyl hydrogen sulfate ethanol sulfuric acid

Application:

it is used industrially in the conversion f various oils into fats e.g. in the preparation of Margarine.

(6). Polymerization reactions.

Also called **self-addition reactions**

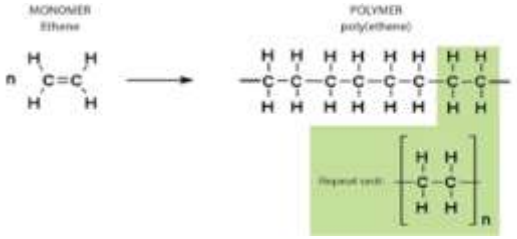
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Examples:

Polymerisation of ethene



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1. High temperatures of about 200^{0}C
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ii) Procedure: Ethene is heated at 200^{0}C and 1000 atm. Pressure over a catalyst.

iii) Observation: Sticky white substance which hardens on cooling is formed. This solid is polyethylene, commonly referred to as polythene.

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Generally

Uses of polythene

1. Used for the manufacture of many domestic articles (bowls, buckets, water cans, and cold water pipes) e.t.c. **Note:** Polythene pipes have a great advantage over metal pipes as they can be welded quickly and do not burst in frosty weather.
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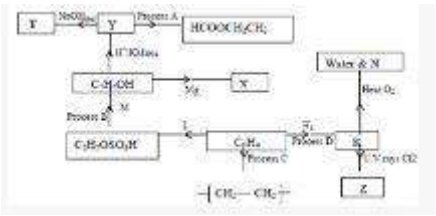
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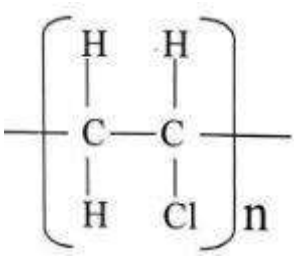
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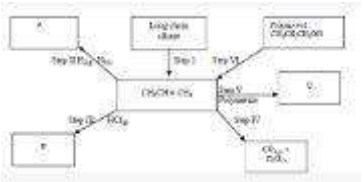
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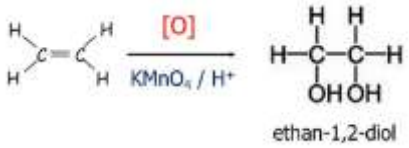
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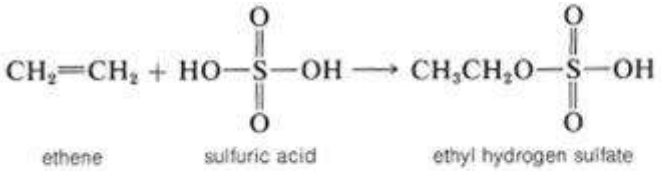
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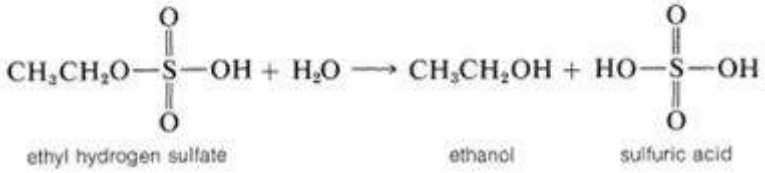
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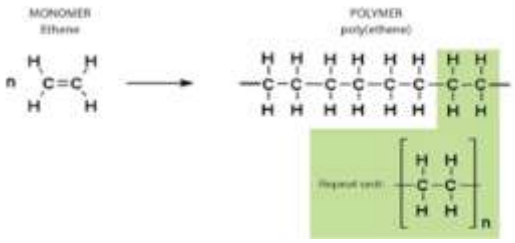
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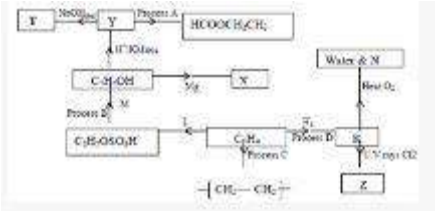
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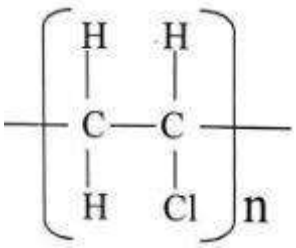
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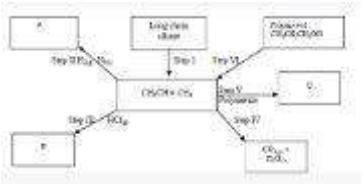
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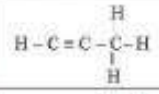
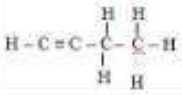
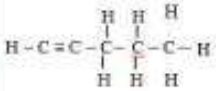
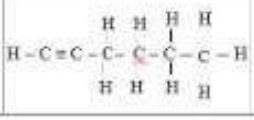
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Reagent;

Condition;

ALKYNES

1. These are hydrocarbons with a general formula
2. CnH2n-2 and triple bond as the functional group .
3. n is the number of Carbon atoms in the molecule.
4. The carbon atoms are linked by at least one triple bond to each other and single bonds to hydrogen atoms.
5. The triple bond in alkyne is the functional group.
6. The functional group is the reacting site of the alkynes.
7. The triple bond in alkyne can easily be broken to accommodate more /four more monovalent atoms.
8. The triple bond in alkynes make it thus unsaturated like alkenes.
9. Most of the reactions of alkynes like alkenes take place at the triple bond.
10. Their name end with a suffix – **yne** after the prefix
11. The alkynes have the prefix **,eth_,prop_** and **but_** to represent 1,2,3 and 4 carbons in the compound. All other use the numeral prefix **pent_,Hex__, hept__**.

Name	Open structure	Condensed structure
Ethyne	H – C ≡ C – H	CH≡CH
Propyne		CH≡C–CH$_3$
Butyne		CH≡C–CH$_2$–CH$_3$
Pentyne		CH≡C–CH$_2$–CH$_2$–CH$_3$
Hexyne		CH≡C–CH$_2$–CH$_2$–CH$_2$–CH$_3$



Ethyne

(Acetylene)

C_2H_2



butyne

C_4H_6



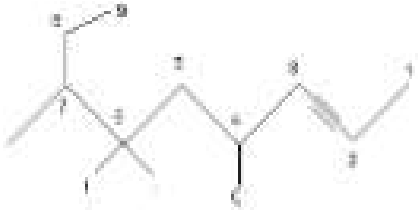
hexyne

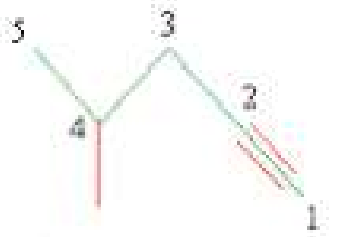
C_6H_{10}

Naming of Alkynes

RULES

1. Identify the longest continuous/straight carbon chain which contains the triple bond to get/determine the parent alkene
2. Number the longest chain form the end of the chain which contains the triple bond so as triple bond get lowest number possible.
3. Indicate the positions by splitting “alk-position-yne” e.g. but-2-yne, pent-1,3-diyne
4. Determine the position, number and type of branches.

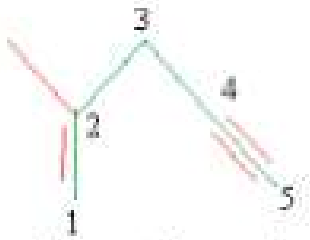




4-methyl-1-pentyne

or

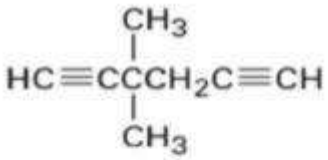
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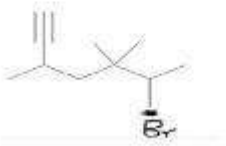


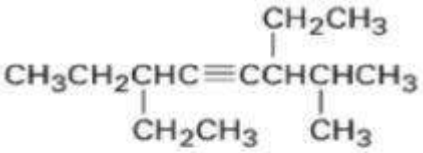
2-methylpent-1-en-4-yne

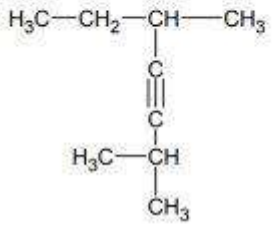
QUESTIONS

(1). Name the following compounds









(2). Name or draw out the following molecules;

a. 4,4-dimeth7yl-2-pentyne

b. 4-penten-1-yne

c. 1-ethyl-3-dimethylnonyne

Isomerism of Alkynes

Just like alkanes, alkynes exhibit isomerism. However since they have a triple bond in their structure, they can exhibit two types of isomerism.

- Positional isomerism
- Structural isomerism.
- **Positional isomers** for alkynes are constitutional isomers that have the same carbon skeleton and the same triple bond but differ from each other in the location of the triple bond in the carbon chain
- **Structural isomers** for alkynes are isomers that have the same triple bond in the same position but differ from each other in the connectivity of the carbon skeleton.

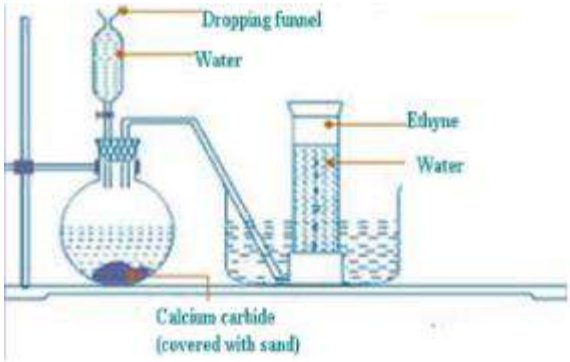
EXCERCISE

a. What is meant by isomerism?

b. Draw and name two isomers of Butyne

Preparation of Alkynes

1. Ethyne is prepared from the reaction of water on calcium carbide.
2. The reaction is highly exothermic and thus a layer of sand should be put above the calcium carbide to absorb excess heat to prevent the reaction flask from breaking. Copper(II)sulphate is used to catalyze the reaction



$$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$$

Calcium carbide — Water — Calcium hydroxide — Ethyne

Properties of alkynes

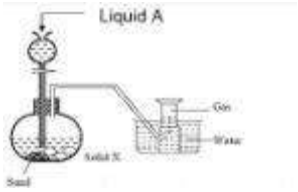
Physical properties

1. Like alkanes and alkenes, alkynes are colorless gases, solids and liquids that are not poisonous.
2. They are slightly soluble in water. The solubility in water decrease as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetra chloromethane and methylbenzene. Ethyne has a pleasant smell when pure.
3. The melting and boiling point increase as the carbon chain increase. This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase

Alkyne	General formula	Melting point(°C)	Boiling point(°C)	State at room(298K) temperature and pressure atmosphere (101300Pa)
Ethyne	CH CH	-82	-84	gas
Propyne	CH_3 C CH	-103	-23	gas
Butyne	CH_3CH_2 CCH	-122	8	gas
Pent-1-yne	$CH_3(CH_2)_2$ CCH	-119	39	liquid
Hex-1-yne	$CH_3(CH_2)_3$C CH	-132	71	liquid

QUESTIONS

The diagram below shows the set up of an apparatus used in the preparation of ethyne gas. Study it and answer the questions that follow.



a. Identify Liquid A and solid X

b. What is the role of sand in the experiment?

c. Write a chemical equation for the reaction.

d. Identify the method of gas collection and give reasons why ethyne is collected by the above method.

e. It is necessary to ensure that the flask is dry before the experiment starts. Give a reason.

f. Identify a catalyst that can be used in this process.

Chemical Properties of Alkynes

Chemical properties

Burning/combustion

1. Alkynes burn with a **yellow**/ luminous very **sooty** flame in **excess** air to form carbon(IV) oxide and water.
2. Alkenes burn with a **yellow**/ luminous very **sooty** flame in **limited** air to form carbon(II) oxide/carbon and water.
3. Burning of alkynes with a **yellow**/ luminous very **sooty**/ smoky flame is a confirmatory test for the **presence** of the triple bond because they have very **high C:H ratio.**

$$2C_2H_2 + O_2 \longrightarrow 4C + 2H_2O$$

Ethyne Oxygen Carbon Water

$$2C_2H_2 + 3O_2 \longrightarrow 4CO + 2H_2O$$

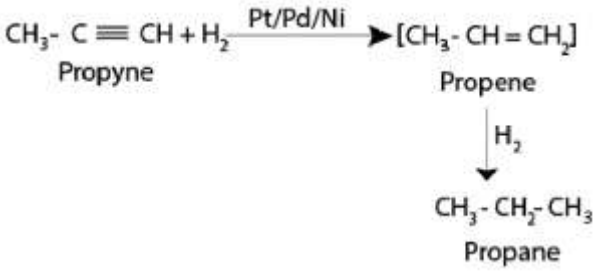
Ethyne Oxygen Carbon monoxide Water

Addition reactions

1. An addition reaction is one which an unsaturated compound reacts to form a saturated compound.
2. Addition reactions of alkynes are also named from the reagent used to cause the addition/convert the triple to single C- C bond

Hydrogenation

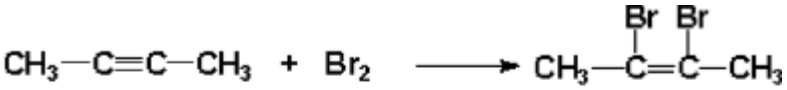
1. Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at 200^{0}C temperatures react with alkynes to form alkenes then alkanes.
2. During hydrogenation, two hydrogen atom in the hydrogen molecule attach itself to one carbon and the other two hydrogen to the second carbon breaking the **triple** bond to **double** the **single**.

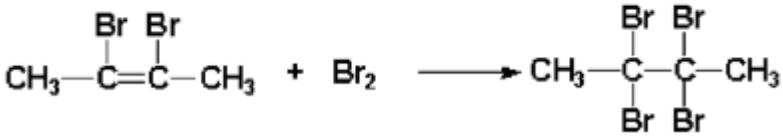


Halogenation.

1. Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alkyne to form an halogenoalkene then halogenoalkane.
2. The reaction of alkynes with halogens is faster than with alkenes. The triple bond in the alkyne break and form a double then single bond.
3. The color of the halogen fades as the number of moles of the halogens remaining unreacted decreases.

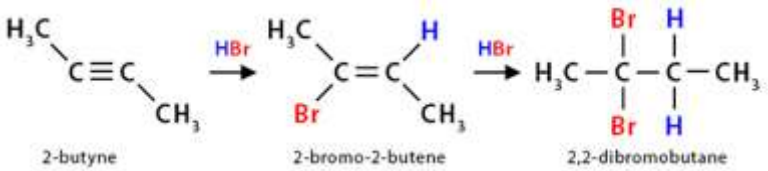
Two bromine atoms bond at the 1st carbon in the triple bond while the other two goes to the 2nd carbon





Hydrohalogenation

1. **Hydrogen halides** reacts with alkyne to form a halogenoalkene then halogenoalkane.
2. The triple bond in the alkyne break and form a double then single bond.
3. The main compound is one which the **hydrogen** atom bond at the carbon with **more hydrogen .**
4. This addition Markovnikov's rule that states, a hydrogen adds to the carbon that has the greater number of hydrogens, and halide adds to the carbon with the fewer hydrogens. Therefore, the end product will always be a 1,1-dibromo or 1,1-dichloro or 1,1-diiodo



Uses of Ethyne

(1). Industrial manufacture of compounds like adhesives and plastics

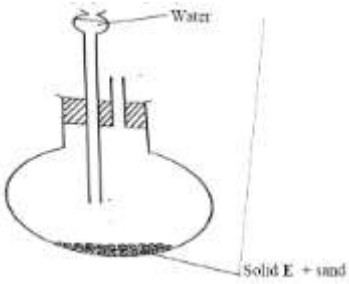
(2). It's used in the oxy-acetylene flame which is used for welding and cutting metals.

QUESTIONS

(1). Illustrate the gradation in physical properties of alkynes.

(2). List the chemical properties of ethyne.

(3). The set-up below was used to prepare ethyne gas



a. Identify solid E

b. Complete the diagram to show how the gas can be collected

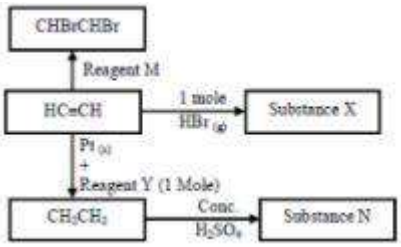
c. Write an equation to show how the gas is formed

d. Complete the equation below:

$C_2H_2 + 2I_2 \rightarrow$

e. What is the role of sand in the experiment?

(4). The scheme below shows some reactions starting with ethyne. Study it and answer the questions that follow.



a. Name substance X and N

b. Name reagent M.

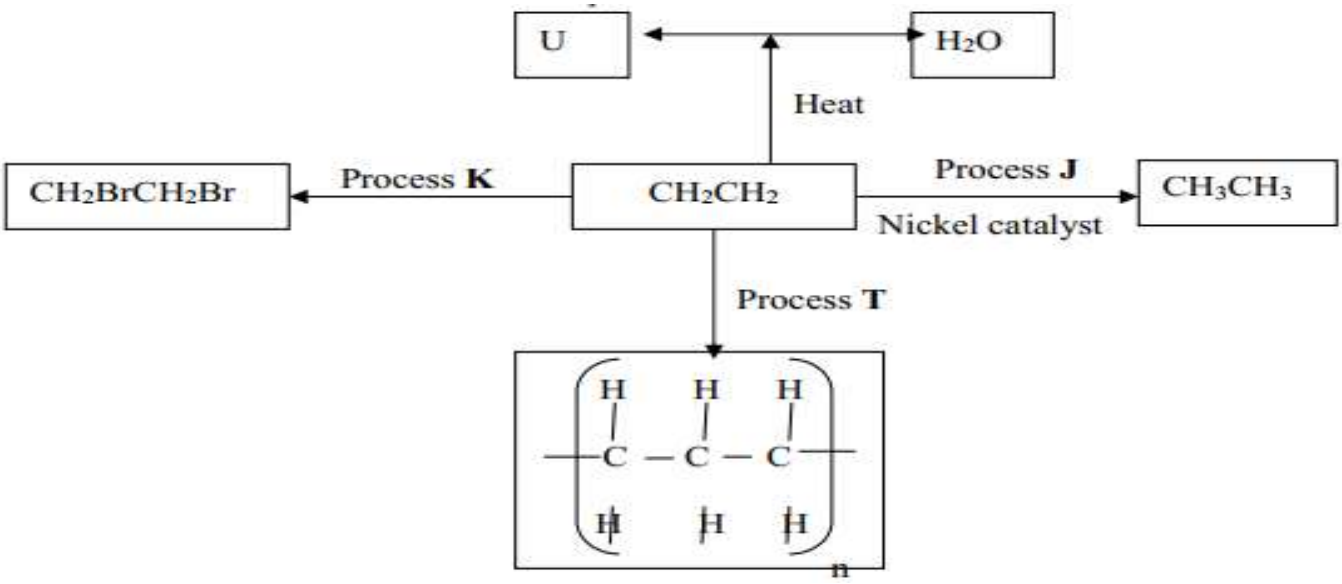
(5). Ethene undergoes polymerization to form a polymer. Give an equation for the reaction and name the product.

(i). Equation;

(ii). Name:

TOPICAL REVISION QUESTIONS

1. Use the flow chart below to answer the questions that follow:



   a. What observation would be made in process **K**?

   b. Name another conditions necessary for process **J** to take place

   c. Give the name of substance V

2. But-z-ene undergoes hydrogenation according to the equation given below
   $CH_3CH = CHCH_{3\ (g)} + H_{2(g)} \rightarrow CH_3CH_2CH_2CH_{3(g)}$

   a. Name the product formed when but-z-ene reacts with hydrogen gas

   b. State one industrial use of hydrogenation

3. Write the structures of the following compounds:-

   a. But-2-yne

   b. 2,2-dimethylpropane

4.

   a. What is meant by Isomerism?

   b. Draw and name two Isomers of butane.

5. Study the information in the table below and answer the questions that follow:

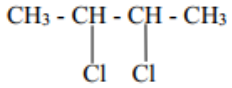
Ion	No. of protons	No. of electrons
P^{3-}	7	10
Q^{+}	19	18
R^{2+}	12	10

6.

   a. Write the electron arrangement of element P.

   b. Give the group and period to which elements Q and R respectively.
      Q .....
      R .....

7. Compound W reacted with chlorine to form compound X only. The structural formula of X is shown below:



   a. Give the structural formula and name of compound W

   b. Name compound X .....

8. In petrol chemical industries, long chain alkanes are broken down in to simpler substances in a process called cracking

   a. Why is cracking necessary?

   b. State the two conditions required in cracking

   c. Draw the structure of 1-chloro-2, 2-dimethylpropane

9. In a reaction an alcohol K was converted to hex-1-ene
   Name reagent and condition necessary for the reaction above to occur

10.

   a. Give the IUPAC systematic names of compounds Q and R
   Q: $CH_2CHClCHClCH_2CH_3$
   R: $CH_3CHClCH_2ClCH3$

   b. The organic compounds **Q** and **R** in (a) above, are formed when one mole of hydrocarbon b reacts with two moles of hydrogen chloride gas;

      i. Structural formula of N

      ii. The IUPAC systematic name of N

11. Distinguish between the isotopes and isomers

12. Polymerisation of ethene takes place as shown in the equation below



   Name the type of polymerisation undergone by ethene in the reaction above

13. State Gay Lussac's law

14. $10cm^3$ of methane (CH_4) gas is exploded with $150cm^3$ of air containing 20% oxygen and 80% nitrogen. The products were allowed to cool to room temperature. What will be the total volume of the gases at the end of the reaction?

15. Give the open structures of:-

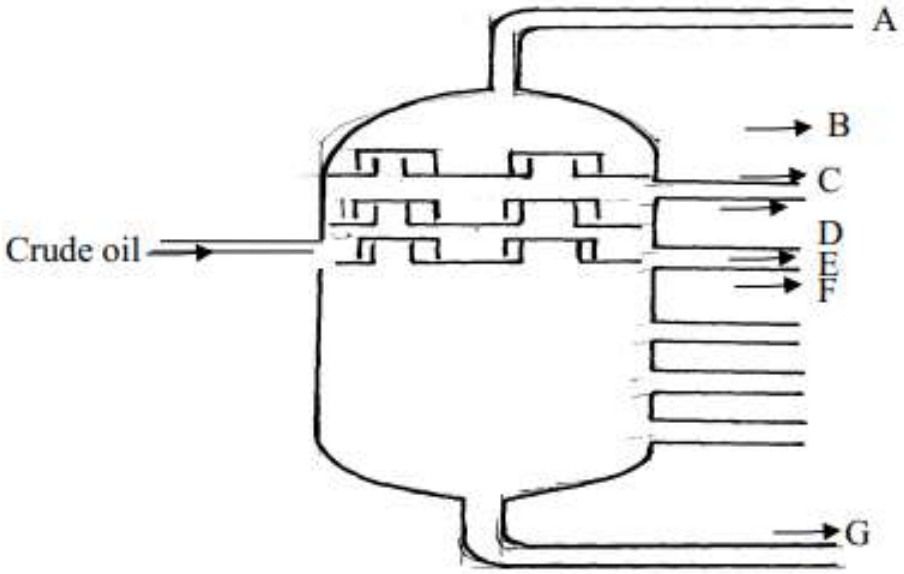
   a. 3-chlorohex-1-yne

   b. CH_3OH

16. A fixed mass of gas occupies $105cm^3$ at $-14^{\circ}C$ and 650mmHg pressure. At what temperature in degrees Celsius will it have a volume of $15cm^3$ if the pressure is adjusted to 690mmHg pressure?

17. Write an equation for the reaction that takes place between ethene and concentrated Sulphuric (VI) acid

18. Petroleum (crude oil) is a mixture of several compounds which are separated in a Changamwe refinery by means of apparatus as shown below:



a.

i. What is the name of the apparatus above

ii. What is the name of the process which is used in separation of crude oil

iii. What physical property of compounds in the mixture does the separation depend

iv. Use the letter A to G to describe where the following could be formed:.

I. The fraction that represents gases

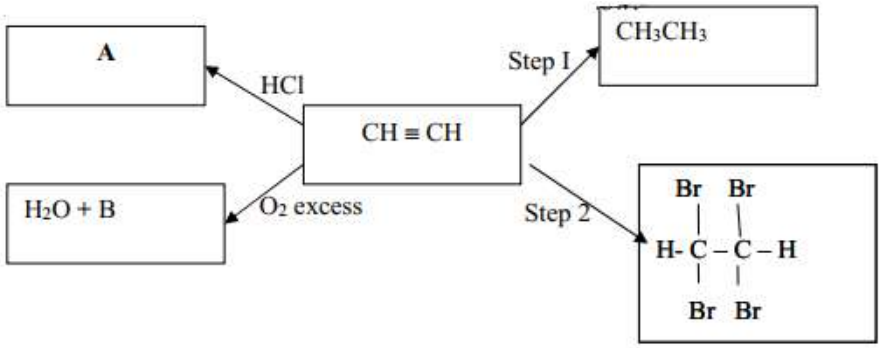
II. The fraction that represents the largest molecules

III. The fraction that represents liquids with the lowest boiling points

b. State the use of product produce at
G.....
C.....

c. Draw apparatus for the separation of the product produce at D and water

19. Study the flow chart below and answer the questions that follow:



a. Give the name of the substance CH ≡ CH .....

b. To which group of hydrocarbons does the substance in (i) above belong?

c. Give two reagents that can be used to prepare the substance named in (i) above

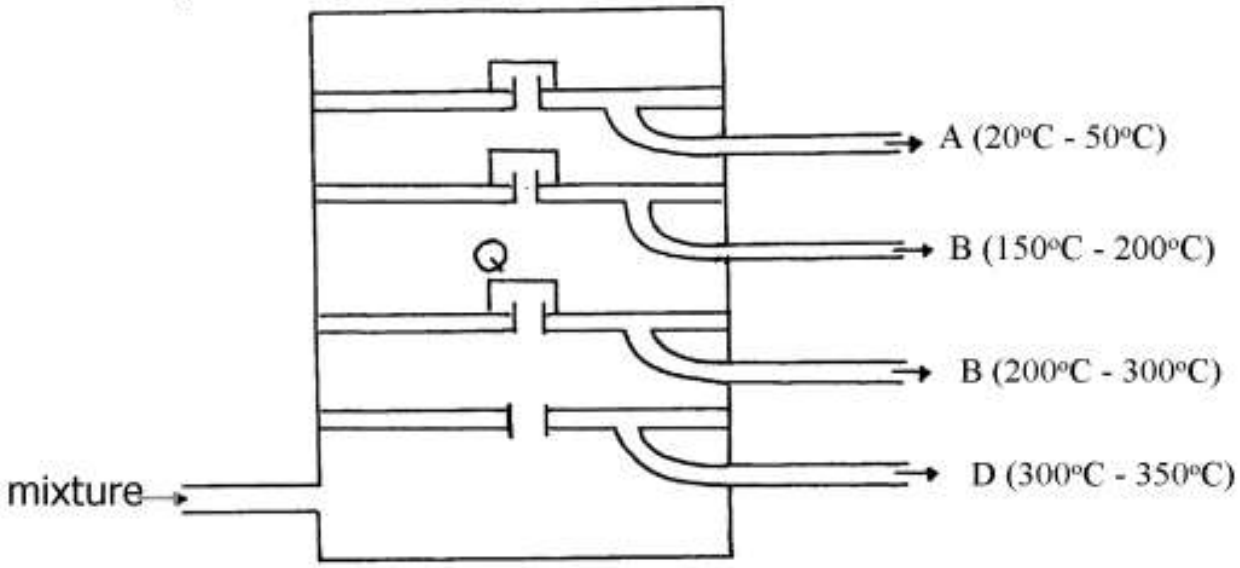
d. State two physical properties of the substances in (i) above

e. Give the names to the process in step I and 2

f. Write an equation to show how substance A is formed

g. Identify substance B .....

20. The diagram below represents a large-scale fractional distillation plant used to separate the components A, B, C and D in a mixture



The components have the following average relative molecular masses not necessarily in that order; 282, 184, 44 and 128.

a. What is the physical state of B at the position marked Q?

b. Which component has an average relative molecular mass of 128? Explain

c. State with a reason whether C is pure or impure

d. Explain how the mixture is separated into its components

e. Name two naturally occurring mixtures that are separated using this process

21.

a. The table below gives information about the major constituents of crude oil. Study it and answer the questions that follow:

Constituent	**Boiling point °C**
Gases	Below 40
Petrol	40 -175
Kerosene	175 - 250
Diesel	250 - 350
Lubricating oil	350 - 400
Bitumen	Above 400

b.

i. Which of the constituents of crude has molecules with the highest number of carbon atoms? Explain

ii. Name the process you would use to separate a mixture of petrol and diesel and explain how the separation takes place

iii. Explain why the constituents of crude oil do not have a sharp boiling point

iv. Name the gas that is likely to be a constituent of crude oil and write its formula

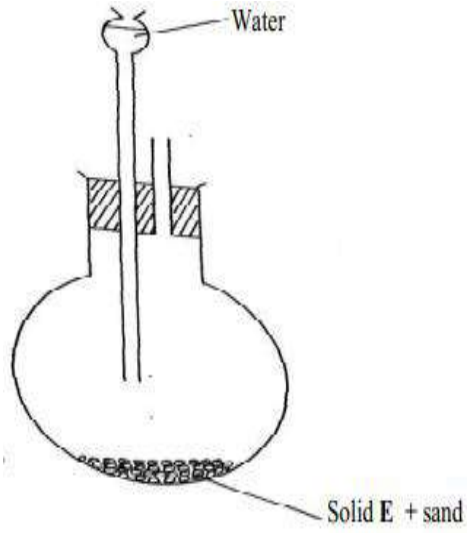
c.

i. What condition could cause a poisonous gas to be formed when kerosene is burnt. Explain

ii. Give one use of bitumen

22.

a. The set-up below was used to prepare ethyne gas



i. Identify solid E

ii. Complete the diagram to show how the gas can be collected

iii. Write an equation to show how the gas is formed

iv. Complete the equation below:
$C_2H_2 + 2I_2 \rightarrow$

v. What is the role of sand in the experiment?

b.

i. Explain the meaning of esterification

ii. Complete the equation below :
$CH_3COOCH_3 + H_2O \rightarrow$

iii. What type of reaction is occurring above

c. Given the reaction:

$$C_8H_{18} \xrightarrow{\text{solid F}} N + C_2H_4$$

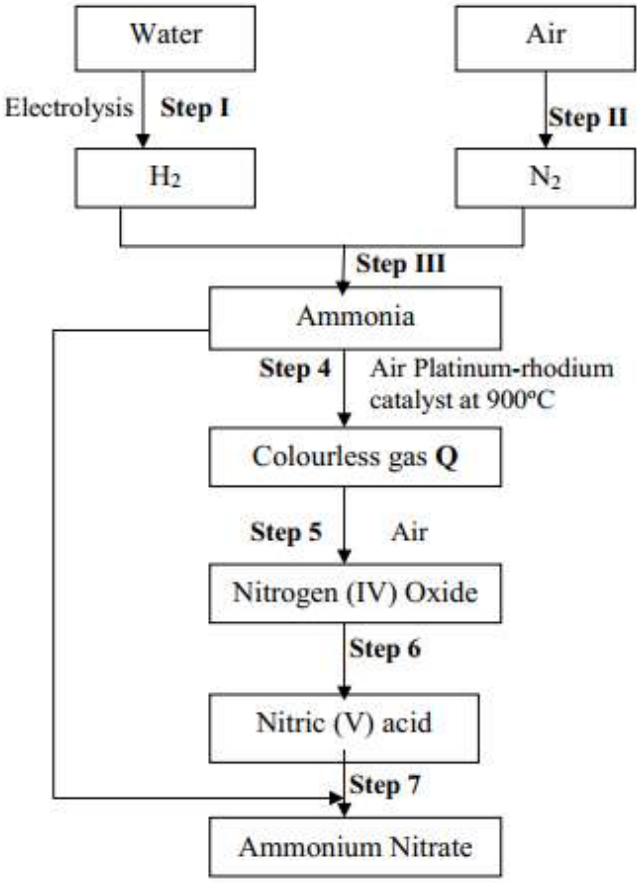
i. Identify substance:
F..... **N**.....

ii. Name the process represented above?

d. Give **one** use of substance **N**

23.

a.



i. Name another source of hydrogen apart from electrolysis of water

ii. What conditions are necessary for step III to occur?

iii. Write the equation for the formation of colourless gas Q

iv. Give one use of nitric (V) acid

b. State and explain the observations that would be made if a sample of copper metal is heated with concentrated nitric (V) acid

24.

a. Give the systematic names of the following compounds:-

i.

$$CH_2 = \underset{\displaystyle Br}{\overset{\displaystyle |}{C}} - CH_3$$

.....

ii. $CH_3CH_2CH_2C \equiv CH$ .....

b. State the observations made when butan-1-ol reacts with:-

i. Acidified potassium dichromate (VI) solution

ii. Potassium metal

c. Ethanol obtained from glucose can be converted to ethene as shown below:-

$C_6H_{12}O_6 \xrightarrow{\text{Step I}} C_2H_5OH \xrightarrow{\text{Step II}} CH_2 = CH_2$

Name and describe the processes that take place in steps I and II

d. Compounds A and B have the same molecular formula $C_3H_6O_2$. Compound **A** librates Carbon (IV) Oxide on addition of aqueous sodium carbonate while compound **B** does not. Compound **B** has a sweet smell. Draw the possible structures of:-

i. Compound A

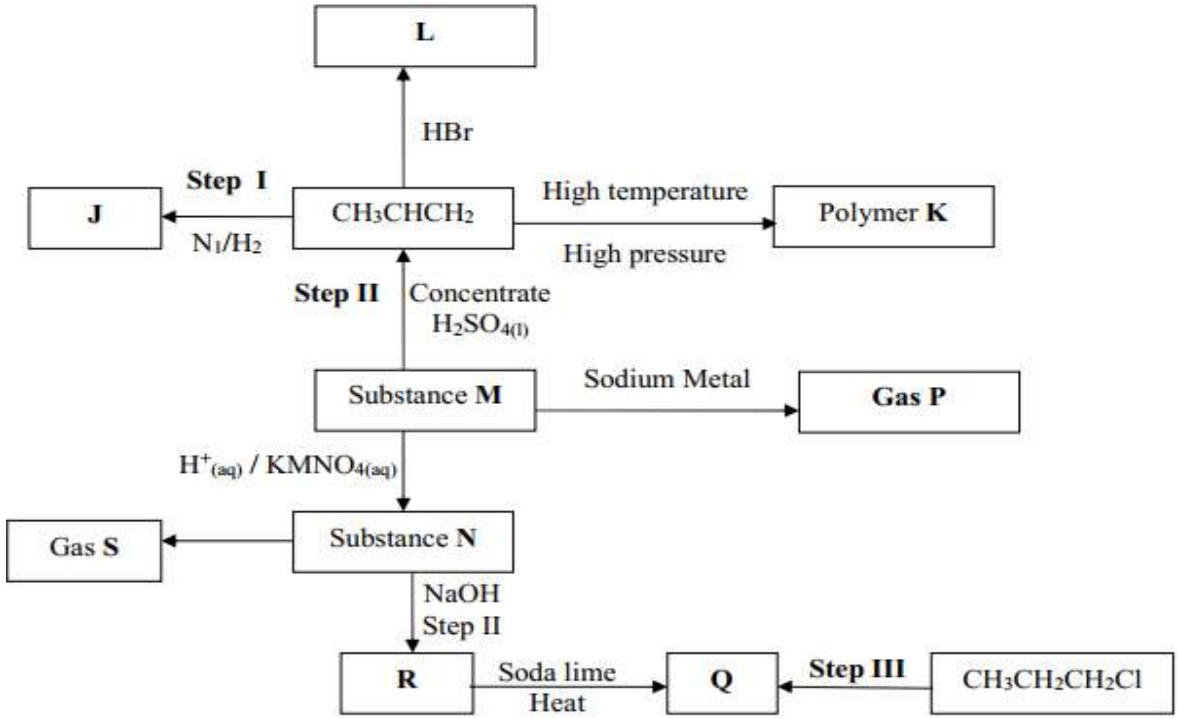
ii. Compound B

e. Give two ways how the disposal of polymers such as polychloroethene by burning pollutes the environment

25.

a. Name the following compounds $(CH_3)_3CCH_2CH_2CH_3$

Use the flow chart below to answer the questions that follow:-



b.

i. Name the following :-

I. Gas S .....

II. Gas P .....

III. J .....

ii. Name the processes involved in the following steps:

I. Step I .....

II. Step II .....

III. Step III .....

iii. Write a chemical equation for the complete combustion of substance M

iv. Name the condition and reagent in step III
Condition ……
Reagent ……

v. Calculate the mass of salt R that would be formed by using 21.9 tonnes of N when it reacts with excess sodium hydroxide (C= 12.0 H= 1.0 Na = 23)

vi.

I. Draw the structure of polymer K

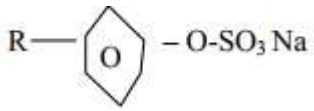
II. State one use of the above polymer
……

c.

i. Name the class to which the following cleansing agents belong:-

I. $R - COONa^+$ ……

II.



……

ii. Which cleaning agent above is not environmental friendly? Explain

26. The molecular formula of a hydrocarbon is C_6H_{14}. The hydrocarbon can be converted into two other hydrocarbon as shown by the equation below:
$C_6H_{14} \rightarrow C_2H_6 + \mathbf{X}$

a. Name and draw the possible structural formula of **X**

b. State and explain the observations that would be made if a few drops of bromine water were added to a sample of **X**

c. Write an equation for the complete combustion of C_3H_8

27.

a. Give the names of the following

i. $CH_3CH_2CH_3$

ii. CH_3CCCH_3

b. Ethene is used in making polyethene bag in a process called polymerization

i. Name the type of polymer that is formed when ethane polymerise

ii. Describe a simple chemical test that can be used to identify ethane gas in the laboratory

c. Study the information in the table below and answer the questions that follow:-

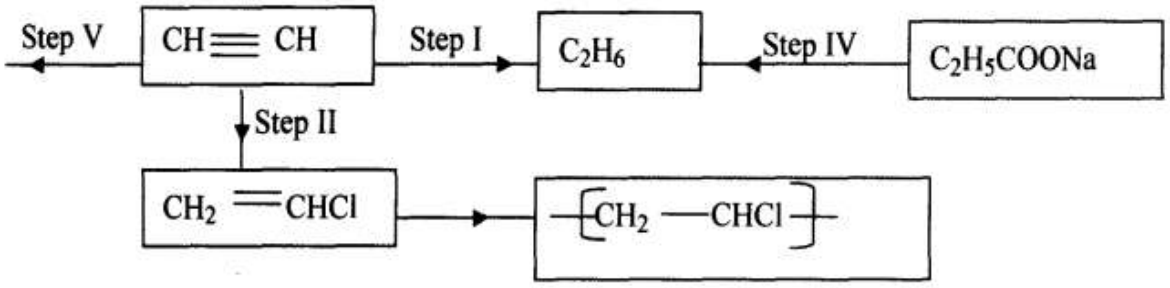
No. of carbon atoms	R.M.M of the Hydrocarbon
2	28
3	42
4	56

d.

i. Write the general formula of the hydrocarbons in the table above

ii. Determine the molecular of a hydrocarbon with 5 carbon atoms and draw its structural formula
Molecular formula
Structural formula

e. Study the scheme below and answer the questions that follow



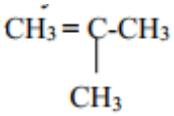
i. Name the reagents in
Step I .....
Step II .....
Step IV .....

ii. Write an equation for the complete combustion of CH≡CH

iii. Give two uses of CH_4

28. Give the systematic names of the following compounds;

a.



b. $CH_3CH_2CH_2C\equiv CH$

29. Study the data given in the following table and answer the questions that follow. The letters are not the actual symbols of elements.

Element	Number of protons	Melting Point	Bpt °C
A	11	98	890
B	12	650	1110
C	13	60	2470
D	14	1410	2360
E	15	442 590	280
F	16	113 119	445
G	17	-101	-35
H	18	-189	-386

30.

a. State and explain the trend in melting point in **A B C**

b. Explain why the melting point and boiling points of element **D** is the highest

c. Explain why the element represented by letter **E** has two melting point values

d. Write down the chemical formula between element **C** and sulphate ions

e. Name the chemical family in which **H** belong and state one use of the element

f. What is the nature of the oxide of the elements represented by letters **C** and **F**?

31. Study the information in the table below and answer the questions that follow

No. of carbon atoms	**R.M.M of the Hydrocarbon**
2	28
3	42
4	56

32.

a. Write the general formula of the hydrocarbons in the table

b. Predict the relative atomic mass of the hydrocarbons with 5 carbon atoms

c. Determine the relative atomic mass of the hydrocarbon in (ii) above and draw its structural formula (H=1.0, C=12.0)

33. Substance "M" with a general formula C_2H_y burnt in chlorine gas with a red flame producing a cloud of black specks and colourless gas G.

a. State the collective name for compounds which 'M' belongs

b. With reason, state the identity of the black specks and colour gas "G".

34.

a. 2.63g of a solution of sodium chloride at 20.0^oC was reacted with silver nitrate. After filtration, washing and drying, 2.36g of silver chloride was obtained. Determine the solubility of sodium chloride at 20.0^oC . (Na=23, Cl= 35.5, Ag = 108)

b. Determine the number of moles of carbon (IV) Oxide gas produced when sodium carbonate reacted with dilute sulphuric (VI) acid (Molar gas volume =24dm^3)

35. Write down all the isomers of but-z-ene and give their IUPAC names

36. A hydrocarbon compound Z decolourizes bromine liquid in the presence of light but does not decolourize acidified potassium manganate (VII). Name and draw the structural formula of the eighth member of this homologous series

37.

a. What is meant by isomerism?

b. Draw and name two isomers of Butyne

Answers

1.

a. Bromine decolorized immediately in ethane gas ✓1

b. Temperature between150°C - 250°C or temperature of 180°C

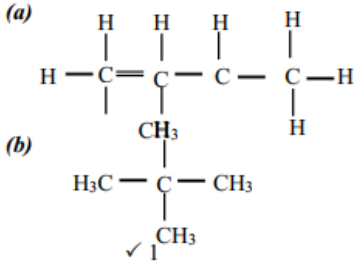
c. arbon (IV) oxide or CO2(g) ✓

2.

a. Butane

b. Manufactures of cooking fats and margarine

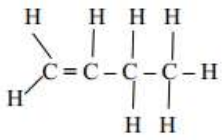
3.



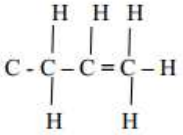
4.

a. Existence of cpds with the same molecular formula but different structural formula/arrangement of atoms

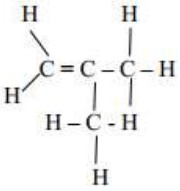
b.



n – butane/ ✓½
1 – butane/
But-1-ene



2 – butane/ ✓½
But-2-ene



2 - methyl
prop-1-ene

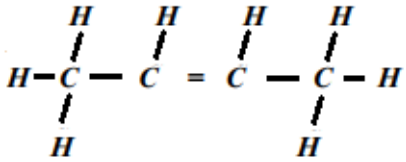
5.

a. 2.5

b. Q Group 1 ✓½, Period 4 ✓½
R Group 2 ✓½, Period 3 ✓½

6.

a.



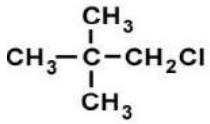
b. 2,3 dichlorobutane

7.

a. To produce simpler hydrocarbons of industrial importance e.g. ethane which is widely used

b. - Elevated temperature / high temperature 900 C
- Catalyst

c.



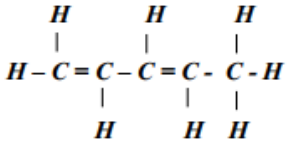
8. Reagent: concentrated sulphuric acid
   Condition: temperature 180^{o} C

9.

   a. $H_2CHClCHClCH_2CH_3$
      Name: 2, 3 dichloropentane

   b.

      i. Structural Formula



      ii. IUPAC name
      pent – 1,3 – diene

10. Isotopes are atoms of the same element with same atomic number but different mass numbers while isomers are compounds with the same molecular formula but different structural formula

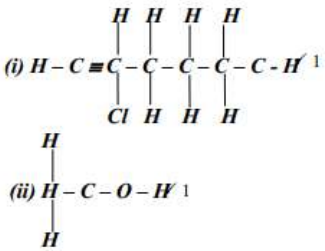
11. Addition polymerization. √1

12.

   a. When gases combine they do so in volume which bear a simple ratio to one another and to the product if gaseous under standard temperature and pressure

13. $CH_4 + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_{2(l)}$
    $10cm^3 \quad 20cm^3 \quad 10cm^3$ √½
    Volume of $O_2 = {}^{20}/_{100} \times 150$
    $= 30cm^3$
    Remaining volume of $O_2 = 30 - 20 = 10\ cm^3$
    Total volume of the gases = 20+10+10
    $= 40cm^3$

14.



15. $T_2 = \dfrac{690 \times 15 \times 259}{650 \times 105}$ √

= 39.3K✓
= - 233.7° C✓

16. $CH_2 = CH_{2(g)} + H_2SO_{4(l)} \rightarrow CH_3CH_2OSO_3H(aq)$ ✓ 1 mark

17.

a.

i. Fractional Column.

ii. fractional distillation.

iii. different boiling points.

iv.

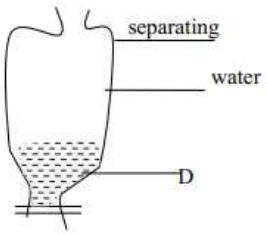
I. A

II. F

III. B

b. G – road making or water proofing
C jet fuel or cooking and lighting.

c.



18.

a. ethyne

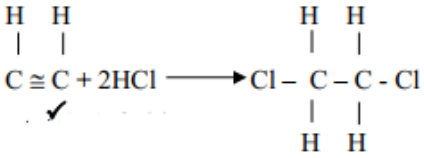
b. Alkynes – because it has triple bond between the two carbon atoms

c. Water and calcium carbide

d. - Colourless, odourless
- less denser than air
- Insoluble in water but soluble in organic solvents

e. Step 1 - Hydrogenation
Step 2 - Halogenations

f.



g. Carbon (IV) Oxide

19.

a. Gas /vapour

b. B - It has the second lowest boiling point thus second lowest molecular mass

c. C is impure since it boils over a range of temperature

d. It is boiled heated and the vapour of the components condense at different temperatures

e. - Liquid air
- Crude oil

20.

a.

i. Bitumen it has the highest boiling point

ii. Fractional distillation; during distillation petrol would distill off at 175^{o}C, while diesel will distill at 35^{0}C

iii. Each component is a mixture of hydrocarbons which have different boiling points

iv. Methane, CH_4, EthaneC_2H_6 propane, C_3H_8, Butane C_4H_{10}

b.

i. Burning in limited amount of air will produce carbon monoxide (carbon (II) Oxide) which is poisonous

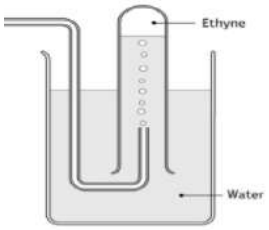
ii. Manufacture of Tar used in road tarmacking sealing of leakages on roofs

21.

a.

i. Calcium carbide – CaC_2

ii. Over water method



iii. $CaC_{2(s)} + 2H_2O_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + C_2H_{2(g)}$

iv. $C_2H_2 + 2I_2 \rightarrow C_2H_2I_2$

v. The reaction if highly exothermic hence sand helps to absorb excess heat.

b.

i. A reaction in which an organism acid reacts with an alkanol to form a sweet smelling compound called ester.

ii. $CH_3COOCH_3 + H_2O \rightarrow CH_3COOH + CH_3OH$

iii. Hydrolysis

c.

i. F – Aluminium oxide – Al_2O_3
N – C_6H_{14} – Hexane

ii. Cracking

d. A fuel

22.

a.

i. Cracking of crude oil fractions. ✓1

ii. Temp – 400 – 500^0c
Pressure – 200 – 500 atmospheric Any 2 = 1
Catalyst – Finally divided iron.

iii. $4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(l)}$

iv. - Manufacture of nitrate fertilizers. ✓1
- Manufacture of explosives.
- Purification of metals.

b. - Red brown gas ✓1 with pungent irritating smell due to reduction of HNO_3 to NO_2
- Blue ✓1 solution due to formation of Cu $(NO_3)_2$

23.

a.

i. 2-bromo propene or 2- bromo prop-i-ene

ii. Pent-i-ene

b.

i. Changes form orange to Green

ii. Effervescence//bubbles of gas produced

c. <u>Step 1</u>
- Fermentation of glucose
Glucose broken down in obscene of oxygen using enzymes
- Dehydration of ethanol; using concentrated sulphuric (VI) acid and high temperature of 170^oC

<u>Step II</u>
- Dehydration of ethanol; using concentrated sulphuric (VI) acid and high temperature of 170^oC

d. Compound A

e. - release chlorine gas which destroy ozone layer
- Chlorine gas combines with vapour in atmosphere to form acid rain which destroy vegetation
- Chlorine gas can cause respiratory diseases

24.

a. 2,2 – dimethyl pentane

b.

i.

I. carbon IV oxide.

II. Hydrogen gas.

III. Propane.

ii.

I. Hydrogenation.

II. Neutralization

III. substitution

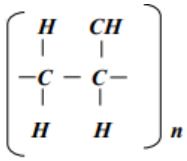
iii. $CH_3CH_2CH_2O_{(l)} + 9O_{2\,(g)} \rightarrow 6\,CO_{2\,(g)} + 8\,H_2O_{(l)}$

iv. Condition Presence of U.V light
Reagents – Chlorine gas

v. $CH_3CH_2COOH + NaOH \rightarrow CH_3CH_2COONa + H_2O_{(l)}$
Mole ratio :
74 tones of acid = 96 tones of salt
21.9 = ?
$\frac{21.9 \times 96}{74}$ = 28.4 tones
Or $\frac{21.9}{74}$ = 0.29 moles of salt
= 0.29 × 96 = 28.4 tones

vi.

I.



II. use in making – Plastic crates plastic boxes plastic ropes

c.

i.

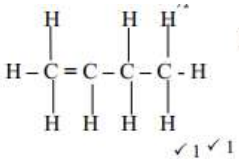
I. soapy detergent

II. Soap less detergent

ii. Soap less Detergent - non biodegradable.

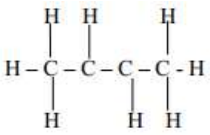
25.

a.



✓1 ✓1

But-1-ene *Or (accept any 1)* *But-2-ene*



✓1 ✓1

b. Bromine water is decolourised because X is unsaturated or has a (-C = C-) double bond.

c. $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

26.

a.

i. Propane

ii. But - 2 - yne

b.

i. Ploythene

ii. Bubble pass ethane gas in acidified KMnO4 or acidified $K_2Cr_2O_7$

c.

i. C_nH_{2n}

ii. C_5H_{10}

d.

i. Step I – hydrogen
Step II – Hydrogen chloride
Step IV – Sodalime

ii. $2CH = CH_{(g)} + O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(l)}$
- A fuel
- Manufacture of methanol
- Manufacture of methanol

27.

a. 2 – Methylprop – 1 - ene

b. Pent – 4 – yne
✓ 1 mark
✓ 1 mark [Total 12 marks]

28.

a. The melting point increases from A to C this is due to increase in number delocalized electron hence increase in the strength of metallic bond.

b. forms a giant structure with strong covalent bonds. Hence high melting.

c. It exhibits allotopy ie may exist as two different form in the same state.

d. $C_2(SO_4)_3$

e. Noble gases or inert
- Used in filament bubls
- Used to produce an inert atmosphere in high temperature inetallurgical processes e.g welding.

f. C is amphoteric oxide
- F acidic it is non –metal oxide.

29.

a. CnH2n, where n = No. of carbon atoms

b. 70

c. C_5H_{10}, $CH_3CH{=}CHCH_2CH_3$
OR $CH_3CH_2CHCH_2{=}CH_2$

30.

a. Hydrocarbon. √1

b. Black specks is carbon
Colourless gas is steam ✓1 3
Hydrocarbon burn in air to form carbon ✓ ½ and water✓½

31.

a. $NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$
Moles of AgCl= $\frac{\text{Mass}}{\text{R.F.M}}$
= $^{2.36}/_{143.5}$
= 0.016446moles
Mole ratio NaCl : AgCl
1 : 1
Moles of NaCl = 0.61446moles
Mass of NaCl = RFM x moles
= 58.5 x 0.016446
= 0.962091g
Mass of solvent (water) = 2.63 – 0.962091
= 1.667909g
1.667909g of water dissolves 0.962091g of NaCl
100g of water dissolves = $\frac{100 \times 0.962091g}{1.667909}$
= $^{57.68}/_{100g}$ of water

b. $24000cm^3$ = 1mol
$80cm^3 = \frac{80x1}{2400}$
= 0.00333moles

32.

a. $CH_3CH = CHCH_3$ – But-2-ene

b. 2- methyl 1 prop-2-ene

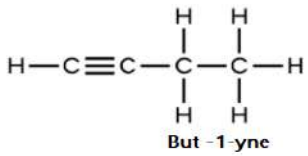
c. $CH_2 = CHCH_2CH_3$ – But-1-ene

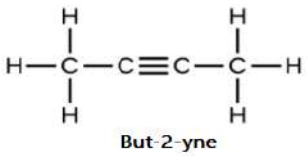
33. Octane
or $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

34.

a. Existence of same molecular formula but different structural formula ✓1

b.


But -1-yne


But-2-yne

