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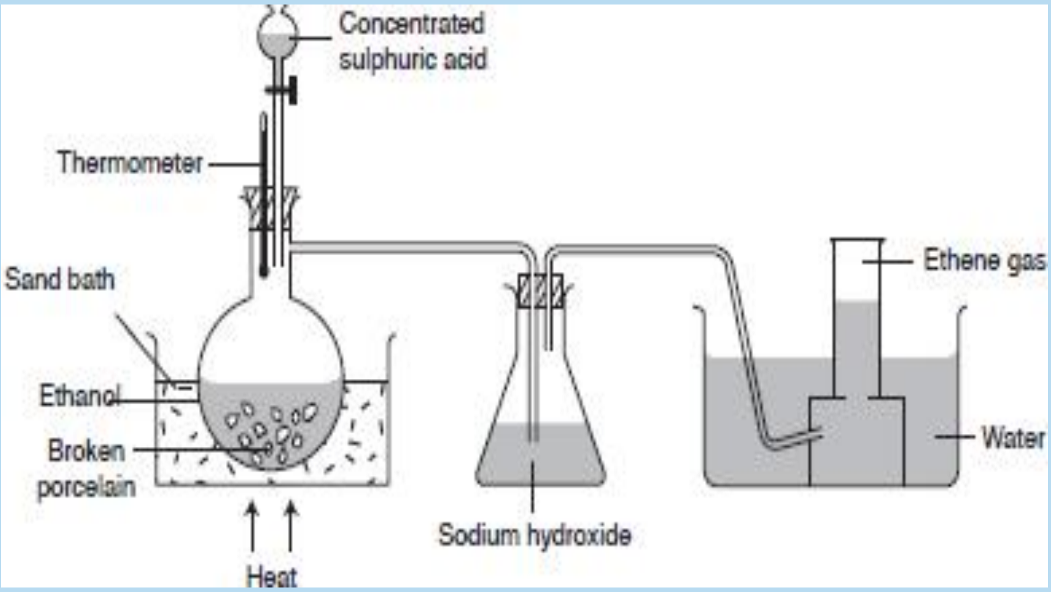
FORM THREE CHEMISTRY HANDBOOK

With well drawn diagrams, solved examples and questions for exercise

(REVISED EDITION)



TR. BENARD CLAIN
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Acknowledgement

First and foremost I thank the Almighty God for the gift of life, energy, knowledge and skills to pursue this work.

I am very grateful to the entire Swafaa Academy Secondary school fraternity for generously supporting me all round as I worked on this material. I must specifically appreciate the H.O.D Sciences Swafaa Academy Mr. Kipchumba Arusei for the support and encouragement.

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The care and best wishes I received from my mother Hellen Clain and my siblings deserve special attention. They were a great source of encouragement.

Lines that influence activities in my life

1. ***God is always there to assist provided you ask for Him.***
2. ***At its best, Chemistry eliminates complexity by revealing underlying simplicity.***
3. ***There is no method of changing your fate except through hard work.***
4. ***Cohesion with immediate neighbours and determination always betters your immediate environment.***

Brief Personal Profile

Benard Clain A. is an experienced teacher of Physics and Chemistry. He is a First Class Honors BSC graduate from Masinde Muliro University Of Science And Technology with an additional PGDE from the said University. Currently taking masters degree. He also has profound knowledge in computer applications and graphics.

TOPIC ONE
GAS LAWS

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Objectives

By the end of this Chapter, the learner should be able to:

(a) State Boyle's and Charles's laws.
(b) Describe experiments to illustrate Boyle's and Charles' laws.
(c) State and use the combined gas law to solve numerical problems.
(d) State Graham's law of diffusion and relate the rate of diffusion to relative molecular mass of a gas.
(e) Explain diffusion in terms of kinetic theory of matter.

Organizer

GAS LAWS

Boyle's Law

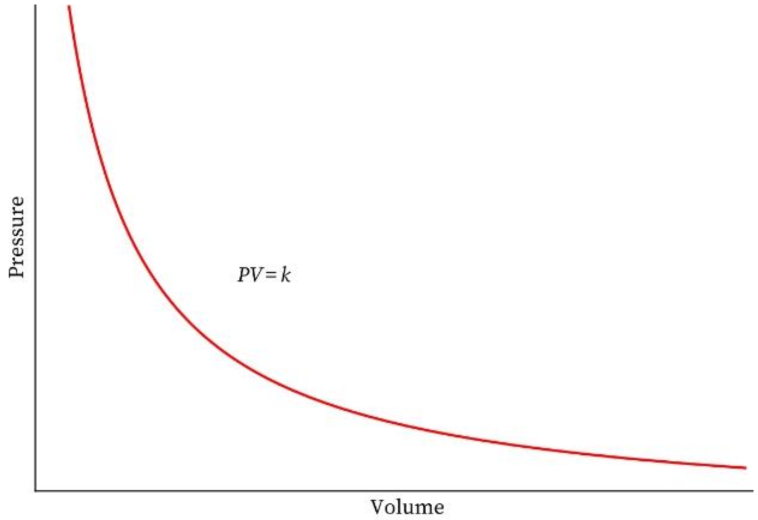
Boyle's law deals with the relationship between pressure and volume of a fixed mass of a gas when temperature is kept constant.

Pressure in a gas is as a result of the **collisions of the gas molecules with the walls of the container.** When the volume of the fixed mass of a gas is **decreased** through compression at constant temperature, the molecules travel a shorter distance to collide with the walls of the container, leading to increased number of collisions per unit time. The pressure of the gas therefore increases with the increased rate of collisions.

Boyle's law states that t*he volume of a given mass of a gas is inversely proportional to its pressure at constant temperature*.

What Boyle's law implies is that as the pressure increases, the volume decreases. The pressure of the gas inside the barrel of a pump is directly proportional to the physical pressure applied to compress the gas. A graph of the physical volume is a curve as shown below:.

The mathematical expression of Boyle's law is:



$$\text{Volume } \alpha \frac{1}{\text{Pressure}}$$

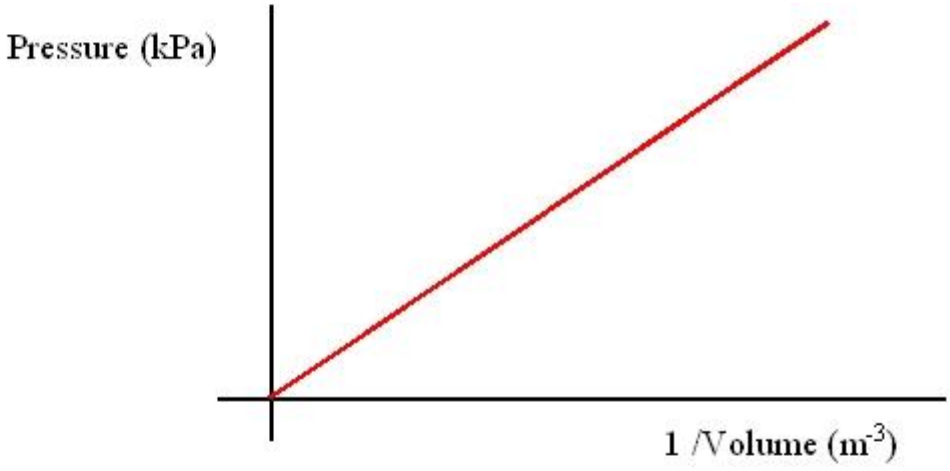
$$V \alpha = \frac{1}{P} \text{ or } P \alpha = \frac{1}{P}$$

Hence, VP = Constant

The expression implies that when the volume of a fixed mass of a gas changes from V_1 to V_2 its pressure also changes from P_1 to P_2. This leads to the general expression:

$$\mathbf{P_1V_1 = P_2V_2}$$

When a graph of pressure of a fixed mass of gas is plotted against the reciprocal of volume, a straight line is obtained.



The SI unit of pressure is the Pascal (Pa). It is equal to one Newton per square metre (NM^{-2}). Other units used to express pressure are atmospheres. One atmospheric pressure is equal to 760 mmHg pressure or 1.01325 × 10^5 Pascals. The SI unit of volume is cubic metres (m^3). One cubic metre is equal to 1.0 × 10^6 cubic centimetres (cm^3).

Worked Examples

1. A volume of 375 cm^3 of a gas has a pressure of 20 atmospheres. What will be its volume if pressure is reduced to 15 atmospheres?

Solution

From Boyle's law, $P_1 V_1 = P_2V_2$

P_1 = 20 atmospheres, P_2 = 15 atmospheres, V_1 = 375 cm^3, V_2 = ?

Substituting for P_1, V_1 and P_2 the equation becomes;

$$20 \times 375 = 15 \times V_2$$

$$V_2 = \frac{20 \times 375}{15}$$

$$V_2 = 500\ \text{cm}^3$$

2. A given mass of gas occupies a volume of 200 cm^3 at a pressure of 5 atmospheres. At what pressure will the gas have a volume of 800 cm^3?

Solution

From Boyle's law, $P_1V_1 = P_2V_2$

P_1 = 5 atmospheres, P_2 = ?, V_1 = 200 cm^3, V_2 = 800 cm^3

$5 \times 200 = P_2 \times 800$

$$P_2 = \frac{5 \times 200}{800}$$

P_2 = 1.25 atmospheres

3. A certain mass of gas occupies 250 cm^3 at 25°C and 750 mmHg. Calculate its volume at 25°C if pressure changes to 760 mmHg in SI Units.

Solution

$V_1 = 250\ cm^3$, $V_1 = ?$, $P_1 = 750$ mmHg, $P_2 = 760$ mmHg

$$P_1V_1 = P_2V_2 \Rightarrow V_2 = \frac{P_1V_1}{2}$$

$$\therefore V_2 = \frac{750 \times 250}{760}$$

$$V_2 = 246.7 \times 10^{-6} m^3 = 2.467 \times 10^{-4} m^3$$

4. At a Constant temperature, a gas at 540 mmHg pressure occupies a volume of 300 litres. The gas is made to expand and occupy a volume of 600 litres. What is the new gas pressure in SI Units?

Solution

$P_1 = 540$ mmHg, $V_1 = 300\ cm^3$, $V_2 = 600\ cm^3$

$540 \times 300 = P_2 \times 600$

$$P_2 = \frac{540 \times 300}{600}$$

$$= 270 \text{ mmHg}$$

$$760 \text{ mmHg} = 1.01325 \times 10^5 \text{ Pascals}$$

$$270 \text{ mmHg} = \frac{1.01325 \times 10^5 \times 270 P_a}{760}$$

$$= 3.6 \times 10^4 \text{ Pascals}$$

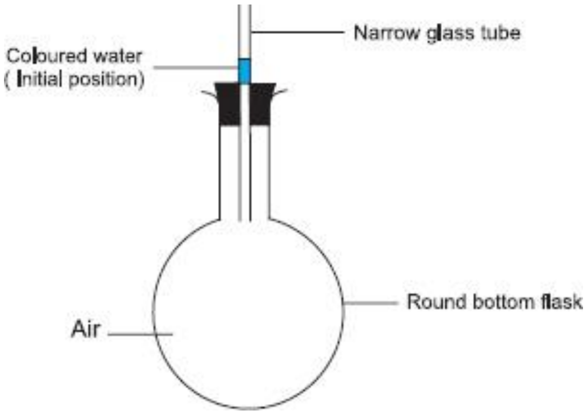
Charles' Law

Charles' Law deals with the relationship between the volume of a fixed mass of a gas with its temperature at constant pressure.

Charles's Law states that: *The volume of a given mass of a gas is directly proportional to its absolute temperature, its pressure being kept constant*.

How does the volume of a fixed mass of gas vary with temperature at constant pressure?

Fit a narrow glass tube into a rubber bung. Loosely fix the bung in a round bottomed flask. Introduce a drop of coloured water into the glass tube. When the drop is half way down the glass tube, firmly stopper the flask making it airtight. Note the position of the coloured water column in the tube. Immerse the flask in a trough of warm water. Observe and record what happens to the water column. Repeat the experiment using ice cold water in a trough. Record your observations.



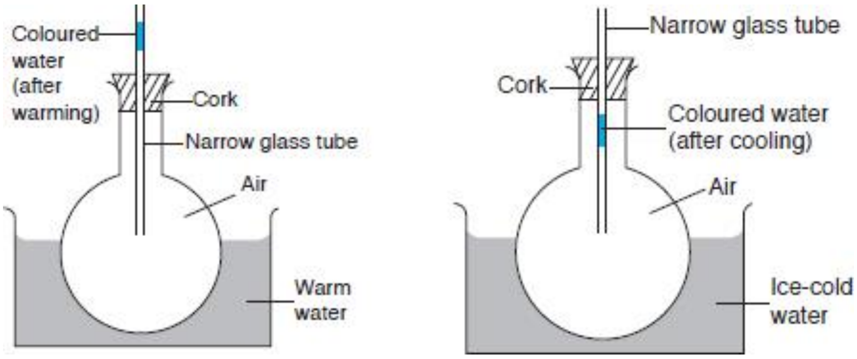
Volume of a given mass of a gas

Questions

1. **What happens to the water column when the flask is warmed or cooled?**

When the flask is immersed in a trough of warm water, the column of coloured water first drops then moves up steadily.

When the flask is immersed in warm water the coloured water column in the capillary tube drops initially because the flask expands before the gas inside absorbs heat to start expanding.



2. How does the change in temperature affect the volume of the fixed mass of gas?

   Charles's Law: ***The volume of a given mass of a gas is directly proportional to its absolute temperature, its pressure being kept constant*.**

3. Which factor is kept constant in this experiment?

   Pressure is kept constant.

4. Explain the observations made during the experiment in terms of the kinetic theory.

When the gas inside heats up, it expands and pushes up the water column in the capillary tube. When a fixed mass of a gas is heated at constant pressure, its volume increases to counter balance the constant pressure. The heat energy increases the kinetic energy of the gas molecules. This leads to increased rate of collisions with the walls of the container causing an increase in the gas pressure.

Mathematically, Charles's Law is expressed as; Volume α absolute temperature. $V \alpha T$

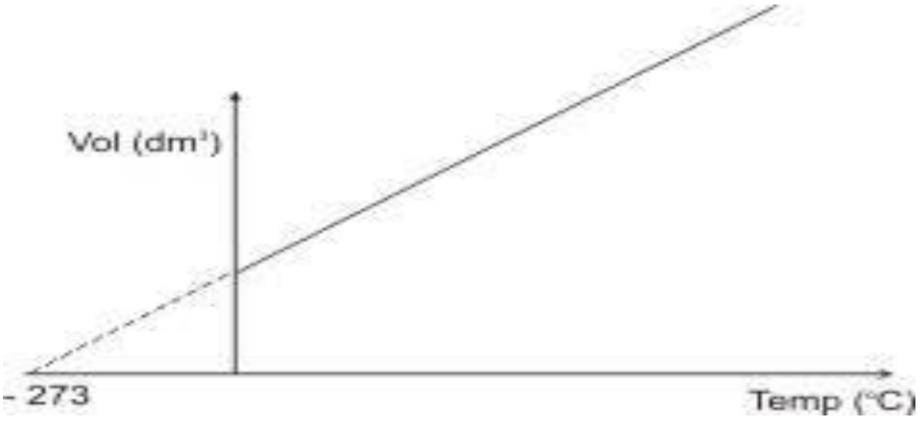
(Where V is the volume of the gas and T its absolute temperature in Kelvin).

$$\text{Thus, } \frac{V}{T} = \text{constant}$$

When the volume of a fixed mass of a gas changes from V_1 to V_2, its absolute temperature changes from T_1 to T_2 leading to the expression:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

A graph of volume of a fixed mass of a gas against temperature is a straight line. The graph intercepts the temperature axis at −273 °C when extrapolated as shown below.



A gas **whose volume is theoretically zero at -273°C** is referred to as an **ideal gas.**

In reality, a gas **would have turned into a liquid** before reaching a temperature of **−273°C.**

The **temperature at which the volume of a gas is assumed to be zero** is called **absolute zero.**

From the absolute zero temperature, the absolute temperature scale is derived. The lowest temperature on the absolute scale is **zero Kelvin (0 K). The Kelvin (K) is the SI unit of the temperature on the absolute scale.** The equivalence of **−273°C** on the Kelvin scale is **0 K.**

Temperature inter conversion from one scale to the other is done as follows:

1 To convert temperature in degrees Celsius (°C) to Kelvin (K), add 273. For example, to convert 204 °C to Kelvin add 273 this;

T = 204 + 273

= 477K

2. To convert temperature in the Kelvin scale to degrees Celsius, subtract 273.

For example, convert 405 K to degrees celsius (°C)

t = 450 – 273

= 177°C.

Worked Examples

1. Convert the temperatures below to the absolute scale:
   (i) 0°C
   (ii) 25°C
   (iii) –20°C

Solution

T = t + 273

(i) 0 + 273 = 273 K
(ii) 25 + 273 = 298 K
(iii) –20 + 273 = 253 K

Note:

Temperature on the Kelvin scale is denoted by T, while on the celsius scale it is denoted by t.

2. Convert the temperature recorded below in Kelvin to temperature in degrees celsius (°C):
   (i) 0 K
   (ii) 250 K
   (iii) 273 K

Solution

t = T – 273

t = T – 273

t = T – 273

(i) 0 – 273 = – 273°C
(ii) 250 – 273 = – 23°C
(iii) 273 – 273 = 0°C

3. A gas occupies 450 cm^3 at 27°C. What volume would the gas occupy at 177°C, if its pressure remains constant?

Solution

From Charles's Law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Converting the temperature to absolute temperature then substituting for V_1, T_1 and T_2 then:

$$\frac{V_2}{450} = \frac{450}{300}$$

$$V_2 = \frac{450 \times 450}{300} = 675 \text{ cm}^3$$

4. At a temperature of 57°C, nitrogen gas occupies a volume of 750 cm^3. At what temperature will the gas occupy 100 cm^3. Express the answer in degrees celsius.

Solution

From Charles's Law: $\frac{V_2}{T_2} = \frac{V_1}{T_1}$

$$\frac{100}{T_2} = \frac{750}{330}$$

$$T_2 = \frac{100 \times 330}{750} = 44 \text{ K}$$

Converting to °C celsius: $t = T - 273$

$$= 44 - 273$$

$$= 229°\text{C}$$

3.Combined Gas Law

The combined gas law deals with the variation in the volume of a fixed mass of a gas with respect to changes in temperature and pressure.

The mathematical expression of Charles's Law, V α T can be combined with that of Boyle's law $V \propto \frac{1}{P}$ to obtain the expression:

$$V \propto \frac{T}{P}$$

Therefore, PV α T

Hence, $\frac{PV}{T}$ = Constant

If a fixed mass of a gas of volume V_1 exerts a pressure P_1 at absolute temperature T_1, the expression may be written as:

$$\frac{P_2V_2}{T_2} = \text{Constant}$$

Suppose the same mass of gas has a volume, V_2, and exerts a pressure, P_2, at absolute temperature, T_2, then the expression becomes:

$$\frac{P_1V_1}{T_1} = \text{Constant}$$

Therefore, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

This is the **ideal gas equation.** It enables the volume of a gas to be obtained under any conditions of temperature and pressure provided its volume under some other conditions of temperature and pressure is known.

Standard Conditions

There are two conditions considered when comparing volumes of gases:

1. **Standard temperature and pressure (s.t.p).** The s.t.p conditions refer to a temperature of 273 K and a pressure of 760 mmHg (1 atmosphere).

2. **Room temperature and pressure (r.t.p).** This refers to a temperature of 298 K and a pressure of 760 mmHg.

It should be seen that from the standard conditions, all computations of temperature should be expressed in Kelvin.

Worked Examples

1. **What will be the volume of a given mass of oxygen at 25°C if it occupies 100 cm^3 at 15°C? (pressure remains constant).**

Solution

Let the initial volume, temperature and pressure be V_1, T_1, P_1 respectively, and the final be V_2, T_2, P_2. From the gas equation;

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Where, $V_1 = 100\ cm^3$, $T_1 = 15 + 273 = 288K$ and $P_1 = P_2$ = constant.

$$V_2 = \frac{P_1V_1}{T_1} \times \frac{T_2}{P_2} \text{ since } P_1 = P_2$$

$$\text{Then, } V_2 = \frac{V_1T_2}{T_1} = \frac{100 \times 298}{288} = 103.5\ cm^3$$

2. **A given mass of a gas occupies 20 cm^3 at 25°C and 670 mmHg pressure. Find out the volume it will occupy at:**
   - **(a) 10°C and 335 mmHg**
   - **(b) 0°C and 760 mmHg**

Solution

(a) From the gas equation,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Where, $P_1 = 670$ mmHg, $V_1 = 20$ cm^3,

$T_1 = 25 + 273 = 298$ K, $T_2 = 10 + 273 = 283$ K, $P_2 = 335$ mmHg, $V_2 = ?$

Therefore, $\frac{670 \times 20}{298} = \frac{335 \times V_2}{283}$ $\qquad$ $V_2 = \frac{670 \times 20 \times 283}{288 \times 335} = 38 \text{ cm}^3$

(b) $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Where, $P_1 = 670$ mmHg, $\qquad V_1 = 20$ cm^3, $\qquad T_2 = 273$ K, $\qquad V_2 = ?$

Therefore, $\frac{670 \times 20}{298} = \frac{760 \times V_2}{273}$ $\qquad$ $V_2 = \frac{670 \times 20 \times 273}{298 \times 760} = 16 \text{ cm}^3$

Some Applications of Gas Laws

The effects of changes in pressure, volume and temperature on a fixed mass of a gas have been used in a wide range of applications which include:

1. Inflating tyres, balls and balloons appropriately depending on the prevailing temperature conditions.
2. Designing of aerosol cans and tear-gas canisters which contain a gas compressed under pressure to act as a propellant of liquid contents in the cans.
3. Regulation of pressure in an aircraft for comfortable in flight environment at high altitude.

Diffusion and the Graham's Law

Diffusion is the process by which **particles spread out from a region of high concentration to regions of low concentration.**

The scent of a strong perfume reaches all corners of a room as soon as the container is opened because of diffusion. The **inter–molecular forces of attraction in gases are very weak due to their large inter-molecular distances. A gas therefore, always spreads out to fill up all the space available. The perfume spreads from an area of high concentration to areas where its concentration is low.**

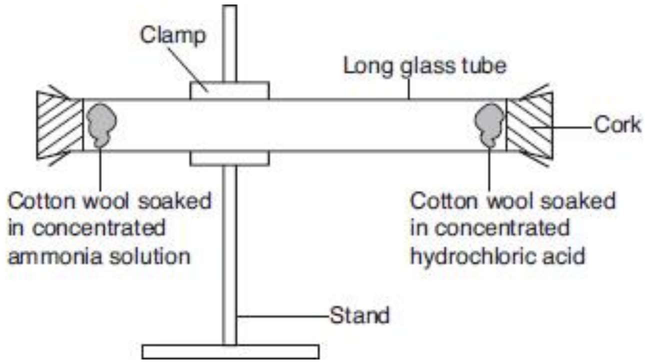
The spreading out of gas particles in air takes a shorter time than solid particles in a solvent. This is because gas particles are far apart and have more kinetic energy than the liquid particles.

Graham's law of diffusion.

Different gases have different rates of diffusion.

According to **Graham's Law**, **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**

To verify this, two gases, ammonia and hydrogen chloride can be allowed to diffuse inside a combustion tube as shown below.



Questions

1. **What observations are made in the glass tube and after how long?**

   A white solid is formed in the tube closer to the end where the cotton wool soaked in concentrated hydrochloric acid was placed after about 5 minutes.

2. **Which gas covered a longer distance?**

   Ammonia

3. **Explain the observations made in the glass tube.**

Concentrated ammonia solution generates ammonia gas while concentrated hydrochloric acid generates hydrogen chloride gas. Ammonia and hydrogen chloride gases diffuse in air in the tube, and when their molecules meet, they react to form white, solid ammonium chloride.

$$NH_3(g) + HCl(g) \xrightarrow{\quad\quad} NH_4Cl(s$$

4. **Determine the molecular masses of ammonia (NH_3) and hydrogen chloride (HCl). N = 14.0, H = 1.0, Cl = 35.5.**

   Ammonia =14+3=17g/mol

   Hydrogen chloride =1+35.5=36.5g/mol

5. **Calculate the rate of diffusion of:**

   (a) Ammonia gas.

The distance covered by ammonia was 12 cm. The rate of diffusion of ammonia gas in air is $\frac{12 \text{ cm}}{5 \text{ minutes}} = 2.4$ **cm/minute**

   (b) Hydrogen chloride gas.

The distance covered by hydrogen chloride was 8 cm within the same time interval of 5 minutes.

The rate of diffusion of hydrogen chloride gas in air is

$\frac{8 \text{ cm}}{5 \text{ minutes}} = 1.6$ **cm/minute.**

6. **What is the relative rate of diffusion of ammonia to hydrogen chloride gas in air?**

The relative rate of diffusion of ammonia gas compared to hydrogen chloride gas in air is:

$$\frac{\text{Rate of diffusion of NH}_3}{\text{Rate of diffusion of HCl}} = \frac{2.4 \text{ cm/minutes}}{1.6 \text{ cm/minuts}} = 1.5$$

This means ammonia diffuses 1.5 times faster than hydrogen chloride. This is because ammonia gas is less dense than hydrogen chloride gas. Therefore, gases with low densities diffuse faster than those with high densities.

The mathematical expression of Graham's Law is:

$$\text{Rate (R)}\ \alpha = \frac{1}{\sqrt{\text{density(d)}}}$$

Meaning that:

$$\text{Rate (R)} = \frac{\text{Constant}}{\sqrt{\text{density(d)}}}$$

$$R = \frac{\text{Constant}}{\sqrt{d_A}}$$

When the rates of diffusion of two gases A and B are compared, the equations are:

$$\text{(i)}\quad \text{Rate of diffusion of gas A} = \frac{\text{Constant}}{\sqrt{\text{density of gas A}}}$$

$$R_A = \frac{\text{Constant}}{\sqrt{d_A}}$$

$$\text{(ii)}\quad \text{Rate of diffusion of gas B} = \frac{\text{Constant}}{\sqrt{\text{density of gas B}}}$$

$$R_B = \frac{\text{Constant}}{\sqrt{d_B}}$$

In equation (i), $R_A\sqrt{d_A} = \text{Constant}$

In equation (ii), $R_B\sqrt{d_B} = \text{Constant}$

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, Graham's law can also be expressed as:

$$\text{Rate} = \frac{\text{Constant}}{\sqrt{\text{molecular mass}}}$$

Therefore, if the rate of diffusion of two gases A and B are compared, then:

$$\frac{R_A}{R_B} = \sqrt{\frac{M_B}{m_A}}$$

Where M_A and M_B are relative molecular masses of the gases A and B respectively. Since Rate is inversely proportional to time, (Rate α $\frac{1}{\text{Time}}$).

It means that it is also possible to compare the time taken for **equal volumes** of two gases to diffuse under similar conditions:

$$R_A = \frac{\text{Constant}}{\text{Time of diffusion of gas A}} \text{ and } R_B = \frac{\text{Constant}}{\text{Time of diffusion of gas B}}$$

This means that:

R_AT_A = Constant, and R_BT_B = Constant

(R = rate, T = time)

Therefore,

$R_AT_A = R_BT_B$ and $\frac{R_A}{R_B} = \frac{T_B}{T_A}$

Where T_A and T_B are the times of diffusion of gases A and B respectively.

$$\text{But } \frac{R_A}{R_B} = \sqrt{\frac{d_B}{d_A}} = \sqrt{\frac{M_B}{M_A}}$$

$$\text{Therefore, } \frac{T_B}{T_A} = \sqrt{\frac{d_B}{d_A}} \text{ and } \frac{T_B}{T_A} = \sqrt{\frac{M_B}{M_A}} \qquad \frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$$

Worked Examples

1. **Equal volumes of carbon(II) oxide and carbon(IV) oxide are allowed to diffuse through the same medium. Calculate the relative rate of diffusion of carbon(II) oxide. (C = 12.0, Oxygen = 16.0)**

Solution

Relative molecular mass (M_r) of CO = 12 + 16 = 28

Relative molecular mass (M_r) of CO_2 = 12 + 32 = 44

$$\frac{R_{CO}}{R_{CO_2}} = \sqrt{\frac{M_rCO_2}{M_rCO}}$$

$$\frac{R_{co}}{R_{CO_2}} = \sqrt{\frac{44}{28}}$$

$$\frac{R_{CO}}{R_{CO_2}} = 1.254$$

Carbon(II) oxide diffuses 1.254 times faster than carbon(IV) oxide.

2. **If it takes 20 seconds for 200 cm^3 of oxygen gas to diffuse across a porous plug. How long will it take an equal volume of sulphur(IV) oxide to diffuse across the same plug? (O = 16.0, S = 32.0).**

Solution

$$\frac{T_{O_2}}{T_{SO_2}} = \sqrt{\frac{M_rO_2}{M_rSO_2}} \quad \Rightarrow \quad \frac{20}{T_{SO_2}} = \sqrt{\frac{32}{64}}$$

$$\frac{20}{T_{SO_2}} = \frac{1}{\sqrt{2}}$$

$$T_{SO_2} = 20\sqrt{2} = 28.3 \text{ seconds}$$

3. **Determine the molecular mass of the gas Y which diffuses $1\frac{1}{2}$ times slower than Oxygen. 0 = 16.0**

$$\frac{R_Y}{R_{O_2}} = \sqrt{\frac{M_rO_2}{M_rY}}$$

$$\frac{1}{1.5} = \sqrt{\frac{32}{M_{r_Y}}}$$

$$\frac{1}{2.25} = \frac{32}{M_{r_Y}}$$

$$M_{r_Y} = 2.25 \times 32$$

$$M_{r_Y} = 72$$

4. **If it takes 30 seconds for 100 cm^3 of carbon(IV) oxide to diffuse across a porous plate. How long will it take 150 cm^3 of nitrogen(IV) oxide to diffuse across the same plate under similar conditions? (C = 12.0, N = 14.0, O = 16.0).**

Solution

$$\frac{R_{CO_2}}{R_{NO_2}} = \sqrt{\frac{M_rNo_2}{M_rCo_2}} \quad \text{But } RCO_2 = \frac{100cm^3}{30s} = 3.33cm^3 \text{ per second}$$

$$\frac{3.33}{R_{NO_2}} = \sqrt{\frac{46}{44}}$$

$$\frac{3.33}{R_{NO_2}} \; 1.0225$$

$$R_{NO_2} = \frac{3.33}{1.0225} = 3.26 \text{ cm}^3 \text{ per second}$$

Therefore, time taken for NO_2 to diffuse is:

$$\frac{150 \text{ cm}^3}{3.26 \text{ cm}^3 \text{ sec}^{-1}} = 46 \text{ seconds.}$$

Alternative Working

Since equal volumes must be compared:
100 cm^3 of CO_2 takes 30 seconds

$$\therefore 150 \text{ cm}^3 \text{ of } CO_2 \text{ will take } \frac{30}{100} \times 150 = 45 \text{ seconds}$$

$$\frac{T_{CO_2}}{T_{NO_2}} = \sqrt{\frac{M_rCO_2}{M_rNO_2}}$$

$$\frac{45}{T_{NO_2}} = \sqrt{\frac{44}{28}} \Rightarrow \frac{45}{T_{NO_2}} = 0.978$$

$$T_{NO_2} = \frac{45}{0.978} = 46 \text{ seconds}$$

5. **Calculate the relative rate of diffusion of ammonia gas compared to that hydrogen chloride gas under the same conditions of temperature and pressure.**

From the expression:

$$\frac{R_A}{R_B} = \sqrt{\frac{M_B}{M_A}}$$

$$\text{Then: } \frac{R_{NH_3}}{R_{HCl}} = \sqrt{\frac{35.5}{17}} = 1.465$$

This ratio implies that ammonia diffuses 1.5 times faster than hydrogen chloride gas.

Review Questions

1. 2006 Q 3 P1
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 partition under the same conditions? (S= 32.0, 0 = 16.0)
(3marks)

2. 2007 Q 12 P1

(a) State the Charles law (1mark)

(b) The volume of a sample of nitrogen gas at a temperature of 291 K and 1.0×10^5 Pascals was $3.5 \times 10^{-2} m^3$. Calculate the temperature at which the volume of the gas would be $2.8 \times 10^{-2} m^3$ at 1.0×10^5 Pascals. (2marks)

3. 2008 Q 1 P1

A small crystal of potassium manganate (VII) was placed in a beaker water. The beaker was left standing for two days without shaking. State and explain the observations that were made. (2marks)

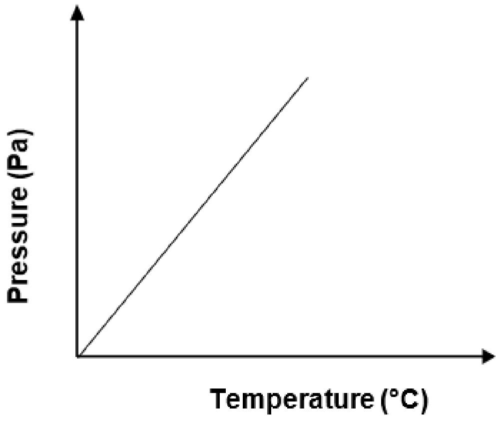
4. 2008 Q 20 P1

(a) State the Graham's law diffusion. (1mark)

(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 $12 cm^3 s^{-1}$, calculate the rate of diffusion of **X** through the same material. (2marks)

5. 2009 Q 26 P1

The graph below shows the relationship between pressure and the temperature of a gas in a fixed volume container.



(a) State the relationship between pressure and temperature that can be deduced from the graph. (1 mark)

(b) Using kinetic theory, explain the relationship shown in the graph. (2 marks)

6. 2010 Q8 P1

The pressure of nitrogen gas contained in a 1 dm^3 cylinder at -196 °C was 10^7 Pascals. Calculate the:

(a) Volume of the gas at 25 °C and 10^5 Pascals. (1½ marks)

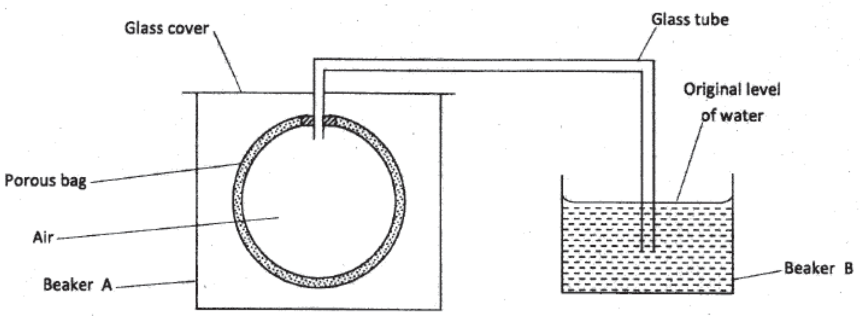
(b) Mass of nitrogen gas.
(Molar volume of gas is 24 dm^3, N = 14.0) (1 ½ marks)

7. 2011 Q6 P1

A certain mass of gas occupies 0.15 dm^3 at 293K and 98,648.5 Pa. Calculate its volume at 101,325 Pa and 273K. (2 marks)

8. 2012 Q19 P1

The set up shown below was used to investigate a property of hydrogen gas.



State and explain the observation that would be made in the glass tube if beaker **A** was filled with hydrogen gas. (3 marks)

9. 2013 Q14 P1

(a) State the Charles' law. (1 mark)

(b) A certain mass of gas occupies 146 dm^3 at 291 K and 98.31 kPa. What will be its temperature if its volume is reduced to 133 dm^3 at 101.325 Pa?

(2 marks)

10. 2014 Q6 P1

100cm^3 of a sample of ethane gas diffuses through a porous pot in 100 seconds. What is the molecular mass of gas Q if 1000 cm^3 of the gas diffuses through the same porous pot in 121 seconds under the same conditions?
(C-12.0, H=1.0) (3 marks)

11. 2015 Q4 P1

(a) State the Boyle's Law. (1 mark)

(b) A gas occupies 500cm^3 at 27 °C and 100,000 Pa. What will be its volume at 0 °C and 101,325 Pa? (2 marks)

12. 2016 Q20 P1

60cm^3 of oxygen gas diffused through a porous partition in 50 seconds. How long would it take 60cm^3 sulphur (IV) oxide gas to diffuse through the same partition under the same condition (S=32.0, O=16.0) (3 marks)

13. 2017 P1 Q6.

(a) State Charles' Law. (1 mark)

(b) Explain why the pressure of a fixed mass of a gas increases, when the volume of the gas is reduced at constant temperature. (2 marks)

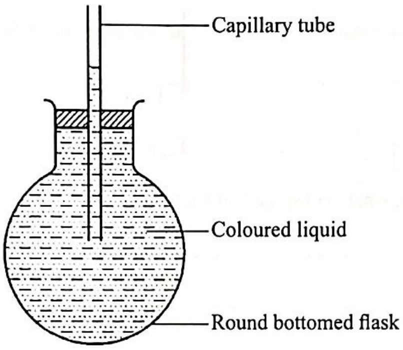
14. 2018 P1 Q 6.

(a) State Graham's law of diffusion. (1 mark)

(b) Explain why a balloon filled with helium gas deflates faster than a balloon of the same size filled with argon gas. (2 marks)

15. 2018 P2 Q3(b)

Use the set-up in **Figure 3** to answer the questions that follow. The flask was covered with a cloth that had been soaked in ice-cold water.



(i) State the observation made on the coloured water. Explain. (2 marks)

(ii) Name the gas law illustrated in **Figure 3.** (1 mark)

16. 2019 P1 Q26.

140cm^3 of nitrogen gas diffuses through a membrane in 70 seconds. How long will it take 200 cm^3 of carbon (IV) oxide gas to diffuse through the same membrane under the same conditions of temperature and pressure. (3 marks)

TOPIC TWO
THE MOLE

Table of Contents

Organizer

Objectives

By the end of this topic, the learner should be able to:

(a) Define the mole and relate it to relative atomic mass.

(b) Convert mass into moles, and moles to mass.

(c) Use data to determine the empirical and molecular formulae of compounds.

(d) Define the terms concentration, molarity and dilution of a solution.

(e) Define and prepare molar solutions.

(f) Carry out titrations and solve problems involving molar solutions.

(g) Write full formulae and ionic equations.

(h) Define molar gas volume and atomicity of gases.

(i) State Avogadros' and Gay Lussac's laws and carry out related calculations.

(j) Carry out calculations on molar gas volumes both at standard temperature and pressure and room temperature and pressure.

THE MOLE

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

In chemistry, substances are called chemicals.

Relative Mass and the Avogadro's constant.

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

The **Carbon-12** isotope is used as a reference for measuring relative atomic mass because it is a stable solid and a very common element.

The relative atomic mass of an element is defined as

$$\text{R.A.M.} = \frac{\text{Average mass of one atom of the element}}{\text{One twelfth } (\frac{1}{12}) \text{ of the mass of one atom of carbon-12}}$$

Since relative atomic mass is a ratio, it has no units.

Element	R.A.M	Element	R.A.M
Hydrogen	1	Magnesium	24
Carbon	12	Sulphur	32
Oxygen	16	Calcium	40
Sodium	23		

Relative Atomic masses of some selected elements

The **number of atoms** in **one relative atomic mass unit in grams of any element** has been established to be **6.023×10^{23}**. This number is referred to as **Avogadro's Constant, 'L'.**

A mole is the **amount of any substance** that contains Avogadro's number of particles **(6.023×10^{23} particles)** .

The unit "Mole" is used to **measure the amount of particles (atoms, molecules, ions, electrons etc.) of any substance.**

The mass in grams of one mole of a substance is referred to as **molar mass**

Inter Conversion of Molecules and Moles

Molecules contain more than one atoms of the same or different elements.

There are two atoms of oxygen in one molecule of oxygen. Therefore, the relative molecular mass of oxygen molecule is, 16 × 2 = 32. This implies that a mole of oxygen molecules has a mass of 32 g. This mass contains 6.023×10^{23} molecules of oxygen.

Since each molecule of oxygen has two atoms, then 1 mole of oxygen molecules would contain 2 × 6.023 × 10^{23} = 1.2046 × 10^{24} atoms of oxygen. Similarly, 0.5 mole of oxygen molecules would contain: 6.023 × 10^{23} × 0.5 = 3.0115 × 10^{23} molecules of oxygen and 3.0115 × 10^{23} × 2 = 6.023 × 10^{23} atoms of oxygen.

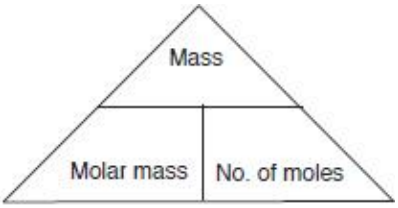
Inter conversion of Mass and Moles

(i) Relationship between mass of an element and the mole.

Given that one mole of an element has a mass equal to its relative atomic mass in grams, then it can be deduced that:

Mass of an element (in grams) = Molar mass × number of moles.

This formula can easily be remembered by using the triangle below:



$$\text{No of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Key: • Horizontal line in triangle represents division sign.

- Vertical line in triangle represents multiplication sign.

Therefore, it would be seen that:

$$\text{Molar mass} = \frac{\text{Mass}}{\text{No of moles}}$$

$$\text{Mass} = \text{No. of moles} \times \text{molar mass.}$$

Worked Examples

1. **Calculate the mass of 0.4 moles of calcium (Ca = 40).**

Mass = molar mass × number of moles

= 0.4 × 40

= 16 g

2. **Determine the number of moles in 13.5 g of aluminium (A1 = 27)**

27 g = 1 mole

$$13.5\text{ g} = \frac{1}{27} \times 13.5$$

$$= 0.5\text{ moles.}$$

(ii) Relationship Between Relative Molecular Mass and the Mole

The **Relative molecular mass (for molecules)** of an element or compound is the **sum of all the relative atomic masses of the atoms in a molecule of the element or compound**.

For example, the relative molecular mass of chlorine Cl_2 is 35.5 × 2 = 71 g/mol

That of water, H_2O is (1×2) + 16 = 18g/mol.

Relative formula mass (for ionic substances) is the **sum of the relative atomic masses of all the atoms in a formula unit of a compound**.

For example, the relative formula mass of potassium sulphate K_2SO_4 is (39 × 2) + 32 + (16 × 4) = 174 g/mol.

NB: **One mole** of any substance has a **mass equal to the relative formula mass of that substance**. The substance could be made of atoms, ions or molecules.

2. Empirical and Molecular Formula

The **empirical formula** of a compound ***shows the simplest whole number ratio in which atoms combine to form the compound*.** The formula is determined from experimental data. It provides useful information about the chemical formula of a compound.

Empirical formula

Experiment to determine the formula of the compound formed when magnesium burns in air.

Weigh a clean dry crucible together with its lid. Clean about 15 cm length of magnesium ribbon thoroughly. Using a pencil, wind the magnesium ribbon into a coil. Place it inside the crucible and replace the lid. Weigh again the crucible together with the magnesium ribbon and lid. Heat the contents strongly for a few minutes, occasionally lifting the lid slightly using a pair of tongs..

When there are no more flare-ups, remove the lid and heat the crucible strongly. Remove the source of heat and allow the crucible to cool. When cold, replace the lid and weigh again.

Repeat the heating and cooling until a constant mass is obtained. Record your results in a table

Discussion Questions.

1. **Why is it necessary to clean the magnesium ribbon?**
   The magnesium ribbon is cleaned at the start of the experiment so as to remove any oxide film on it.
2. **Why was the lid kept on the crucible at first?**
   It is important to keep the lid in place to prevent any solid from escaping.
3. **What was the purpose of the occasional lifting of the lid?**
   It is necessary to lift the lid from time to time to allow in air.
4. **Explain why it is necessary to heat the crucible and its contents until there is no further change in mass.**
   The purpose of heating until a constant mass is obtained ensures that all the magnesium has reacted.
5. **What masses of magnesium and oxygen combine to form magnesium oxide?**

Sample Results

Mass of empty crucible + lid, (A)	19.52 g
Mass of crucible + lid + magnesium, (B)	20.36 g
Mass of crucible + lid + magnesium oxide, (C)	20.92 g

Mass of magnesium, (D) = (B – A)	0.84 g
Mass of magnesium oxide, (E) (C – A)	1.40 g
Mass of oxygen, (F = E – D)	0.56 g

6. **How many moles of atoms of each element reacted in this experiment? Calculate the mole ratio of magnesium to oxygen in magnesium oxide.**

The number of moles of atoms of magnesium, and oxygen which combine can be found by dividing the reacting masses of these elements by their respective relative atomic masses.

The formula of magnesium oxide can now be determined as follows:

Reacting elements:	**Magnesium**	**Oxygen**
Reacting masses (g):	0.84	0.56
Relative atomic mass:	24	16
No of moles =	$\frac{0.84}{24}$	$\frac{0.56}{16}$
Mole ratio	0.35	0.35
(divide by the smallest)	1	:1

7. **What is the simplest formula of magnesium oxide?**

It follows that, one mole of magnesium atoms combines with one mole of oxygen atoms. Therefore, The simplest formula of magnesium oxide is **MgO.**

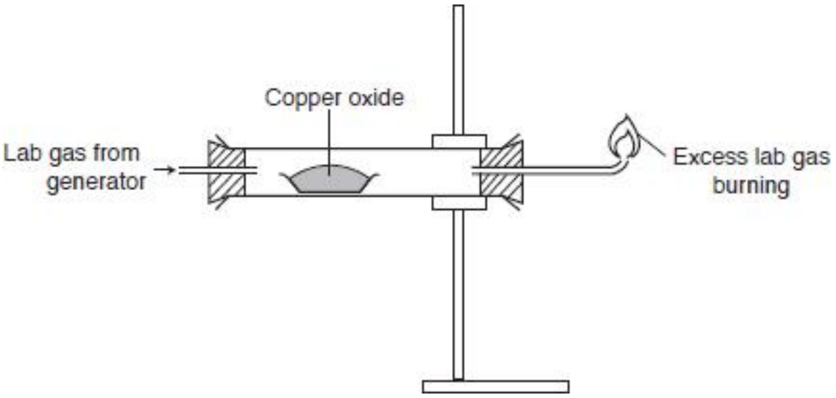
In this case, it also happens to be the chemical formula..

8. **State the possible sources of error in this experiment.**

Any variation from this ratio is due to experimental errors and side reaction such as the formation of magnesium nitride.

Experiment to determine the formula of the compound formed when copper combines with oxygen.

Weigh an empty porcelain boat. Place a small amount of copper(II) oxide into the porcelain boat and weigh again. Place the porcelain boat and its contents in a combustion tube as shown below. Pass a stream of laboratory gas through the tube for a short time. Light the gas at the end.



Start to heat the copper(II) oxide and record your observations. Continue heating the tube until there is no further change. Remove the source of heat but keep the lab gas flowing. When the tube has cooled, turn off the gas supply. Carefully remove the porcelain boat with the residue from the combustion tube and weigh. Repeat the procedure until a constant mass is obtained. Record the results in a table.

Discussion Questions

1. **Why is it necessary to ensure that all the air in the combustion tube is driven out before heating?**

Before heating, lab gas was passed through the tube to remove any traces of air in the tube

1. **How would you know that the reaction is complete?**

When all the black copper(II) oxide has been reduced to reddish brown metallic copper.

2. **What is the colour of the residue?**

In this experiment, black copper(II) oxide is reduced to reddish brown metallic copper using hydrogen gas. The oxygen combines with lab gas to form water and carbon(IV) oxide.

3. **Explain why it is necessary to keep the stream of lab gas on as the product is cooling.**

Metallic copper was cooled in a stream of lab gas to prevent re-oxidation of the hot metal by air.

5. **What other reducing agents can be used to remove oxygen from copper(II) oxide? Explain why some of these agents could not be used in this experiment.**

Ammonia gas can be used as the reducing agent in place of lab gas. Other reducing agents such as carbon and metals that are more reactive than copper can also be used to remove oxygen from copper(II) oxide but they are normally not used because it would be difficult to isolate the copper.

4. **Determine the formula of the copper oxide from these results.**

Mass of empty porcelain boat, (A)	15.6 g
Mass of porcelain boat + copper(II) oxide, (B)	19.1 g
Mass of porcelain boat + residue, (C)	18.4 g
Mass of copper(II) oxide, (D) = (B – A)	3.5 g
Mass of copper, (E) = (C – A)	2.8 g
Mass of oxygen, (F) = (D – E)	0.7 g

The formula can be determined as follows:

Element	**Copper**	**Oxygen**
Mass:	2.8 g	0.7 g
Relative atomic mass:	63.5	16
No. of moles:	$\frac{2.8}{63.5} = 0.044$	$\frac{0.7}{16} = 0.044$
Mole ratio	1	:1

One mole of copper combines with one mole of oxygen. The simplest formula is CuO. This is also the chemical formula of copper(II) oxide.

7. Determine the empirical formula using the percentage by mass of the combining elements.

Chemical formula is also derived from percentage composition of the constituent elements. Using the data for copper(II) oxide experiment,

$$\text{Percentage composition by mass of copper is: } \frac{\text{Mass of copper}}{\text{Mass of copper(II) oxide}} \times 100$$

$$= \frac{2.8}{3.5} \times 100 = 80\%$$

$$\text{Percentage composition by mass of oxygen is: } \frac{\text{Mass of oxygen}}{\text{Mass of copper oxide}} \times 100$$

$$= \frac{0.7}{3.5} \times 100 = 20\%$$

The percentages are then taken to represent the actual masses of the elements in the compound.

The empirical formula is then determined as follows:

Element	**Copper**	**Oxygen**
Percentage composition:	80	20
Relative atomic mass:	63.5	16
No. of moles:	$\frac{80}{63.5}$ = 1.25	$\frac{20}{16}$ = 1.25
Mole ratio:	1	:1

The empirical formula is CuO

Worked Example 1

An oxide of silicon was found to contain 47% by mass silicon. What is the empirical formula of the oxide? (Si = 28, O = 16).

Solution

Element	**Silicon (Si)**	**Oxygen (O)**
Percentage composition:	47	100 – 47 = 53
Relative atomic mass:	8	16
$\text{No. of moles} = \frac{\text{Composition \%}}{\text{Relative atomic mass}}$	$\frac{47}{28} = 1.68$	$\frac{53}{16} = 3.31$
Mole ratio:	$\frac{1.68}{1.68}$ = 1	: $\frac{3.31}{1.68}$:= 1.94
Simple whole No. ratio	1	:2

Therefore, the empirical formula of silicon oxide is SiO_2.

Worked Example 2

The percentage composition by mass of an oxide of iron is 70% iron and 30% oxygen. Determine its empirical formula. (Fe = 56, O = 16).

Solution

Element	**Iron(Fe)**	**Oxygen(O)**
Composition %:	70	**30**
Relative atomic mass:	56	**16**
No. of moles:	$\frac{70}{56} = 1.25$	$\frac{30}{16} = 1.875$
	$\frac{1.25}{1.25}$	$\frac{1.875}{1.25}$
Mole ratio:	1	:1.5

To make it a whole number, ratio multiply by 2

2 : 3

The empirical formula of the iron oxide is $Fe_2 O_3$.

(b) Molecular Formula

Molecular formula shows the **actual number of each kind of atoms present in a molecule of the compound.**

If the empirical formula is known, then the molecular formula can be determined by the relationship.

(Mass of empirical formula)$_n$ = Molecular mass,

Where n is a whole number.
The molecular mass is **always a multiple of the empirical formula mass.**

Worked Example 1

A hydrocarbon was found to contain 92.3% carbon, and the remaining is hydrogen. If its molecular mass is 78, determine its molecular formula. (C = 12, H = 1).

Solution

Element	**Carbon(C)**	**Hydrogen(H)**
Percentage composition:	92.3	7.7
R.A.M:	12	1
No. of moles:	$\frac{92.3}{12}$	$\frac{7.7}{1}$
	= 7.7	:= 7.7
Mole ratio:	1	:1

The empirical formula is CH.

The molecular formula is determined as follows:
$$\begin{aligned} (CH)_n &= 78 \\ (12 + 1)n &= 78 \\ 13n &= 78 \\ n &= 6 \end{aligned}$$

Therefore, the molecular formula is C_6H_6

Worked Example 2

A compound of carbon, hydrogen and oxygen contains 54.55% carbon, 9.0% hydrogen, and 36.6% oxygen. If its relative molecular mass is 88, what is its molecular formula? (C = 12, O = 16, H = 1).

Solution

First obtain the empirical formula as follows:

Element	**Carbon (C)**		**Hydrogen (H)**		**Oxygen (O)**
Composition percentage:	54.55		9.09		36.36
R.A.M:	12		1		16
No. of moles:	$\frac{54.55}{12} = 4.55$		$\frac{9.09}{1} = 9.09$		$\frac{36.36}{16} = 2.27$
Mole ratio:	$\frac{4.55}{2.27}$	:	$\frac{9.09}{2.27}$	:	$\frac{2.27}{2.27}$
	2.004	:	4.004	:	1
Simple whole No. ratio:	2	:	4	:	1

The simplest formula is C_2H_4O.

Since the molecular mass is 88, then

$$(C_2H_4O)_n = 88$$

$$(12 \times 2) + (4 \times 1) + (16 \times 1)_n = 88$$

$$44n = 88$$

$$n = 2$$

The molecular formula is $C_4H_8O_2$.

Worked Example 3

When a certain hydrocarbon is burnt completely in excess oxygen, 5.28 g of carbon (IV) oxide and 2.16 g of water were formed. If the molecular mass of the hydrocarbon is 84, determine the molecular formula of the hydrocarbon.

Solution

This is an example where the mole ratio of the products can be used to determine the formula of the reactant.

Products:	**CO2**	**H20**
Mass:	5.28	2.16
Formula mass:	44	18
No. of moles:	$\frac{5.28}{44} = 0.12$	$\frac{2.16}{18} = 0.12$
Mole ratio:	1	:1

Therefore, it implies that only one mole of the carbon(VI) oxide and one mole of water were produced.

The empirical formula is obtained by working out the masses of carbon and hydrogen in carbon(IV) oxide and water respectively.

Products:	**Carbon (CO_2)**	**Water (H_2O)**
Mass:	$\frac{12}{44} \times 5.28 = 1.44$	$\frac{2}{18} \times 2.16 = 0.24$
No. of moles:	$\frac{1.44}{12} = 0.12$	$\frac{0.24}{1} = 0.24$
Mole ratio:	$\frac{0.12}{0.12}$	$:\frac{0.24}{0.12}$
	1	:1

Therefore, the empirical formula is CH_2.

Since the molecular mass is 84, then

$(CH_2)_n = 84$

$(12 + 2)_n = 84$

$14n = 84$

$n = 6$

The molecular formula of the hydrocarbon, is C_6H_{12}.

3. Molar Solutions

A molar solution is a solution that **contains one mole of a solute in one litre of the solution**.

Concentration of a Solution

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

The concentration of a solution may be expressed in terms of **mass of solute in grams per given volume or number of moles of the solute per given volume**.

For **uniformity**, concentration is normally expressed in either **grams per litre** of solution **(g/dm^3)** or in **moles per litre of solution (mol/dm^3).**

Note: 1 litre = 1000 cm^3 = 1 dm^3

Example

Suppose 4.0 g of sodium hydroxide were dissolved in 400 cm^3 of distilled water then made up to 500 cm^3 of solution, determine the concentration in:

- **(a) grams.**
- **(b) moles.**

Solution

(a) The concentration is 4.0 g/500 cm^3 of solution.

Mass in 500 cm^3 = 4.0 g

$$\text{Mass in } 1000\text{ cm}^3 = \frac{4.0\text{ g}}{500\text{ cm}^3} \times 1000\text{ cm}^3$$
$$= 8.0\text{ g}$$

In grams per litre, the concentration of the solution would be expressed as 8.0 g/dm^3

(b) The concentration is 4.0g/40g/mol. = 0.1 mole/500 cm^3 of solution.

Moles in 500 cm^3 = 0.1 mole

$$\text{Moles in 1000 cm}^3 = \frac{0.1\ \text{mole}}{500\ \text{cm}^3} \times 1000\ \text{cm}^3$$
$$= 0.2\ \text{mole.}$$

Thus, the concentration in moles per dm-3 is 0.2 mole/dm^3.

It can be shown that 1 cm^3 of 4 g/500 cm^3 solution and 1 cm^3 of 8.0 g/1000 cm^3 contain the same mass of the solute.

(i) Mass in 500 cm^3 = 4.0 g

Mass in 1 cm^3 = $\frac{4.0\ \text{g}}{500\ \text{cm}^3}$ = 0.008 g cm^{-3}

(ii) Mass in 1000 cm^3 = 8.0 g

(Mass in 1 cm^3 = $\frac{8.0\ \text{g}}{1000\ \text{cm}^3}$ cm^3 = 0.008 gcm^{-3}

Thus, the concentration of a 500 cm^3 solution containing 4 g of solute is the same as the concentration of a 1000 cm^3 solution containing 8 g of solute.

Molarity of a Solution

The **molarity** of a solution is the **number of moles of the solute per litre of solution (moles/1000 cm^3)**. This is normally expressed as **mol dm^{-3}.**

Example

Suppose 71g of sodium sulphate are dissolved in enough water then made to one litre of solution. Determine the molarity of the solution formed. (Na = 23, S = 32, O = 16).

Mass of Na_2SO_4 dissolved = 71.0 g

Mass of 1 mole of Na_2SO_4 = 142 g

No. of moles in 1 dm^3 = $\frac{71\ \text{g}}{142\ \text{g mol}^{-1}}$ = 0.5 mole dm^{-3}

Molarity is 0.5 M Na_2SO_4.

Preparation of Molar Solutions

The apparatus used to prepare molar solutions include volumetric flasks, and measuring cylinders.

Preparing a solution involves dissolving the required mass of the solute in a little water say, 100 cm^3 of distilled water in a beaker. When all the solute has dissolved, the solution is transferred to the volumetric flask. The beaker is rinsed and the solution is transferred into the flask.

The solution is then made up to the mark on the narrow neck of the flask. The flask is then stoppered. The neck of the flask is narrow to ensure high accuracy of the instrument. The flask is then shaken to obtain a uniform solution.

Experiment: To prepare a molar solution of sodium hydroxide.

Weigh accurately 40.0 g of sodium hydroxide pellets, put them in 200 cm^3 of distilled water in a 250 cm^3 beaker and stir to dissolve. When all the solid has dissolved, transfer the solution into a one-litre volumetric flask.

Rinse the beaker with some distilled water, and transfer the solution into the volumetric flask. Using distilled water in a wash bottle, make up the solution to the mark. Repeat the procedure using exactly 20.0 g and 10.0 g, of sodium hydroxide, prepare the solutions in 500 cm^3 and 250 cm^3 volumetric flasks respectively.

Questions

1. **Calculate the number of moles of sodium hydroxide in the:**
   - **(a) 1 litre solution.**

Mass dissolved = 40.0 g

Molar mass of NaOH = 40.0 g mol^{-1}

$$\text{No. of moles in 1 litre solution} = \frac{40.0\ g}{40.0\ g\ mol^{-1}}$$

$$= 1.0\ \text{mole}$$

(b) 500 cm^3 solution.

Mass dissolved = 20.0 g

Mass of NaOH = 40.0 g

$$\text{No. of moles in } 500\ cm^3 = \frac{20.0\ g}{40.0\ g\ mol^{-1}}$$

$$= 0.5\ \text{mole.}$$

(c) 250 cm^3 solution.(Na = 23, O = 16, H = 1).

Mass of NaOH dissolved = 10.0 g

Molar mass of NaOH = 40.0 g

$$\text{No. of moles in } 250\ cm^3 = \frac{10.0\ g}{40.0\ g\ mol^{-1}}$$

$$= 0.25\ \text{mole}$$

2. Determine the molarity of the solutions of sodium hydroxide in (1) (a), (b) and (c). Comment on the results.

(a) No. of moles in 1000 cm^3 = 1.0 mole dm^{-3}

Molarity is 1.0 M NaOH.

(b) No. of moles in 500 cm^3 = 0.5 mole

No. of moles in 1000 cm^3

$$= \frac{0.5\ g}{500\ cm^3} \times 1000\ cm^3 = 1.0\ \text{mole}\ dm^{-3}$$

Molarity is 1.0 M NaOH.

(c) No. of moles in 250 cm^3 = 0.25 mole

$$\text{No. of moles in } 1000\ cm^3 = \frac{0.25\ \text{mole} \times 1000\ cm^3}{250\ cm^3} = 1.0\ \text{mole}\ dm^{-3}$$

Molarity is 1.0 M NaOH.

Dilution of a Solution

Dilution is a **process by which the concentration of a solution is lowered by adding more solvent into the solution.** During dilution, the **amount of solute remains the same as the volume of the solution increases.**

Experiment to dilute 2 M HCl

Measure accurately 25 cm^3 of 2 M 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. Repeat the process using a 500 cm^3 volumetric flask instead of the 250 cm^3 flask.

Questions

1. **Determine the number of moles in the 25 cm^3 of the 2.0 M HCl.**

$$\begin{aligned}\text{Moles of HC1 in 1000 cm}^3 &= 2.0\text{ moles}\\ \text{Moles of HC1 in 25 cm}^3 &= \frac{2.0\text{ Moles}}{1000\text{ cm}^3}\times 25\text{ cm}^3\\ &= 0.05\text{ moles}\end{aligned}$$

2. **Determine the molarity of the solution contained in the:**
   - **(i) 250 cm^3 flask.**
   - **(ii) 500 cm^3 flask.**

$$\begin{aligned}\text{(a) Moles of HC1 in 250 cm}^3 &= 0.05\text{ moles}\\ \text{Moles in 1000 cm}^3 &= \frac{0.05\text{ Moles}\times 1000\text{ cm}^3}{250\text{ cm}^3}\\ &= 0.2\text{ mole}\end{aligned}$$

Molarity of the solution is 0.2M HC1.

$$\begin{aligned}\text{(b) Moles of HC1 in 500 cm}^3 &= 0.05\text{ moles}\\ \text{Moles in 1000 cm}^3 &= \frac{0.05\text{ Moles}\times 1000\text{ cm}^2}{500\text{ cm}^2}\\ &= 0.1\text{ mole}\end{aligned}$$

Molarity of the solution is 0.1M HC1.

(iii) Comment on the results.

Although equal volumes of solution containing equal moles were diluted, the molarities of the resulting solutions are different. The solution to which less water was added is more concentrated than the one to which more water was added.

Solutions of different molarities may be prepared by adding water to equal volumes of solutions of the same concentration and making up to different volumes.

3. **Derive a general formula of preparing a less concentrated solution from a more concentrated solution.**

The molarity of the dilute solution (M_2) is obtained by multiplying the molarity of the concentrated solution (M_1) with the volume of the concentrated solution (V_l), and dividing by the final volume of the dilute solution (V_2);

Thus: $M_2 = \dfrac{M_1V_1}{V_2}$

Therefore, $M_2V_2 = M_1V_1$

The following worked examples show how this relationship is applied in calculations.

Example 1

Calculate the volume of a 5.0 M H_2SO_4 solution that will be required to make a 1000 cm^3 solution of 0.05 M H_2SO_4.

Solution

$M_1V_1 = M_2V_2$

$M_1 = 5.0\ M$, $V_1 = ?$ $M_2 = 0.05\ M$, $V_2 = 1000\ cm^3$

$$V_1 = \frac{M_2V_2}{M_1}$$

Substituting

$$= \frac{0.05 \times 1000\ cm^2}{5.0}$$

$$= 10\ cm^3$$

Example 2

When 50 cm^3 of 2M potassium hydroxide solution was diluted, the final concentration was O.1M. Calculate the volume of the diluted solution.

Solution:

$$M_1V_1 = M_2V_2$$

$$M_1 = 2,\ V_1 = 50\ cm^3$$

$$M_2 = 0.1,\ V_2 = ?$$

Substituting

$$V_2 = \frac{2 \times 50}{0.1}$$

$$V_2 = 1000\ cm^3$$

Example 3

Calculate the volume of 18 M sulphuric(IV) acid, H_2SO_4, that will be required to prepare 3.6 litres of 0.2 M sulphuric(VI) acid.

Solution

From $M_1V_1 = M_2V_2$

$$V_1 = \frac{M_2V_2}{M_1}$$

But, $M_2 = 0.2\ M$

$V_2 = 3.6$ litres

$M_1 = 18\ M$

$$\Rightarrow V = \frac{0.2 \times 3.6}{18}$$

$= (0.2 \times 0.2)$ litres

$= 0.04$ litres (Convert into cm^3)

$= 40\ cm^2$

Example 4

A label on a bottle containing sulphuric(IV) acid has the following information:

- **Density = 1.836 g/cm^3**
- **Percentage purity = 98%**

- **Relative formula Mass = 98**

Calculate:

(a) The concentration of the acid.

(b) The volume of the concentrated sulphuric(VI) acid that should be diluted to produce 2 litres of 2 M sulphuric(VI) acid.

Solution:

(a) $$\text{Mass of } H_2SO_4 \text{ in } 1\text{cm}^3 = \frac{1.836 \text{ g}}{1 \text{ cm}^3} \times 1\text{cm}^3$$

$$= 1.836 \text{ g}$$

$$\text{Mass in } 100 \text{ cm}^3 = 1.836 \times 1000$$

$$= 1836 \text{ g}$$

$$\text{Mass of pure acid} = 1836 \text{ g} \times \frac{98}{100}$$

$$= 1799.28 \text{ g}$$

$$\text{Molarity of acid} = \frac{\text{Mass per litre}}{\text{Relative formula Mass}} = \frac{1799.28}{98}$$

$$= 18.36 \text{ M}$$

(b) From the equations,

$$M_1V_1 = M_2V_2 \qquad V1 = \frac{M_2V_2}{M_1} \qquad \text{Where, } M_2 = 2,\ V_2 = 2, \text{ and } M_1 = 18.36$$

$$\text{Substituting:} \quad V_1 = \frac{2 \times 2}{18.36}$$

$$V_1 = 0.217864923 \text{ litres} = 218 \text{ cm}^3$$

4. Stoichiometry of Chemical Equations

Stoichiometry is a **quantitative relationship between reactants and products in a chemical reaction.**

When particles combine in a chemical reaction, they do so in specific proportions. The proportion in which they combine is referred to as **mole ratio.**

In a balanced chemical equation, the quantities of reactants and products are in a whole number ratio. For example in the equation;

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s),$$

The reactants and product are in the ratio 2:1:2.

A chemical equation in which the reactants and products are in whole number ratios is called a **stoichiometric equation.**

Experiment to determine the equation for the reaction between iron metal and copper(II) sulphate solution.

Weigh accurately 0.56 g of freshly acquired iron filings and put them in a clean dry weighed beaker.

Take about 20 cm^3 of 2 M copper(II) sulphate solution and transfer it into a clean boiling tube, and heat it until it is nearly boiling. Add the hot copper (II) sulphate solution to the iron filings in the beaker and stir the mixture to cool and record your observations.

Carefully decant as much of the liquid as possible, ensuring that no solid is lost. Wash the solid twice with distilled water. Place the beaker with the solid residue on a sand bath. When the solid is sufficiently dry remove the beaker from the sand bath and allow to cool. Weigh the beaker again with its contents. Record your results in a table.

Mass of iron filings	0.56 g
Mass of copper displaced	0.64 g

Discussion Questions

1. **What observations were made when iron filings and copper(II) sulphate reacted?**

When iron reacts with copper ions in solution, the products formed are brown copper metal and iron(II) sulphate in solution.

Copper(II) sulphate + Iron metal $\longrightarrow$ Iron(II) sulphate + Copper metal.

2. **Why was it necessary to use excess copper(II) sulphate?**

To ensure that all the iron metal used is reacted, blue copper(II) sulphate is used in excess, hence the solution remains blue when the reaction is completed.

3. **Calculate the:**

   (i) Number of moles of iron metal reacted.

Moles of iron used = $\frac{0.56\text{ g}}{56\text{ g mol}^{-1}}$ = 0.01 mole

   (ii) Number of moles of copper metal displaced from the solution. (Fe = 56, Cu = 63.5).

Moles of copper metal displaced = $\frac{1.64\text{ g}}{63.5\text{ g mole}}$ = 0.01007 mole

1. **Write the equation of the reaction between iron and copper ions in solution.**

From the masses of iron metal used, and the copper metal displaced from the solution, the reacting mole ratio of iron metal and copper ions in solution can be calculated.

Thus, Fe: Cu^{2+}

0.01: 0.01007
1.0: 1.007

Whole number mole ratio 1:1

Hence, one mole of iron atoms reacts with one mole of copper ions to produce one mole of atoms of copper metal. Since one mole of copper sulphate contains one mole of copper atoms, the equation of the reaction is

$$Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$$

Experiment to determine the equation for the reaction between lead(II) nitrate and potassium iodide.

Take six test-tubes and label them, 1 to 6. Run 5 cm^3, of 1.0 M potassium iodide solution from a burette into each one of them. Add 1.0 cm^3 of 1.0 M lead(II) nitrate solution to the test-tube labelled 1, and stir the mixture well with a glass rod. Add about 5 drops of ethanol to the mixture, stir, and place it in a test-tube rack.

Add 1.5 cm^3, 2.0 cm^3, 2.5 cm^3, 3.0 cm^3 and 3.5 cm^3 of the 1.0 M lead(II) nitrate to the test-tubes labelled 2, 3, 4, 5, and 6 respectively. Add about 5 drops of ethanol to each test-tube, stir and allow to settle. Measure the height of the precipitate in each tube in (mm) and record the measurements in a table. Plot a graph of the heights of the precipitate against the volume of lead(II) nitrate solution added.

Sample results

Test tube number	1	2	3	4	5	6
Volume of 1 M $Pb(NO_3)_2$ (cm^3)	1.0	1.5	2.0	2.5	3.0	3.5
Height of precipitate (mm)	8	12	16	20	20	20

Discussion Questions

1. **What was observed on mixing the two solutions?**

When lead(II) nitrate solution reacted with potassium iodide solution, a bright-yellow precipitate of lead(II) iodide was formed. The insoluble lead(II) iodide settled at the bottom of the test-tube when allowed to rest.

2. **What was the purpose of adding ethanol to the mixture?**

Ethanol was added to the mixture to speed up the settling of the precipitate. Warming the mixture also hastens the settling process.

3. **Calculate:**

(a) Number of moles of KI in 5 cm^3 of 1.0 M KI solution.

Moles of KI in 5.0 cm3 = $\frac{1.0 \times 5\ cm^3}{1000\ cm^3}$ = 0.005 moles

(b) Number of moles of $Pb(NO_3)_2$ which reacted completely with 5.0 cm^3 of 1.0 M KI.

Moles of $Pb(NO_3)_2$ that reacted completely, $\frac{1.0\ mole \times 2.5\ cm^3}{1000\ cm^3}$ = 0.0025 moles

4. **The heights of the precipitate remained constant in the test-tubes labelled 4, 5 and 6. Explain.**

The heights of the precipitate in test-tubes labeled 4 to 6, remained constant because lead(II) ions were in excess and all the iodide ions had reacted.

2. **How many moles of KI would react with one mole of lead(II) nitrate?**

Thus, lead ions and iodide ions reacted in the ratio 0.0025 : 0.005

i.e., $Pb^{2+} : I^-$

0.0025 : 0.005

Whole number ratio is 1 : 2

6. **Write:**

(a) Balanced chemical equation.

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI(s) + 2KNO_3(aq)$$

(b) Ionic equation for the reaction between lead(II) nitrate and potassium iodide.

Both the lead(II) ions (Pb^{2+}) and iodide ions (1^-) have undergone a chemical change. However nitrate ions and the potassium ions remained unchanged after the reaction. The ionic equation for the reaction can therefore be represented as:

$$Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$$

Experiment: To determine the equation for the reaction between barium ions and carbonate ions.

Measure exactly 25.0 cm^3 of 0.2 M barium chloride solution, and place it in a beaker. Add 25 cm^3 of 0.2 M sodium carbonate solution. Stir the mixture and allow it to settle. Filter the solution, and dry the solid residue between a filter paper. Do not dry by heating. Transfer the dry solid onto a filter paper which you have weighed and recorded its weight. Record your results.

Discussion Questions

1. **What ions are present in:**
   - **(a) barium chloride solution?**
     Ba2+ and Cl– ions
   - **(b) sodium carbonate solution?**
     Na+ and CO_3^{2-} ions

3. **What observations were made when the two solutions were mixed? Explain.**

A White precipitate is formed.

On mixing the solutions, barium ions combine with carbonate ions to form insoluble barium carbonate. Sodium and chloride ions remain in solution unchanged.

4. **The solid residue is not dried by heating. Explain.**

The barium carbonate obtained should not be dried by heating because the compound easily decomposes on heating.

$$BaCO_3(s) \xrightarrow{\text{heat}} BaO(s) + CO_2(g)$$

5. **Calculate the number of moles of barium ions and carbonate ions used in the reaction. What assumptions have you made?**

Assuming that 25.0 cm^3 of 0.2M sodium carbonate reacted completely with 25.0 cm^3 of 0.2 M barium chloride., then

(a) Moles of barium chloride used

$$= 25\text{ cm}^3 \times \frac{25\text{ cm}^3 \times 0.2\text{ moles}}{1000\text{ cm}^3} = 0.005\text{ mole}$$

$$\text{(b) Moles of sodium carbonate used} = \frac{25\text{ cm}^3 \times 0.2\text{ moles}}{1000\text{ cm}^3}$$

$$= 0.005\text{ mole.}$$

6. **Calculate the number of moles of barium carbonate formed.**

Suppose 0.985 g of barium carbonate were formed, then:

$$\text{Moles of BaCO}_3\text{(s) formed} = \frac{0.985\text{ g}}{197\text{ mol}^{-1}}$$

$$= 0.005\text{ mole.}$$

7. **What is the whole number ratio of barium carbonate formed to that of barium ions and carbonate ions used?**

1:1. One mole of barium chloride reacts with one mole of sodium carbonate to produce one mole of barium carbonate.

7. **Write the:**
   (i) **stoichiometric equation.**

The equation for the reaction is:

$$BaCl_2(aq) + Na_2CO_3(aq) \longrightarrow BaCO_3(s) + 2NaCl(aq)$$

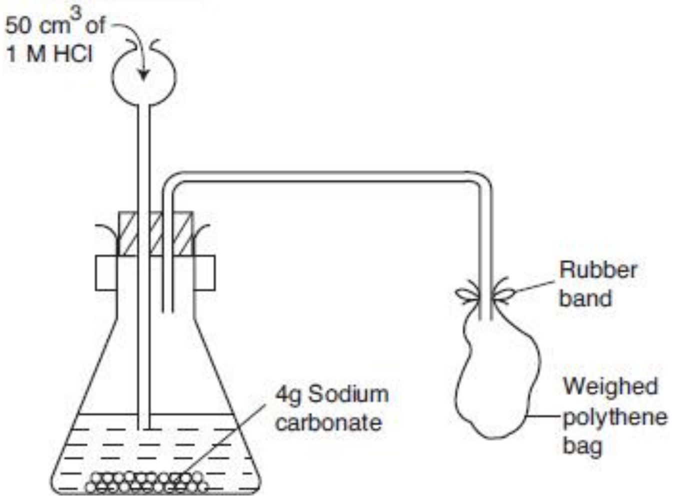
   (ii) **ionic equation for the reaction.**

The ionic equation for the reaction is:

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow BaCO_3(s).$$

Experiment to determine the equation of reaction between hydrochloric acid and sodium carbonate.

Weigh a polythene bag and a rubber band and record their masses. Place 2.65 g of sodium carbonate in a conical flask. Stopper the flask with a cork fitted with a thistle funnel and delivery tube. Connect the delivery tube to the weighed polythene bag and tie the mouth of the bag tightly with the rubber band as shown below.



Measure exactly 50.0 cm^3 of 1M hydrochloric acid solution and transfer the acid to the thistle funnel. Run the acid from the thistle funnel into the conical flask. Collect the gas evolved in the bag. Remove the bag when the reaction has stopped. Weigh the bag with the gas again and record the mass.

Discussion Questions.

1. **Calculate:**
   (a) **Moles of the acid used and Moles of the gas collected.**
2. **Write the equation for the reaction between hydrochloric acid and sodium carbonate.**

$$\text{Moles of } H^+(aq) \text{ from HC1} = \frac{50 \text{ cm}^3 \times 1.0 \text{ mole}}{1000 \text{ cm}^3} = 0.05 \text{ mole.}$$

$$\text{Moles of } CO_3^{2-}(aq) \text{ from } Na_2CO_3 = \frac{2.65 \text{ g}}{106 \text{ g mol}^{-1}} = 0.025 \text{ mole.}$$

Assuming all the acid and carbonate reacted and the mass of carbon(IV) oxide collected was found to be 1.1g,

then, moles of carbon(IV) oxide $= \frac{1.1\text{ g}}{44\text{ mol}^{-1}}$

= 0.025 mole.

Thus, 0.025 moles of CO_3^{2-}(aq) reacted with 0.05 moles of H^+(aq) to form 0.025 moles of carbon(IV) oxide.

The whole number mole ratio is:	H^+	:	CO_3^{2-}	:	CO_2
	0.05	:	0.025	:	0.025
	2	:	1	:	2

The ionic equation is:

$$2H^+(aq) + CO_3^{2-}(aq) \longrightarrow CO_2(g) + H_2O(l)$$

and the stoichiometric equation is

$$2HCl(aq) + Na_2CO_3(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H2O(l)$$

5. Volumetric Analysis

Volumetric analysis, also called **titration,** is a method of quantitative chemical analysis in which the amount of a substance is determined by measuring volumes of solution. One solution of known concentration and volume is reacted carefully with another of unknown concentration to determine the reacting volumes. This data is used to determine the concentration of the said solution.

A solution whose exact concentration is known is called a **standard solution.**

The apparatus used in volumetric analysis are **pipettes** and **burettes.**

A pipette is designed to deliver a definite volume of a solution, e.g., 10.0 cm^3, 20.0 cm^3 or 25.0 cm^3.

Before use, the pipette **must be rinsed with the solution to be drawn.** It is then filled by sucking the liquid to **a few centimeters above the calibration mark.**

The solution is then allowed to **drain slowly under gravity until the meniscus is at the same level with the mark.** Sucking can be done using the mouth or a pipette filler. Pipette fillers **must** be used especially when solutions are toxic.

A burette is designed to deliver variable volumes of solution as needed during a titration. It is graduated from 0.0 to 50.0 cm^3 with unit intervals of 0.1cm^3.

Before use, a burette must be **rinsed with the solution to be used in it.** It is then filled carefully **beyond** the 0.0 mark and the level of the solution is **adjusted until the bottom of the meniscus** is at the same level with the graduation mark.

Titration

Titration is a quantitative analysis process using solutions. A solution of known concentration is added gradually to another solution of unknown concentration until the reaction between the two solutions is complete.

The point at which the reaction is complete is called the **end point**. An indicator is used to identify the **end point.**

The volume of the solution that run out of the burette in every titration experiment is known as a **titre.**

After the titration the data should be arranged in tabular form. The volume of the pipette used should be shown always.

The volume of pipette used .... cm^3.

Table of results of a titration experiment

Titration	**1**	**2**	**3**
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume (dilute) used (cm^3)			

Average volume used = ... cm^3.

Average titre = ... cm^3.

The titre should be recorded in **one or two decimal places. When in two decimal places, the last digit should be zero or five.**

The volumes used for the three experiments should be **consistent and have an accuracy of ± 0.2 cm^3.** The average titre is given by **adding the volumes of the consistent titres used divided by number of experiments done.**

Indicators are useful in pointing out the precise end-point of a titration. The choice of an indicator for every titration experiment therefore depends on its properties as shown below.

Indicator	**Colour in Acid solution**	**Colour in Alkali solution**	**Colour in neutral solution**
Phenolphthalein	Colourless	Pink	Colourless
Methyl orange	Pink	Yellow	Orange
Screened Methyl orange	Red	Green	Grey

Choice of Indicators

The table bellow is a guide on how to choose an appropriate indicator for different acid base titrations.

Titration		**Suitable Indicator for use**
Strong acid	Strong alkali	Any indicator
Strong acid	Weak base	Screened methyl orange or methyl orange
Weak acid	Strong base	Phenolphthalein

The three basic Titrations are direct, back and redox Titrations.

Direct Titration

This is a type of titration that involves the addition of a standard solution to a fixed volume of the analyte (substance being anlaysed), or vice versa with the sole aim of determining either:

(a) Concentration.

(b) Relative atomic mass of the elements of the sample analyte or,

(c) The percentage composition of the constituent elements.

A good example is **a netralisation reaction**.

Experiment 1: What is the equation for the reaction between sodium hydroxide and hydrochloric acid?

Using a pipette, transfer 25 cm^3 of 0.1 M sodium hydroxide solution into a conical flask. Add 2 drops of phenolphthalein indicator. Fill a clean burette with 0.1 M hydrochloric acid and read the initial level of the acid accurately.

Run the acid solution in about 1 cm^3 portions, swirl the liquid in the flask after each addition of the acid until the colour of the solution changes to colourless permanently.

Record the final reading of the burette. Rinse the conical flask and repeat the experiment. Record the results in the table below.

The volume of the pipette used is .... cm^3.

Titration	**1**	**2**	**3**
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume acid used (cm^3)			

Average volume of acid used cm^3.

Discussion Questions

1. **What was the volume of hydrochloric acid that reacted with 25 cm^3 of 0.1 M sodium hydroxide solution?**

   Take the average titre as 25.0 cm3

2. **How many moles of hydrochloric acid are there in the average titre?**

Taking the average titre as 25.0 cm^3, therefore the number of moles of hydrochloric acid used are:

$$\frac{25}{1000} \times 0.1 = 0.0025$$

3. **How many moles of sodium hydroxide are there in 25 cm^3 of 0.1 M sodium hydroxide solution?**

Number of moles of sodium hydroxide used are:

$$\frac{25}{1000} \times 0.1 = 0.0025$$

4. **Write down the equation for the reaction between hydrochloric acid and sodium hydroxide.**

The mole ratio of acid to base is: 1:1, therefore, the equation is written as:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(1).$$

Mole ratio: 1:1.

- The reaction between an acid and an alkali forms a salt and water only. This kind of reaction is called **neutralisation.**
- The ionic equation is

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

Experiment 2: What is the equation for the reaction between sodium hydroxide and sulphuric(VI) acid?

Pipette 25.0 cm^3 of 0.1 M sodium hydroxide solution in to a conical flask. Add 2 drops of phenolphthalein indicator and titrate with 0.1 M sulphuric(VI) acid from a burette. Record the results as shownm

Volume of pipette used .... cm^3

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume (dilute) used (cm^3)			

Average volume acid used (cm^3)

Discussion Questions

1. **What volume of sulphuric(VI) acid neutralised 25 cm^3 of 0.1 M sodium hydroxide solution?**
   Taking the average titre as 12.5 cm3 of sulphuric(VI) acid
2. **How many moles of sulphuric(VI) acid are there in the average titre?**
   Moles of sulphuric(VI) acid = $\frac{125}{1000} \times 0.1 = 0.00125$
3. **How many moles of sodium hydroxide are there in 25 cm^3 of 0.1 M sodium hydroxide?**
   Moles of sodium hydroxide = $\frac{25}{1000} \times 0.1 = 0.0025$
4. **How many moles of sodium hydroxide react with 1 mole of sulphuric(VI) acid? Write the equation for the reaction.**

The mole ratio of sulphuric(VI) acid: sodium hydroxide is 0.00125: 0.0025 which is 1: 2. Therefore, equation for reaction is:

$$H_2SO_4(aq) + NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(aq)$$

Ionic equation:

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

5. **From experiment (1), and b2) compare the volumes of 0.1 M sulphuric(VI) acid and 0.1 M hydrochloric acid which neutralise 0.1 M sodium hydroxide solution in each case.**

Comparing experiment 1 and 2, the volume of 0.1 M hydrochloric acid is twice the volume of 0.1 M sulphuric acid that reacted with 0.1 M sodium hydroxide solution. This means that sulphuric acid is able to release twice the number of hydrogen ions as compared to hydrochloric acid.

- The number of hydrogen ions which can be produced by a molecule of an acid on ionising is referred to as the **basicity** of the acid. Thus, the basicity of HCl and HNO_3 is one (1) while that of H_2SO_4 and (HOOC–COOH) ethanedioic acid is two(2).

Name of acid	Formula	Number of replaceable hydrogen	Basicity
Hydrochloric acid	HCl	1	Monobasic
Nitric acid	HNO_3	1	Monobasic
Ethanoic acid	CH_3COOH	1	Monobasic
Sulphuric acid	H_2SO_4	2	Dibasic
Ethanedioic acid	HOOC–COOH	2	Dibasic
Phosphoric acid	H_3PO_4	3	Tribasic

For example:

$$HCl(aq) \xrightarrow{water} H^+(aq) + Cl^-(aq)$$

$$H_2SO_4(l) \xrightarrow{water} 2H^+(aq) + SO_4^{2-}(aq)$$

$$CH_3COOH(aq) \xrightarrow{water} H^+(aq) + CH_3COO^-(aq)$$

- In ethanoic acid only one hydrogen atom is replaceable, the one attached to O–H, the others are not replaceable.

Experiment 3: How can hydrochloric acid solution be standardised using sodium carbonate?

Put about 100 cm^3 of distilled water in 250 cm^3 volumetric flask. Measure 2 cm^3 of concentrated hydrochloric acid in a 10 ml measuring cylinder and transfer it to the flask. Add distilled water to make up to the 250 cm^3 mark. Rinse the burette with the acid. Pour the acid solution into the burette.

Weigh exactly 1.325 g of anhydrous sodium carbonate, dissolve in water and make up to 250 cm^3 in a flask. Transfer 25 cm^3 of the sodium carbonate solution using a pipette into a conical flask. Add 2 or 3 drops of methyl orange indicator and titrate with the hydrochloric acid solution until the indicator changes colour from yellow to pink. Record your data.

Volume of pipette used is **25 cm^3**

Titration	1	2	3
Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of acid used (cm^3)			

Discussion Questions

1. **What is the meaning of the term 'standardisation'?**
   Standardisation is the process by which an unknown concentration of a solution is determined by use of a standard solution in a titration experiment
2. **Determine the average volume of acid used**.
   Suppose the average volume of the acid used is 22.5 cm3
3. **Write the equation for the reaction between sodium carbonate and hydrochloric acid.:**

$$Na_2CO_3(aq) + 2HCl\ (aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

4. **Determine:**

   (a) The molarity of the sodium carbonate solution.

Relative formula mass of Na_2CO_3 is 106 g.

If 1.325 g of Na_2CO_3 was dissolved in 250 cm^3, then in 1000 cm^3, the mass would be:

$$\frac{1.325}{250} \times 1000 = 1.325 \times 4 = 5.3g$$

Hence, molarity $= \dfrac{\text{Mass in grams/litre}}{\text{Molar mass}} = \dfrac{5.3}{106} = 0.005\ M$

   (b) The number of moles of sodium carbonate that reacted with the acid.

Moles of Na_2CO_3 in 25 cm^3 $= \dfrac{25}{1000} \times 0.005 = 0.00125$ mole.

4. **Determine the number of moles of acid used.**

Mole ratio Na_2CO_3: HCl is 1: 2.

Therefore, moles of HCl in 22.5 cm3 = 2 × 0.00125 = 0.0025 mole

6. **Calculate the concentration of the acid in:**

   (a) mol $dm.^{-3}$

moles of HCl in 1000 cm^3 $= \dfrac{0.0025 \times 1000}{22.5} = 0.11$ mole.

Molarity of hydrochloric acid is 0.11 M.

   (b) g $dm.^{-3}$

Molar mass of HCl is 36.5 g. Concentration of HCl in grams/litre is:

36.5 × 0.11 = 4.015 grams/litre = 4.0 g/L.

Back Titration

Back titration is a method of volumetric analysis used to **determine the concentration (or amount) of a reactant**.

It involves reacting quantities of the substance being analysed with an excess amount of a suitable reagent whose volume and concentration is known. The quantity of the excess is then measured through titration. The amount of the reagent that reacted with the analyte is determined by subtracting the excess amount from the initial. The data is then used to determine the concentration (or amount) of the analyte.

Experiment 1: What is the atomic mass of divalent metal M?

Weigh exactly 0.50 g of a divalent metal carbonate, MCO_3. Put the weighed carbonate in a conical flask. Add to it 30.00 cm^3 of 0.50 M hydrochloric acid. Add two to three drops of phenolphthalein in the resulting solution. Titrate the solution against a 1.0 M sodium hydroxide solution until the colour of the solution just turns pink permanently. Record your results in a table. Repeat the experiment three times to complete the table.

Test	1	2	3	4
Final burette reading (cm^3)	5.5	10.6	15.6	20.5
Initial burette reading (cm^3)	0.0	5.5	10.6	16.6
Volume of base used (cm^3)	5.5	5.1	5.0	4.9

Discussion Questions

1. **Calculate the average volume of base used to neutralise the excess acid.**

The average titre (base) used is determined by considering the values which are consistent to within ± 0.20 range only.

$$\text{Average volume of base used} = \frac{5.1 + 5.0 + 4.9}{3} \text{ cm}^3$$
$$= 5.0 \text{ cm}^3$$

2. Determine:

(a) The number of moles of base used.

Volume of base that neutralised excess acid = 5.0 cm^3. Therefore, 1000 cm^3 of base contains 1.0 mole.

$$\therefore\ 5 \text{ cm}^3 \text{ of the base would contain} = \frac{1.0 \text{ mole}}{1000 \text{ cm}^3} \times 5 \text{ cm}^3$$
$$= 0.005 \text{ mole.}$$

(b) (i) The volume of acid that was neutralised by the base.

The equation for the neutralisation reaction is:

$$\text{HCl (aq)} + \text{NaOH (aq)} \longrightarrow \text{NaCl (aq)} + H_2O \text{ (1)}$$

Thus, the reacting mole ratio of acid to base, HCl: NaOH is 1: 1. Hence, moles of excess acid neutralised was 0.005 mole. But, the molarity of the acid was 0.50 M HCl and 0.50 moles was contained in a 1000 cm^3.

$$0.005 \text{ mole would be contained in} = \frac{1000 \text{ cm}^3 \times 0.005 \text{ moles}}{0.50 \text{ mole}}$$
$$= 10.0 \text{ cm}^3$$

Thus, volume of excess hydrochloric acid after reaction with the carbonate is 10.0 cm^3.

This means only 20.0 cm^3 of the 30.0 cm^3 of acid added to the carbonate to completion.

(ii) The number of moles of acid that reacted with the carbonate.

1000 cm^3 of acid contained = 0.50 mole.

$$20 \text{ cm}^3 \text{ of acid would contain} = \frac{0.50 \text{ mole} \times 20 \text{ cm}^3}{1000 \text{ cm}^3}$$
$$= 0.01 \text{ mole.}$$

3. (i) Write the equation for the reaction between the acid and carbonate.

$$MCO_3\text{ (s)} + 2HCl\text{(aq)} \longrightarrow MCl\text{(aq)} + CO_2\text{ (g)} + H_2O\text{ (l)}$$
$$1 \quad : \quad 2$$

(ii) Determine the number of moles of the carbonate in the sample.

The reacting mole ratio of the carbonate to the acid is 1:2

MCO3:2HCl

1:2

Thus moles of the carbonate that reacted with 0.10 mole of the acid is $\frac{1}{2}$ of 0.10 mole.

$$\text{Mole of carbonate} = \frac{\text{moles of acid}}{2}$$

$$= \frac{0.10 \text{ mole}}{2} = 0.005 \text{ mole.}$$

(iii) Determine the relative formula mass of the carbonate.

Mass of 0.005 mole = 0.50 g

$$\text{Mass of 1.0 moles} = \frac{0.50 \text{ g}}{0.005 \text{ mole}} = 100 \text{ g}$$

$$\text{R.F.M. } (MCO_3) = 100 \text{ g}$$

(iv) Determine the relative atomic mass of M.

$$\text{R.F.M. } (MCO_3) = 100 \text{ g}$$

$$M + 12 + 48 = 100 \text{ g}$$

$$\text{R.A.M. of M} = (100 - 60) \text{ g} = 40 \text{ g}$$

Alternatively,

$$\text{R.F.M. of carbonate} \times 0.005 \text{ mole} = 0.50 \text{ g}$$

$$0.005 \times (M + 12 + 48) = 0.50 \text{ g}$$

$$0.005\ M + 0.30 \text{ g} = (0.50 - 0.30)$$

$$0.005\ M = 0.20 \text{ g}$$

$$M = \frac{0.20}{0.005} = 40$$

$$\text{R.A.M. of M} = 40\text{g}$$

Redox Titrations.

Titrations involving redox reactions do not require indicators because the colour of some of the reagents change when their oxidation states change.

Examples of redox titrations involve the use of potassium manganate(VII) and potassium dichromate(VI). Manganate(VII) ions are purple in colour. When reduced to manganese(II), the solution becomes colourless.

$$\underset{\textit{Purple}}{Mn^{7+}(aq) + 5\ e^-} \longrightarrow \underset{\textit{Colourless}}{Mn^{2+}\ (aq)}$$

Dichromate(VI) ions are orange in colour. When reduced to chromium(III) ions, the solution becomes green.

$$\underset{\textit{Orange}}{Cr^{6+}\ (aq) + 3\ e^-} \longrightarrow \underset{\textit{Green}}{Cr^{3+}\ (aq)}$$

Redox titrations are used to standardise solutions and to determine the purity of compounds. Potassium manganate(VII) and potassium dichromate(VI) solutions **should be acidified before use in titrations. This is done to ensure complete reduction to manganese(II) and chromate(III) ions, respectively.**

Experiment 1: How can potassium manganate(VII) solution be standardised using iron(II) salt?

Weigh accurately 9.8 g of the iron(II) salt, $FeSO_4.(NH_4)_2SO_4.6H_2O$) and dissolve in distilled water that has been boiled and cooled to remove air. Add more distilled water to make up to 250 cm^3.

Transfer 25.0 cm^3 of the solution into a conical flask using a pipette. Add 10 cm^3 of dilute sulphuric acid and titrate with potassium manganate(VII) from a burette until a persistent pink colour appears. Record the volume of potassium manganate(VII) used. Repeat the experiment twice and record your results in a table

Titration	1	2	3
Final burette readings (cm^3)			
Initial burette reading (cm^3)			
Volume of $KMnO_4$ used (cm^3)			

Discussion Questions

1. **Determine the average volume of potassium manganate solution.**

Suppose the average titre is 24.0 cm3.

2. Calculate the molarity of the iron(II) salt solution.

Molar mass of the iron(II) salt = 392 g.

Number of moles of iron(II) ions in 250 cm^3 solution = $\frac{9.8\text{ g}}{392\text{ mol}^{-}} = 0.025\text{ mole}$

Moles in 1000 cm^3 = $\frac{0.025\text{ mole} \times 1000\text{ cm}^3}{250\text{ cm}^3}$ = 0.1 mole

Thus molarity of the iron(II) salt is 0.1 M.

3. **Calculate the number of moles of iron(II) ions used.**

Number of moles of iron(II) ions in 25 cm^3.

Solution, = $\frac{25.0\text{ g} \times 0.1}{1000}$ = 0.0025 mole.

4. **Given the equation of the reaction is:**

$$MnO_4^-(aq) + 8H^+(aq) + 5\,Fe^{2+}(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

Determine:

(a) The number of moles of MnO_4^- ions in 24.0 cm^3 solution.

Reacting mole ratio $MnO_4^- : Fe^{2+}$

1 : 5

Therefore, moles of MnO_4^- in 24.0 cm^3, = 0.0025 × $\frac{1}{5}$

$= 0.0005\text{ mole.}$

(b) **The molarity of the $KMnO_4$(aq) solution**.

$$\text{Molarity of } KMnO_4 = \frac{0.005\text{ moles} \times 100\text{ cm}^3}{24.0\text{ cm}^3}$$

$$= 0.020833\text{ M} = 0.02\text{ M}$$

Experiment 2: What is the amount of water of crystallisation in ammonium iron (II) sulphate?

Ammonium iron(II) sulphate crystals have the following formula $(NH_4)_2SO_4$. $FeSO_4$. nH_2O. Weigh accurately 8.8 g of the salt and dissolve in 50 cm^3 of 2.0 M sulphuric(IV) acid. Make up to 250 cm^3 of solution with distilled water. Pipette 25.0 cm^3 of this solution and pour it into a conical flask. Titrate this against the acidified manganate(VII) used in experiment 1 until the solution becomes colourless. Record your results in a table.

Titration	**1**	**2**	**3**
Final burette readings (cm^3)			
Initial burette reading (cm^3)			
Volume of $KMnO_4$ used (cm^3)			

Discussion Questions

1. Calculate the:

(a) Average volume of $KMnO_4$ solution used.

Suppose the average volume of the 0.02 M KMnO4 solution used was 22.5 cm3

(b) Number of moles of $KMnO_4$ that reacted.

Moles in 1000 cm^3 = 0.02 mole.

$$\text{Number of moles in } 22.5\ cm^3 = \frac{0.020\text{ mole}}{1000\ cm^3} \times 22.5\ cm^3$$
$$= 0.00045\text{ mole.}$$

2. Using the following ionic equation determine the number of moles of iron(II) salt in 25 cm^3 of the solution used.

$$MnO_4^-(aq) + 8H^+(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l).$$

But, from the ionic equation, the reacting mole ratio of MnO_4^- : Fe^{2+} is 1 : 5.

thus, moles of iron(II) ions reacted = 0.00045 × 5 mole

= 0.00225 mole.

3. Determine the molarity of the iron(II) salt solution.

0.00225 mole were contained in 25 cm^3.

That is, 25 cm^3 contained = 0.00225 mole.

$$250\ cm^3 \text{ contained} = \frac{0.00225 \times 250}{25}$$
$$= 0.0225\text{ mole.}$$

Hence, molarity of iron(II) salt is determined as follows:

250 cm^3 contains 0.0225 mole

$$1000\ cm^3 \text{ contains} = \frac{0.0225\text{ mole} \times 1000\ cm^3}{250\ m^3} = 0.09\text{ mole.}$$

4. Determine:

(a) The value of 'n' in the formula, $(NH_4)_2SO_4$. $FeSO_4$. nH_2O.

250 cm^3 solution contains 8.8 g

$$1000 \text{ solution contains } \frac{8.8\text{ g} \times 1000\ cm^3}{250\ cm^3} = 35.2\text{ g}$$

Mass of 0.09 mole = 35.2 g

$$\text{Mass of 1.0 mole} = \frac{35.2\ \text{g} \times 1.0\ \text{mole}}{0.09\ \text{mole}}$$

$$= 391.1\text{g}$$

$$\text{To find 'n'}, 36 + 32 + 64 + 56 + 32 + 64 + 18n = 391.1$$

$$284+18n = 391.1$$

$$18n = 391.1–284$$

$$n = \frac{107.11}{18} = 5.9506 \simeq 6.0$$

(b) The formula of the iron(II) salt.

The formula of the salt is: $(NH_4)_2SO_4.FeSO_4.6H_2O$

6. Atomicity and Molar Gas Volume

Atomicity

Atomicity is the number of atoms in one molecule of an element.

For example, oxygen (O_2), hydrogen (H_2) and nitrogen (N_2) have two atoms per molecule, thus they are **diatomic.**

Ozone (O_3) has three atoms in a molecule and is therefore **triatomic.**

The noble gases are said to have an atomicity of one (**monoatomic**) even though they are not molecular.

Atomicity of some selected gaseous elements.

Gas	Chemical formula	Atomicity
Helium	He	1
Argon	Ar	1
Neon	Ne	1
Nitrogen	N2	2
Hydrogen	H2	2
Oxygen	O2	2
Chlorine	Cl2	2
Ozone	O3	3

Bromine	Br2	2

Molar Gas Volume

The volume occupied by one mole of any gas at given temperature and pressure is called the **molar gas volume.** Its value at s.t.p is **22.4 dm^3**. At **r.t.p**, the molar gas volume is **24.0 dm^3**.

Combining Volumes of Gases

The relationship between reacting volumes of gases is summarised by Gay **Lussac's** Law: *When gases react, they do so in volumes that bear a simple ratio to one another and to the volumes of the product if gaseous, temperature and pressure remaining constant*.

For example:

$$\begin{array}{lll} \text{Nitrogen} + \text{Hydrogen} & \longrightarrow & \text{Ammonia Gas} \\ 10\ \text{cm}^3 \quad 30\ \text{cm}^3 & & 20\ \text{cm}^3 \\ 1\ \text{vol} \quad\ \ 3\ \text{vol} & & 2\ \text{vol} \end{array}$$

Example 2

$$\text{Carbon(II) oxide gas} + \text{Oxygen gas} \longrightarrow \text{Carbon(IV) oxide gas}$$

$$\begin{array}{lll} 2CO(g) + O_2(g) & \longrightarrow & CO_2(g) \\ 20\ \text{cm}^3 \quad 10\ \text{cm}^3 & & 10\ \text{cm}^3 \\ 2\ \text{vol} \quad\ \ 1\ \text{vol} & & 1\ \text{vol} \end{array}$$

Example 3

$$\text{Hydrogen gas} + \text{Oxygen gas} \longrightarrow \text{Steam}$$

$$\begin{array}{llll} 2H_2(g) & + O_2(g) & \longrightarrow & 2H_2O(g) \\ 30\ \text{cm}^3 & 15\ \text{cm}^3 & & 30\ \text{cm}^3 \\ 2\ \text{vol} & 1\ \text{vol} & & 2\ \text{vol} \end{array}$$

Avogadros' Law also states that ***Equal volumes of gases will contain equal number of molecules.***

The number of molecules per mole of any gas is **6.02 × 10^{23}** and occupies **a volume of 22.4 dm^3 at s.t.p**. The number **6.02 × 10^{23} is** known **as the Avogadro's number.**

Avogadro's law implies that volume ratios of reacting gases can be used interchangeably with mole ratios when gases react. This means that reacting volumes of gases can be used to determine the equation of reacting gases provided the volumes are measured under the same conditions of temperature and pressure.

Worked Examples

1. **In an experiment, 20 cm^3 of sulphur(II) oxide are found to react completely with 10 cm^3 of oxygen to produce 30 cm^3 of sulphur(IV) oxide. Determine the equation for the reaction.**

Sulphur(IV) oxide +	Oxygen →	Sulphur(VI) oxide
gas	gas →	gas
$SO_2(g)$	$+ O_2(g)$ →	$SO_3(g)$
$20\ cm^3$	$10\ cm^3$	$20\ cm^3$
2 vol	2 vol	2 vol
2 moles	1 mole	2 moles

So the balanced equation is:

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

2. **When 100 cm^3 of a gaseous hydrocarbon(C_xH_y) burn in 400 cm^3 of oxygen, 100 cm^3 of oxygen is unused, 200 cm^3 of carbon(IV) oxide and 200 cm^3 of steam are formed. Deduce the equation for the reaction and the formula of the hydrocarbon.**

Solution

$C_xH_y(g)$ +	$O_2(g)$ →	$CO_2(g)$ +	$H_2O(g)$
$100\ cm^3$	$300\ cm^3$	$200\ cm^3$	$200\ cm^3$
1 vol	3 vol	2 vol	2 vol
1 mol	3 mol	2 mol	2 mol

The formula deduced from reacting volumes is worked out as follows:

$$C_xH_y\,(g) + O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

Writing a balanced equation from the above:

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

The formula of the hydrocarbon is derived on the basis of the following facts:

(i) In a balanced chemical equation, the number of each kind of atoms on the reactant side must be equal to that in the product side.

(ii) Hydrocarbons are compounds consisting of carbon and hydrogen only.

Review Exercise 2

Revision Questions

1. 2006 Q 8 (P1)

When 94.5g of hydrated barium hydroxide, $Ba(OH)_2.nH_2O$ were heated to constant mass, 51.3g of anhydrous barium hydroxide were obtained. Determine the empirical formula of the hydrated barium hydroxide. (3 marks)

2. 2007 Q 20

An alkanol has the following composition by mass: hydrogen 13.5%,
oxygen 21.6% and carbon 64.9%.
Determine the empirical formula of the alcohol (C=12.0; H=1.0) =16.0).

(2 marks)

3. 2007 Q 22

6.84g of aluminium sulphate were dissolve in 150cm^3 of water. Calculate the molar concentration of the sulphate ions in the solution.

(Relative formula mass of aluminium sulphate is 342) (3 marks)

4. 2008 Q 2

When a hydrated sample of calcium sulphate $CaSO_4.\mathbf{x}H_2O$ was heated until all the water was lost, the following data recorded;

Mass of crucible = 30.296 g
Mass of crucible +hydrated salt = 33.111 g
Mass of crucible + anhydrous salt = 32.781 g

Determine the empirical formula of the hydrated salt.

(Relative formula mass of $CaSO_4$ =136, H_2O =18). (3 marks)

5. 2008 Q 5

Phosphoric acid is manufactured from calcium phosphate according to the following equation.

$$Ca_3(PO_4)_2(s) + 3H_2SO_4(l) \longrightarrow 2H_3PO_4(aq) + 3CaSO_4(s)$$

Calculate the mass (in kg) of phosphoric acid that would be obtained if 155 kg of calcium phosphate reacted completely with the acid

(Ca=40, P=31, S=32, O=16, H=1) (2 marks)

6. 2008 Q 27

In an experiment to determine the percentage of magnesium hydroxide in an anti-acid, a solution containing 0.50 g of the anti-acid was neutralized by 23.0 cm^3 of 0.010M hydrochloric acid. (Relative formula mass of magnesium hydroxide =58)

Determine the:

(a) Mass of magnesium hydroxide in the anti-acid; (2 marks)

(b) Percentage of magnesium hydroxide in the anti-acid. (1 mark)

7. 2009 Q 13 P1, 2016 Q7 P1

When 8.53g of sodium nitrate were heated in an open tube, the mass of oxygen gas produced was 0. 83g.Given the equation of the reaction is:

$$2NaNO_3(s) \xrightarrow{heat} 2NaNO_2(s) + O_2(g)$$

Calculate the percentage of sodium nitrate that was converted to sodium nitrite.

(Na= 23.0, N = 14.0, O = 16.0) (3 marks)

8. 2010 Q 6 P1

Aluminium oxide reacts with both acids and bases.

(a) Write an equation for the reaction between aluminium oxide and hydrochloric acid. (1 mark)

(b) Using the equation in (a) above, calculate the number of moles of hydrochloric acid that would react completely with 153.0g of aluminium oxide. (Al = 27.0, O= 16.0) (2 marks)

9. 2010 Q 8 P1

The pressure of nitrogen gas contained in a 1 dm^3 cylinder at -196 °C was 10^7 Pascals. Calculate the:

(a) Volume of the gas at 25 °C and 10^5 Pascals. (1½ marks)

(b) Mass of nitrogen gas. (Molar volume of gas is 24 dm^3, N = 14.0) (1 ½ marks)

10. 2010 Q 17 P1

Analysis of a compound showed that it had the following composition:
69.42% carbon, 4.13% hydrogen and the rest oxygen.

(a) Determine the empirical formula of the compound. (C = 12.0, H = 1.0, O = 16.0) (2 marks)

(b) If the mass of one mole of the compound is 242, determine its molecular formula. (1 mark)

11. 2010 Q 3 P2, 2016 Q1 P2

Use the information in the table below to answer the questions that follow. The letters do not represent the actual symbols of the elements.

Element	**Atomic number**	**Melting point (°C)**
R	11	97.8
S	12	650.0
T	15	44.0
U	17	-102
V	18	-189
W	19	64.0

(a) Give the reasons why the melting point of:
- (i) **S** is higher than that of **R**; (1 mark)
- (ii) **V** is lower than that of **U**. (2 marks)

(b) How does the reactivity of **W** with chlorine compare with that of **R** with chlorine? Explain. (2 marks)

(c) Write an equation for the reaction between T and excess oxygen. (1 mark)

(d) When 1.15g of **R** were reacted with water, 600cm^3 of gas was produced. Determine the relative atomic mass of R. (Molar gas volume = 24000 cm^3) (3 marks)

(e) Give **one** use of element **V**. (1 mark)

12. 2011 Q 7 P1

When lead (II) nitrate is heated, one of the products is a brown gas.

(a) Write the equation of the reaction that occurs. (1 mark)

(b) If 0.290 dm^3 of the brow n gas was produced, calculate the mass of the lead (II) nitrate that was heated. (R.F.M of lead (II) nitrate = 331; Molar gas volume = 24 dm^3).

(2 marks)

13. 2011 Q 25 P1

(a) State the Gay Lussac's Law. (1 mark)

(b) 10 cm^3 of a gaseous hydrocarbon, C_2H_x required 30cm^3 of oxygen for complete combustion. If steam and 20 cm^3 of carbon (IV) oxide were produced, what is the value of **x**?

(2 marks)

14. 2011 Q 26

The data given below was recorded when metal **M** was completely burnt in air. **M** is not the actual symbol of the metal. (R.A.M; M=56, O=16)

Mass of empty crucible and lid =10.240g
Mass of crucible, lid and metal M =10.352g
Mass of crucible, lid and metal oxide = 10.400g

(a) Determine the mass of:

(i) Metal M (½ mark)

(ii) Oxygen (½ mark)

(b) Determine the empirical formula of the metal oxide. (2 marks)

15. 2012 Q8 P1

10cm^3 of concentrated sulphuric (VI) acid was diluted to 100 cm^3. 10 cm^3 of the resulting solution was neutralised by 36 cm^3 of 0.1M sodium hydroxide solution. Determine the mass of sulphuric (VI) acid that was in the concentrated acid. (S = 32.0; H= 1.0; O = 16.0).

(3 marks)

16. 2012 Q11 P1

The empirical formula of **A** is CH_2Br. Given that 0.470g of **A** occupies a volume of 56 cm^3 at 546K and 1 atmosphere pressure, determine its molecular formula.

(H = 1.0, C= 12.0, Br = 80.0, molar gas volume at STP = 22.4 dm^3) (3 marks)

17. 2012 Q23 P1

Describe how the percentage by mass of copper in copper carbonate can be determined.

(3

marks)

18. 2013 Q23 P1

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

(a) What volume of oxygen was used during the reaction? (1 mark)
(b) Determine the molecular formula of the hydrocarbon (2 marks)

19. 2013 Q24 P1

A solution was made by dissolving 8.2g of calcium nitrate to give 2 litres of solution. (Ca= 40.0; N=14.0; O= 16.0)
Determine the concentration of nitrate ions in moles per litre. (3 marks)

20. 2013 Q1 P2

The grid given below represents part of the periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbol of the element.

M				N	P	T		
R								

(a) Select a letter which represents an element that losses electrons most readily. Give a reason for your answer. (2 marks)

(b) Explain why the atomic radius of P is found to be smaller than that of N (2 marks)

(c) Element M reacts with water at room temperature to produce 0.2 dm^3 of gas. Determine the mass of M which was reacted with water.
(Molar gas volume at room temperature is 24 dm^3; Relative atomic mass of M=7) (3 marks)

21. 2014 Q13 P1

100cm^3 of 0.05 M sulphuric (VI) acid were placed in a flask and a small quantity of anhydrous sodium carbonate added. The mixture was boiled to expel all the carbon (IV) oxide. 25cm^3 of the resulting solution required 18 cm^3 of 0.1 M sodium hydroxide solution to neutralize it. Calculate the mass of sodium carbonate added.
(Na = 23.0; O=16.0; C=12.0) (3 marks)

22. 2014 Q25 P1

An organic compound had the following composition 37.21% carbon, 7.75% hydrogen and the rest chlorine. Determine the molecular formula of the compound, given that the molecular mass of the compound is 65.
(C=12.0; H=1.0; Cl=35.5) (3 marks)

23. 2015 Q5 P1

Calculate the mass of Zinc oxide that will just neutralize dilute nitric (V) acid containing 12.6 g of nitric (V) acid in water. (Zn = 65.0; O =16.0, H = 1.0, N = 14.0). (3 marks)

24. 2015 Q26 P1

A hydrocarbon contains 14.5% of hydrogen. If the molar mass of the hydrocarbon is 56, determine the molecular formula of the hydrocarbon.
(C = 12.0; H = 1.0) (3 marks)

25. 2015 Q1b P2

Zinc oxide can be obtained by heating zinc nitrate. A student heated 5.76 g of zinc nitrate.

(a) Write an equation for the reaction that occurred. (1 mark)

(b) Calculate the total volume of gases produced. (Molar gas volume is 24 dm^3; Zn = 65.4; O = 16.0; N = 14.0). (4 marks)

(c) Identify the element that is reduced when zinc nitrate is heated. Give a reason. (2 marks)

26. 2017 P1 Q4.

The empirical formula of lead(II) oxide was determined by passing excess dry hydrogen gas over 6.69g of heated lead(II) oxide.

(a) What was the purpose of using excess dry hydrogen gas? (2 marks)

(b) The mass of lead was found to be 6.21g. Determine the empirical formula of the oxide. (Pb = 207.0 0 = 16.0) (2 marks)

27. 2017 P1 Q10.

20 cm^3 of ethanoic acid was diluted to 400 cm^3 of solution. Calculate the concentration of the solution in moles per litre. (C = 12.0; H = 1.0; 0 =16.0) (Density of ethanoic acid = 1.05 g/cm^3)
(3 marks)

28. 2017 P2 Q4.

W is a colourless aqueous solution with the following properties:

I. It turns blue litmus paper red.
II. On addition of cleaned magnesium ribbon, it gives off a gas that burns with a pop sound.
III. On addition of powered sodium carbonate, it gives off a gas which forms a precipitate with calcium hydroxide solution.

IV. When warmed with copper(II) oxide powder, a blue solution is obtained but no gas is given off.

V. On addition of aqueous barium chloride, a white precipitate is obtained.

(a) (i) State what properties (I) and (III) indicate about the nature of W. (1 mark)
(ii) Give the identity of W. (1 mark)
(iii) Name the colourless solution formed in (II) and (III). (2 marks)
(iv) Write an ionic equation for the reaction indicated in (V). (1 mark)

(b) Element V conducts electricity and melts at 933 K. When chlorine gas is passed over heated V, it forms a vapour that solidifies on cooling. The solid chloride dissolves in water to form an acidic solution. The chloride vapour has a relative molecular mass of 267 and contains 19.75% of V. At a higher temperature, it dissociates to a compound of relative molecular mass 133.5. When aqueous sodium hydroxide is added to the aqueous solution of the chloride, a white precipitate is formed which dissolves in excess alkali. (V = 27.0; Cl = 35.5)

(i) Determine the:
I. empirical formula; (2 marks)
II. molecular formula. (2 marks)

(ii) Draw the structure of the chloride vapour and label the bonds. (1 mark)

(iii) Write an equation for the reaction that form a white precipitate with sodium hydroxide. (1 mark)

29. 2017 P2 Q5

(a) When 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure, 50 cm^3 of hydrogen gas was collected. (Mg = 24.0; Molar gas volume = 24.0 dm^3)

(i) Draw a diagram of the apparatus used to carry out the experiment described above. (3 marks)

(ii) Write the equation for the reaction. (1 mark)

(iii) Calculate the volume of hydrogen gas produced. (2 marks)

(iv) Calculate the volume of 0.1 M hydrochloric acid required to react with 0.048 g of magnesium. (3 marks)

30. 2018 P1 Q 7.

30.0 cm^3 of aqueous sodium hydroxide containing 8.0 g per litre of sodium hydroxide were completely neutralised by 0.294 g of a dibasic acid. Determine the relative formula mass of the dibasic acid. (Na = 23.0; O = 16.0; H 1.0) (3 marks)

31. 2018 P1 Q28.

Distinguish between empirical and molecular formula of a compound. (1 mark)

32. 2019 P1 Q7.

A solution contains 40.3g of substance XOH per litre. 250.0 cm^3 of this solution required 30.0cm^3 of 0.3M sulphuric(VI)acid for complete neutralisation.

(a) Calculate the number of moles of XOH that reacted. (½ mark)
(b) Determine the relative atomic mass of X. (1½ mark)

33. 2019 P1 Q 13.

5 g of calcium carbonate was strongly heated to a constant mass.
Calculate the mass of the solid residue formed. (Ca = 40.0; C = 12.0; 0 = 16.0). (2 marks)

34. 2019 P1 Q 29.

Name an appropriate apparatus that is used to prepare standard solutions in the laboratory. (1 mark)

TOPIC THREE

ORGANIC CHEMISTRY I

(HYDROCARBONS)

By the end of this chapter, the learner should be able to:

(a) Define organic chemistry.
(b) Define hydrocarbons and classify them according to their structure and bonding.
(c) Write molecular and structural formulae of alkanes, alkenes and alkynes.
(d) Name and draw the structures of alkanes, alkenes and alkynes.
(e) Define an isomer, name and draw the structures of the isomers of hydrocarbons with up to five carbon atoms in their molecules.
(f) Describe the general methods of preparing different hydrocarbons.
(g) Explain the gradual change in properties among the members of a homologous series.
(h) State some uses of hydrocarbons.

Organic chemistry is the study of carbon compounds excluding the oxides of carbon and carbonates.

Carbon forms a very large number of compounds. This is due to the fact that carbon:

(a) Uses all its valence electrons to form four covalent bonds with carbon atoms and with atoms of other elements.
(b) Can form single, double and triple bonds with other carbon atoms.
(c) Atoms can join each other to form long chains.

HYDROCARBONS.

Hydrocarbons are compounds which are made up of carbon and hydrogen atoms only.

They are classified into three groups depending on the type of bond that exists between individual carbon atoms in a molecule. The groups are alkanes, alkenes and alkynes.

1. ALKANES

Alkanes are a group of hydrocarbons whose carbon atoms are linked by single covalent bonds.

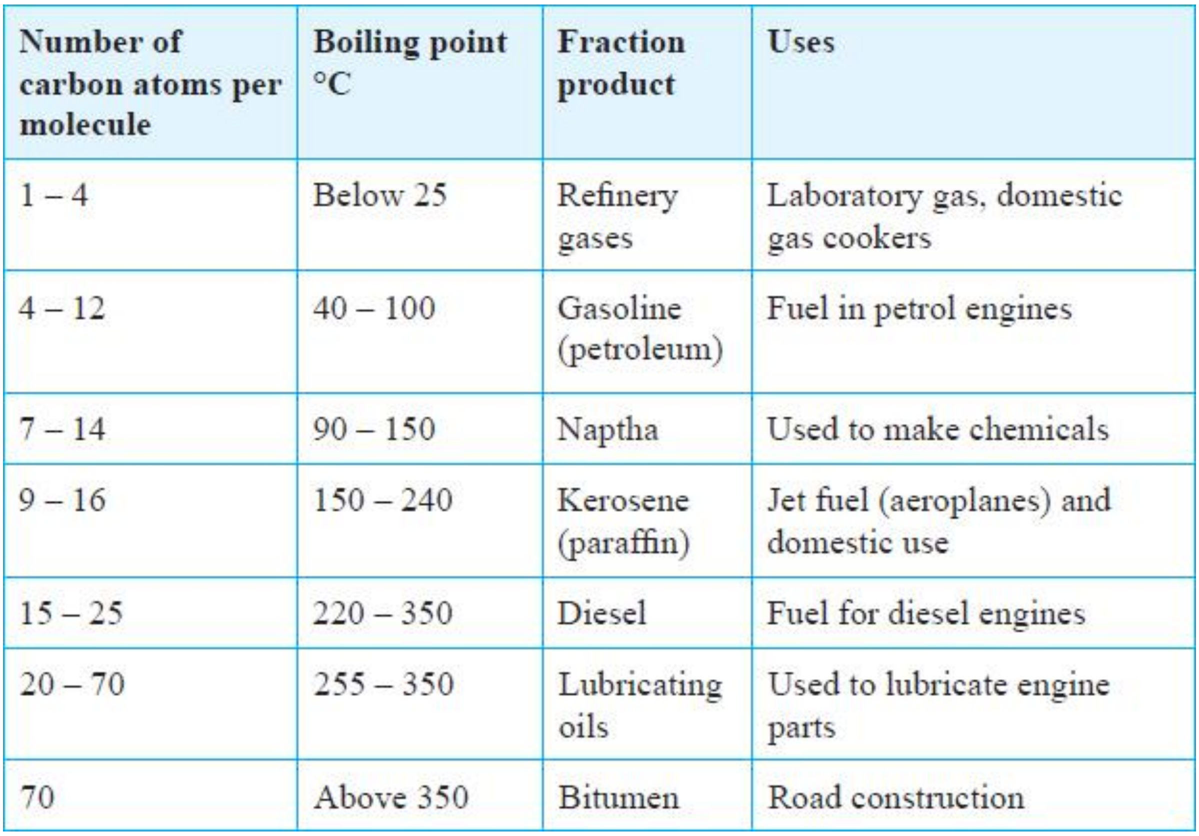
Hydrocarbons in which only single covalent bonds are present in the molecules are said to be **saturated**.

Sources of Alkanes

Alkanes occur naturally as natural gas, biogas and in crude oil.

Crude oil is the major natural source of alkanes. It contains a range of alkanes which can be separated into various components by fractional distillation. This is due to the fact that the different alkanes have different boiling points. The fractions with short carbon chains have low boiling points. As the boiling point of the fractions increase, the viscosity and the intensity of colour of the fractions increase while flammability decreases.

Crude oil is separated into different components through fractional distillation. Each component separates at different temperature at different parts of the fractionating column. The different fractions have different uses.



Number of carbon atoms per molecule	Boiling point °C	Fraction product	Uses
1 – 4	Below 25	Refinery gases	Laboratory gas, domestic gas cookers
4 – 12	40 – 100	Gasoline (petroleum)	Fuel in petrol engines
7 – 14	90 – 150	Naptha	Used to make chemicals
9 – 16	150 – 240	Kerosene (paraffin)	Jet fuel (aeroplanes) and domestic use
15 – 25	220 – 350	Diesel	Fuel for diesel engines
20 – 70	255 – 350	Lubricating oils	Used to lubricate engine parts
70	Above 350	Bitumen	Road construction

Cracking of Alkanes

Long chain alkanes are broken up through a process called **cracking of alkanes.**

During the cracking process short chain alkanes, alkenes and hydrogen are produced.

The general equation for cracking is:

Long chain alkane $\xrightarrow{400-700^\circ C}$ Smaller chain alkane + Alkenes + Hydrogen For example, when propane is irradiated with high energy radiation, the following reaction occurs.

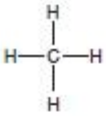
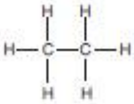
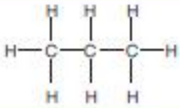
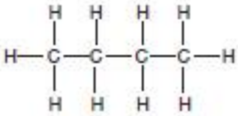
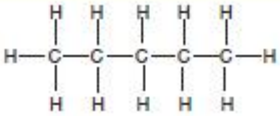
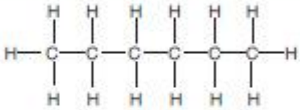
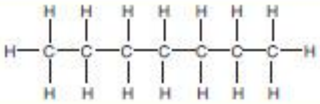
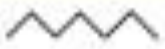
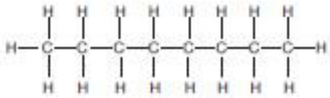
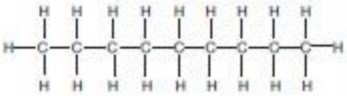
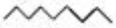
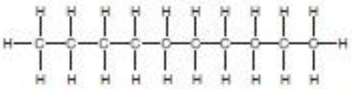
$$2CH_3CH_2CH_3 \xrightarrow{600-700^\circ C} CH_4 + CH_2=CH_2 + CH_3CH=CH_2 + H_2$$

There are two ways of cracking of alkanes; thermal cracking **and** catalytic cracking.

- Thermal cracking takes place at a very high temperature and only involves heating the long chain alkanes.
- Catalytic cracking takes place at a relatively low temperature and involves heating the long chain alkanes in the presence of a catalyst and at low pressure.

Nomenclature

Alkanes have names which end with the suffix, **– ane; and a** prefix numeral used to indicate the number of carbon atoms in the chain. The first four members have their prefixes as **meth-, eth-, prop -but-** with the next members being , **pent-** (5), **hex-** (6), **and hept-** (7). The simplest alkane is methane where n = 1 and its formula is CH_4.

Number of carbon atoms	Name	Molecular formula	Open Structural formula	Skeletal formula	Condensed structural formula
1.	Methane	CH_4			CH_4
2.	Ethane	$C_2 H_6$			$CH_3 CH_3$
3.	Propane	$C_3 H_8$			$CH_3 CH_2 CH_3$
4.	Butane	$C_4 H_{10}$			$CH_3 (CH_2)_2 CH_3$
5.	Pentane	$C_5 H_{12}$			$CH_3 (CH_2)_3 CH_3$
6.	Hexane	$C_6 H_{14}$			$CH_3 (CH_3)_4 CH_3$
7.	Heptane	$C_7 H_{16}$			$CH_3 (CH_2)_5 CH_3$
8.	Octane	$C_8 H_{18}$			$CH_3 (CH_2)_6 CH_3$
9.	Nonane	$C_9 H_{20}$			$CH_3 (CH_2)_7 CH_3$
10.	Decane	$C_{10} H_{22}$			$CH_3 (CH_2)_8 CH_3$

In the skeletal formula, only bonds between carbon atoms are shown. The bonds are drawn in a zig-zag manner. A straight section of a zig-zag line represents a covalent bond between two carbon atoms in the structure.

A **molecular formula** shows the elements present in a compound and the proportions of their atoms in the compounds.

A **structural formula** shows how the atoms of the different elements are arranged in the compound.

The consecutive members of the alkane homologous series differ by a – CH_2 unit, and conform to the **general formula C_nH_{2n+2} where n is 1, 2, 3…**

A homologous series is a *group of compounds with similar chemical properties, chemical formulae, and they exhibit a steady gradual change in physical properties.*

The homologous series of alkanes has the following characteristics:

(i) All members conform to a **general molecular formula of C_nH_{2n+2}.**
(ii) All members show similar chemical properties.
(iii) The physical properties of the members change gradually along the series.
(iv) The general methods of preparation can be applied to any member of the series.

Isomerism in Alkanes

Isomers *are compounds that have the same molecular formula but different structural formulae.*

The isomers of the same alkane differ in their physical properties such as boiling points, melting points and density but their chemical reactions are similar.

If any hydrogen atom attached to an alkane chain is removed, an alkyl group is formed. The suffix **-ane** in the alkane is replaced by **-yl**, for example:

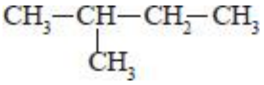
Alkane	***Alkyl group***
Methane, CH_4	*Methyl, CH_3–*
Ethane, CH_3CH_3	*Ethyl, CH_3CH_2–*
Butane, $CH_3CH_2CH_2CH_3$	*Butyl, $CH_3CH_2CH_2CH_2$–*

Methane, ethane, and propane do not have isomers. Isomerism in alkanes starts from butane.

If any hydrogen atom attached to any middle carbon atom is replaced by an alkyl group, a branched alkane is obtained.

Rules for naming of the branched alkanes

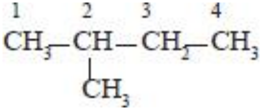
1. Identify the longest continuous carbon chain to determine the name of the parent alkane. For example, in the structure,



$$CH_3-\underset{\substack{|\\ CH_3}}{CH}-CH_2-CH_3$$

The longest chain has four carbon atoms hence the parent name is butane.

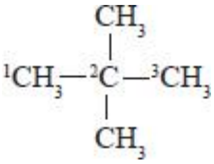
2. Number the longest chain from the end of the chain that is near the branching. For example, in the structure above, numbering should start from the left, thus:



$$\overset{1}{CH_3}-\underset{\displaystyle CH_3}{\overset{2}{CH}}-\overset{3}{CH_2}-\overset{4}{CH_3}$$

The branching is therefore on carbon 2 of the longest chain and the substituent group is a methyl.

3. In case there are two or more similar substituent groups in the chain, they are indicated by the prefixes; ***di-, tri-, tetra-***, … Comas are used to separate the numbers for example, in the following:

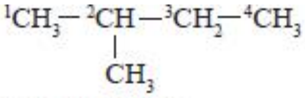


$$^{1}CH_3-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{^{2}C}}-{}^{3}CH_3$$

There are two methyl groups attached to carbon number two of the longest chain hence, 2, 2 - dimethyl. When naming the Isomers of an alkane, the position of the substituent group is written first followed by the name of the substituent group and finally the name of the parent alkane.

Examples

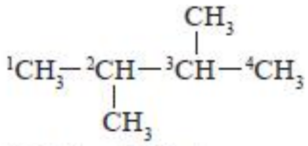
1.



$$^{1}CH_3-\underset{\displaystyle CH_3}{^{2}CH}-{}^{3}CH_2-{}^{4}CH_3$$

2-methylbutane

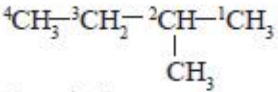
3.



$$^{1}CH_3-\underset{\displaystyle CH_3}{^{2}CH}-\overset{\displaystyle CH_3}{^{3}CH}-{}^{4}CH_3$$

2, 3-dimethylbutane

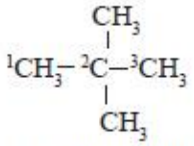
2.



$$^{4}CH_3-{}^{3}CH_2-\underset{\displaystyle CH_3}{^{2}CH}-{}^{1}CH_3$$

2-methylbutane

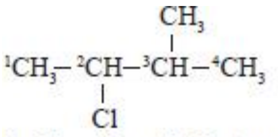
4.



$$^{1}CH_3-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{^{2}C}}-{}^{3}CH_3$$

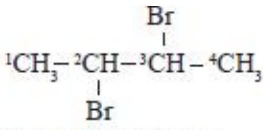
2, 2-dimethylpropane

4. When the substituent groups are a halogen and an alkyl group, the halogen is always placed before the alkyl group. For example;



$$^{1}CH_3-{}^{2}\underset{Cl}{CH}-{}^{3}\overset{CH_3}{CH}-{}^{4}CH_3$$

2-chloro-3-methylbutane



$$^{1}CH_3-{}^{2}\underset{Br}{CH}-{}^{3}\overset{Br}{CH}-{}^{4}CH_3$$

2, 3-dibromobutane

Preparation of Alkanes

In the laboratory, alkanes are prepared by the reaction between sodalime (mixture of sodium hydroxide and calcium oxide) and an appropriate alkanoate.

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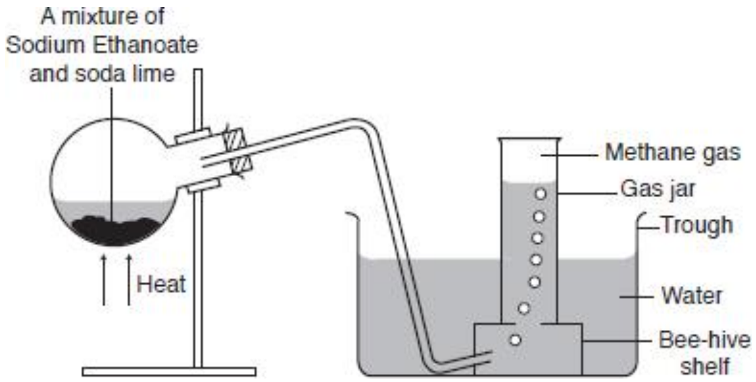


When a mixture of sodium ethanoate and soda lime (sodium hydroxide and calcium oxide) is heated, methane gas is obtained.

$$CH_3\,COONa(s) + NaOH(aq) \xrightarrow{heat} CH_4(g) + Na_2CO_3(s)$$

(sodium ethanoate) (from soda lime) (methane) (sodium carbonate)

Similarly, when a mixture of sodium propanoate, and soda lime is heated ethane gas is obtained.

$$CH_3CH_2COONa(s) + NaOH(aq) \xrightarrow{heat} CH_3CH_3(s) + Na_2CO_3(s)$$

(sodium propanoate) (from sodalime) (ethane) (sodium carbonate)

Physical Properties

- Methane is a non-poisonous, colourless gas.
- It is slightly soluble in water and hence can be collected over water. However, it is quite soluble in organic solvents such as ethanol and tetrachloromethane. It is less dense than air.

The general physical properties of the first ten alkanes

	Name of Alkane	Formula	State at room Temperature	Melting Point (K)	Boiling Point	Density g/cm^3
1	Methane	CH_4	Gas	90	112	0.424
2	Ethane	C_2H_6	Gas	91	184	0.546
3	Propane	C_3H_8	Gas	85	231	0.501
4	Butane	C_4H_{10}	Gas	138	273	0.579
5	Pentane	C_5H_{12}	Liquid	143	309	0.626
6	Hexane	C_6H_{14}	Liquid	178	342	0.657
7	Heptane	C_7H_{16}	Liquid	182	372	0.684
8	Octane	C_8H_{18}	Liquid	216	399	0.703
9	Nonane	C_9H_{20}	Liquid	219	424	0.708
10	Decane	$C_{10}H_{22}$	Liquid	243	447	0.730

- The melting and boiling points of alkanes **increase with increase in the number of carbon atoms**. An increase in number of carbon atoms results in an increase in the strength of intermolecular (van der Waals forces).
- The first four straight chain alkanes are gases, the next six are liquids, and the rest are solids. Density increases with an increase in molecular mass.
- Generally, the solubility of alkanes decreases as the molecular size increases.

Chemical Properties

Combustion

Methane burns in excess air with a pale blue flame to form carbon(IV) oxide and water.

$$CH_4g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

In a limited supply of air the flame is luminous due to incomplete combustion.

$$CH_4(g) + O_2(g) \rightarrow CO(g) + 2H_2O(l)$$

Other alkanes behave in a similar manner.

Substitution reactions

A substitution reaction is a reaction in which one atom or group of atoms in a molecule is replaced by another.

Alkanes undergo substitution reactions with halogens in the presence of sunlight because the halogen molecules are supplied with energy from light necessary to split them into atoms. These free halogen atoms are very reactive hence they replace the hydrogen atoms in the methane molecule.

All the alkanes undergo substitution with chlorine and bromine in the presence of sunlight.

When a substituent, like a halogen, is present in the alkane, the resulting name **must** contain the halogen. The name of the **halogen becomes the prefix, and the parent alkane, the suffix.** The name of the halogen is therefore written as:

Iodo- to represent iodine.

***Bromo-* to represent bromine.**

***Chloro-* to represent chlorine.**

***Fluoro-* to represent fluorine.**

Chlorine reacts with methane in the presence of ultra violet light. Chlorine atoms repeatedly substitute the hydrogen atoms in the methane molecule until all the hydrogen atoms have been replaced as shown below.

(i) $CH_4(g) + Cl_2(g) \xrightarrow{\text{ultra violet light}} CH_3\,Cl(g) + HCl(g)$

Chloromethane

(ii) $CH_3\,Cl(g) + Cl_2(g) \xrightarrow{\text{ultra violet light}} CH_2Cl_2(g) + HCl(g)$

Dichloromethane

(iii) $CH_2\,Cl_2(g) + Cl_2(g) \xrightarrow{\text{ultra violet light}} CHCl_3(g) + HCl(g)$

Trichloromethane

(iv) $CHCl_3(g) + Cl_2(g) \xrightarrow{\text{ultra violet light}} CCl_4(g) + HCl(g)$

Tetrachloromethane

If bromine is used, it likewise substitutes hydrogen but the reaction is slow. The mixture of bromine and methane gets decolourised when left in sunlight. The equation for the reaction of methane and bromine is:

$$CH_4(g) + Br_2(g) \rightarrow CH_3\,Br(g) + HBr.$$

Uses of Alkanes

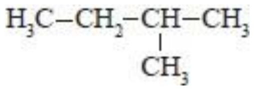
1. Gaseous alkanes such as methane, propane, butane are used in homes as fuels.
2. They are used in the manufacture of carbon black which is a component of printers' ink, and paint.
3. They are used in the manufacture of methanol, methanal, and chloromethane which are useful industrial chemicals.
4. As a source of hydrogen during the cracking process.

Review Questions: Alkanes

1. Name each of the following compounds:

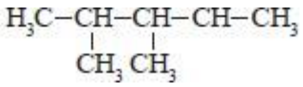
(a) $H_3C—CH_2—CH_2—CH_3$

(b)



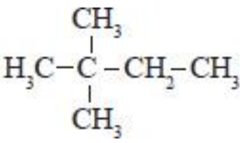
$$H_3C-CH_2-\underset{\begin{array}{c}|\\CH_3\end{array}}{CH}-CH_3$$

(c)



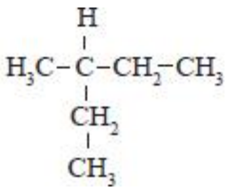
$$H_3C-\underset{\begin{array}{c}|\\CH_3\end{array}}{CH}-\underset{\begin{array}{c}|\\CH_3\end{array}}{CH}-CH-CH_3$$

(d)



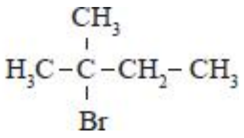
$$H_3C-\overset{\begin{array}{c}CH_3\\|\end{array}}{\underset{\begin{array}{c}|\\CH_3\end{array}}{C}}-CH_2-CH_3$$

(e)



$$H_3C-\overset{\begin{array}{c}H\\|\end{array}}{\underset{\begin{array}{c}|\\CH_2\\|\\CH_3\end{array}}{C}}-CH_2-CH_3$$

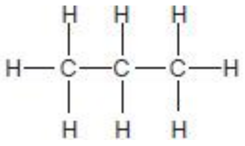
(f)



$$H_3C-\overset{\begin{array}{c}CH_3\\|\end{array}}{\underset{\begin{array}{c}|\\Br\end{array}}{C}}-CH_2-CH_3$$

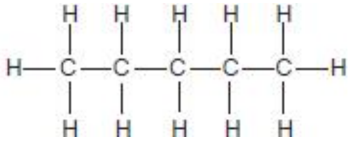
2. Give the names and molecular formula of the following.

(a)



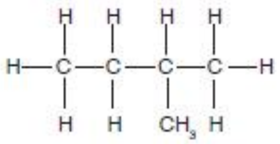
$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

(b)



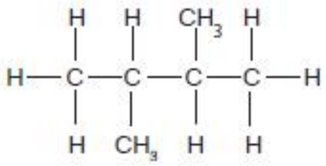
$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array}$$

(c)



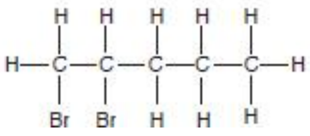
$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & CH_3 & & H & & \end{array}$$

(d)



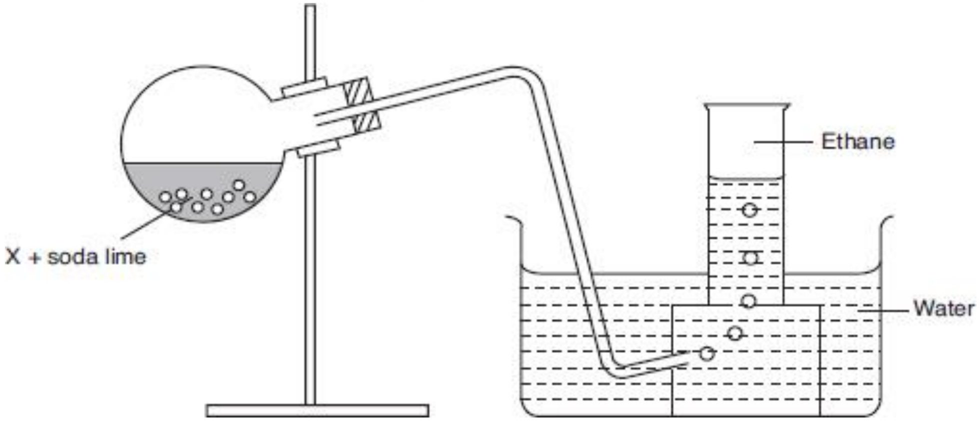
$$\begin{array}{ccccccccccc} & & H & & H & & CH_3 & & H & & \\ & & | & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & | & & \\ & & H & & CH_3 & & H & & H & & \end{array}$$

(e)

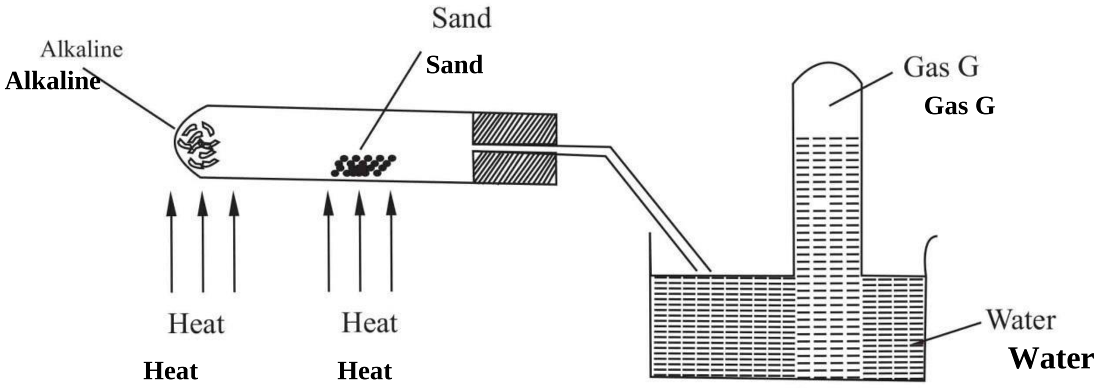


$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & | & & | & & \\ & & Br & & Br & & H & & H & & H & & \end{array}$$

3. Draw the structural formula for each of the following compounds:
   - (a) 2-methylheptane.
   - (b) 3-ethylhexane.
   - (c) 2, 2, 4-trimethylpentane.
   - (d) 2, 3-dichlorobutane.
4. Draw and name the structures of all the isomers of:
   - (a) Butane.
   - (b) Pentane.
   - (c) Hexane.
5. Distinguish between thermal and catalytic cracking of alkanes.
6. The following set up was used to prepare ethane in the laboratory.



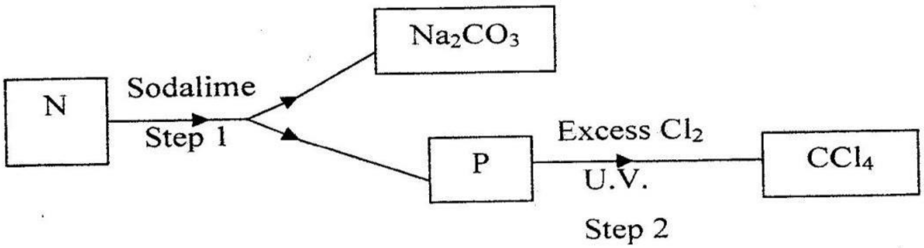
   - (a) Identify a condition missing in the set up.
   - (b) Name substance X and write its chemical formula.
   - (c) Name the product produced alongside ethane in the reaction.
7. Draw and name all the possible isomers of an alkane with the molecular formula C_6H_{14}.
8. State three uses of alkanes.
9. The figure below represents the set up that was used to crack an alkane.



   i What was the purpose of the sand?

ii After some time, a colourless gas G was collected in the test-tube. Describe a chemical test and the observation that would be made in order to identify the class of compounds to which gas G belongs

10. Give the name and draw the structural formula of the compound formed when one mole of ethane reacts with one mole of chlorine gas.
11. Study the flow chart below and answer the questions that follow



(a) Identify N and P

(b) What name is given to the type of halogenations/ chlorination reaction in step 2?

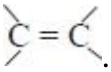
12. (a) What is meant by isomerism? (1 mark)

(b) Draw and name two isomers of butane. (2 marks)

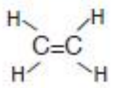
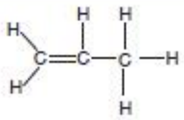
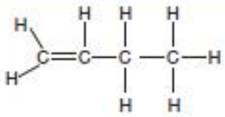
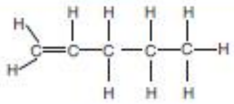
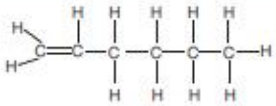
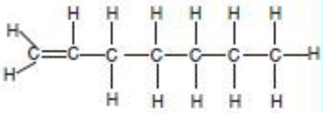
13. What name is given to a compound that contains carbon and hydrogen only? (½ mark)

14. Hexane is a compound containing carbon and hydrogen.
    a. What method is used to obtain hexane from crude oil?
    b. State one use of hexane

2. ALKENES

Alkenes are hydrocarbons which contain at least **one double bond** between two carbon atoms in a molecule . The first member of the series is ethene C_2H_4. because at least two carbon atoms are necessary for a double bond to be formed.

The First members of the alkene series

Number of carbon atoms (n)	Name	Molecular Formula	Open structural formula	Skeletal Structure	Condensed structural formula
2	Ethene	C_2H_4			CH_2CH_2
3	Propene	C_3H_6			CH_2CHCH_3
4	Butene	C_4H_8			CH_2CHCH_3
5	Pentene	C_5H_{10}			$CH_2CHCH_2CH_2CH_3$
6	Hexene	C_6H_{12}			$CH_2CH(CH_2)_3CH_3$
7	Heptene	C_7H_{14}			$CH_2CH(CH_2)_4CH_3$

Each alkene differs from the next one by a $–CH_2$. group, and conforms to the **general formula C_nH_{2n} where n represents the number of carbon atoms in a molecule (n = 2, 3, 4, 5…).**

The presence of a double bond in alkenes results in unsaturation. The double bond in alkenes is easily broken to accommodate more atoms. The C = C double bond determines the chemical properties of alkenes.

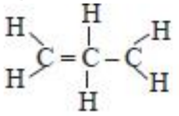
A functional group is an atom or group of atoms which is responsible for the characteristic reactions of a compound.

Nomenclature

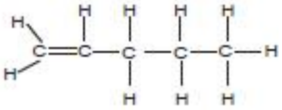
The longest straight chain in alkenes is that which contains the C– C double bond. All alkenes have names ending with – **ene.** To name them, a prefix indicating the number of carbon atoms in the longest straight chain is followed by the suffix – **ene.**

Naming of alkenes is based on the following rules:

1. To determine the parent name of the alkene identify the longest carbon chain containing the double bond.

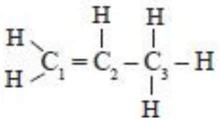


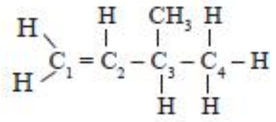
Propene



Butene

2. Number the carbon atoms in the longest carbon chain starting from the end nearer to the double bond. The double bond should be given the **lowest possible number**.

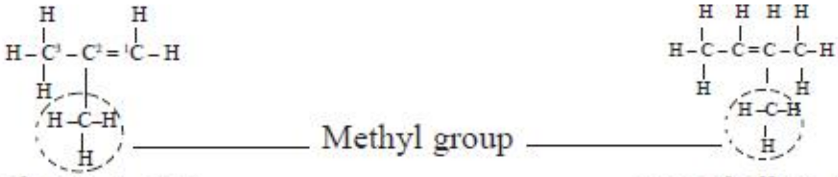




In the two cases, the double bond is between carbon 1 and carbon 2. In naming the lower position is used.

3. Indicate the positions of the substituent groups by showing the position of the carbon atom to which they are attached. For example:

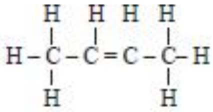
The structure should be:



2–methylprop–1–ene *2–methylbut–2–ene*

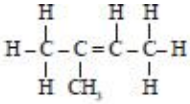
Examples

1.



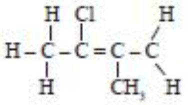
But-2-ene

2.



2-methylbut–2–ene

3.



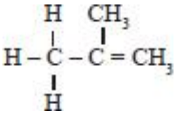
2 – chloro –3–methylbut –2–ene

Isomerism in Alkenes

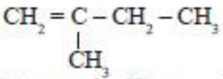
Alkenes show **branching,** and **positional isomerism.**

Branching isomerism occurs when a substituent group is attached to one of the carbon atoms in the longest chain containing the double bond.

For example, the branched isomer of but-1-ene is:

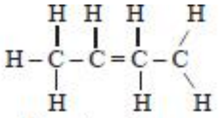


Pentene has two branched isomers:



2–methylbut – 1– ene

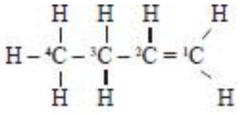
and



Pent – 2– ene

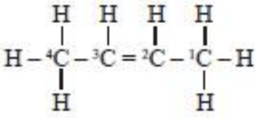
Positional isomerism occurs when **the position of the double bond in an alkene changes.** For example:

(a)



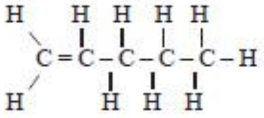
But – 1– ene

and



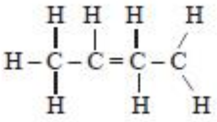
But – 2– ene

(b)



Pent – 1– ene

and



Pent – 2– ene

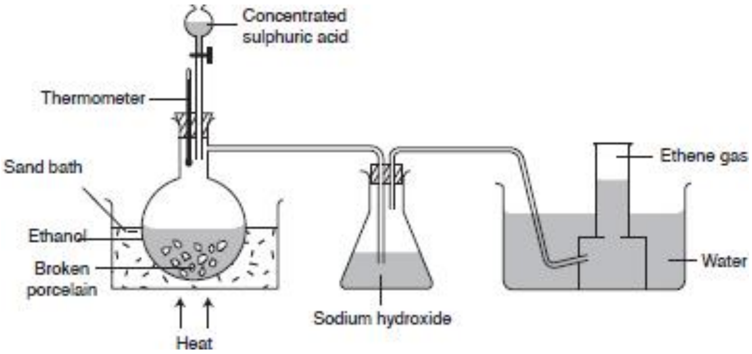
Laboratory preparation of alkenes.

Alkenes can be prepared by **dehydrating their corresponding alkanols** using concentrated sulphuric acid or aluminum oxide.

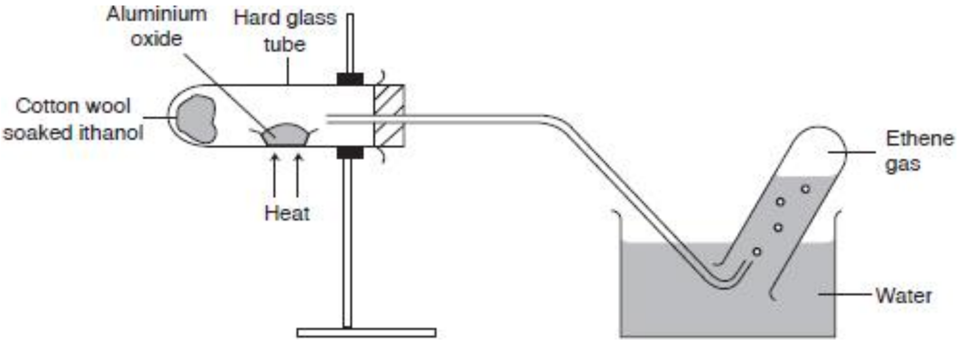
For example, ethene can be prepared by dehydrating ethanol using concentrated sulphuric acid or aluminium oxide. The mixture of ethanol and concentrated sulphuric acid or aluminum oxide must be heated to a temperature of 170°C.

1. Dehydratingethanol with concentrated sulphuric acid.

The ratio of ethanol to concentrated sulphuric acid is 1:2 respectively.



2. Dehydrating ethanol with aluminum oxide



Discussion Questions.

1. Write an equation for the reaction between ethanol and concentrated sulphuric acid.

Equation

$$\underset{(Ethanol)}{C_2H_5OH(l)} \xrightarrow[Heat]{Conc\ H_2SO_4} \underset{(Ethene)}{CH_2CH_2(g)} + H_2O(l)$$

The removal of water molecules from ethanol is a dehydration process. Conc. sulphuric acid is a dehydrating agent.

2. What is the purpose of the broken porcelain in the reaction flask?

Broken porcelain or sand is used to prevent bumping which may result in cracking of the flask.

3. Why was sodium hydroxide solution used in the experiment?

The gas produced is passed through sodium hydroxide to remove sulphur(IV) oxide and carbon(IV) oxide which are formed when concentrated sulphuric acid and ethanol decompose respectively due to heat.

4. What property of ethene gas enables one to use the above method of collecting it?

Ethene is slightly soluble in water hence is collected over water.

5. Why is it preferable to use a sand bath instead of heating the round-bottomed flask directly.

To prevent bumping which may result in cracking of the flask.

6. Explain how aluminium oxide is used to prepare ethene.

The aluminium oxide acts both as a catalyst as well as a dehydrating agent.

$$C_2H_5OH(l) \xrightarrow[Heat]{Al_2O_3} CH_2CH_2(g) + H_2O(l)$$

Physical Properties

- Ethene is a colourless and oduorless gas.
- It is slightly soluble in water hence is collected over water.
- The solubility of alkenes decreases with increase in molecular mass. Ethene is very soluble in organic solvents such as methylbenzene, and tetrachloromethane.
- The melting point, and boiling points of alkenes generally **increase with increase in the number of carbon atoms** due to an **increase in inter-molecular forces**, hence, high boiling and melting points.

Name of Alkene	Molecular formula	M.pt °C	B.pt °C	Density (g/cm^3
Ethene	$CH_2{=}CH_2$	–169	–104	–
Propene	$CH_3CH{=}CH_2$	–185	–47.7	–
But-l-ene	$CH_3CH_2CH{=}CH_2$	–185	–62	
Pent-l-ene	$CH_3(CH_2)_2CH{=}CH_2$	–138	30.0	0.640
Hex-l-ene	$CH_3(CH_2)_3CH{=}CH_2$	–98	63.9	0.674

Chemical Properties

Alkenes are more reactive than alkanes due to the presence of the double bond. The double bond is the reactive site of alkenes.

(i) Combustion

All alkanes burn in air with a yellow sooty flame because they are **unsaturated.** They have a higher carbon to hydrogen ratio than alkanes.

Since alkenes burn with a sooty flame, they are not preferred for use as fuels.

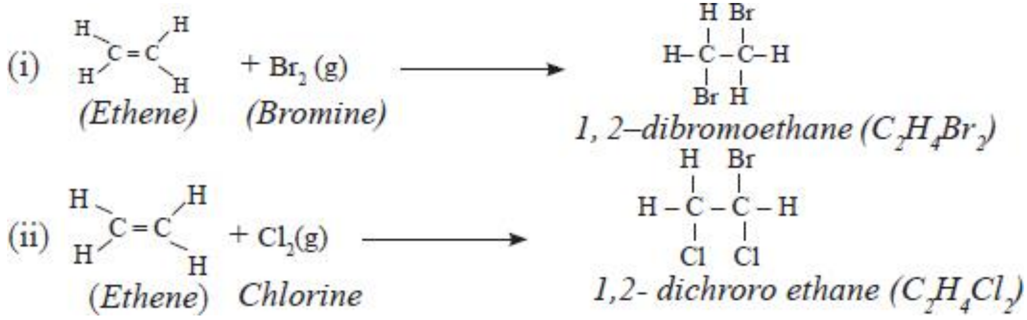
$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

(ii) Addition reactions

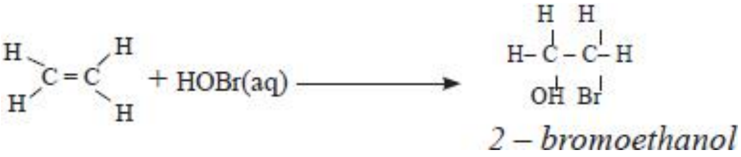
Ethene undergoes addition reactions because of the double bond. **An addition reaction is one in which one molecule adds to another to form a single product.**

(a) Halogenation

This is the addition of halogen atoms across a double bond. Chlorine and bromine are decolourised immediately when ethene is added. Equations for the reactions are:



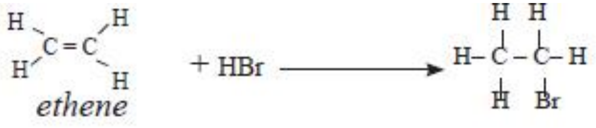
(iii) When bromine is dissolved in water, and reacted with ethene, the following reaction takes place:

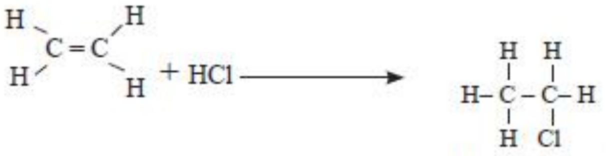


Bromine water is decolourised. This is a test for alkenes.

(b) Hydrohalogenation

Addition of hydrogen halides such as hydrogen bromide, and hydrogen chloride is as shown:





chloroethane (C_2H_5Cl)

(c) Hydrogenation

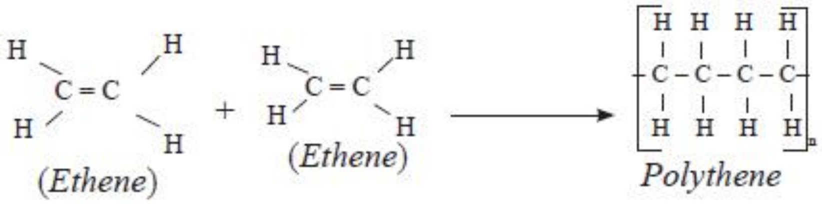
Hydrogen gas reacts with ethene at high temperature in the presence of palladium or nickel catalyst to form ethane.

$$\underset{\text{Ethane}}{CH_2{=}CH_2} + H_2(g) \xrightarrow[150°C-250°C]{\text{Nickel catalyst}} \underset{\text{Ethane}}{CH_3CH_3}$$

When hydrogen gas is passed through liquid vegetable and animal oil heated to a temperature of 180°C, in the presence of a nickel catalyst, solid fat is formed. Therefore this process of hydrogenation is used in the manufacture of margarine to make the oils solid.

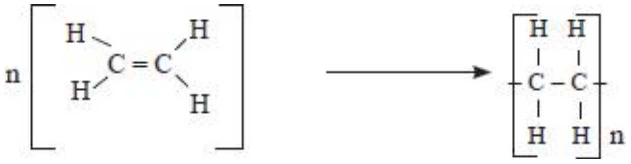
(ii) Self-addition reaction (polymerisation)

Ethene molecules have the ability to react with each other to form a larger molecule which has a higher molecular mass. Each molecule of ethene is known as a **monomer**. When many monomers are joined together, they form a **polymer**. The following equation shows how a polymer is formed:



The process in which several monomers combine to form a polymer is referred to as polymerisation.

When ethene molecules join with each other this way, they form a polymer known as **polyethene.** This polymerisation process can be represented by a general equation;



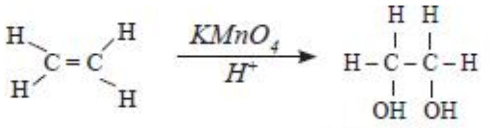
Where n = 2, 3, 4, 5..

(iii) Addition reactions with oxidising agents

(a) Potassium manganate(VII), $KMnO_4$

When ethene is bubbled into acidified potassium manganate(VII) solution, **the colour of the solution turns from purple to colourless.** The manganate(VII) is an oxidising agent and it **adds oxygen at the double bond.** The manganate(VII) ion is reduced to manganese(II) ions and water.

This is also another test for alkenes. The equation is:



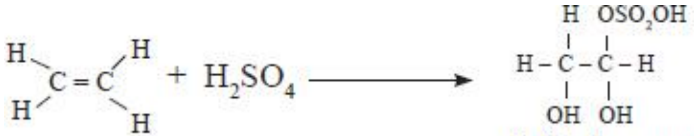
Ethian – 1, 2, – diol

(b) Acidified potassium chromate(VI) ($K_4Cr_2O_7$)

When the potassium dichromate(VI) is used in the reaction the orange colour of the chromate(VI) changes to green. Potassium chromate(VI) acts as an oxidising agent adding oxygen at the double bond. The chromate(VI) ions are reduced to chromium(III) ions(Cr^{3+}).

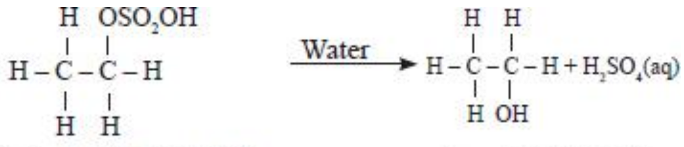
(c) Reaction with concentrated sulphuric acid

Ethene reacts with concentrated sulphuric acid at room temperature to form ethyl hydrogen sulphate.



ethylhydrogensulphate

When ethyl hydrogen sulphate is added to water, and warmed, the product formed is ethanol.



ethylhydrogensulphate $CH_3CH_2SO_4H$ *ethanol* CH_3CH_2OH

Ethyl hydrogen sulphate is hydrolysed to ethanol. This process is referred to as *hydrolysis*.

Hydrolysis is the reaction of a compound with water such that the hydroxyl group of the water remains intact.

Generally, the hydrogen atom is added first to ethene molecule to break the double bond, then the hydroxyl group is added.

Tests for Alkenes

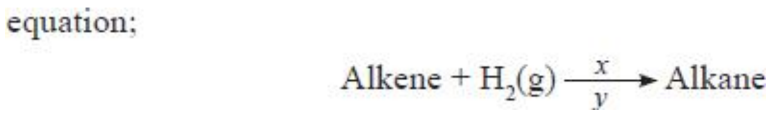
The addition reactions of alkenes with bromine water, acidified potassium manganate (VII) or acidified potassium(VI) dichromate can be used to test for the presence of a double bond.

Uses of Alkenes

1. In the manufacture of plastics.
2. In the manufacture of ethanol through hydrolysis reactions.
3. In the ripening of fruits (ethene).
4. In the manufacture of detergents.
5. In the laboratory preparation of ethan –1,2 –diol (glycol) which is used as in coolant (especially as an engine coolant).

Review Questions: Alkenes

1. Name the following compounds:
   (a) $H_3C - CH= CH - CH_3$
   (b) $H_2C = CH - CH_2 - CH_2 - CH_3$
   (c) $H_2C = CH - CH = CH - CH_3$
   (d) $H_2C = CH - CH = CBr - CH_3$
2. Draw the structural formula of the following compounds:
   (a) 2-methylpent – 2-ene.
   (b) 3-methylbut-1-ene.
   (c) 1-chloro – 2 -methylpentane.
   (d) 1 – bromo – 3 – ethylpentene.
3. Define the following terms:
   (a) monomer.
   (b) polymer.
4. Alkenes undergo hydrogenation to form alkanes as shown by the following equation;



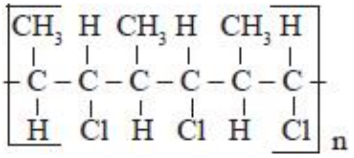
equation;

$$\text{Alkene} + H_2(g) \xrightarrow[y]{x} \text{Alkane}$$

Identify catalyst x and condition y.

5. State four uses of alkenes.
6. State a chemical test carried out to show evidence for the unsaturation in hydrocarbons.

7. The following structure represents part of a polymer. Given that the molecular mass of the monomer is 76.5. Answer the questions below.



(a) Identify and draw the structure of the repeat unit.

(b) Name the monomer.

(c) Write the molecular formula of the monomer.

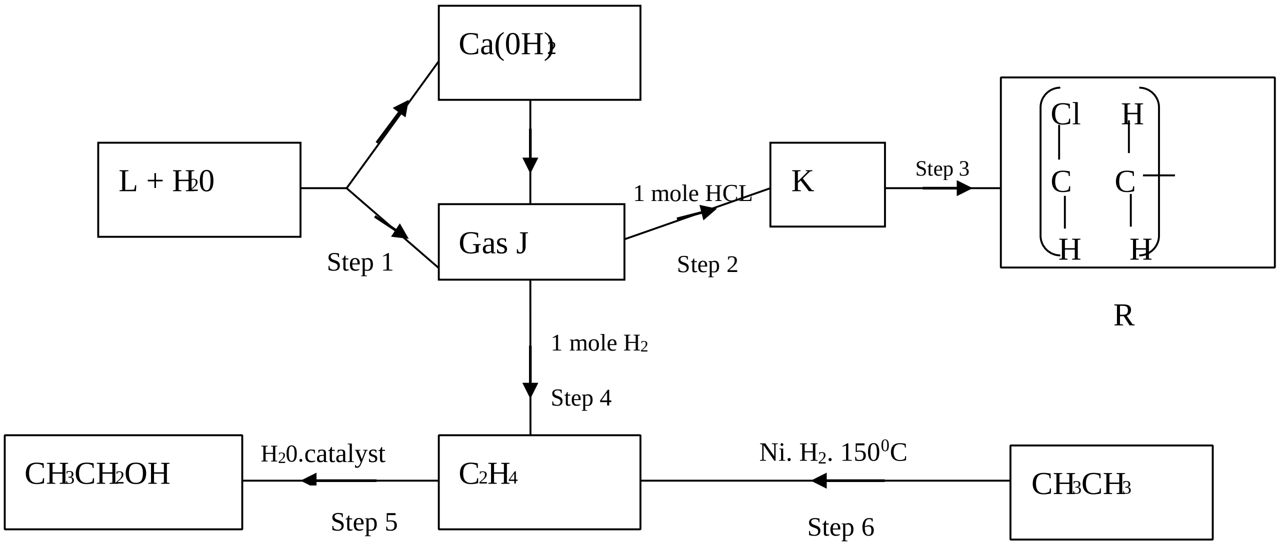
8. The following equation represents the reaction for the preparation of ethene from ethanol;

$$C_2H_5OH(l) \xrightarrow[Heat]{A} C_2H_4(g) + H_2O(l)$$

(a) Name reagent A.

(b) Name the process by which ethene is produced in the reaction represented by the equation.

9. Study the flow chart below and answer the questions that follow.



(a) Identify reagent L.

(b) Name the catalyst used in step 5.

(c) Draw the structural formula of gas J.

(d) What name is given to the process that takes place in step 5?

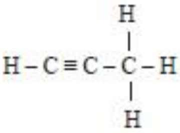
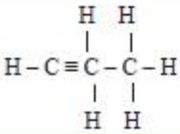
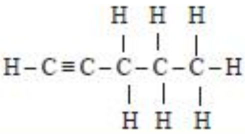
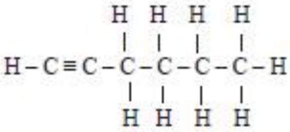
(e) Propane can be changed into methane and ethane as shown in the equation below;

$$CH_3CH_2CH_{3(g)} \xrightarrow{\text{High temperature}} CH_{4(g)} + C_2H_{4(g)}$$

Name the process undergone by propane.

3. ALKYNES

Alkynes are hydrocarbons which **contain a carbon – carbon triple bond (–C≡C–) in the molecule.**

Number of Carbon atoms(n)	Name	Molecular formula	Open structural formula	Condensed structural formula	Skeletal structure
2	Ethyne	C_2H_2	$\mathrm{H-C\equiv C-H}$	CHCH	
3	Propyne	C_3H_4	 $\mathrm{H-C\equiv C-\overset{H}{\underset{H}{C}}-H}$	$CHCCH_3$	
4	Butyne	C_4H_6	 $\mathrm{H-C\equiv C-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H}$	$CHCCH_2CH_3$	
5	Pentyne	C_5H_8	 $\mathrm{H-C\equiv C-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H}$	$CHCCH_2CH_2CH_3$	
6	Hexyne	C_6H_{10}	 $\mathrm{H-C\equiv C-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H}$	$CHC(CH_2)_3CH_3$	

Each alkyne differs from the next by a $-CH_2$ group and conforms to the **general formula, C_nH_{2n-2}, where 'n' represents the number of carbon atoms in a molecule, (n = 2, 3, 4…).** The first alkyne member is n = 2 because at least two carbon atoms are necessary for the formation of a triple bond. The presence of the triple bond results in unsaturation.

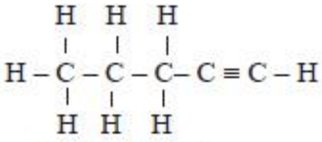
Nomenclature

All alkynes have names ending in –**yne**.

To name them, a prefix indicating the number of carbon atoms in the longest straight chain is followed by the suffix –**yne**.

The longest continuous carbon chain must contain the carbon-carbon triple bonds. This chain is numbered such that the carbon atoms having the triple bond have the lowest possible value as shown in the examples below.

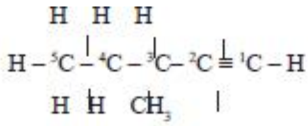
(i)



$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-C\equiv C-H$

pent –1–yne

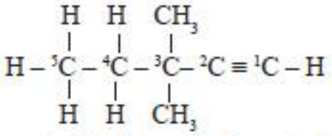
(ii)



$H-{}^5\underset{H}{\overset{H}{C}}-{}^4\underset{H}{\overset{H}{C}}-{}^3\underset{CH_3}{\overset{H}{C}}-{}^2C\equiv{}^1C-H$

3–methylpent–1–yne

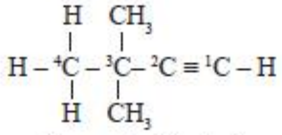
(iii)



$H-{}^5\underset{H}{\overset{H}{C}}-{}^4\underset{H}{\overset{H}{C}}-{}^3\underset{CH_3}{\overset{CH_3}{C}}-{}^2C\equiv{}^1C-H$

3, 3–demethylpent–1–yne

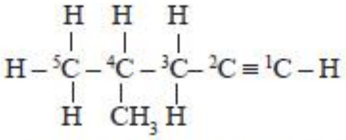
(iv)



$H-{}^4\underset{H}{\overset{H}{C}}-{}^3\underset{CH_3}{\overset{CH_3}{C}}-{}^2C\equiv{}^1C-H$

3–methylbut–1–yne

(v)



$H-{}^5\underset{H}{\overset{H}{C}}-{}^4\underset{CH_3}{\overset{H}{C}}-{}^3\underset{H}{\overset{H}{C}}-{}^2C\equiv{}^1C-H$

4–methylepent–1–yne

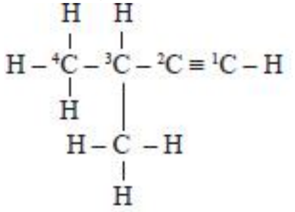
Isomerism in Alkynes

Alkynes show branching isomerism, and positional isomerism.

(a) Branching Isomerism

This occurs when a substituent group is attached to the longest chain containing the carbon - carbon triple bond.

For example the branched isomer of pent-l-yne is:



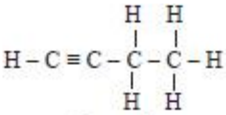
$H-{}^4\underset{H}{\overset{H}{C}}-{}^3\underset{H-\underset{H}{C}-H}{\overset{H}{C}}-{}^2C\equiv{}^1C-H$

3–methylbut-1-yne.

(b) Positional Isomerism

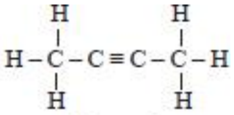
The position of the triple bond in an alkyne molecule can change. This results in the formation of two or more compounds with the same molecular formula but different structural formula.

For example:



$H-C\equiv C-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$

But–1–yne

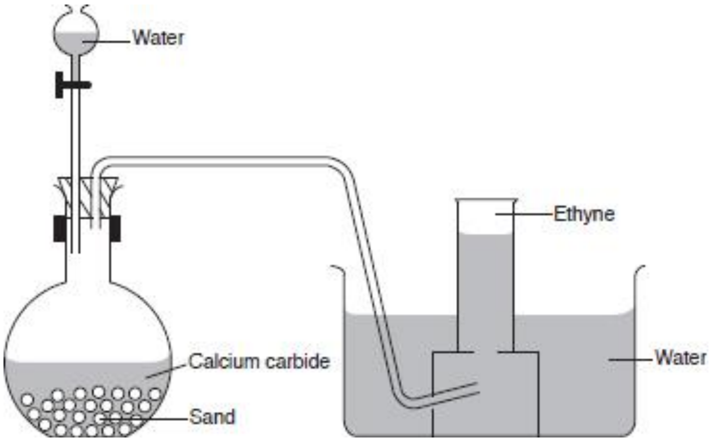


$H-\underset{H}{\overset{H}{C}}-C\equiv C-\underset{H}{\overset{H}{C}}-H$

But–2–yne

Ethyne and propyne do not show positional isomerism.

Laboratory Preparation of Ethyne



- Calcium carbide reacts with water at room temperature to form ethyne and calcium hydroxide.

$$CaC_2(s) + 2H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

- A layer of sand is placed at the base of the flask because the reaction is highly exothermic. The sand **absorbs** the excess heat and therefore **prevents the flask from breaking.**
- The flask must be dry before the start of the experiment **to avoid formation of the gas** before the set-up is complete.

Physical Properties

Ethyne is a colourless gas and has a pleasant smell when pure. It is slightly soluble in water, and therefore can be collected over water. The solubility of alkynes is higher in non-polar solvents. The table below gives some properties of alkynes.

Name	Molecular formula	Melting point (°C)	Boiling point (°C)
Ethyne	C_2H_2	–82	–84
Propyne	C_3H_4	–103	–23
But-l-yne	C_4H_6	–122	8
Pent-l-yne	C_5H_8	–90	39
Hex-l-yne	C_6H_{10}	–132	71

Alkynes with lower molecular mass are **gases** at room temperature. While those with a higher molecular mass are **solids**.

Chemical Properties

Ethyne burns in air, and also undergoes addition reactions.

(a) Combustion

Ethyne like any other hydrocarbon will burn in air to form an oxide of carbon, and water. This reaction is usually accompanied by production of a lot of heat, therefore, it is preferably used in oxy-acetylene flames.

In a limited supply of air, ethyne undergoes incomplete combustion to form a mixture of carbon and carbon(II) oxide. The yellow sooty flame observed is due to the **unburnt** carbon.

$$C_2H_2(g) + O_2(g) \xrightarrow{\textit{Limited supply of oxygen}} CO(g) + C(s) + H_2O(1)$$

In excess air, ethyne burns **completely** to form carbon(IV) oxide and water.

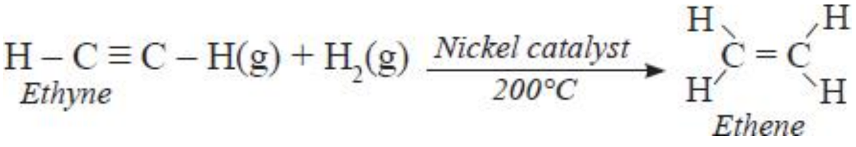
$$2C_2H_2(g) + 5O_2(g) \xrightarrow{\textit{Excess oxygen}} 4CO_2(g) + 2H_2O(1)$$

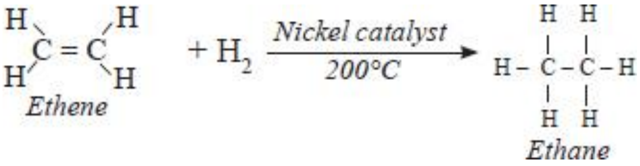
(b) Addition Reactions

Addition reactions in alkynes are faster than in alkenes due to the presence of the triple bond.

(i) Reaction with hydrogen (hydrogenation)

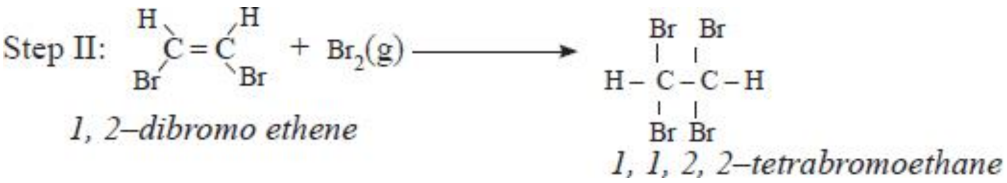
Ethyne reacts with hydrogen in the presence of a nickel catalyst to form first ethene, then ethane.





(ii) Reaction with halogens (halogenation)

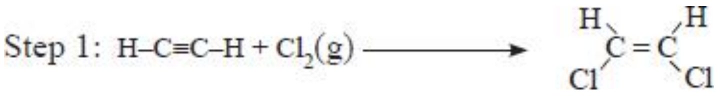
When ethyne reacts with red-brown bromine vapour, **the bromine vapour is decolourised**. The decolourisation process is faster in ethyne than in ethene. The bromine atoms are added to the carbon-carbon triple bond to form 1, 1, 2, 2,-tetrabromoethane. This reaction takes place at **room temperature** and it is a two-step reaction. Thus:

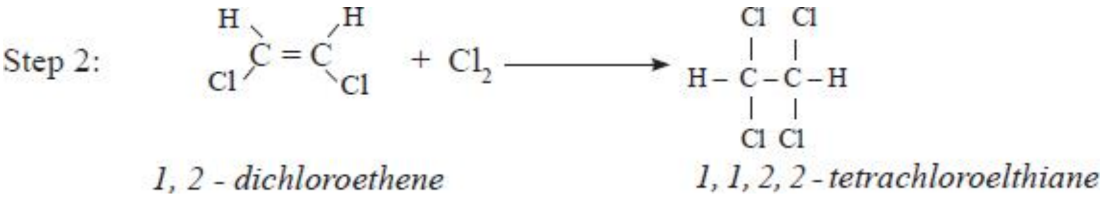


Pure chlorine reacts with ethyne with a violent explosion forming carbon, and hydrogen chloride.

$C_2H_2 + Cl_2(g) \rightarrow 2C(s) + 2HCl(g)$

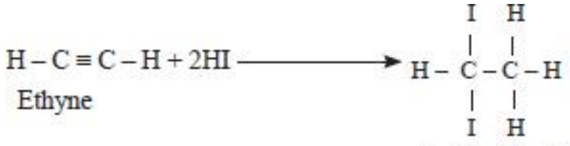
When diluted with an inert gas, chlorine reacts with ethyne to from 1, 1, 2, 2-tetrachloroethane in a two step reaction.





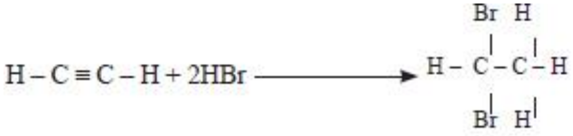
(iii) Reaction with hydrogen halides (Hydrohalogenation)

Ethyne does not react with all the hydrogen halides. Hydrogen iodide reacts readily at room temperature, hydrogen bromide reacts when warmed while hydrogen chloride reacts slowly.



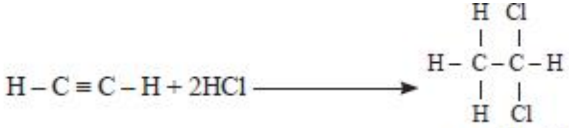
1, 1–diiodoethane

(ii) Hydrogen bromide reacts with ethyne to produce 1,1–dibromoethane (CH_3CHBr_2)



1, 1 dibromoethane

(iii) The reaction between hydrogen chloride gas is slow.



1,1 dichloroethane

Test for Alkynes

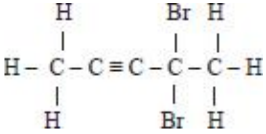
The test for an alkynes is similar to the tests for alkenes in which the oxidising agents are decolourised. However, the reaction is **faster** in alkynes than in alkenes.

Uses of Alkynes

1. They are used in the industrial manufacture of compounds such as adhesives, and plastics.
2. In the manufacture of synthetic fibres such as rayon.
3. In the production of important chemical reagents and solvents, in which it is used as a starting material.
4. They are used in the oxy-acetylene flame which is used for welding, and cutting metals.

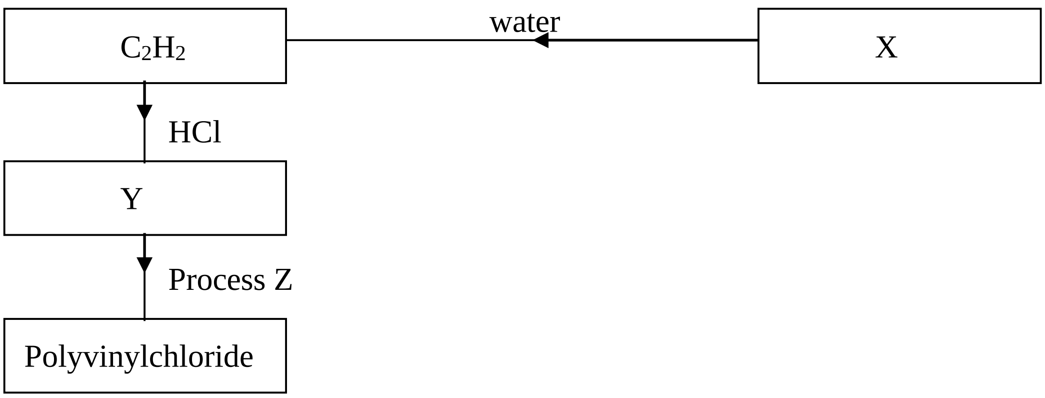
Review Questions: Alkynes

1. (a) Write the molecular formula of butyne.
   (b) Draw the open structural formula of butyne.
   (c) Butyne burns with a yellow sooty flame in a limited supply of air. Explain.
   (d) Write the equation for the reaction in (c).
2. Name the compound represented by the following structure.



$H-C-C\equiv C-C-C-H$

3. Draw the open structure of the compound with the following molecular structure: C_6H_{10}
4. The reaction between bromine vapour and ethyne is faster than with ethene. Explain.
5. The relative formula mass of a hydrocarbon is 58. Draw and name two possible structures of the hydrocarbon (C=12.0; H=1.0
6. Alkanes, alkenes and alkynes can be obtained from crude oil. Draw the structure of the second member of the alkyne homologous series.
7. Draw and name the structure of the compound formed when one Mole of ethyne reacts with one mole of hydrogen bromide.
8. Draw the structures of the alkynes whose molecular formula is C_4H_6
9. Study the flowchart below and answer the questions that follow:

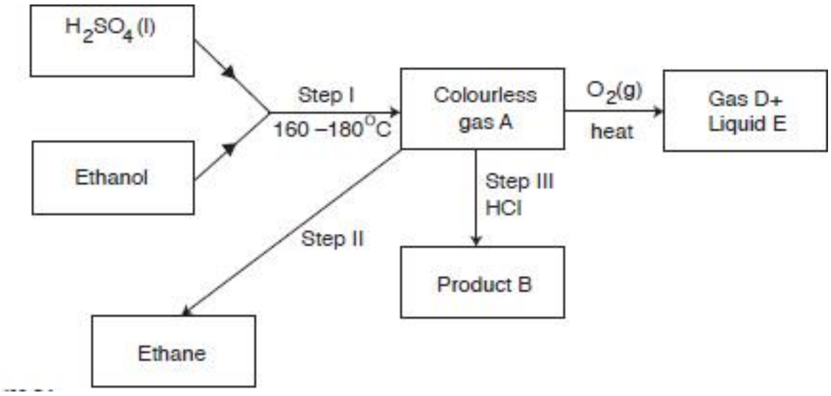


(a) Identify X and Y
(b) State two uses of polyvinylchloride
10. Draw and name the isomers of butyne

10. State one use of polystyrene

Revision Exercise: HYDROCARBONS

1. Crude oil is the main source of organic compounds such as hydrocarbons. The hydrocarbons in the crude oil have to be separated.
   (a) Name four important hydrocarbons obtained from crude oil.
   (b) Give the uses of the four hydrocarbons named in (a) above.
   (c) Explain with the help of a suitable diagram, the principles used in separating hydrocarbons in crude oil.
2. What do you understand by the following terms:
   (a) Catalytic cracking.
   (b) Thermal cracking.
3. Study the following reaction scheme and answer the questions that follow:



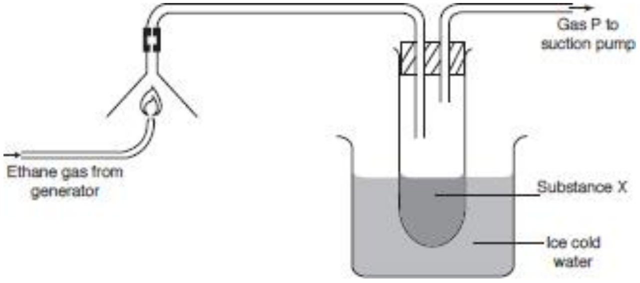
   (a) Name:
      (i) Colourless gas A.
      (ii) Product B.
      (iii) Gas D.
      (iv) Liquid E.
   (b) Write balanced equations for each of the reactions forming the products in (a).
   (c) Name the type of reactions taking place in Step I and II.
   (d) State the importance of the reaction taking place in Step II.

4. Explain why an organic compound with formula C_3H_6 burns with a more sooty flame than C_3H_8
5. Butane and bromine react as shown below:

$CH_3,CH_2.CH_2CH_3 + Br_2 \rightarrow CH_3CH_2.CH_2Br + HBr$

   (a) Name the type of reaction taking place in the equation above.
   (b) State the condition under which the above reaction takes place. Explain.
6. A hydrocarbon Q, was found to decolourise potassium manganate(VII) solution. When two moles of Q are burnt completely, six moles of carbon(IV) oxide and six moles of water were formed.
   (a) Write the structural formula of Q.
   (b) Name the homologous series to which Q belongs.
   (c) Name one industrial source of Q.

7. The diagram below shows the combustion of ethane gas.



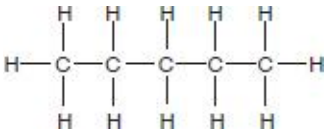
(a) Identify substance x.

(b) Write the equation for the complete combustion of ethane gas.

(c) What is the purpose of ice cold water in the experiment?

(d) The pH of substance X was found to be less than 7. Explain this observation.

8. One mole of hydrogen bromide reacts with an organic compound N to give a single product with the structural formula shown below:



$CH_3-CH_2-CH_2-CH_2-CH_3$

(a) Give the name of the hydrocarbon.

(b) Draw the possible structure of N.

TOPIC FOUR

NITROGEN AND ITS COMPOUNDS

Table of Contents

Objectives

By the end of the chapter the learner should be able to:

(a) Describe how nitrogen is isolated from air.
(b) Describe how nitrogen is prepared in the laboratory and state its properties and uses.
(c) Describe the laboratory preparation and state the properties and uses of the oxides of nitrogen.
(d) Describe the laboratory preparation of ammonia and state its properties and uses.
(e) Explain the differences in chemical reactions of ammonia gas and its aqueous solution.
(f) Describe the industrial manufacture of ammonia.
(g) Calculate the percentage of nitrogen in nitrogen containing fertilisers.
(h) Describe the laboratory preparation and manufacture of nitric(V) acid and state its uses.
(i) Explain the reactions of both dilute and concentrated nitric(V) acid.
(j) Identify the products formed when different nitrates are heated.
(k) State and explain the pollution effect of nitrogen compounds in the environment

NITROGEN AND ITS COMPOUNDS

Nitrogen is a non metal in group V of the periodic table. It has an atomic number of seven and has an electron arrangement of 2.5. Nitrogen exists in air as a gaseous element consisting of diatomic molecules. About 78% by volume of the atmosphere is composed of nitrogen making it the most important source of nitrogen.

Nitrogen also occurs combined in Compounds such as sodium nitrate, potassium nitrate and proteins

Nitrogen

Isolation of Nitrogen from Air

Laboratory Isolation

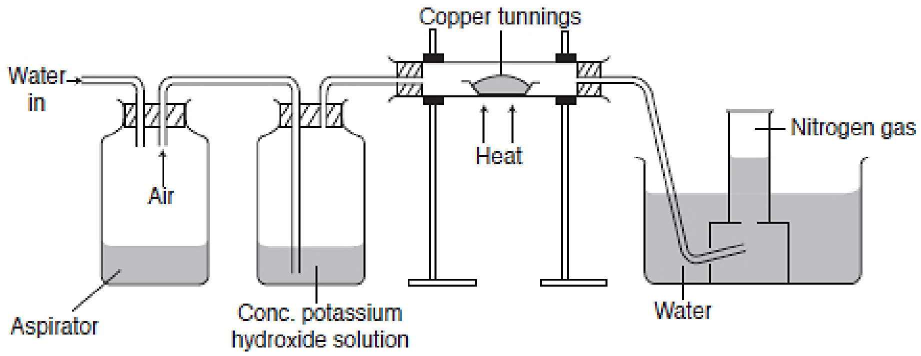
In the laboratory, nitrogen is isolated from air by first passing the air through concentrated potassium hydroxide or sodium hydroxide which **absorbs carbon(IV) oxide.**

$$KOH(aq) + CO_2(g) \longrightarrow KHCO_3(aq)$$

The remaining part of air is then passed over heated copper turnings to remove oxygen.

$$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$$

The residual gas is then collected as shown below.



The nitrogen so obtained **contains noble gases as impurities.**

Fractional distillation of liquefied air

Nitrogen is obtained on a large scale by fractional distillation of liquid air.

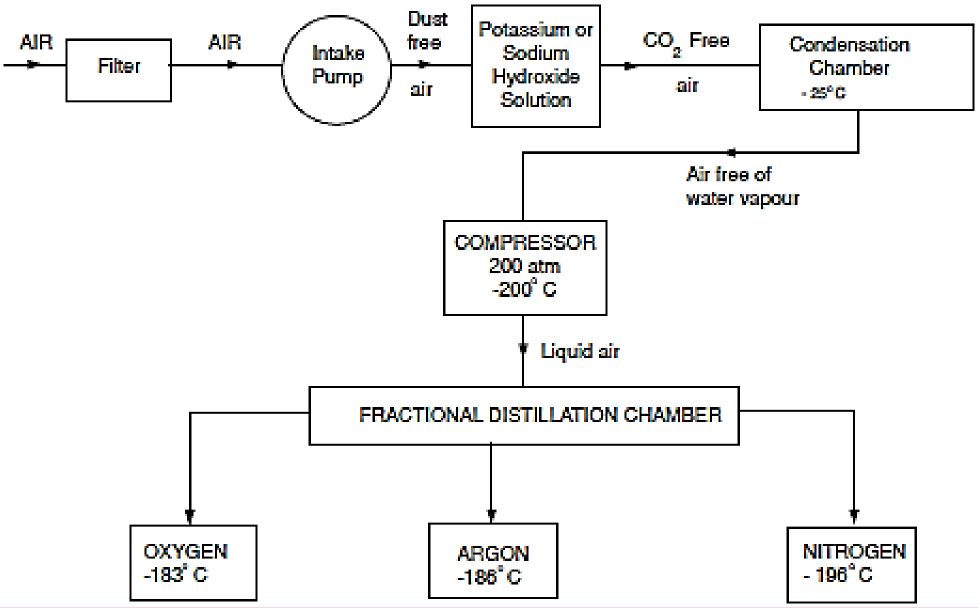
Dust particles are first removed from air through **electrostatic precipitation.**

The dust-free air is then passed through **concentrated sodium hydroxide solution to remove carbon(IV) oxide gas.**

The remaining air is passed into a **condensation chamber and cooled to –25°C. Water vapour separates out as ice.** The residual gas is compressed to a pressure of 200 atmospheres.

Repeated expansion and contraction of the air cools it to a liquid at a temperature of –200°C. The main constituents of liquid air at this stage are nitrogen and oxygen whose boiling points are –196°C and –183°C respectively. **Nitrogen has a lower boiling point and therefore distills out first followed by oxygen.**

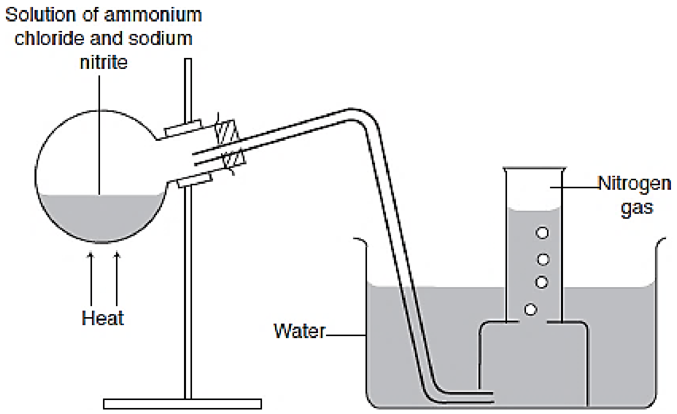
The flow chart below summarises the major steps in the isolation of nitrogen.



Laboratory preparation of nitrogen.

Nitrogen can be prepared by heating a mixture of sodium nitrite and ammonium chloride Ammonium chloride and sodium nitrite react to form ammonium nitrite and sodium chloride.

Ammonium nitrite is unstable and decomposes to form nitrogen and steam.



Ammonium chloride and sodium nitrite react to form ammonium nitrite and sodium chloride.

$$NaNO_2(aq) + NH_4Cl(aq) \longrightarrow NH_4NO_2(aq) + NaCl(aq)$$

Ammonium nitrite is unstable and decomposes to form nitrogen and steam.

$$NH_4NO_2(aq) \longrightarrow N_2(g) + 2H_2O(g)$$

Side Notes

- Nitrogen is slightly soluble in water hence it is collected over water.
- Nitrogen obtained from this reaction is less dense than that isolated from air because it does not contain impurities.

Properties of nitrogen.

Nitrogen is a colourless, oduorless and tasteless gas that is slightly less dense than air. The boiling point is –96°C. It is slightly soluble in water.

Nitrogen neither burns nor supports combustion. It is neutral and has no effect on moist litmus paper and does not react readily with other elements.

- Nitrogen gas does not react with sulphur and phosphorus. It is chemically unreactive under ordinary conditions. A molecule of nitrogen has a triple covalent bond, N≡N, which is very strong and requires a lot of energy to break. The presence of nitrogen in the air slows down the rate of burning due to its inert nature.
- At **higher temperatures**, it reacts with metals in group I and II to form their nitrides.

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

When some water is added to the nitride, ammonia gas and magnesium hydroxide are formed.

$$Mg_3N_2(s) + 6H_2O(l) \longrightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$$

- Nitrogen is distinguished from other gases by its negative results with all the tests used to identify common gases.

Uses of Nitrogen

- Manufacture of ammonia in the Haber process.
- In light bulbs because of its inert nature it cannot react with the hot filament of the bulb.
- As a refrigerant, e.g., in the storage of semen for artificial insemination.

Oxides of Nitrogen

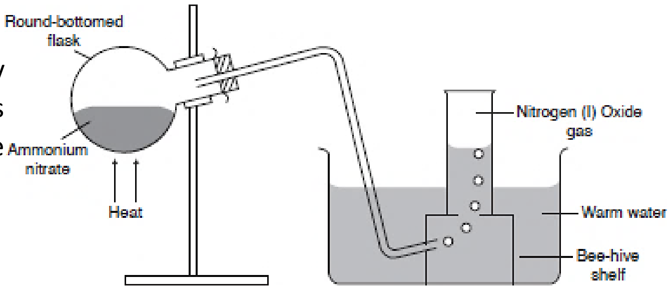
Several oxides of nitrogen exist. These include:

1. Nitrogen(I) oxide N_2O.
2. Nitrogen(II) oxide NO.
3. Nitrogen(IV) oxide, NO_2.

(a) Nitrogen(I) oxide (Dinitrogen oxide)

Laboratory preparation of nitrogen (I) oxide.

Nitrogen (I) oxide can be prepared in the laboratory by heating ammonium nitrate. Ammonium nitrate melts and decomposes on heating to form nitrogen(I) oxide and steam.

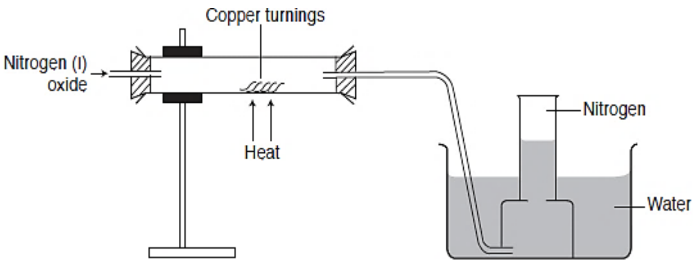


$$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$$

Properties of nitrogen (I) oxide.

- Nitrogen(I) oxide is a **colourless gas with a pleasant smell**. The gas causes insensitivity when inhaled.
- It is **slightly less dense than air**. It is fairly soluble in cold water but insoluble in warm water. For this reason it is usually collected over warm water.
- Nitrogen(I) oxide gas is **not reactive at room temperature**. However, it relights a glowing splint. The heat from the glowing splint dissociates the gas producing nitrogen and oxygen gas. The oxygen produced relights the splint.
- When nitrogen(I) oxide is passed over heated copper, a black residue of copper (II) oxide and nitrogen are formed.

$Cu(s) + N_2O(g) \xrightarrow{\qquad\qquad} CuO(s) + N_2(g)$



- Sulphur burns brilliantly in nitrogen(I) oxide to form sulphur(IV) oxide and nitrogen.

$S(s) + N_2O(g) \xrightarrow{\qquad\qquad} SO_2(s) + N_2(g)$

Uses of Nitrogen(I) Oxide

- It was formerly used as an anaesthetic during dental surgery. Patients recovering from it laugh hysterically hence the name 'laughing gas'.
- It is used as a food additive.
- It is used as an oxidiser in racing car engines and rockets.
- It is used to produce flames for analytical work.

(b) Nitrogen(II) Oxide (Nitrogen Monoxide)

Laboratory preparation of nitrogen (II) oxide

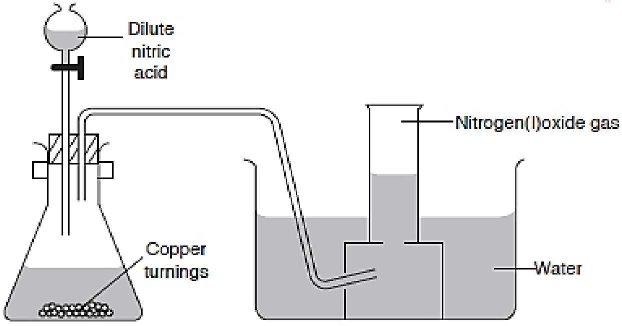
Nitrogen(II) oxide is produced when copper and dilute nitric(V) acid react.

$$3Cu(s) + 8HNO_3(aq) \xrightarrow{\qquad\qquad} 3Cu(NO_3)_2(aq) + 4H_2O(1) + 2NO(g)$$

Nitrogen(II) oxide is oxidised by oxygen in the flask to form red-brown fumes of nitrogen(IV) oxide.

$$2NO(g) + O_2\ (g) \xrightarrow{\qquad\qquad} 2NO_2\ (g)$$

The fumes gradually disappear as the air in the flask is exhausted. Nitrogen(IV) oxide formed dissolves in the water in the trough.



Properties of Nitrogen(II) Oxide (Nitrogen Monoxide)

Nitrogen(II) oxide is a colourless gas. It is slightly soluble in water and has no effect on moist litmus paper.

Nitrogen(II) oxide neither burns nor supports combustion. However it oxidises some strongly heated elements to their oxides.

Magnesium continues to burn in nitrogen(II) oxide. The heat produced decomposes the gas into nitrogen and oxygen. The oxygen produced enables the hot element to continue burning.

$$2Mg(s) + 2NO(g) \longrightarrow 2MgO(s) + N_2(g)$$

$$2Cu(s) + 2NO(g) \longrightarrow 2CuO(s) + N_2(g)$$

$$P_4(s) + 10NO(g) \longrightarrow 2P_2O_5(g)\ 5N_2(g)$$

Iron(II) sulphate solution turns dark brown when nitrogen(II) oxide is bubbled through it. This is due to the formation of iron(II) sulphate – nitrogen(II) oxide complex, $FeSO_4.NO$.

$$\underset{\textit{Pale green}}{FeSO_4(aq) + NO(g)} \longrightarrow \underset{\textit{dark brown}}{FeSO_4.NO\ (aq)}$$

When exposed to air, nitrogen(II) oxide is readily oxidised by oxygen to form red brown fumes of nitrogen(IV) oxide. This reaction is used as **the test for nitrogen(II) oxide.**

(c) Nitrogen(IV) Oxide (Nitrogen Dioxide).

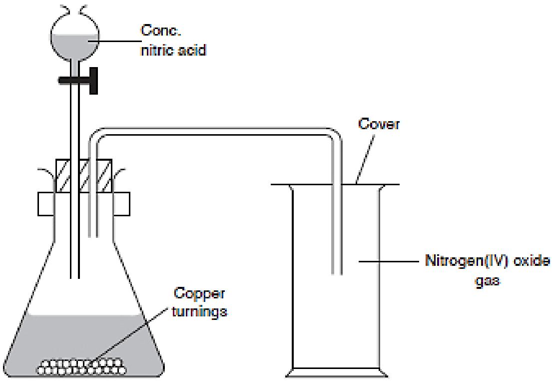
Preparation of nitrogen (IV) oxide

Nitrogen (IV) Oxide can be prepared by the reaction of concentrated nitric(V) acid on copper turnings.

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Nitrogen(IV) oxide may also be prepared by thermal decomposition of nitrates of metals below sodium in the reactivity series. However, lead(II) nitrate is the most suitable because it is not hydrated.

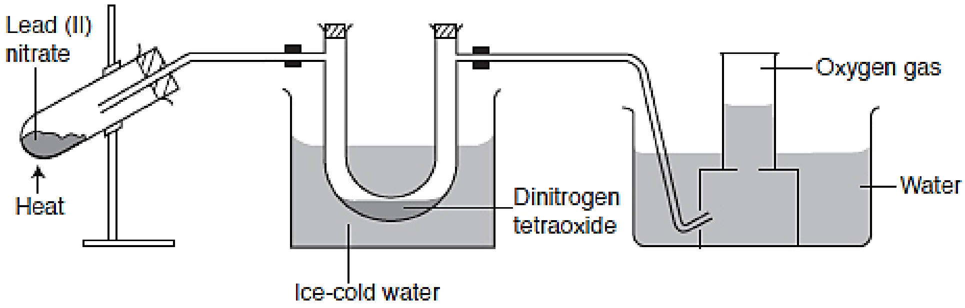
When concentrated nitric(V) acid is added to copper turnings, a vigorous reaction takes place and red-brown fumes of nitrogen(IV) oxide are evolved.



$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Nitrogen(IV) oxide may also be prepared by thermal decomposition of nitrates of metals below sodium in the reactivity series. However, lead(II) nitrate is the most suitable because it is not hydrated.

When heated, lead(II) nitrate decomposes to form lead(II) oxide, nitrogen(IV) oxide and oxygen. The set up below can be used to prepare nitrogen(IV) oxide from metal nitrates.



$$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$$

The gases produced are passed through a U-tube surrounded by ice-cold water. On cooling, nitrogen(IV) oxide condenses to form dinitrogen tetraoxide, N_2O_4, which is a pale-yellow liquid. Oxygen gas is collected over water.

$$\underset{\textit{Reddish-brown}}{2NO_2(g)} \xrightarrow{\textit{cooling}} \underset{\textit{pale-yellow}}{N_2O_4\,(l)}$$

Properties of Nitrogen(IV) Oxide

- Nitrogen(IV) oxide gas is easily liquefied. At room temperature nitrogen(IV) oxide collected exists as an equilibrium mixture of itself and dinitrogen tetraoxide.

$$\underset{\textit{Reddish-brown}}{2NO_2(g)} \rightleftharpoons \underset{\textit{pale-yellow}}{N_2O_4\,(g)}$$

- When heated, nitrogen(IV) oxide dissociates to form a colourless mixture of nitrogen (II) oxide and oxygen.

$$\underset{\textit{Reddish-brown}}{2NO_2(g)} \longrightarrow \underset{\textit{colourless}}{2NO\,(g) + O_2\,(g)}$$

- Nitrogen(IV) oxide is a reddish-brown gas, with an irritating pungent smell. The gas is poisonous. It is denser than air and is therefore collected by downward delivery. Nitrogen(IV) oxide is soluble in water. It dissolves in water to form both nitric(III) acid and nitric(V) acid.

$$2NO_2(g) + H_2O(l) \longrightarrow HNO_2(aq) + HNO_3(aq)$$

- It neutralises alkalis forming a mixture of their corresponding nitrates and nitrites. Magnesium ribbon continues to burn in nitrogen(IV) oxide to form white fumes of magnesium oxide and nitrogen.

$$4Mg(s) + 2NO_2(g) \longrightarrow 4MgO(s) + N_2(g)$$

- Burning phosphorus continues to burn in nitrogen(IV) oxide to form phosphorus (V) oxide.

$$2P_4(s) + 10NO_2(g) \longrightarrow 4P_2O_5(g) + 5N_2(g)$$

The heat produced by the burning element decomposes nitrogen(IV) into nitrogen and oxygen. Oxygen liberated combines with the hot element.

Uses of Nitrogen (IV) Oxide

- In the manufacture of nitric(V) acid.
- As an intermediate in the manufacture of explosives, nylon and plastics.
- As an oxidising agent in the lead chamber during the manufacture of sulphuric (VI) acid.

Ammonia

Ammonia, NH_3, is a compound of nitrogen and hydrogen. It is a gas at room temperature.

Laboratory preparation of Ammonia

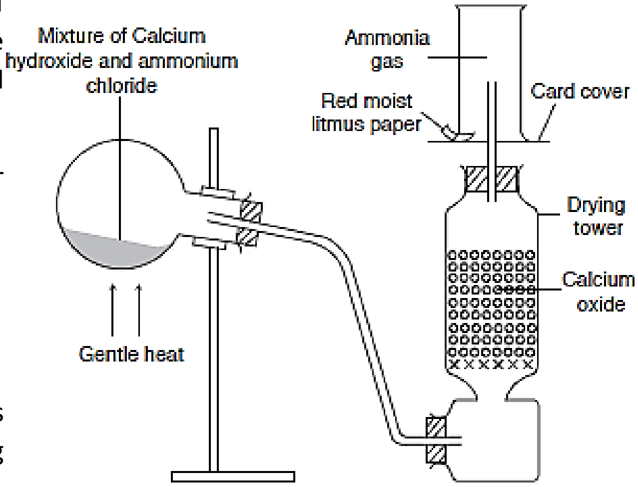
Ammonia is prepared in the laboratory by heating a mixture of an ammonium salt and an alkali. Ammonia is a more volatile base and is displaced by a less volatile base, e.g., NaOH, KOH and $Ca(OH)_2$.

$$Ca(OH_2)(s) + 2NH_4Cl(s) \longrightarrow CaCl_2(s) + 2H_2O(l) + 2NH_3(g)$$

The ionic equation for the reaction is:

$$NH_4^+(aq) + OH^-(aq) \longrightarrow NH_3(g) + H_2O(l)$$

Calcium oxide (quicklime) is used as a drying agent. This is because ammonia reacts with the other common drying agents.



Side Notes

- The flask containing the mixture is set in a slanting position to prevent water which condenses on the cooler parts of the apparatus, from running back into the flask and causing it to crack.
- A moist litmus paper placed at the mouth of the gas jar turns blue indicating the gas jar is full of ammonia.
- Calcium oxide (quicklime) is used as a drying agent. This is because ammonia reacts with the other common drying agents.

  Concentrated sulphuric (VI) acid reacts with ammonia to form ammonium sulphate.

$$2NH_3(g) + H_2SO_4(l) \longrightarrow (NH_4)_2SO_4(aq)$$

  Calcium chloride forms a complex compound with ammonia.

$$CaCl_2(s) + 4NH_3(g) \longrightarrow CaCl_2.4NH_3(s)$$

Properties of Ammonia

Ammonia is a colourless gas, with a characteristic **choking pungent smell**. It is less dense than air and therefore it is collected by upward delivery. The gas is **very soluble** in water.

Ammonia turns moist red litmus paper blue showing that it is alkaline. This is the **confirmatory test for ammonia.**

When a glass rod dipped in concentrated hydrochloric acid is brought to the mouth of a gas jar full of ammonia, hydrogen chloride fumes react with ammonia to form white fumes of ammonium chloride.

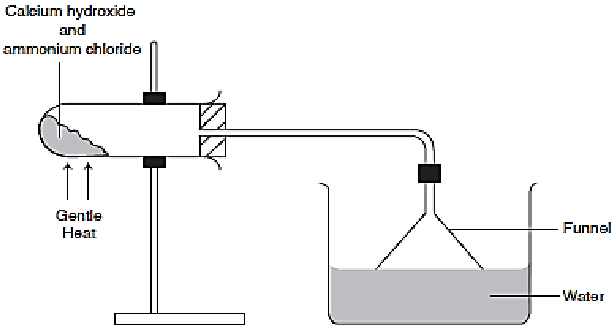
$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(g)$$

Preparation of aqueous ammonia.

Ammonia dissolves in water to form an alkaline solution. To prepare aqueous ammonia,

Practically Speaking

Place a mixture of calcium hydroxide and ammonium chloride in a boiling tube and heat. Allow the gas collected to dissolve in water using a funneled delivery tube.



Side Note

A funnel is used to prevent the water from 'sucking back' into the flask by providing a large surface area for absorption of ammonia.

Discourse

Ammonia dissolves in water to form an alkaline solution. A little of the dissolved gas combines with water to form ammonium ions and hydroxide ions.

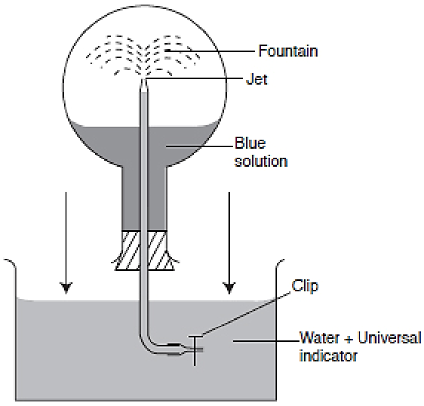
$$NH_3\,(g) \xrightarrow{H_2O\ (l)} NH_3\,(aq)$$

$$NH_3\,(aq) + H_2O\ (l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

The hydroxide ion, OH^-, gives the solution its alkaline properties.

The fountain experiment

This experiment illustrates the high solubility of ammonia gas in water. Universal indicator is added to enhance the visibility of the "fountain".



When a drop of water gets to the jet, it dissolves a lot of the ammonia gas in the flask creating a partial vacuum. When the clip is opened the second time, the water is forced into the flask by the atmospheric pressure leading to the formation of a fountain.

Reaction of aqueous ammonia with metal ions.

Aqueous ammonia **precipitates many metal hydroxides** from solutions containing ions of the metal. This property is applied in **qualitative analysis of metal ions**. These tests can be done practically as follows:

Practically Speaking

Place about 2 cm^3 of solutions containing ions of calcium, magnesium, aluminium, zinc, iron, lead and copper in separate test-tubes. To each solution add aqueous ammonia drop by drop. Shake the mixture after each drop. Continue adding aqueous ammonia until it is in excess. Record your observations.

Observations

Metal ion in solution	Few drops of aqueous ammonia	Excess aqueous ammonia
Ca^{2+}	No white precipitate	No white precipitate
Mg^{2+}	White precipitate	White precipitate does not dissolve.
Al^{3+}	White precipitate	White precipitate does not dissolve.
Zn^{2+}	White precipitate	White precipitate dissolves to form a colourless solution.
Fe^{2+}	Green precipitate	Green precipitate does not dissolve.
Fe^{3+}	Brown Precipitate	Brown precipitate does not dissolve.
Pb^{2+}	White precipitate	White precipitate does not dissolve.
Cu^{2+}	Pale-blue precipitate	Pale-blue precipitate dissolves to form a deep-blue solution.

Discussion

Aqueous ammonia is a **weak alkali and will not precipitate the hydroxide in water if it is slightly soluble,** for example, **calcium hydroxide**. Ionic equations representing precipitation of insoluble metal hydroxides are:

$$Mg^{2+}(aq) + 2OH^-(aq) \longrightarrow Mg(OH)_2(s)$$

White precipitate

$$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)$$

White precipitate

$$Zn^{2+}(aq) + 2OH^-(aq) \longrightarrow Zn(OH)_2(s)$$

White precipitate

$$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$$

Green precipitate

$$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$$

Red brown precipitate

$$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$$

White precipitate

$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$$

Pale-blue precipitate

Zinc hydroxide and copper(II) hydroxide dissolve in excess ammonia due to the **formation of complex ions.**

$$Zn(OH)_2(s) + 4NH_3(aq) \longrightarrow [Zn(NH_3)_4]^{2+} + 2OH^-(aq)$$

White precipitate *Tetramine zinc ion (colourless)*

$$Cu(OH)_2(s) + 4NH_3(aq) \longrightarrow [Cu(NH_3)_4]^{2+} + 2OH^-(aq)$$

White precipitate *Tetramine copper (II) ion (deep blue)*

Calcium ions form a white precipitate when **2 M aqueous ammoni**a is used.

Reaction of aqueous ammonia with dilute acids.

Ammonia gas dissolves in water to form aqueous ammonia. Some of the dissolved ammonia molecules react with water molecules to form **ammonium hydroxide.**

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4OH(aq)$$

Therefore, **aqueous ammonia neutralises acids to form ammonium salts and water**.

$$2NH_4OH(aq) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq) + H_2O(l)$$

$$NH_4OH(aq) + HCl(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$$

$$NH_4OH(aq) + HNO_3(aq) \longrightarrow NH_4NO_3(aq) + H_2O(l)$$

If dry ammonia is bubbled through dilute acids, ammonium salts are formed.

$$2NH_3(g) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)$$

$$NH_3(g) + HCl(aq) \longrightarrow NH_4Cl(aq)$$

$$NH_3(g) + HNO_3(aq) \longrightarrow NH_4NO_3(aq)$$

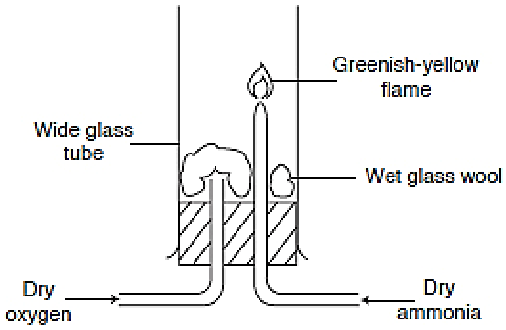
Does ammonia burn in air?

Ammonia does not burn in air. However, it burns with a greenish-yellow flame in air enriched with oxygen to form nitrogen and steam.

The set up below can be used to investigate this property of Ammonia.

Ammonia burns with a greenish-yellow flame in air enriched with oxygen to form nitrogen and steam.

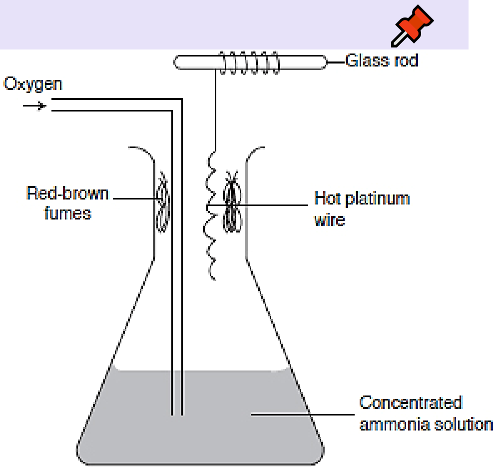
$$4NH_2(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(g)$$



Catalytic oxidation of Ammonia

Ammonia is oxidised to nitrogen(II) oxide in the presence of a platinum catalyst. This can be demonstrated practically.

Practically Speaking



Place concentrated ammonia solution in a conical flask. Arrange the apparatus as shown. Heat a platinum wire until it is red-hot. Pass oxygen or air through the concentrated ammonia. Place the hot platinum or copper in the flask. Record your observations.

Observation and Discussion

- The hot platinum wire glows on coming into contact with the fumes of ammonia from the concentrated ammonia solution.

  Ammonia is oxidised to nitrogen(II) oxide in the presence of a platinum catalyst.

$$4NH_3(g) + 5O_2(g) \xrightarrow{\textit{platinum catalyst}} 4NO_2(g) + 6H_2O(g)$$

- Reddish-brown fumes of nitrogen (IV) oxide are produced due to further oxidation of nitrogen (II) oxide.

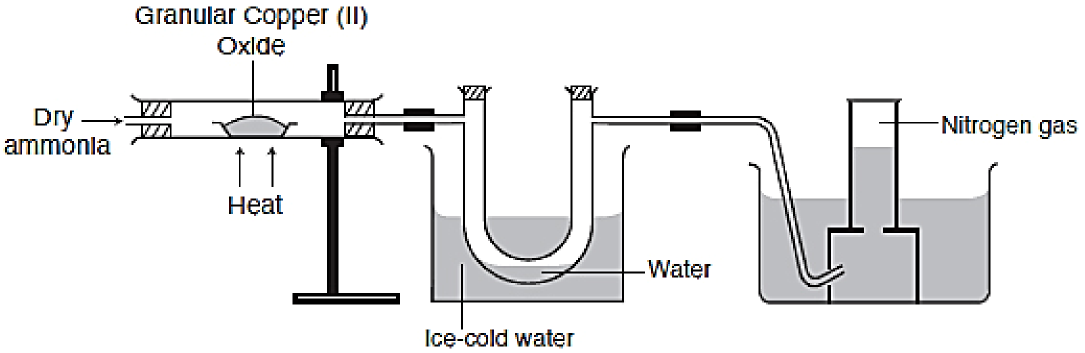
$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

Side notes

- A lot of heat is produced during the reaction that enables the platinum coil to continue glowing.
- The platinum wire acts as a catalyst. Copper wire may also be used as a catalyst.

Reaction of ammonia with copper(II) oxide.

Ammonia gas can reduce copper(II) Oxide to copper.



When ammonia is passed over heated copper(II) oxide, the black copper(II) oxide turns into a brown solid as copper(II) oxide is reduced to copper metal by ammonia. Water and nitrogen gas is also obtained.

$$\underset{\textit{Black}}{3CuO(s)} + 2NH_3(g) \longrightarrow \underset{\textit{brown}}{3Cu(s)} + N_2(g) + 3H_2O(l)$$

The **colourless liquid** collected in the U-tube turns cobalt(II) chloride paper pink. This test confirms the presence of **water**.

The gas collected has **no effect on a moist litmus paper and lime water. It extinguishes a burning splint. The gas is nitrogen.**

Large Scale Manufacture of Ammonia, the Haber Process

Ammonia is prepared on a large scale by the Haber Process.

The raw materials

Nitrogen- obtained by the fractional distillation of liquid air.

Hydrogen - obtained from natural gas or as a by-product of cracking long chain alkanes.

The process

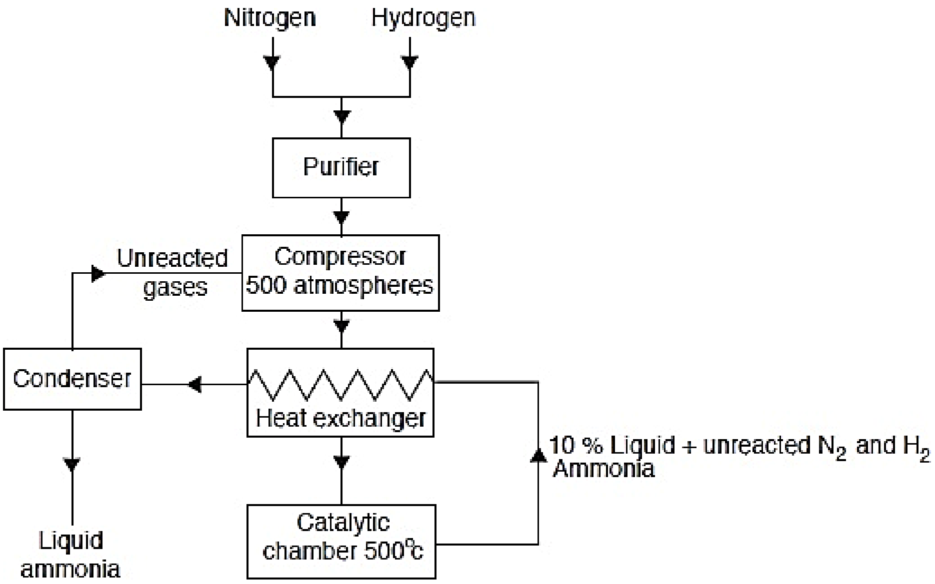
Nitrogen and hydrogen gases are mixed in the ratio of 1:3 and passed through a purifier to remove impurities which **would otherwise poison the catalyst**. The impurities removed are; **carbon(IV) oxide, sulphur(IV) oxide and dust particles.**

The mixture of nitrogen and hydrogen is then compressed to between 200 and 500 atmospheres and passed into a catalytic chamber maintained at temperatures of between 400°C and 500°C. **Iron is used as the catalyst.** The reaction is highly exothermic and sustains the temperature in the catalytic chamber.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) + \text{Heat}$$

The hot mixture of ammonia and gases which have not reacted are passed into the heat exchanger where cooling takes place. The heat is used to heat the incoming mixture of nitrogen and hydrogen, this makes the process economical.

The ammonia produced is liquefied in the condenser and stored in cylinders. The gases which have not reacted are recycled therefore reducing wastage.



The yield of ammonia is high at higher pressures and low temperature. However, it is expensive to maintain higher pressure. At low temperatures the reaction is slow.

To reduce the production cost **optimum conditions of 500°C, 200 atmospheres and iron catalyst** are employed.

Uses of Ammonia

- As a fertiliser.
- Manufacture of nitrogenous fertilisers.
- As a refrigerant, e.g., in large scale refrigerating plants such as ships and warehouses.
- Softening water.
- Removal of greasy stains.
- Manufacture of hydrazine that is used as rocket fuel.

4. Nitrogenous Fertilisers

(i) Ammonium sulphate, $(NH_4)_2SO_4$

It is the most widely used nitrogenous fertiliser. It is prepared by reacting ammonia and sulphuric acid.

$$2NH_3(g) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(s)$$

The percentage of nitrogen in ammonium sulphate fertilisers is determined as follows.

$$\text{Percentage of nitrogen} = \frac{\text{relative atomic mass of nitrogen}}{\text{relative formula mass}} \times 100$$

$$= \frac{28}{132} \times 100$$

$$= 21.\ 21\%$$

(ii) Ammonium nitrate, NH_4NO_3

It is used widely due to a high content of nitrogen. It is prepared by neutralising nitric acid with ammonia.

$$NH_3(aq) + HNO_3(aq) \longrightarrow NH_4NO_3(aq)$$

The percentage composition of nitrogen in ammonium nitrate is determined as:

$$\text{Percentage of nitrogen} = \frac{\text{relative atomic mass of nitrogen}}{\text{relative formula mass}} \times 100$$

$$= \frac{28}{80} \times 100$$

$$= 35.0\%$$

(iii) Ammonium phosphate $(NH_4)_3PO_4$

Apart from nitrogen, ammonium phosphate also contains phosphorus which is of vital importance for plant growth. It is prepared by neutralising phosphoric acid with ammonia.

$$3NH_3(g) + H_3PO_4(aq) \longrightarrow (NH_4)_3\ PO_4(aq)$$

$$\text{Percentage of nitrogen in } (NH_4)_3PO_4 = \frac{42}{149} \times 100$$

$$= 28.19\%$$

(iv) Urea $(NH_2)_2$ CO

Urea is prepared by passing carbon(IV) oxide through ammonia solution.

$$CO_2(g) + 2NH_3(aq) \longrightarrow CO(NH_2)_2(aq) + H_2O(l)$$

$$\text{Percentage of nitrogen in urea} = \frac{28}{60} \times 100$$

$$= 46.67\%$$

(v) Calcium ammonium nitrate (CAN)

Calcium ammonium nitrate is a mixture of ammonium nitrate and calcium nitrate.
The percentage composition of nitrogen in calcium ammonium nitrate is about 27%.

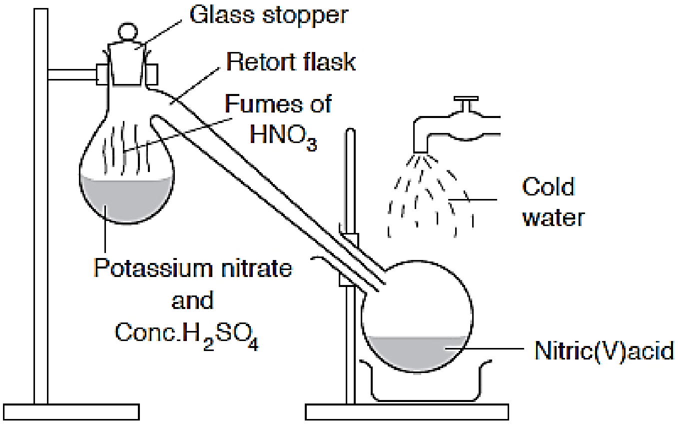
From the examples given, urea is a better nitrogenous fertiliser because it contains a higher content of nitrogen.

5. Nitric(V)Acid

Laboratory preparation of nitric(V) acid.

Nitric(V) acid is prepared by reacting concentrated sulphuric(VI) acid with a nitrate. Potassium nitrate is commonly used as it does not contain water of crystallisation.

When a mixture of potassium nitrate and concentrated sulphuric(IV) acid is heated, fumes of nitric(V) acid are produced. Nitric(V) acid is more volatile and is readily displaced from nitrates by the less volatile sulphuric(VI) acid.



$$KNO_3(s) + H_2SO_4(l) \longrightarrow KHSO_4(s) + HNO_3(aq)$$

- The nitric(V) acid fumes are **brown due to the presence of nitrogen(IV) oxide** formed by thermal decomposition of nitric(IV) acid as in the equation below:

$$4HNO_3(g) \rightleftharpoons 4NO_2(g) + O_2(g) + 2H_2O(g)$$

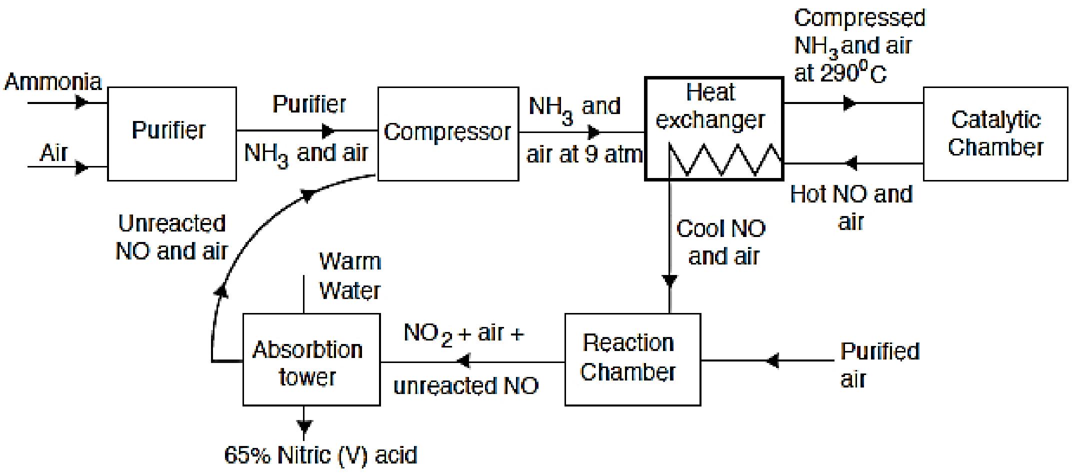
colourless *reddish-brown*

- The nitric(V) acid collected is **yellow in colour** due to the presence of dissolved nitrogen(IV) oxide. The yellow colour **can be removed by bubbling air through the acid.**
- The **acid salt, $KHSO_4$** is formed during the preparation of the nitric(V) acid **instead of the normal salt, K_2SO_4. The normal salt is not formed because it requires high temperatures** that cannot be achieved in a classroom situation.

Industrial Manufacture of Nitric(V) Acid (Ostwald process)

Nitric(V) acid is manufactured by a process that relies on the **catalytic oxidation of ammonia.**

The raw materials for this process are **ammonia, air and water**.



A mixture of ammonia and air is first **purified to remove dust particles and other impurities that would otherwise 'poison' the catalyst.**

The mixture is then compressed to a pressure of 9 atmospheres and passed on to the catalytic chamber via the heat exchanger.

As it passes through the heat exchanger, the temperature of the ammonia-air mixture rises to 230°C.

In the catalytic chamber, the temperature of the mixture is raised further to 900°C. It is then passed over a **platinum-rhodium catalyst.** The ammonia combines with oxygen to form nitrogen(II) oxide. The reaction is **highly exothermic**.

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

The hot gaseous products from the catalytic chamber are cooled in the heat exchanger. Nitrogen(II) oxide is mixed with more air and is oxidised to nitrogen(IV) oxide.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

This reaction is carried out at a **temperature of below 45°C to ensure a high yield of nitrogen(IV) oxide.**

The mixture of nitrogen(IV) oxide, air and unreacted nitrogen(II) oxide is passed up the absorption tower against a downward flow of hot water. Nitrogen(IV) oxide and water combine to form nitric(V) acid and nitric(III) acid (nitrous acid).

$$2NO_2\,(g) + H_2O\,(l) \longrightarrow \underset{\textit{Nitric (V) acid}}{2HNO_3(aq)} + \underset{\textit{Nitric (III) acid}}{2HNO_2(g)}$$

The nitric(III) acid formed is oxidised to nitric(V) acid by excess air.

$$\underset{\textit{Nitric (III) acid}}{2HNO_2(g)} + O_2\,(g) \longrightarrow \underset{\textit{Nitric (V) acid}}{2HNO_3(aq)}$$

The mixture that results is 65% nitric(V) acid and 35% water.

A **more concentrated acid may be obtained by careful distillation of the solution over phosphorus(V) oxide or concentrated sulphuric(VI) acid that acts as a dehydrating agent.**

Concentrated nitric(V) acid is a colourless fuming liquid with a boiling point of 83°C. Commercial nitric(V) acid is 68% pure and has a density of 1.42 g/cm^3.

Reactions of Dilute Nitric(V) Acid

Reaction with metals.

Dilute nitric(V) acid does not **produce hydrogen with some metals as expected. This is because any hydrogen produced is at once oxidised by the nitric(V) acid to water. The acid is reduced to one or more of the oxides of nitrogen.**

Very dilute nitric(V) acid has little oxidising power and magnesium will displace hydrogen from it.

$$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$$

With zinc, no hydrogen is produced as the hydrogen is immediately oxidised to water. The acid is reduced to nitrogen(II) oxide.

$$3Zn(s) + 8HNO_3(aq) \longrightarrow 3Zn(NO_3)_2(aq) + 4H_2O(l) + 2NO(g)$$

The nitrogen(II) oxide is colourless but on exposure to air it is oxidised to red-brown nitrogen(IV) oxide.

$$\underset{\text{colourless}}{2NO(g)} + O_2(g) \longrightarrow \underset{\text{red-brown}}{2NO_2(g)}$$

50% dilute nitric(V) acid **does not behave like a typical acid** because it **reacts with copper to produce copper(II) nitrate, nitrogen(II) oxide and water.**

$$3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + H_2O(l)$$

Reaction with carbonates

Dilute nitric(V) acid **behaves like a 'typical' acid by reacting with carbonates and hydrogen carbonates to produce a nitrate, carbon(IV) oxide and water.**

$$Na_2CO_3(s) + 2HNO_3(aq) \longrightarrow 2NaNO_3(aq) + CO_2(g)+ H_2O(l)$$

$$ZnCO_3(s) + 2HNO_3(aq) \longrightarrow Zn(NO3)_2(aq) + CO_2(g) + H_2O(l)$$

$$CuCO_3(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

$$NaHCO_3(s) + HNO_3(aq) \longrightarrow NaNO_3(aq) + CO_2(g) + 2H_2O(l)$$

Reaction with metal hydroxides and oxides

Metal hydroxides and oxides are **neutralised** by dilute nitric (V) acid to form a metallic nitrate and water only.

$$NaOH(aq) + HNO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$$

$$\underset{\text{White solid}}{CaO(s)} + 2HNO_3(aq) \longrightarrow \underset{\text{colourless solution}}{Ca(NO_3)(aq)} + H_2O(l)$$

$$\underset{\text{Black solid}}{CuO(s)} + 2HNO_3(aq) \longrightarrow \underset{\text{colourless solution}}{Cu(NO_3)(aq)} + H_2O(l)$$

$$\underset{\text{Yellow solid}}{PbO(s)} + 2HNO_3(aq) \longrightarrow \underset{\text{colourless solution}}{Pb(NO_3)(aq)} + H_2O(l)$$

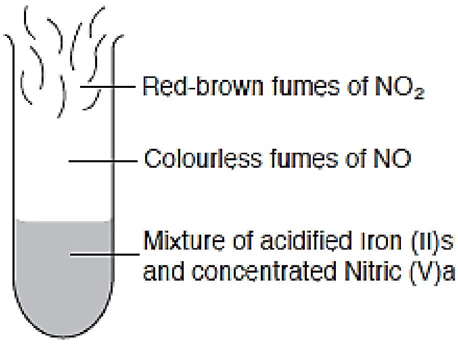
Reactions of Concentrated Nitric(V) Acid

Reaction with iron(II) sulphate, sulphur and copper.

Concentrated nitric(V) acid is a **powerful oxidising agent.** It oxidises iron(II) salts to iron(III) salts and is itself reduced to nitrogen(II) oxide. The pale-green iron(II) sulphate solution is oxidised to yellow iron(III) sulphate.

$$6FeSO_4(aq) + 3H_2SO_4(aq) + 2HNO_3(l) \longrightarrow 3Fe_2(SO_4)_3\,(aq) + 2NO(g) + 4\,H_2O(l)$$

Nitrogen(II) oxide is colourless. As it moves up the test-tube it is quickly oxidised by oxygen in the air to red-brown fumes of nitrogen(IV) oxide.



Hot concentrated nitric(V) acid oxidises sulphur to sulphuric(VI) acid and water while it is reduced to nitrogen(IV) oxide.

$$S(s) + 6HNO_3(aq) \longrightarrow H_2SO_4(aq) + 6NO_2(g)\ 2H_2O(l)$$

Copper is oxidised to copper(II) nitrate as the nitric(V) acid is reduced to water and nitrogen(IV) oxide.

$$3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 4H_2O(l) + 2NO_2(g)$$

Uses of Nitric(V) Acid

- Manufacture of fertilisers, e.g., ammonium nitrate.
- Manufacture of explosives, e.g., Trinitrotoluene.
- Manufacture of dyes and drugs.
- Purification of metals such as silver and gold.
- Etching designs on some metals.

Nitrates

Nitrates are salts that are derived from nitric(V) acid. They are usually prepared by the action of dilute nitric(V) acid on either a base or a carbonate. Some nitrates can be obtained by direct reaction of a metal with concentrated nitric(V) acid.

Action of Heat on Nitrates

Hydrated metallic nitrates produce **water vapour when gently heated. On strong heating, they all decompose.** Sodium nitrate and potassium nitrate melt and decompose on strong heating to produce oxygen gas and the nitrite of the metal.

$$\mathrm{2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)}$$

$$\mathrm{KNO_3(s) \longrightarrow KNO_2(s) + O_2(g)}$$

Nitrates of **metals below sodium** up to copper in the reactivity series decompose on heating to give the **metal oxide, nitrogen(IV) oxide and oxygen.** Thus:

$$\mathrm{2Zn(NO_3)_2\,(s)} \xrightarrow{\textit{heat}} \underset{\substack{\textit{yellow when hot}\\ \textit{white when cold}}}{\mathrm{2ZnO(s)}} \mathrm{+ 4NO_2(g) + O_2(g)}$$

$$\underset{\textit{white}}{\mathrm{2Pb(NO_3)_2\,(s)}} \xrightarrow{\textit{heat}} \underset{\textit{yellow when cold}}{\mathrm{2PbO(s)}} \mathrm{+ 4NO_2\,(g) + O_2(g)}$$

$$\underset{\textit{blue}}{\mathrm{2Cu(NO_3)_2\,(s)}} \xrightarrow{\textit{heat}} \underset{\textit{black}}{\mathrm{2CuO(s)}} \mathrm{+ 4NO_2\,(g) + O_2(g)}$$

Note: When lead(II) nitrate is heated a cracking sound is produced. The cracking sound is due to the air trapped inside the crystal which on expansion causes the crystals to burst producing the sound.

Nitrates of metals below copper decompose on heating to produce the metal, nitrogen(IV) oxide and oxygen. For example silver nitrate decomposes to produce silver, nitrogen(IV) oxide and oxygen gas.

$$\mathrm{2AgNO_3(s) \longrightarrow 2Ag(s) + 2NO_2(g) + O_2(g)}$$

The ease with which metal nitrates decompose **increases down the reactivity series of the metals.**

Ammonium nitrate when heated **decomposes to nitrogen(I) oxide and water.**

$$\mathrm{NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)}$$

Dry nitrogen(I) oxide relights a glowing splint. However, the steam **produced does not allow a glowing splint to relight.**

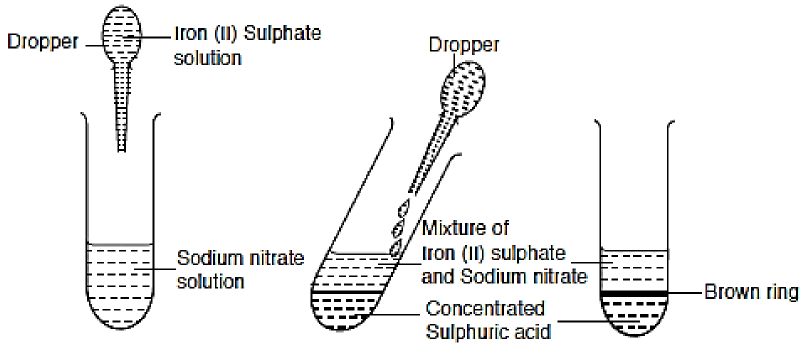
Test for Nitrates

The brown ring test for nitrates

Concentrated sulphuric acid is added to a mixture of sodium nitrate and iron (II) sulphate solution. A brown ring is formed at the junction of the two layers that are formed. This is the confirmatory test for nitrates. All nitrates give the same result.

Practically Speaking

Test:Place one spatulaful of sodium nitrate into a clean test-tube. Add 2 cm3 of distilled water and shake the mixture. Add 1 cm3 of freshly prepared iron (II) sulphate and shake the mixture again. Add concentrated sulphuric acid slowly and carefully along the wall of the test-tube.



Discussion

When concentrated sulphuric acid is added to the mixture of sodium nitrate and iron (II) sulphate solution, the **acid sinks to the bottom**. This is because the acid is **denser that the solution**. The reaction produces heat. **A brown ring is formed at the junction of the two layers. The brown ring is due to the formation of iron(II) sulphate–nitrogen (II) oxide complex, $FeSO_4.NO$.**

A safer method of testing for the presence of nitrate ions is by warming a mixture of concentrated sulphuric(VI) acid and the suspected nitrate in the presence of copper turnings.

Test: Place a spatulaful of sodium nitrate in a dry test-tube. Add a few drops of concentrated sulphuric acid. Warm the mixture and add a few pieces of copper turnings.

When sodium nitrate is mixed with concentrated sulphuric acid and the mixture warmed in the presence of copper turnings, brown fumes of nitrogen(IV) oxide are produced. This reaction takes place in two stages.

$$NaNO_3(s) + H_2SO_4(l) \longrightarrow NaHSO_4(s) + HNO_3(l)$$

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

Pollution Effects of Nitrogen Compounds in the Atmosphere

Oxides of nitrogen such as nitrogen(II) oxide and nitrogen(IV) oxide are atmospheric pollutants. Nitrogen(II) oxide is produced in internal combustion engines when nitrogen and oxygen from the air combine at high temperatures.

$$N_2(g) + O_2(g) \longrightarrow 2NO(g)$$

In the atmosphere, nitrogen(II) oxide is quickly oxidised to nitrogen(IV) oxide.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

The nitrogen(IV) oxide reacts with water in the atmosphere to produce nitric(III) and nitric(V) acids.

$$2NO_2(g) + O_2(g) \longrightarrow HNO_2(aq) + \quad HNO_3(aq)$$

(Nitric (III) acid) (Nitric (V) acid)

The nitric(III) acid is further oxidised to nitric(V) acid.

$$2HNO_2(aq) + O_2(g) \longrightarrow 2HNO_3(aq)$$

The nitric(V) acid eventually reaches the ground as acid rain. Acid rain causes the loss of chlorophyll (chlorosis) from plant leaves. This leads to the death of the affected plants.

Acid rain also corrodes stone buildings and metallic structures thereby weakening them and destroying their beauty. Rainwater containing dissolved nitric(V) acid leaches vital minerals from soils. These are converted into soluble nitrates and washed away from the top soil. This leads to poor crop yields.

Nitrogen(IV) oxide undergoes chemical reactions in the air to produce one of the major components of smog. Smog reduces visibility for motorists, irritates the eyes and causes breathing problems.

When nitrate ions from nitrogen containing fertilisers find their way into water bodies they encourage rapid growth of algae. This eventually leads to the reduction in the oxygen content in the water. As a result, aquatic animals such as fish die.

The presence of nitrate ions in drinking water may also cause ill health to humans. This is because in the body, the nitrate ions are converted into carcinogenic compounds.

Reducing Environmental Pollution by Nitrogen Compounds

There is an urgent need to reduce environmental pollution. Some of the measures that can be taken to reduce environmental pollution by nitrogen compounds include:

(a) Recycling unreacted gases in the manufacture of nitric acid to prevent release into the environment.
(b) Treating sewage and industrial effluents to remove nitrogen compounds before releasing them into rivers and lakes.
(c) Fitting the exhaust systems of vehicles with catalytic converters which convert nitrogen oxides into harmless nitrogen gas.
(d) Adding lime to lakes and soils in the surrounding regions to reduce acidity.
(e) Applying fertilisers at the right time and in the correct proportion to prevent them from being washed into water masses.

Review Exercises

1. 2006 Q17

The first step in the industrial manufacture of nitric acid is the catalytic oxidation of ammonia gas.

(a) What is the name of the catalyst used? (1 mark)
(b) Write the equation for the catalytic oxidation of ammonia (1 mark)
(c) Nitric acid is used to make ammonium nitrate, state two uses of ammonium nitrate. (1 mark)

2. 2007 Q6 P1

In an experiment, a few drops of concentrated nitric acid were added to aqueous iron(II) sulphate in a test-tube. Excess sodium hydroxide solution was then added to the mixture.

(a) State the observations that were made when:
  (i) Concentrated nitric acid was added to aqueous iron (II) sulphate (1mark)
  (ii) Excess sodium hydroxide was added to the mixture. (1mark)
(b) Write an ionic equation for the reaction which occurred in (a) (ii) above. (1mark)

3. 2007 Q6a (p2)

The elements nitrogen, phosphorous and potassium are essential for plant growth.

(i) Potassium in fertilizers may be in the form of potassium nitrate. Describe how a sample of a fertilizer may be tested to find out if it contained nitrate ions. (2 marks)

(ii) Calculate the mass of nitrogen present if a 25kg bag contained pure ammonium phosphate, $(NH_4)_2HPO_4$. (N = 14.0, H=1.0, P = 31.0, O = 16.0) (2 marks)

4. 2008 Q10 P1

When magnesium was burnt in air, a solid mixture was formed. On addition of water to the mixture a gas which turned moist red litmus paper blue was evolved. Explain these observations.

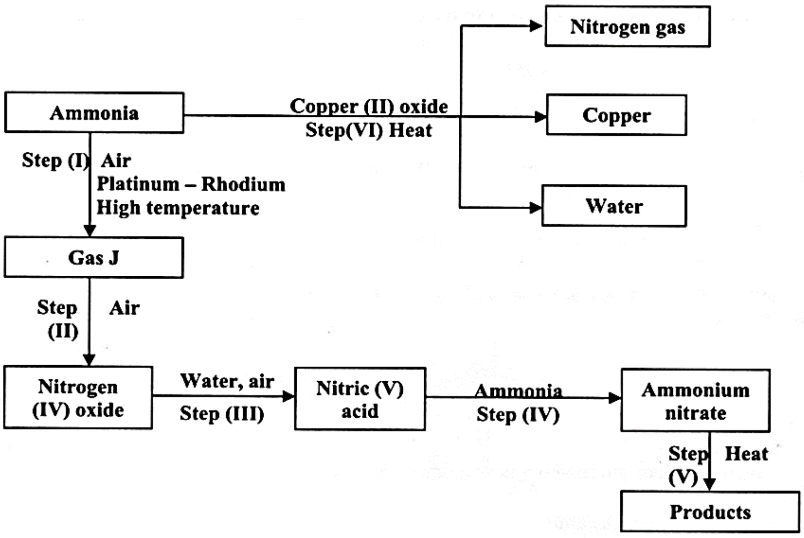
(2 marks)

5. 2008 Q3 (P2), 2016 P2

(a) Describe the process by which Nitrogen is obtained from air on a large scale.

(4 marks)

(b) Study the flow chart below and answer the questions that follow.



(i) Identify gas J. (1 mark)

(ii) Using oxidation numbers, show that ammonia is the reducing agent in step (VI)

(2 marks)

(iii) Write the equation for the reaction that occurs in step (V). (1 mark)

(iv) Give one use of ammonium nitrate. (1 mark)

(c) The table below shows the observations made when aqueous ammonia was added to cations of elements E, F and G until in excess.

Cation of	**Addition of a few drops of Aqueous ammonia.**	**Addition of excess aqueous ammonia.**
E	White precipitate	Insoluble
F	No precipitate	No precipitate
G	White precipitate	Dissolves

(i) Select the cation that is likely to be Zn^{2+}. (1 mark)

(ii) Given that the formula of the cation of element E is E^{2+}, write the ionic equation for the reaction between E^{2+} (aq) and aqueous ammonia. (1 mark)

6. 2009 Q13

When 8.53g of sodium nitrate were heated in an open tube, the mass of oxygen produced was 0.83g. Given the equation of the reaction as

$$2NaNO_3(s \longrightarrow 2NaNO_2(s) + O_2(g)$$

Calculate the percentage of sodium nitrate that was converted to sodium nitrite

(Na= 23.0, N = 14.0, O = 16.0) (3 marks)

7. 2010 Q16

A sample of fertilizer is suspected to be calcium ammonium nitrate. Describe chemical tests for each of the following ions in the sample:

(a) Calcium ions; (2 marks)

(b) Ammonium ions. (1 mark)

8. 2010 Q20

In an experiment to prepare nitrogen (I) oxide, ammonium nitrate was gently heated in a flask.

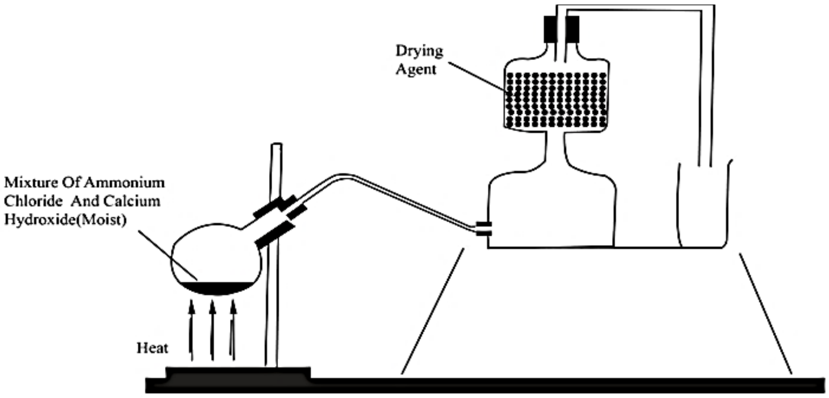
(a) Write the equation for the reaction that took place in the flask. (1 mark)

(b) State and explain how the gas was collected. (1 mark)

(c) A sample of the gas was tested with damp blue and red litmus papers. What observations were made? (1 mark)

9. 2010 Q5 (P2)

(a) A student set up the apparatus as shown in the diagram below to prepare and collect dry ammonia gas.



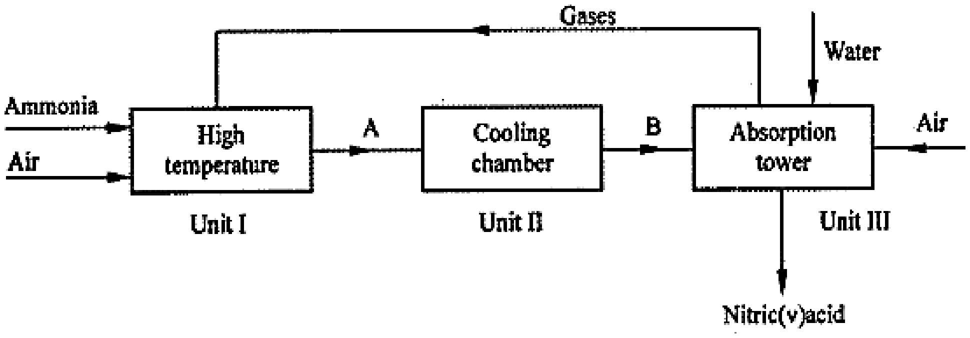
(i) Identify two mistakes in the set up and give a reason for each mistake. (3 marks)

(ii) Name a suitable drying agent for ammonia (1 mark)

(iii) Write an equation for the reaction that occurred when a mixture of ammonium chloride and calcium hydroxide was heated. (1 mark)

(iv) Describe one chemical test for ammonia gas. (1 mark)

(b) Ammonia gas is used to manufacture nitric (V) acid, as shown below.



(i) This process requires the use of a catalyst. In which unit is the catalyst used? (1 mark)

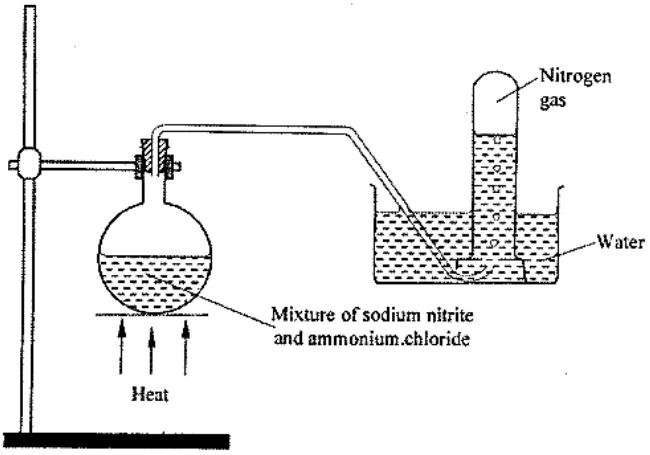
(ii) Identify compound **A** and **B**. (1 mark)

(iii) Using oxidation number, explain why the conversion of ammonia to nitric(V) acid is called catalytic oxidation of ammonia. (2 marks)

(c) Ammonia and nitric (V) acid are used in the manufacture of ammonium nitrate fertilizer. Calculate the amount of nitric (V) acid required to manufacture 1000 kg of ammonium nitrate using excess ammonia. (3 marks)

10. 2011 Q4 P1

The set-up below shows how nitrogen gas is prepared in the laboratory.



(a) Describe how nitrogen gas is formed on the flask. (2 marks)

(b) Nitrogen is inert. State one use of the gas based on this property. (1 mark)

11. 2011 Q 19 P1

50g of ammonium sulphate, $(NH_4)_2SO_4$, and 30 kg of urea, $CO(NH_2)_2$, fertilizers were applied in two equal sizes of plots A and B to enrich their nitrogen content.

Show, by working, which plot was more enriched with nitrogen.

(N = 14; S = 32; O = 16; C 12; H = 1) (3 marks)

12. 2011 Q4 P2

(a) When excess calcium metal was added to 50 cm^3 of 2M aqueous copper (II) nitrate in a beaker, a brown solid and bubbles of gas were observed.

(i) Write two equations for the reactions which occurred in the beaker. (2 marks)

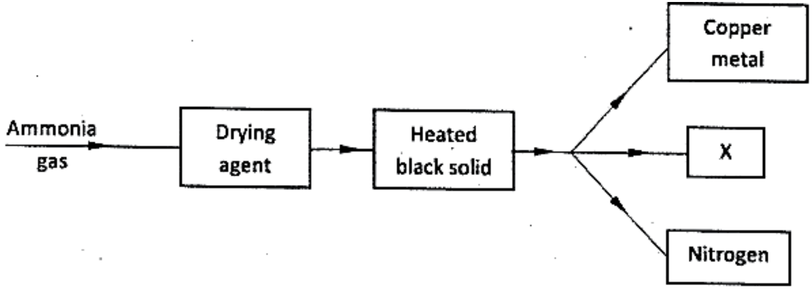
(ii) Explain why it is not advisable to use sodium metal for this reaction. (2 marks)

(iii) Calculate the mass of calcium metal which reacted with copper (II)nitrate solution. (Relative atomic mass of Ca = 40) (2 marks)

(b) The resulting mixture in (a) above was filtered and aqueous sodium hydroxide added to the filtrate dropwise until in excess. What observations were made? (1 mark)

13. 2012 Q12 P1

Study the flow chart below and answer the questions that follow.



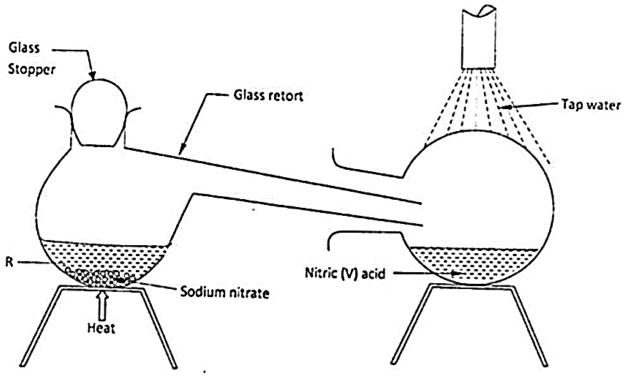
(i) Name a suitable drying agent for ammonia. (1 mark)

(ii) Describe one chemical test for ammonia (1 mark)

(iii) Name X (1 mark)

14. 2012 Q3 P2

(a) In the laboratory, small quantities of nitric (V) acid can be generated using the following set up. Study it and answer the questions that follow.



(i) Give the name of substance R. (1 mark)

(ii) Name one other substance that can be used in place of sodium nitrate. (1 mark)

(iii) What is the purpose of using tap water in the set up above? (1 mark)

(b) Explain the following

(i) It is **not** advisable to use a stopper made of rubber in the set-up. (1 mark)

(ii) The reaction between copper metal with 50% nitric (V) acid in an open test-tube produces brown fumes. (1 mark)

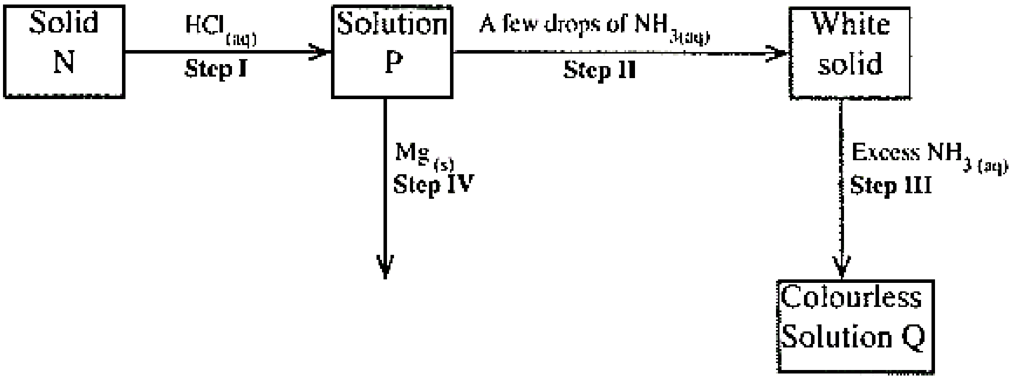
(c) (i) Nitrogen is one of the reactants used in the production of ammonia, name two sources of the other reactant. (2 marks)

(i) A factory uses nitric (V) acid and ammonia gas in the preparation of a fertilizer. If the daily production of the fertilizer is 4800kg; calculate the mass of ammonia gas used in kg. (N = 14.0; O = 16.0; H=1.0) (3 marks)

(ii) State two other uses of nitric (V) acid other than the production of fertilizers. (2 marks)

15. 2013 Q13 P1

The scheme below shoes some reaction sequence starting with slid N. Study it and answer the questions that follows.



(a) Write the formula of the complex ion in solution Q. (1 mark)

(b) Write an equation for the reaction in step IV. (1 mark)

16. 2013 Q17 P1 ,2016 Q14 P1

When fuels burn in the internal combustion engine at high temperature, one of the products formed is nitrogen (II) oxide.

(a) Write the equation for the formation of nitrogen (II) oxide. (1 mark)

(b) Give a reason why nitrogen (II) oxide is not formed at room temperature. (1 mark)

(c) Describe how formation of nitrogen (II) oxide in the internal combustion engine leads to gaseous pollution. (1 mark)

17. 2013 Q21 P1

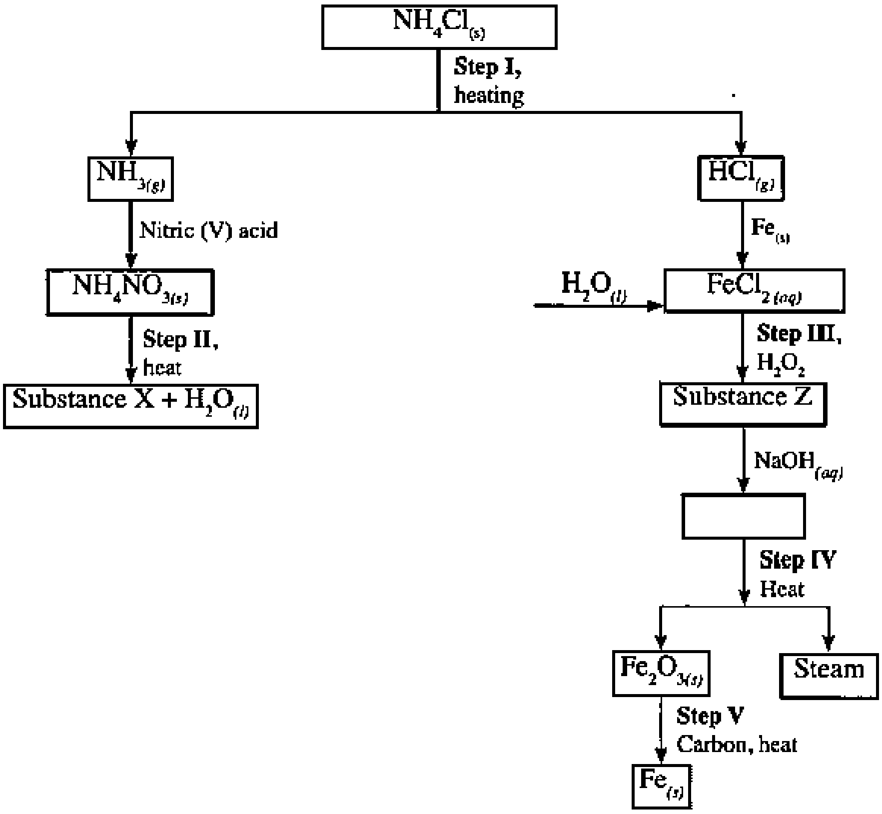
On heating a pale green solid **K**, carbon (IV) oxide gas and a black solid **M** were formed.
On reacting **K** with dilute hydrochloric acid, carbon (IV) oxide gas and green solution **S** were formed.
When excess aqueous ammonia was added to solution **S**, a deep blue solution was formed.

(a) Identify the cation in solid **K.** (1 mark)

(b) Identify the two onions in solution **S**. (2 marks)

18. 2013 Q5(c) P2

Study the flow chart below and answer the questions that follow:



(i) Write an equation for the reaction in:

I. step II; (1 mark)

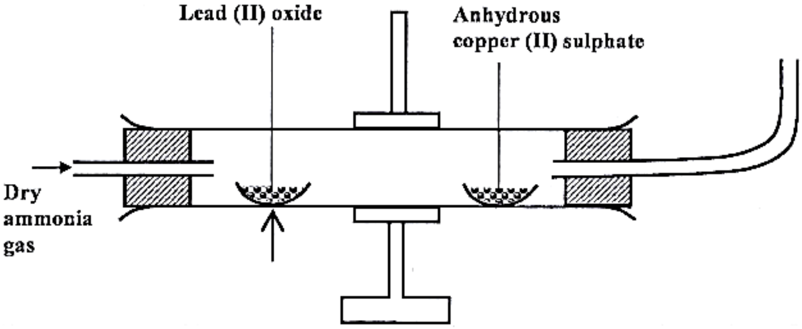
II. step IV. (1 mark)

(ii) State the observation made in step III. Explain. (2 marks)

(iii) Name another substance that can be used in step V. (1 mark)

19. 2014 Q3 P1

Dry ammonia gas was passed over heated lead (II) oxide and the product passed over anhydrous Copper (II) sulphate as shown in the diagram below.



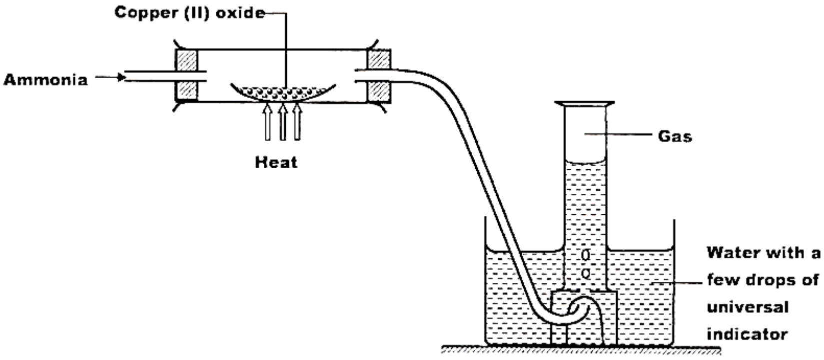
State:

(a) Two observations made in the combustion tube. (2 marks)

(b) The property of ammonia gas shown in this experiment. (1 mark)

20. 2014 Q15 P1

Study the set up below and answer the questions that follow

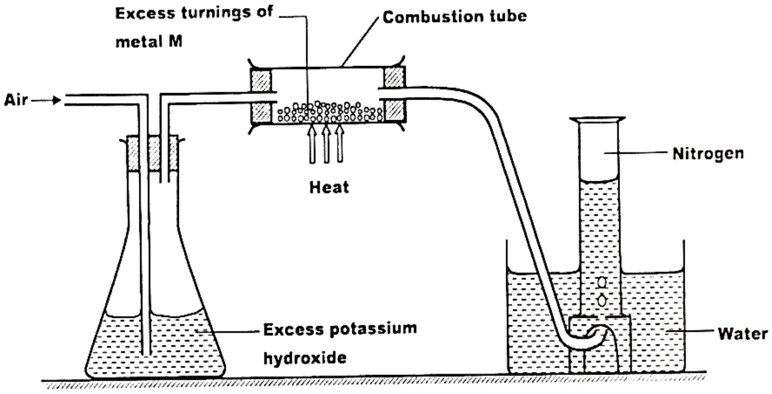


(a) Write an equation for the reaction between ammonia and copper (II) oxide. (1 mark)

(b) During the experiment, the colour of the contents in the water trough changed. State the colour change observed and give an explanation. (2 marks)

21. 2014 Q28 P1

A student used the set up below to prepare a sample of nitrogen gas.



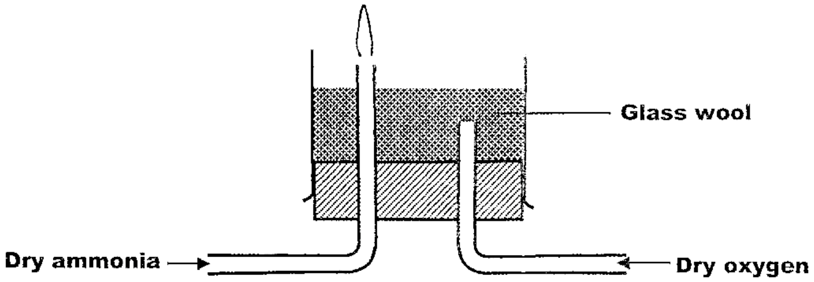
(a) State the function of potassium hydroxide in the set up. (1 mark)

(b) Give a suitable metal M for use in the combustion tube. (1 mark)

(c) Give a reason why the nitrogen gas obtained is not pure. (1 mark)

22. 2015 Q13 P1

Dry ammonia and dry oxygen were reacted as shown in the diagram below



(a) What is the purpose of the glass wool? (1 mark)

(b) What products would be formed if red hot platinum was introduced into a mixture of ammonia and oxygen? (1 mark)

23. 2017 P1 Q12.

Potassium nitrate liberates oxygen gas when heated. Draw a diagram of a set-up that shows heating of potassium nitrate and collection of oxygen gas. (3 marks)

24. 2017 P1 Q16.

In an experiment, concentrated nitric(V) acid was reacted with iron(II) sulphate. State and explain the observations made. (2 marks

25. 2017 P1 Q17.

The flow chart in **Figure 3** shows the process of obtaining a sample of nitrogen gas. Study it and answer the questions that follow.

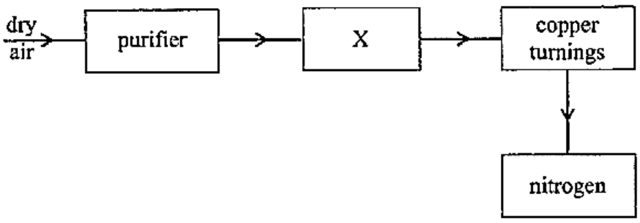


Figure 3

(a) Identify **X** (1 mark)

(b) Write an equation for the reaction with heated copper turnings. (1 mark)

(c) Name an impurity in the sample of nitrogen gas. (1 mark)

26. 2017 P1 Q18.

The set-up in **Figure 4** can be used to prepare nitrogen(II) oxide. Use it to answer the questions that follow.

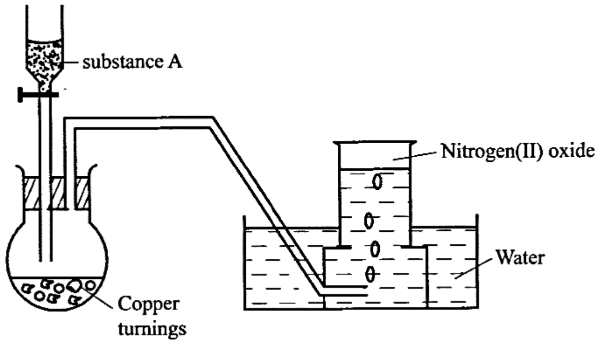


Figure 4

(a) Name substance **A** (1 mark)

(b) When the gas jar containing nitrogen (II) oxide is exposed to air, a brown colour is observed. Explain. (1 mark)

(c) Write an equation for the reaction which occurred in the flask. (1 mark)

27. 2018 P1 Q 8.

Study the flow chart in Figure 1 and answer the questions that follow.

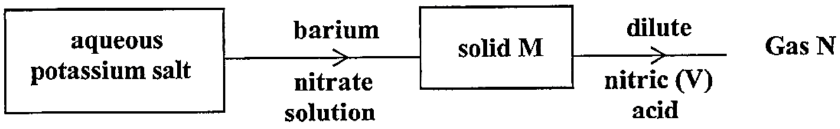


Figure 1

Gas N forms a while suspension with aqueous calcium hydroxide.

(a) Name the anion present in the potassium salt. (1 mark)

(b) Write an ionic equation for the formation of solid M. (1 mark)

(c) Give one use of gas N. (1 mark)

28. 2018 P1 Q27.

$(NH_4)_2HPO_4$ is a fertiliser used by farmers to boost their crop production.

(a) Calculate the mass of phosphorus in a 20 kg packet of $(NH_4)_2HPO_4$. (N = 14.0; H = 1.0; P = 31.0; O = 16.0) (2 marks)

(b) State one advantage of this fertilizer, $(NH_4)_2HPO_4$, over urea ($CO(NH_2)_2$) (1 mark)

29. 2019 P1 Q27.

When burning magnesium ribbon is introduced into a gas jar full of nitrogen, it continues to burn producing a greenish yellow powder.

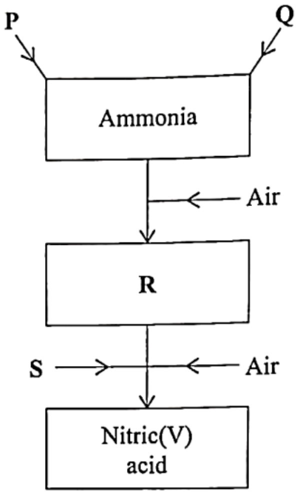
(a) Write an equation for the reaction between nitrogen and magnesium. (1 mark)

(b) Explain why magnesium continues to bum in nitrogen but sulphur does not. (2 marks)

(c) State one use of nitrogen. (1 mark)

30. 2019 P2 Q3.

Figure 3 is a flow chart that shows the process that occurs in the manufacture of nitric(v)acid.



(a) Name substance **P**, **Q**, **R** and **S**.

P.......................... (1 marks)
Q............................. (1 marks)
R............................... (1 marks)
S................................ (1 marks)

(b) To obtain substance R, ammonia is heated at 900 °C in the presence of air and a catalyst. The product is then cooled in air.

(i) Name the catalyst for the reaction. (1 marks)

(ii) Write the equations for the two reactions described in (b). (2 marks)

(iii) Other than nitric(V) acid, name another product that is formed. (1 mark)

(c) When ammonia is reacted with nitric(V) acid, it produces a nitrogenous fertiliser.

(i) Explain why fertilisers play a major role in food production.

(2 marks)

(ii) State two problems associated with the use of nitrogenous fertilisers.

(2 marks)

TOPIC FIVE
SULPHUR AND ITS COMPOUNDS

Table of Contents

Organizer

Objectives

By the end of the Chapter, the learner should be able to:

(a) Name the sources of sulphur.

(b) Describe the extraction of sulphur.

(c) Describe the preparation of the allotropes of sulphur.

(d) State the properties and uses of sulphur.

(e) Name and describe the preparation of the oxides of sulphur.

(f) State the properties and uses of the oxides of sulphur.

(g) Describe the contact process for the manufacture of sulphuric acid.

(h) Describe the properties and state the uses of sulphuric acid.

(i) Describe the preparation and state the properties of hydrogen sulphide.

(j) Explain pollution effects of sulphur containing compounds.

SULPHUR AND ITS COMPOUNDS

Sulphur is the second member of group VI elements. It is placed just below oxygen in the periodic table. It has an atomic number 16 hence its electron arrangement is 2.8.6.

Sulphur occurs naturally as an element in deposits in places such as Texas and Louisiana in U.S.A, Sicily in Italy, and various places in Japan.

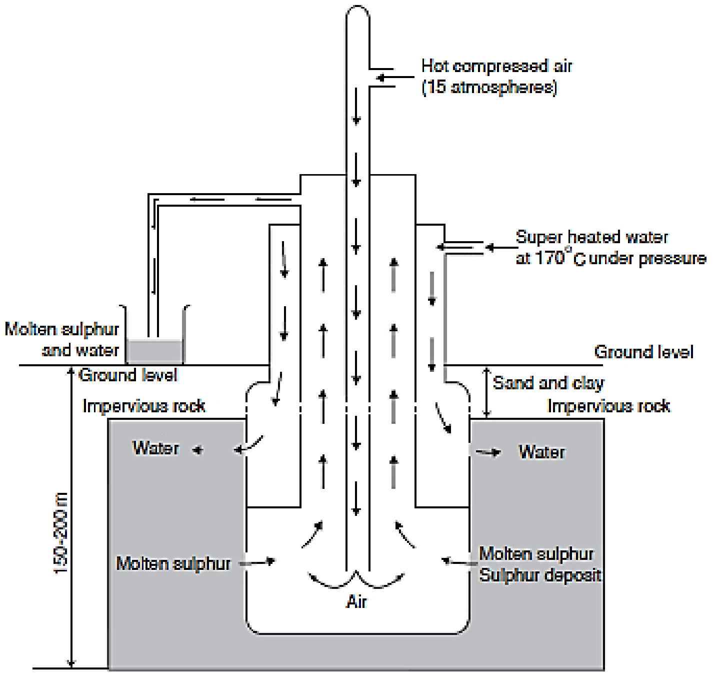
Sulphur also occurs in combination with other elements as sulphides and sulphates. The sulphide ores include copper pyrites ($CuFeS_2$) and iron pyrites (FeS_2).

The sulphate ores include gypsum ($CaSO_4.2H_2O$)and anhydrite ($CaSO_4$). Petroleum gas and coal mines contain sulphur in the form of hydrogen sulphide gas.

Sulphur and its allotropes

Extraction Of Sulphur: The Frasch process

The Frasch process is employed in the extraction of Sulphur. The process is based on the low melting point of sulphur which ranges between 113 °C and 119°C. In the Frasch process, three concentric pipes of different diameters 2 cm, 8 cm, and 15 cm are sunk into the sulphur deposits underground.



Water is heated to about **170°C** under pressure of about **10 atmospheres** and is forced down the outermost pipe. **The pressure ensures that the water remains in liquid state at such high temperature.**

Hot air at a pressure of 15 atmospheres is forced down the innermost pipe. **This produces a light froth consisting of a mixture of molten sulphur and water. The high pressure forces the mixture up the middle pipe**.

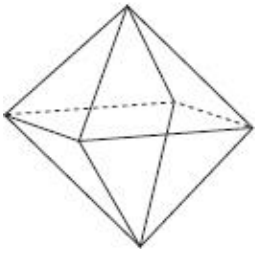
The mixture is run into large tanks on the surface where the sulphur solidifies at 115°C, Separates from the water and is stored. Sulphur obtained this way in usually over 99% pure.

Allotropes of Sulphur

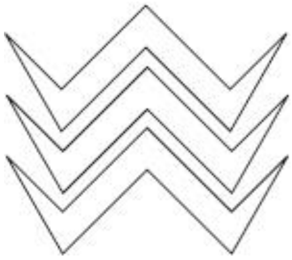
Sulphur exhibits allotropy. The main allotropes are rhombic and monoclinic sulphur

Rhombic Sulphur

Rhombic sulphur is also referred to as **octahedral or α-sulphur**. It is a bright yellow crystalline solid with an **octahedral shape.**



(a) A crystal of rhombic sulphur.



(b) Packing of rhombic sulphur molecules in a crystal

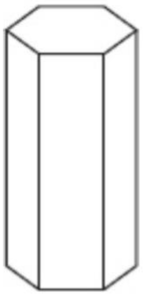
How to prepare Rhombic sulphur:

Place two spatulafuls of powdered sulphur in a boiling tube containing 10 cm^3 of carbon(IV) sulphide. Stir and filter the contents of the tube into a dry beaker using a dry filter paper. Allow the filtrate to evaporate slowly. Use a hand lens to examine the crystals formed.

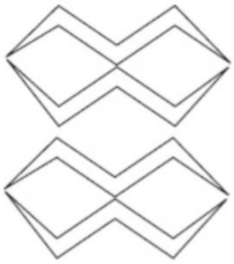
Rhombic sulphur melts at 113°C and has a density of 2.06 g/cm^3. It is the stable allotrope below 96°C. Above this temperature, it slowly changes into monoclinic sulphur.

Monoclinic Sulphur

Monoclinic sulphur is also referred to as prismatic or *ß* - sulphur. It is a pale yellow crystalline solid. The crystals appear needle-like when observed using a hand lens. The actual shape of the crystal is a hexagonal prism.



(a) A crystal of monoclinic Sulphur.

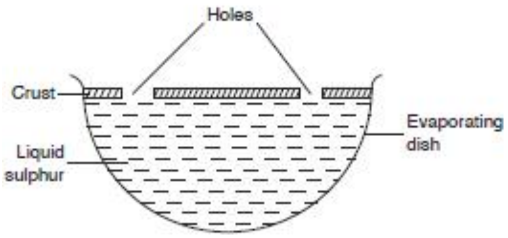


(b) Packing of monoclinic Sulphur molecules in a crystal

How to prepare monoclinic sulphur

Place some powdered sulphur in an evaporating dish and heat gently until the sulphur melts. Using a glass rod, stir and add more sulphur a little at a time until the dish is almost full of molten sulphur.

Remove the source of heat and allow the sulphur to cool and form a crust on the surface. Using a thin glass rod, carefully pierce two holes on widely separated points in the crust as shown below .



Immediately pour out the molten sulphur. Remove the crust by cutting round the edge of the dish with a knife. Use a hand lens to observe the crystals that have formed on the underside of the crust.

Monoclinic sulphur has a **melting point of 119°C and a density of 1.98 g/cm^3. Below 96°C monoclinic sulphur gradually changes to rhombic sulphur. This temperature of 96°C is the transition temperature for rhombic and monoclinic sulphur.**

The temperature at which one allotrope of an element changes to another is called the transition temperature.

Non-crystalline – (amorphous) forms of Sulphur

Include plastic, colloidal, and powdery sulphur.

Plastic sulphur is prepared by heating powdered sulphur until it boils. The boiling sulphur is then poured in a thin continuous stream into a beaker containing cold water. Long elastic yellow threads of plastic sulphur are formed.

This form of sulphur is insoluble in carbon(IV) sulphide. It turns into hard rhombic sulphur if left for a few days.

When dilute hydrochloric acid is added to a test-tube containing a dilute solution of sodium thiosulphate, $Na_2S_2O_3$, a yellow precipitate of **colloidal sulphur** is deposited.

$$N_2S_2O_3,(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$$

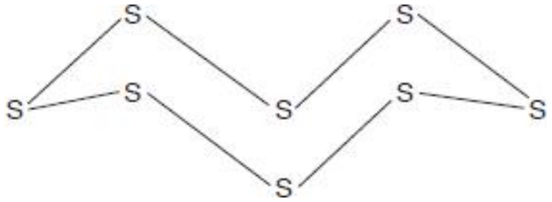
Powdery sulphur may be prepared by saturating distilled water with hydrogen sulphide. The solution is then exposed to air. A white powder is deposited.

$$H_2S(g) + \text{Water} \longrightarrow H_2S(aq)$$

$$2H_2S(aq) + O_2(g) \longrightarrow 2H_2O(l) + 2S(s)$$

Physical Properties of Sulphur

Sulphur is a yellow non-metallic element. A molecule of sulphur consists of a puckered ring of eight atoms of sulphur joined together by strong covalent bonds as shown below. .

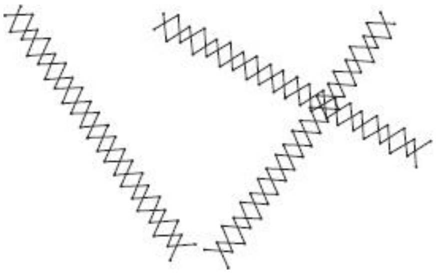


Sulphur is soluble in organic solvents such as **benzene, methyl benzene, and carbon(IV) sulphide. It does not dissolve in water.**

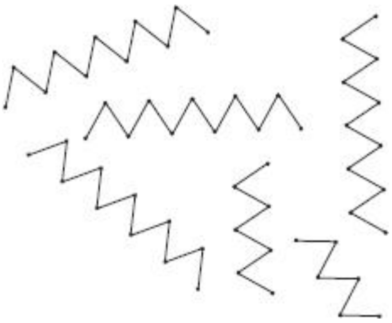
Effect of heat on sulphur.

When the yellow powder is gently heated, it melts at 113°C to a clear amber liquid. At this temperature molten sulphur has a low viscosity, and flows easily. This liquid is made up of rings of sulphur molecules consisting of eight atoms of sulphur, S_8.

On further heating the liquid gradually darkens. At **160°C, it becomes reddish-brown, and very viscous** such that the test-tube may be inverted without the liquid sulphur pouring out. These changes are **due to the breaking of the S_8 rings which join to form long chains with over 100,000 atoms of sulphur.** As the chains **entangle with one another**, the viscosity of the liquid increases.



Long chains of sulphur atoms.



Shorter chains of sulphur atoms

Above 160°C, the liquid **darkens further and becomes almost black**. Near the boiling point the liquid becomes **more mobile.** This is due to the breaking of the long chains to shorter chains.

The liquid boils at 444°C and forms a reddish-brown vapour consisting of S_8, S_6 and S_2 molecules which cools to form a yellow sublimate. This sublimate is known as **"flowers of sulphur"** and consists mainly of S_8 rings.

Chemical Properties of Sulphur

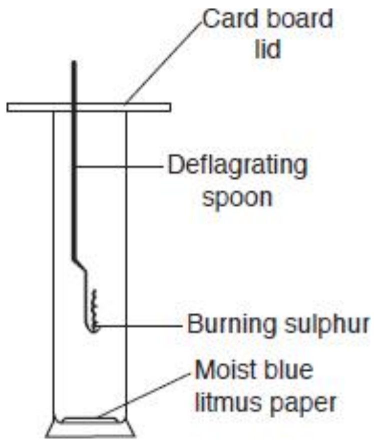
Reaction of sulphur with oxygen, iron powder and copper

Sulphur burns in oxygen with a **bright blue flame** forming misty fumes with a choking smell.

The gas is mainly sulphur(VI) oxide (sulphur dioxide) SO_2, with traces of sulphur(IV) oxide (sulphur trioxide, SO_3).

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

$$2S(s) + 3O_2(g) \longrightarrow 2SO_3(g)$$



The **moist blue litmus paper turns re**d indicating that the oxides produced when sulphur burns are acidic.

Sulphur combines directly with some elements to form sulphides.

- When a mixture of sulphur, and iron powder is heated, a **highly exothermic reaction occurs.** Once the reaction has started at one point, the glow spreads through the mixture without further heating, forming iron(II) sulphide. This is because the **heat produced during the reaction is high enough to sustain the reaction.**

$$Fe(s) + S(s) \longrightarrow \underset{\textit{Black}}{FeS(s)}$$

- Hot powdered copper similarly combines with heated sulphur forming copper(I) sulphide.

$$2Cu(s) + S(s) \longrightarrow \underset{\textit{Black}}{Cu_2S(s)}$$

- Sulphur also combines with some non-metals such as carbon and hydrogen forming non-metallic sulphides.

  Carbon combines with sulphur at high temperatures to form carbon(IV) sulphide. Hydrogen combines with sulphur to form hydrogen sulphide.

$$C(s) + 2S(s) \longrightarrow CS_2(l)$$

$$H_2(g) + S(s) \longrightarrow H_2S(g)$$

Reaction of sulphur with concentrated acids

Sulphur is easily oxidised by both concentrated nitric(V) and sulphuric(VI) acids.

- When warmed with concentrated nitric(V) acid, sulphur is oxidised to sulphuric(VI) acid. The **sulphate (SO_4^{2-}) ion in the acid forms a white precipitate with barium ions in the solution. The nitric(V) acid itself is reduced to red brown nitrogen(IV) oxide and water.**

$$S(s) + 6HNO_3(aq) \longrightarrow H_2SO_4(aq) + 6NO_2(g) + 2H_2O(l)$$

- Concentrated sulphuric(VI) acid **oxidises sulphur to sulphur(IV) oxide** while the acid is **reduced to sulphur(IV) oxide, and water.**

$$S(s) + 2H_2SO,(l) \longrightarrow 3SO_2(g) + 2H_2O(l)$$

Concentrated hydrochloric acid does not react with sulphur because it is not an oxidising agent.

Uses of Sulphur

- Manufacture of sulphuric(VI) acid.
- As a fungicide.
- In the manufacture of bleaching agent used to bleach wood pulp in the paper industry.
- Vulcanisation (hardening) of rubber.
- Used in smaller quantities in the manufacture of dyes, and fireworks..

The Compounds of Sulphur

Sulphur forms several compounds. The common compounds include the oxides, sulphuric acid and hydrogen sulphide

Oxides of Sulphur

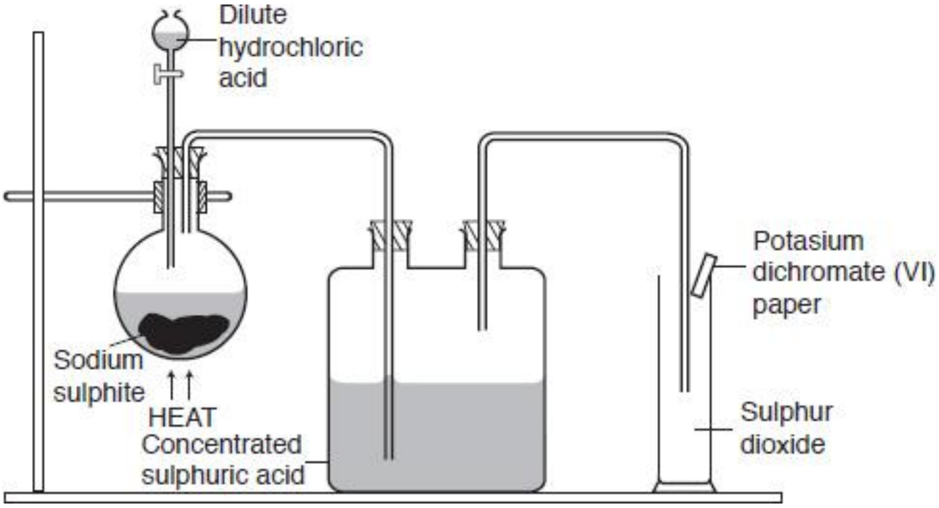
Sulphur forms two oxides namely sulphur(IV) oxide, SO_2, and sulphur(VI) oxide SO_3.

Sulphur(IV) Oxide, SO_2

Laboratory preparation of sulphur (IV) oxide.

Sulphur(IV) oxide is prepared in the laboratory by the action of dilute hydrochloric acid on a suitable sulphite such as sodium sulphite.

$$Na_2SO_3(s) + 2HCl(aq) \longrightarrow SO_2(g) + 2NaCl(aq) + H_2O(l)$$



The gas may also be prepared by the action of concentrated sulphuric(VI) acid on copper turnings. The reaction should be carried out in a fume cupboard.

$$Cu(s) + 2H_2SO_4(aq) \longrightarrow SO_2(g) + 2H_2O(l) + CuSO_4(aq)$$

Sulphur(IV) oxide is dried by passing it through concentrated sulphuric(IV) acid and collected by downward delivery. The gas jar is confirmed to be full of the gas when the paper soaked in orange potassium chromate(VI) turns green.

Physical Properties of Sulphur(IV) Oxide

Sulphur(IV) oxide is a colourless poisonous gas with a characteristic irritating, and choking smell.

It has a boiling point of –10°C and is readily liquefied under pressure.

Discussion Questions

1. **Give a reason why sulphur(IV) oxide is collected by downward delivery.**

   It is denser than air, hence can be collected by downward delivery.

2. **Explain the observation made when a test-tube full of sulphur(IV) oxide is inverted in a trough of water.**

   When a test-tube full of the gas is inverted in a trough of water, the water level rises rapidly inside the test-tube. This shows that the gas is soluble in water.

3. **Explain the observation made when:**

   (a) A damp litmus paper was dropped into a test-tube containing sulphur(IV) oxide gas.

   When a moist blue litmus paper is dropped into a test-tube containing sulphur(IV) oxide, it turns red showing that the gas is acidic. The gas has no effect on dry litmus. The solution of the gas in water is sulphuric(IV) acid. This is a weak dibasic acid.

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

The acid is responsible for the change in colour of the moist litmus.

(a) Sodium hydroxide solution is added to sulphur(IV) oxide.

When sodium hydroxide solution is added to sulphur(IV) oxide gas, neutralisation occurs. The sulphuric(IV) acid formed reacts with sodium hydroxide forming a salt and water. During the reaction between the acid and sodium hydroxide, two types of salts are formed, a normal and an acid salt.

$$2NaOH(aq) + H_2SO_3(aq) \longrightarrow \underset{\text{Normal salt}}{Na_2SO_3} + H_2O(l)$$

$$NaOH(aq) + H_2SO_3(aq) \longrightarrow \underset{\text{Acid salt}}{NaHSO_3(aq)} + H_2O(l)$$

Chemical Properties of Sulphur (IV) Oxide

Bleaching Action of Sulphur(IV) Oxide

Sulphur(IV) Oxide is a bleaching agent. It bleaches by reduction. This property is applied in paper industries to bleach paper.

Discussion Questions

1. **State and explain what is observed when coloured flowers are dropped in a gas jar containing Sulphur (IV) oxide.**

   When coloured flowers are dropped into a gas jar containing sulphur(VI) oxide, the flowers are bleached. Sulphur(IV) oxide combines first with the moisture, forming sulphuric(IV) acid. The sulphuric(IV) acid then combines with oxygen from the dye to form sulphuric(VI) acid. When the dye loses oxygen it becomes colourless. In this reaction, the dye undergoes reduction while the sulphuric(IV) acid is oxidised.

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

$$H_2SO_3(aq) + \text{Dye} \longrightarrow H_2SO_4(aq) + \text{Colourless material.}$$

2. **Give a reason why newsprint paper turns brown after some time.**

   During the manufacture of paper, reducing agents such as sulphuric(IV) acid are used to bleach the materials. When such paper is exposed to the atmosphere in the presence of sunlight, the oxygen removed during bleaching is restored. This explains why newsprint paper turns brown after sometime.

Reducing Action of Sulphur(IV) Oxide

Sulphur(IV) oxide is a strong reducing agent. The reducing property is only displayed when the gas is in aqueous state.

Discussion Questions

1. **Explain the observations made when sulphur(IV) oxide gas is reacted with:**

   (i) Acidified potassium chromate(VI) solution.

Acidified potassium dichromate(VI) turns from orange to green when reacted with sulphur(IV) oxide. The chromium(VI) ion in the dichromate ($Cr_2O_7^{2-}$) is reduced to chromium(III) ion.

$$3SO_2(g) + Cr_2O_7^{2-}(aq) + 2H^+(aq) \longrightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + H_2O(aq)$$

orange *green*

This is a test for Sulphur (IV) Oxide

(ii) Acidified potassium manganate(VII) solution

The colour of acidified potassium manganate(VII) turns from purple to colourless when reacted with sulphur(IV) oxide. This is because the manganese(VII) ion in the manganate ion (MnO^-_4) is reduced to manganese(II) ion.

$$5SO_2(g) + 2MnO^-_4(aq) + 2H_2O(l) + H^+(aq) \longrightarrow 5SO^{2-}_4(aq) + 2Mn^{2+}(aq) + 5H^+(aq)$$

Purple *Colourless*

(iii) Acidified bromine water followed by a few drops of barium chloride solution.

Red brown acidified bromine water is decolourised when reacted with sulphur(IV) oxide. The bromine water is reduced to hydrobromic acid as the sulphur(IV) oxide is oxidised to sulphate.

$$Br_2(aq) + 2H_2O(l) + SO_2(g) + 2H^+(aq) \longrightarrow 2HBr(aq) + SO^{2-}_4(aq) + 4H+(aq)$$

2. **Explain the observations made when:**

   (i) Concentrated nitric(V) acid is added to a test-tube full of sulphur(IV) oxide, followed by barium chloride solution.

When concentrated nitric(V) acid is added to a test-tube containing sulphur(IV) oxide gas brown fumes of nitrogen(IV) oxide are given off.

$$2HNO_3(aq) + SO_2(g) \longrightarrow 2NO_2(g) + H_2SO_4(aq)$$

The nitric(V) acid is reduced to nitrogen(IV) oxide while the sulphuric(IV) acid is oxidised to sulphuric(VI) acid. When a solution of barium chloride is added to the mixture, a white precipitate of barium sulphate is formed indicating the presence of sulphate, SO_4^{2-} ions.

(ii) Sulphur(IV) oxide is reacted with a hot solution of iron(II) chloride.

When warm iron(III) chloride solution is added to sulphur(IV) oxide the yellow colour changes to green. Sulphur(IV) oxide reduces yellow iron(III) ions, $Fe^{3+}(aq)$ to pale green iron(II), $Fe^{2+}(aq)$ ions.

$$\underset{Yellow}{2Fe^{3+}(aq)} + 2H_2O(l) + SO_2(g) \longrightarrow \underset{pale\ green}{2Fe^{2+}(aq)} + SO^{2-}{}_4(aq) + 4H^+(aq)$$

(iii) Sulphur (IV) Oxide is reacted with hydrogen peroxide

Similarly, hydrogen peroxide is reduced to water.

$$2H_2O_2(l) + SO_2(g) + H_2O(l) \longrightarrow 2H_2O(l) + SO_4^{2-}(aq) + 2H^+(aq)$$

(Iv) A burning splint is lowered into a gas jar containing Sulphur (IV) oxide.

When a burning splint is lowered into a test-tube containing sulphur(IV) oxide, it is put off showing that the gas does not support combustion or burn. However, in the presence of a catalyst, the gas is oxidised to sulphur(VI) oxide.

Oxidising Action of Sulphur(IV) Oxide

Sulphur (IV) Oxide also acts as an Oxidising agent. It oxidises burning magnesium to magnesium oxide and hydrogen sulphide to Sulphur.

Discussion *Questions*

Explain the observations made when burning magnesium is lowered into a gas jar of sulphur(IV) oxide.

When burning magnesium is lowered into a gas jar of sulphur(IV) oxide, it continues to burn for some time. White fumes of magnesium oxide, and yellow specks of sulphur are formed.

Burning magnesium continues to burn in sulphur(IV) oxide because the heat produced by the burning magnesium decomposes the sulphur(IV) oxide to sulphur and oxygen. The magnesium combines with the oxygen to form magnesium oxide.

$$SO_2(g) \xrightarrow{heat} S(s) + O_2(g)$$

$$2Mg(s) + O_2(g) \xrightarrow{heat} 2MgO(s)$$

In this reaction, sulphur(IV) oxide is an oxidising agent, supplying oxygen to magnesium.

Explain the observations made when a gas jar of dry hydrogen sulphide gas is inverted over a test-tube containing Sulphur(IV) Oxide and a few drops of water added.

When a gas jar of dry hydrogen sulphide gas is inverted over a test-tube containing dry sulphur(IV) oxide, there is no observable change. When a few drops of water are added into each gas jar and the mixture is shaken, a yellow deposit of sulphur is produced.

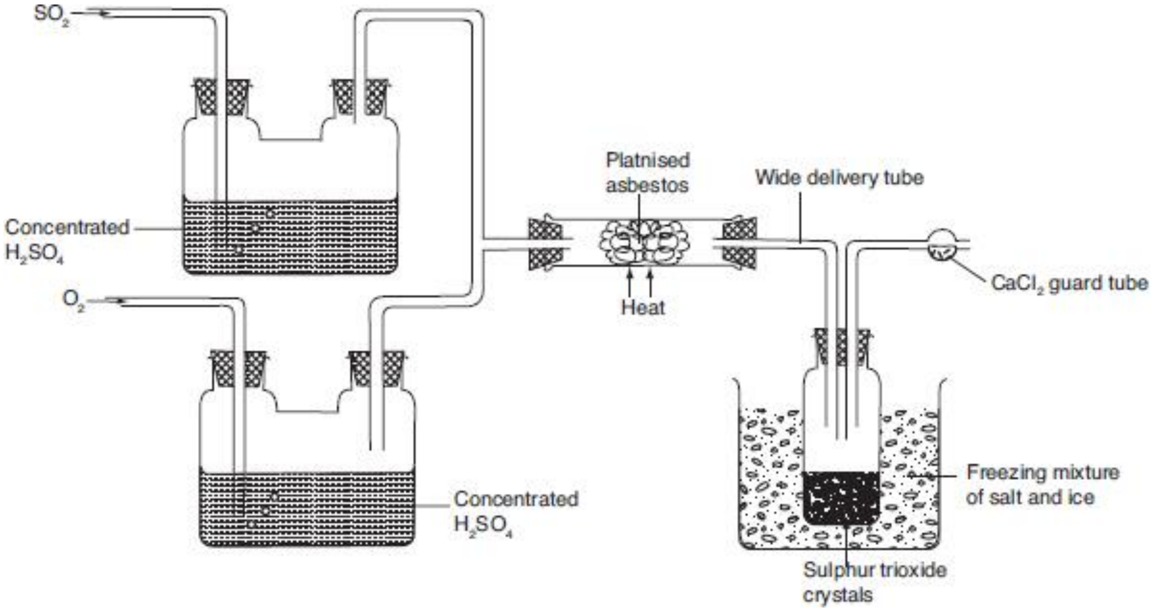
$$2H_2S(g) + SO_2(g) \longrightarrow 3S(s) + 2H_2O(l)$$

The reaction only takes place when the gases are moist. Sulphur(IV) oxide acts as an oxidising agent.

Sulphur (VI) Oxide, SO_3

Preparation of Sulphur (VI) Oxide (SO_3)

In the laboratory, sulphur(VI) oxide is prepared by reacting sulphur(IV) oxide with oxygen in the presence of a platinum catalyst..



Lab Preparation of Sulphur Trioxide, SO_3

Equation for the reaction taking place,

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(s)$$

Other Methods

Decomposing of sodium hydrogen sulphate.

$$2NaHSO_4(aq) \xrightarrow{heat} Na_2SO_4(s) + H_2O(g) + SO_3(g)$$

Discussion Questions

1. **State the purpose of the following in the set up.**
   (a) Concentrated sulphuric(VI) acid.
   Sulphur(IV) oxide and oxygen gases are dried by passing them through concentrated sulphuric(VI) acid
   (b) Freezing salt ice mixture.
   The sulphur(VI) oxide is cooled by the freezing salt-ice mixture and it solidifies.
   (c) Anhydrous calcium chloride.
   Since sulphur(VI) oxide is deliquescent, calcium chloride is used to keep it free from moisture
2. **Write an equation for the reaction that takes place in the combustion tube.**

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(s)$$

3. **Comment on the method of collection of sulphur(VI) oxide.**

   The sulphur(VI) oxide is cooled by the freezing salt-ice mixture and it solidifes. This allows it to be collected as illustrated.

Test for Sulphate and Sulphite ions.

The addition of barium chloride on sodium sulphate and sodium sulphite is used to test for sulphate and sulphite ions.

Test

To 2 cm^3 of sodium sulphate in a test-tube:

(i) Add 2 cm^3 of barium chloride(or nitrate) solution.

(ii) To the mixture in(i) above add 2 cm^3 of dilute hydrochloric acid (or dilute nitric(V) acid).

Repeat procedure (i) and (ii) using 2 cm^3 solution of sodium sulphite.

Discussion Questions

1. **State the observations made when barium chloride solution is added to:**
   - **(a) Sodium sulphate solution and barium sulphite solutions.**

When barium chloride solution is added to sodium sulphate and sodium sulphite solutions, a white precipitate is formed in each case. The white precipitates are barium sulphate and barium sulphite respectively.

$$Ba^{2+}(aq) + SO^{2-}_4(aq) \longrightarrow BaSO_4(s)$$

$$Ba^{2+}(aq) + SO^{2-}_3(aq) \longrightarrow BaSO_3(s)$$

2. **Why is the acid added to the mixture?**

The acid is added to **distinguish between the sulphate and sulphite ions.**

When dilute hydrochloric acid is added to the mixture containing barium sulphate precipitate, the precipitate does not dissolve.

However, when dilute hydrochloric acid is added to the mixture containing barium sulphite precipitate, the precipitate dissolves. This is the test for sulphite ions.

$$Ba^{2+}(aq) + SO^{2-}_3(aq) \longrightarrow BaSO_3(s)$$

$$BaSO_3(s) + 2HCl(aq) \longrightarrow BaCl_2(aq) + SO_2(g) + H_2O(l)$$

$$BaSO_3(s) + 2H^+(aq) \longrightarrow Ba^{2+}(aq) + SO_2(g) + H_2O(l)$$

Uses of Sulphur(VI) Oxide

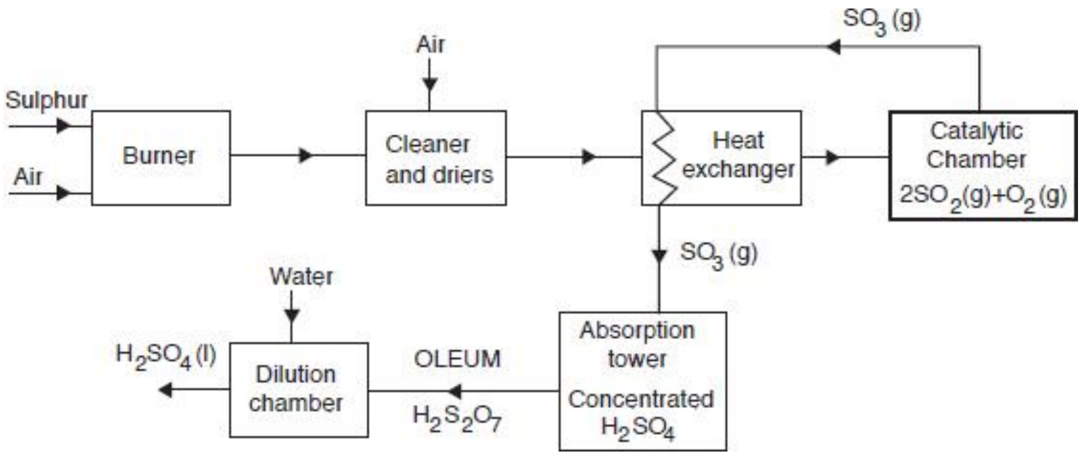
Sulphur(VI) oxide is used:

- As a major raw material in the large scale manufacture of sulphuric(VI) acid.
- To make calcium hydrogen sulphite, $Ca(HSO_3)_2$, used to bleach wood pulp in the manufacture of paper.
- As a fumigant.
- As a preservative in jam and fruit juices.

Sulphuric (VI) acid

Large Scale Manufacture of Sulphuric(VI) Acid

Sulphuric(IV) acid is manufactured by the **contact process shown below.**



The raw materials for its manufacture are **sulphur(IV) oxide and air.**

Sulphides or sulphur are burnt in air to produce sulphur(IV) oxide.

- **Burning of sulphur is the most convenient method of producing sulphur(IV) oxide.**

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

The sulphur(IV) oxide contains **dust particles** as **impurities.**

- **The impurities reduce the surface area of the catalyst, thus impairing its efficiency.** These impurities are said to **poison the catalyst.**
- The dust impurities are removed by **electrostatic precipitation**.

The mixture of gases is passed through **concentrated sulphuric(VI) acid in order to dry it.**

After purification, the gaseous mixture is **pre-heated** in the heat exchanger to **attain suitable reaction temperature** before being passed into the catalytic chamber.

In the catalytic chamber, sulphur(IV) oxide and oxygen react to form sulphur(VI) oxide.

$$2SO_2(g) + O_2(g) \xrightleftharpoons{Catalyst} 2SO_3(g) + \text{Heat}$$

Either **platinum or vanadium(V) oxide** may be used as catalyst.

- However, the **vanadium(V) oxide(V_2O_5) is preferred because it is cheaper and less easily poisoned.**

For maximum yield of sulphur(VI) oxide in the contact process, **low temperature and high pressure** are necessary.

- However, at **low temperature the reaction is slow** while **high pressure is expensive to maintain**. A **temperature of 450°C and pressure of 2 to 3 atmospheres** are used and are referred to as **optimum conditions**.

The sulphur(VI) oxide is not dissolved directly in water because the excessive heat generated could boil the acid to produce a mist of fine droplets of sulphuric(VI) acid in air. Hence, the gas is dissolved in concentrated sulphuric(VI) acid to form oleum ($H_2S_2O_7$).

$$SO_3(g) + H_2SO_4(l) \longrightarrow \underset{\text{(oleum)}}{H_2S_2O_7(l)}$$

- Sulphuric(VI) acid is obtained by **diluting oleum with water.**

$$H_2S_2O_7(1) + H_2O(l) \longrightarrow 2H_2SO_4(l)$$

During the contact process, not all the sulphur(IV) oxide is converted into sulphur (VI) oxide: **only 98% conversion of sulphur(IV) oxide to sulphur(VI) oxide occurs.**

- Sulphur(IV) oxide is a **pollutant**, passing the exhaust gases through chimneys lined with calcium hydroxide reduces the amount of sulphur(IV) oxide released to the atmosphere. This is referred to as **scrubbing the gas.**

$$Ca(OH)_2(aq) + SO_2(g) \longrightarrow CaSO_3(s) + H_2O(l)$$

- In some industries, filters fitted with strong alkalis are installed to remove any traces of acid or mist from exhaust gases.
-

Properties of Sulphuric(VI) acid

Concentrated sulphuric(VI) acid is a colourless oily liquid.

It has a density of 1.84 g/cm^3 and boils at 338°C.

It is very soluble in water. It dissolves with evolution of heat (exothermic reaction)

Concentrated sulphuric(VI) acid is hygroscopic. This property makes the acid a suitable **drying agent** for gases which do not react with it. The acid readily removes water from hydrated salts.

Concentrated sulphuric(VI) acid is a **strong dehydrating agent.**

> A **dehydrating agent** *is a substance which is capable of removing chemically combined water or the elements of water from a compound*.
>
> The process of removing water or its elements from a compound is called **dehydration.**

The acid dehydrates alcohols to to alkenes, methanoic acid to carbon(II) oxide and sugar to carbon. The sugar crystals are charred to a black mass when concentrated sulphuric (VI) acid is added to it.

Hot concentrated sulphuric(VI) acid is a strong **oxidising agent.** It oxidises metals such as copper and non metals such as carbon and sulphur.

It is also a **less volatile acid and displaces more volatile acids** from their salts.

Discussion Questions

1. **Explain why the acid should be added to water and not water to acid.**
   If water is added to the acid, **fumes** are produced since the reaction is quite **exothermic**. For this reason, dilution of the concentrated acid should always be carried out by adding small portions of the acid slowly to a large volume of water with constant stirring.
2. **Explain the observation made when concentrated sulphuric(VI) acid was added to:**
   (a) Copper(II) sulphate crystals.
   When the acid is added to blue crystals of copper(II) sulphate pentahydrate, $CuSO_4.5H_2O$, a white powder of the anhydrous salt is formed.

$$\underset{\text{blue crystals}}{CuSO_4.5H_2O(s)} \xrightarrow{Conc.H_2SO_4} \underset{\text{white powder}}{CuSO_4(s) + 5H_2O(l)}$$

(b) Sugar crystals.

When concentrated sulphuric(VI) acid is added to sugar crystals, the crystals are charred to a black mass.

$$\underset{\text{sugar}}{C_{12}H_{22}O_{11}(aq)} + \xrightarrow[H_2SO_4]{Conc.} 12C(s) + 11H_2O(l)$$

(c) Methanoic acid and ethanol

Methanoic acid is dehydrated to form carbon(II) oxide.

$$\underset{\text{(Methanoic acid)}}{HCOOH(s)} \xrightarrow[H_2SO_4]{Conc.} CO(g) + H_2O(l)$$

The acid also dehydrates alcohols to alkenes.

$$\underset{\text{(Ethanol)}}{C_2H_5OH(ag)} + \xrightarrow[160°C]{Conc\ H_2SO_4} \underset{\text{(ethene)}}{C_2H_4} + H_2O(l)$$

(c)

3. Explain the observations made when concentrated sulphuric(IV) acid was reacted with:

(a) Copper and zinc metals

When copper or zinc metal is added to concentrated sulphuric(IV) acid, a gas is produced which turns acidified potassium chromate(VI) solution from orange to green and decolourises acidified potassium manganate(VII) solution.

$$Cu(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$$

$$Zn(s) + 2H_2SO_4(l) \longrightarrow ZnSO_4(aq) + SO_2(g) + 2H_2O(l)$$

(b) Carbon and Sulphur powder.

Hot concentrated sulphuric(VI) acid **oxidises** non-metals such as **sulphur** and **carbon**.

$$S(s) + 2H_4SO_4(l) \longrightarrow 3SO_2(g) + 2H_2O(l)$$

$$C(s) + 2H_2SO_4(l) \longrightarrow 2SO_2(g) + CO_2\ (g) + 2H_2O(l)$$

4. Explain the observation made when concentrated sulphuric(VI) acid was added to Potassium nitrate crystals and Sodium chloride.

Concentrated sulphuric(IV) acid is **a less volatile acid.** It displaces more volatile acids from their salts.

$$KNO_3(s) + H_2SO_4(l) \longrightarrow HNO_3(g) + KHSO_4(s)$$

$$NaCl(s) + H_2SO_4(l) \longrightarrow HCl(s) + NaHSO_4(s)$$

Reactions of Dilute Sulphuric(VI) Acid

Reactions with metals

When dilute sulphuric(VI) acid is added to magnesium or zinc, an **effervescence occurs as a colourless gas which produces a 'pop' sound when a burning splint is applied is produced.** The reaction between magnesium and dilute sulphuric acid is more vigorous than that of zinc and dilute sulphuric(VI) acid.

$$Mg(s) + H_2SO_4\ (aq) \longrightarrow MgSO_4(aq) + H_2(g)$$

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Copper is below hydrogen in the reactivity series of the metals. It does not therefore displace hydrogen from dilute sulphuric(VI) acid.

The reactions between metals high in the reactivity series such as potassium and sodium with dilute acids are very **violent** and **should never be attempted**. The vigour of the reaction **decreases** as you go **down the reactivity series.**

Reactions with carbonates.

Effervescence occurs when dilute sulphuric(VI) acid is added to **the carbonates of zinc, sodium and copper.** The colourless gas produced forms a **white precipitate with lime water showing that it is carbon(IV) oxide.**

$$ZnCO_3(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$$

$$Na_2CO_3(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$

$$CuCO_3(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l) + CO_2(s)$$

The reaction between **calcium carbonate and dilute sulphuric (VI) acid stops soon after it starts. This is because the calcium sulphate produced during the reaction is insoluble.**

$$CaCO_3(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

The insoluble calcium sulphate forms a **coating** on the **sulphate calcium preventing further contact with the acid. As a result the reaction stops. Lead(II) carbonate behaves in a similar manner.**

Reaction with metal oxides and hydroxides

When dilute sulphuric(VI) acid reacts with a metal oxide or hydroxide, a **salt and water** are formed. However, those metal oxides whose sulphates are **insoluble** react only for a short while.

Thus, the reaction between dilute sulphuric(VI) acid, and lead(II) oxide stops almost immediately. This is due to the formation of an insoluble layer of lead(II) sulphate which effectively prevents further contact between the acid and the oxide.

$$\underset{\textit{(white solid)}}{MgO(s)} + H_2SO_4(aq) \longrightarrow \underset{\textit{(colourless solution)}}{MgSO_4(aq) + H_2O(l)}$$

$$\underset{\textit{(white solid)}}{ZnO(s)} + H_2SO_4(aq) \longrightarrow \underset{\textit{(colourless solution)}}{ZnSO_4(aq) + H_2O(l)}$$

$$\underset{\textit{(black solid)}}{CuO(s)} + H_2SO_4(aq) \longrightarrow \underset{\textit{(blue solution)}}{CuSO_4(aq) + H_2O(l)}$$

$$\underset{\textit{(white solid)}}{PbO(s)} + H_2SO_4(aq) \longrightarrow \underset{\textit{(Reaction stops immediately.)}}{PbSO_4(s) + H_2O(l)}$$

$$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Uses of Sulphuric(VI) acid

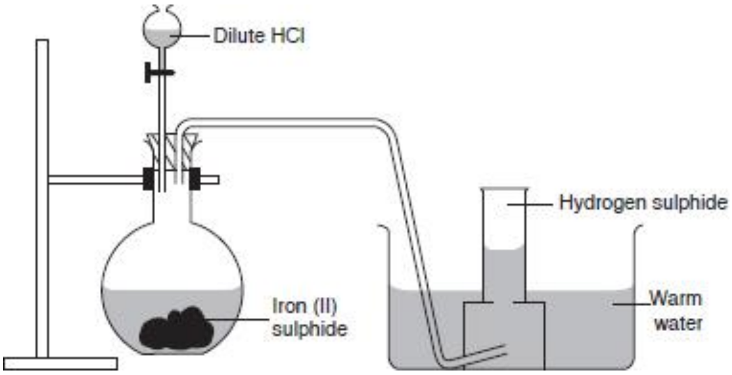
1. Manufacture of fertilisers.
2. Processing of metal ores.
3. Manufacture of detergents.
4. Manufacture of plastics.
5. Manufacture of dyes and paints.
6. Used in lead acid accumulators

Hydrogen Sulphide (H_2S)

Hydrogen sulphide is a gaseous compound of sulphur which is very poisonous. It occurs naturally in some deposits mixed with natural gas.

Preparation and properties of hydrogen sulphide

The set-up below can be used to prepare hydrogen sulphide.



.

Hydrogen sulphide is formed when dilute hydrochloric acid is added to iron(II) sulphide.

$$FeS(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2S(g)$$

However, any metal sulphide, and dilute acid can be used to prepare hydrogen sulphide gas.

The gas is collected over warm water since it dissolves in cold water.

The gas can be dried by passing it through a U-tube packed with **anyhydrous calcium chloride.**

The gas cannot be dried using concentrated sulphuric(VI) acid because it would be oxidised to sulphur.

$$3H_2S(g) + H_2SO_4(l) \longrightarrow 4S(s) + 4H_2O(l)$$

Physical properties of hydrogen sulphide

Hydrogen sulphide is a colourless gas with a characteristic smell of rotten eggs. It is very poisonous. It is slightly soluble in cold water. It is denser than air.

Chemical properties of hydrogen sulphide

(a) Reaction with water

Hydrogen sulphide dissolves in water to form a weak acidic solution of $H_2S(aq)$.

$$H_2S(aq) \rightleftharpoons 2H^+(aq) + S^{2-}(aq)$$

Aqueous hydrogen sulphide is a weak dibasic acid. It forms two types of salts, the hydrogen sulphides, and the sulphides, e.g., sodium hydrogen sulphide (NaHS) and sodium sulphide (Na_2S).

(b) Reaction with Oxygen

Hydrogen sulphide burns in air with a pale blue flame. In a limited supply of air, sulphur and water are formed.

$$2H_2S(g) + O_2(g) \longrightarrow 2S(s) + 2H_2O(g)$$

In excess air, sulphur(IV) oxide and water are formed.

$$2H_2S(g) + 3O_2(g) \longrightarrow 2SO_2(g) + 2H_2O(g)$$

(c) Reactions of hydrogen sulphide as a reducing agent

Hydrogen sulphide is a strong reducing agent and is readily oxidised to sulphur, which is precipitated as a yellow solid. Red-brown bromine water is reduced by hydrogen sulphide forming colourless hydrogen bromide solution and a yellow precipitate of sulphur.

$$Br_2(aq) + H_2S(g) \longrightarrow 2HBr(aq) + S(s)$$

Yellow iron(III) chloride in solution is reduced to green iron(II) chloride.

$$H_2S(g) + 2FeCl_3(aq) \longrightarrow 2FeCl_2(aq) + S(s) + 2HCl(aq)$$

Acidified purple potassium manganate(VII) solution is reduced to a colourless manganese(II) solution.

$$2MnO^-_4(aq) + 5H_2S(g) + 6H^+(aq) \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5S(s)$$

Similarly, acidified orange chromate(VI) is reduced to green chromium(III) ions solution.

$$Cr_2O_7^{2-}(aq) + 3H_2S(g) + 8H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3S(s)$$

A solution of hydrogen peroxide is reduced to water, and a yellow precipitate of sulphur is formed.

$$H_2O_2(aq) + H_2S(g) \longrightarrow 2H_2O(l) + S(s)$$

Brown fumes of nitrogen(IV) oxide are produced and pale yellow sulphur is deposited when hydrogen sulphide is bubbled into dilute nitric(V) acid.

$$2HNO_3(aq) + H_2S(g) \longrightarrow 2NO_2(g) + 2H_2O(l) + S(s)$$

Dilute sulpharic(VI) acid is reduced to sulphur by hydrogen sulphide.

$$H_2SO_4(aq) + 3H_2S(g) \longrightarrow 4S(s) + 4H_2O(T)$$

(d) Reaction with aqueous metallic ious

Hydrogen sulphide reacts with some metal ions in solution to form precipitates of metal sulphides. When bubbled through aqueous copper(II) sulphate, a black precipitate of copper(II) sulphide is formed.

$$Cu^{2+}(aq) + S^{2-}(aq) \longrightarrow CuS(s)$$

Other metal sulphides are precipitated as shown by the following ionic equations.

$$Zn^{2+}(aq) + S^{2-}(aq) \longrightarrow ZnS(s)$$

$$Pb^{2+}(aq) + S^{2-}(aq) \longrightarrow PbS(s)$$

$$Fe^{2+}(aq) + S^{2-}(aq) \longrightarrow FeS(s)$$

Most sulphides are insoluble in water except those of sodium, potassium and ammonium. When equal volumes of equimolar hydrogen sulphide and sodium hydroxide solutions are reacted, sodium hydrogen sulphide, an acid salt is formed.

$$NaOH(aq) + H_2S(aq) \longrightarrow NaHS(aq) + H_2O(l)$$

When excess sodium hydroxide is used sodium sulphide, a normal salt is formed.

$$2NaOH(aq) + H_2S(aq) \longrightarrow Na_2S(aq) + 2H_2O(l)$$

Pollution of the Atmosphere by Compounds of Sulphur

Sulphur compounds especially sulphur(IV) oxide and hydrogen sulphide, are among the major atmospheric pollutants. Sulphur(IV) oxide is usually emitted into the atmosphere when sulphur containing fuels are burnt. Some sulphur(IV) oxide is also emitted during the extraction of metals such as copper and in the manufacture of sulphuric(VI) acid. In the atmosphere sulphur(IV) oxide dissolves in water to form sulphuric(IV) acid.

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

The sulphuric(IV) acid is then oxidised by atmospheric oxygen to sulphuric(VI) acid, which comes down as acid rain or acid fog. These have serious environmental effects. These include:

- Stunted growth in plants due to loss of chlorophyll from plants' leaves.
- Death of plants as a result of defoliation.
- Destruction of aquatic life in acidified lakes.
- Corrosion of stone work on buildings.
- Corrosion of metallic structures.
- Leaching of minerals in the soil.
- Irritation of the respiratory system

Review Exercises

1. 2006 Q 6

In an experiment to study the properties of concentrated sulphuric acid, a mixture of the acid and wood charcoal was heated in a boiling tube.

(a) Write the equation of the reaction that took place in the boiling tube. (1 mark)

(b) Using oxidation numbers, show that reduction and oxidation reactions took place in the boiling tube. (2 marks)

2. 2006 Q 16

When hydrogen sulphide gas was bubbled into an aqueous solution of iron (III) chloride, a yellow precipitate was deposited.

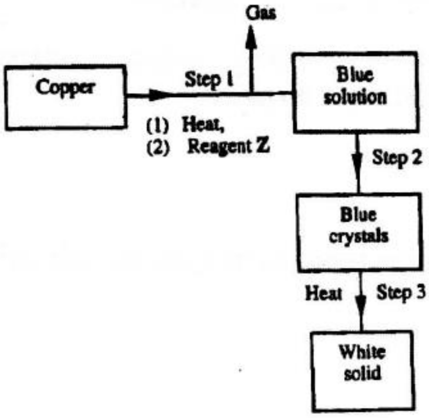
(a) State another observation that was made. (1 mark)

(b) Write an equation for the reaction that took place. (1 mark)

(c) What type of reaction was undergone by hydrogen by hydrogen sulphide in this reaction? (1 mark)

3. 2006 Q 27 P1

Study the flow chart below and answer the questions that follow.



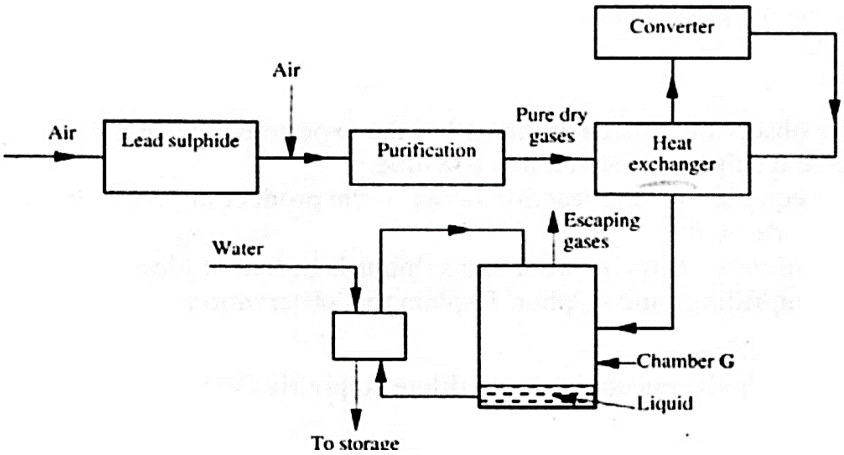
(a) Name reagent z. (1mark)

(b) Describe the process which takes place in step 2. (1mark)

(c) Identify the white solid. (1mark)

4. 2006 Q 4 P2

(a) The diagram below shows some processes that take place during the industrial manufacture of sulphuric acid.



(i) Write the equation for the reaction in which sulphur dioxide gas is produced. (1 mark)

(ii) Why is it necessary to keep the gases pure and dry? (1 mark)

(iii) Describe the process that takes place in chamber G. (1 mark)

(iv) Name the gases that escape into the environment. (1 mark)

(v) State and explain the harmful effect on the environment of one of the gases named in (iv) above (1 mark)

(vi) Give one reason why it is necessary to use a pressure of 2-3 atmospheres and not more. (1 mark)

(b) (i) Complete the table below to show the observations made when concentrated sulphuric acid is added to the substances shown. (2 marks)

Substance	Observation
Iron fillings	
Crystals of white sugar	

(ii) Give reasons for the observations made using:

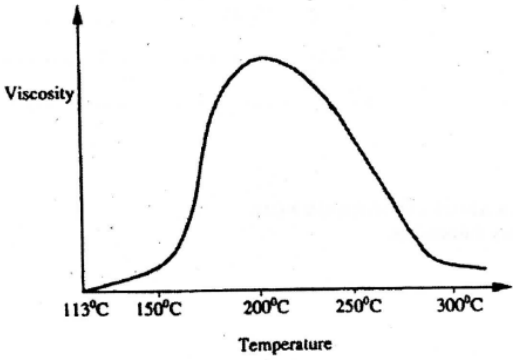
I. iron fillings (1 mark)
II. Crystals of white sugar. (1 mark)

(c) Name one fertilizer made from sulphuric acid. (1 mark)

(d) Suggest a reason why $BaSO_4$(a pigment made from sulphuric acid) would be suitable in making paint for cars. (1 mark)

5. 2007 Q 30 P1

Below is a sketch of a graph showing the change in viscosity? (Ease of flow) with temperature when solid sulphur is heated.



Describe what happens to the sulphur molecules when sulphur is heated from 150 °C to about 200 °C. (2 marks)

6. 2008 Q 8 P1

(a) State the observation made at the end of the experiment when a mixture of iron powder and sulphur is heated in a test tube. (1 mark)

(b) Write an equation for the reaction the product in (a) above and dilute hydrochloric acid. (1 mark)

(c) When a mixture of iron powder and sulphur is heated, it glows more brightly than that of iron fillings and sulphur. Explain this observation. (1 mark)

7. 2008 Q 9 P1

Zinc reacts with both concentrated and dilute sulphuric (VI) acid. Write equations for the two reactions. (2 marks)

8. 2008 Q 30 P1

Crude oil contains sulphur. What would be the effect to the environment of using fuel containing sulphur? (1 mark)

9. 2009 Q 22 P1

A student added very dilute sulphuric (VI) acid to four substance and recorded the observations shown in the table below.

Test	**Substance**	**Gas given off**
1	Sodium	Yes
2	Iron	No
3	Carbon	Yes
4	Copper	No

For which tests are the observations wrong? Explain. (3 marks)

10. 2010 Q 11 P1 ,2016 Q26 P1

Hydrogen sulphide is a highly toxic and flammable gas. It is normally prepared in a fume chamber.

(a) Name two reagents that can be used to prepare hydrogen sulphide in the laboratory. (1 mark)

(b) One of the uses of hydrogen sulphide is to produce sulphur as shown in the following equation;

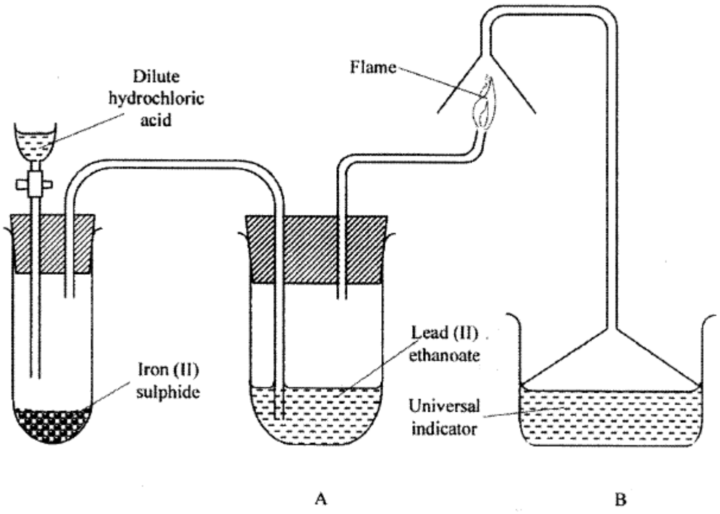
$$2H_2S(g) + SO_2(g) \longrightarrow 3S(s) + 2H_2O(l)$$

Identify the reducing agent in this reaction and give a reason for your answer. (1 mark)

(c) Other than production of sulphuric (IV) acid, state one commercial use of sulphur. (1 mark)

11. 2011 Q 17 P1

The set up below was used to prepare a gas and study some of its properties.
Study it and answer the questions that follow:



(a) State and explain the observations made in the.

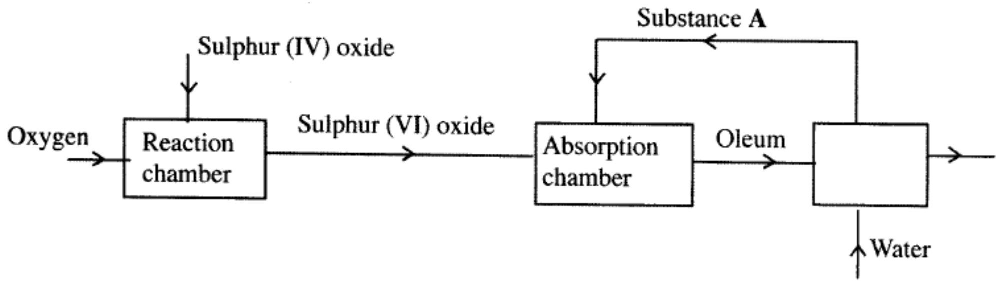
I. tube labelled A; (1 mark)

II. beaker labelled B. (1 mark)

(b) State **one** precaution that should be taken when carrying out this experiment. (1 mark)

12. 2011 Q 1 P2

The flow chart below shows some of the processes involved in large scale production of sulphuric (VI) acid. Use it to answer the questions that follow.



(a) Describe how oxygen is obtained from air on a large scale. (3 marks)

(b) (i) Name substance **A**. (1 mark)

(ii) Write an equation for the process that takes place in the absorption chamber (1 mark)

(c) Vanadium (V) Oxide is a commonly used catalyst in the contact process.

(i) Name another catalyst which can be used for this process. (1 mark)

(ii) Give **two** reasons why vanadium (V) Oxide is the commonly used catalyst. (2 marks)

(d) State and explain the observation made when concentrated sulphuric acid is added to crystals of copper (II) sulphate in a beaker. (2 marks)

(e) The reaction of concentrated sulphuric (VI) acid with sodium chloride produces hydrogen chloride gas. State the property of concentrated sulphuric (VI) acid illustrated in this reaction. (1 mark)

(f) Name **four** uses of sulphuric (VI) acid. (2 marks)

13. 2012 Q18 P1

Acidified potassium manganate (VII) solution is decolourised when sulphur (IV) oxide is bubbled through it. The equation for the reaction is given below.

$2H_2O (l) + 5SO_2(g) + 2KMnO_4(aq \longrightarrow K_2SO_4(aq) + 2MnSO_4(a + 2H_2SO_4(aq)$

(a) Which reactant is oxidised? Explain. (2 marks)

(b) Other than the manufacture of sulphuric (VI) acid, state other use of sulphur (IV) oxide (1 mark)

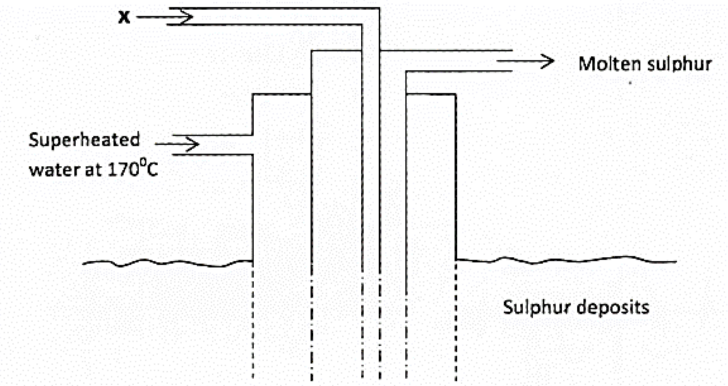
14. 2013 Q12 P1

(a) What would be observed if sulphur (IV) oxide is bubbled through acidified potassium manganate(VII)? (1 mark)

(b) In an experiment, sulphur (IV) oxide was dissolved in water to form solution **L**.
  (i) What would be observed if a few drops of barium nitrate solution were **immediately** added to solution **L**? (1 mark)
  (ii) Write an ionic equation for the reaction that occurred between solution **L** and aqueous barium nitrate in (b) (i) above. (1 mark)

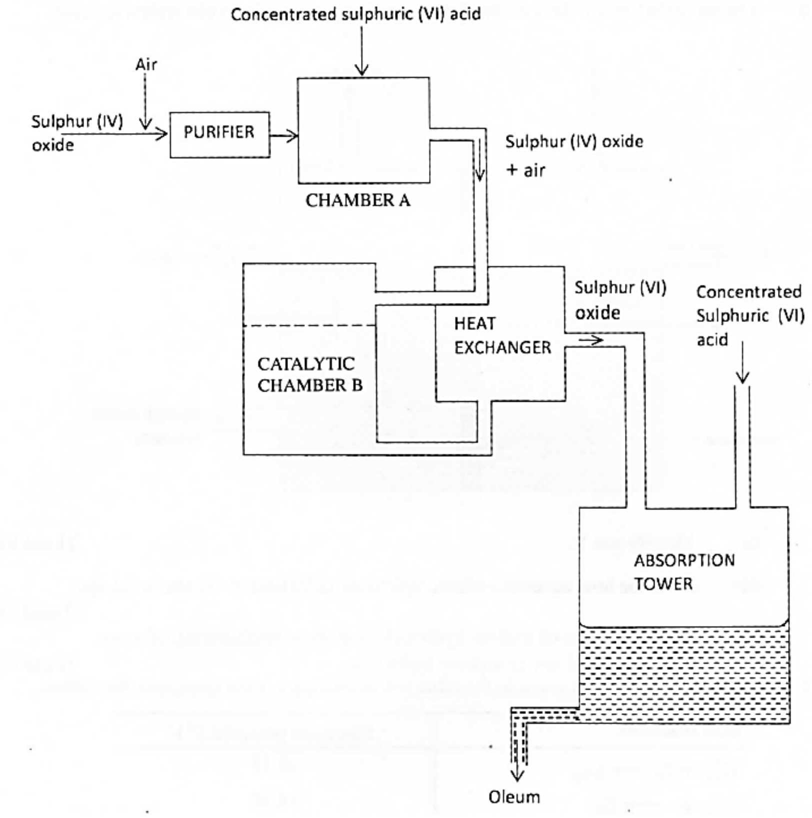
15. 2013 Q3 P2

(a) The diagram below shows the Frasch process used for extraction of sulphur. Use it to answer the question that follows.



(i) Identify X. (1 mark)
(ii) Why is it necessary to use superheated water in this process? (1 mark)
(iii) State two physical properties of sulphur that makes it possible for it to be extracted by this method. (2 marks)

(b) The diagram below shows part of the process in the manufacture of sulphuric (VI) acid. Study it and answer the questions that follow.



(i) Write an equation for the formation of sulphur (IV) oxide from sulphur. (1 mark)

(ii) What is the role of concentrated sulphur (VI) acid in chamber A? (1 mark)

(iii) Name two catalysts that can be used in the catalytic chamber B. (2 marks)

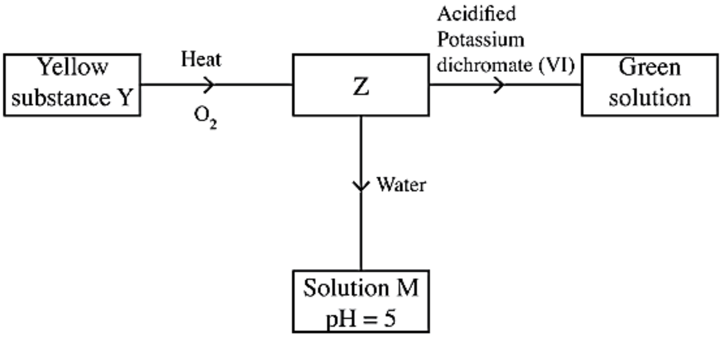
(iv) State two roles of the heat exchanger. (1 mark)

(c) Explain one way in which sulphur (IV) oxide is a pollutant (1mark)

(d) What observation will be made when a few drops of concentrated sulphuric (VI) acid are added to crystals of sugar? Explain your answer. (1 mark)

16. 2014 Q11 P1

Study the flow chart below and answer the questions that follow.



Identify Z and M. (2 marks)

17. 2014 Q20 P1

In the contact process, during the production of sulphur (VI) oxide, a catalyst is used. Give two reasons why vanadium (V) oxide is preferred to platinum.

(2 marks)

18. 2015 Q17 P1

(a) One of the allotropes of sulphur is rhombic sulphur, name the other allotrope.

(1 mark)

(b) Concentrated sulphuric (VI) acid reacts with ethanol and copper.

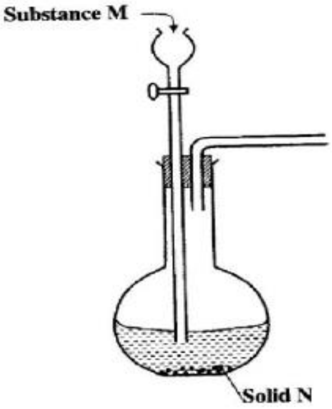
State the property of the acid shown in each case. (2 marks)

(i) Ethanol

(ii) Copper ....

19. 2015 Q5 P2

(a) The set up below can be used to generate a gas without heating. This occurs when substance M reacts with solid N.



(i) Complete the table below giving the names of substance M and solid N if the gasses generated are chlorine and sulphur (IV) oxide. (2 marks)

	Chlorine	Sulphur (IV) Oxide
Substance M		
Solid N		

(ii) Complete the diagram above to show how a dry sample of sulphur (IV) oxide can be collected. (2 marks)

(b) Describe two chemical methods that can be used to test the presence of sulphur (IV)oxide.

(3 marks)

(c) Other than the manufacture of sulphuric (VI) acid, state two uses of sulphur (IV) oxide.

(2 marks)

20. 2017 P1 Q7.

A sample of water is suspected to contain sulphate ions. Describe an experiment that can be carried out to determine the presence of sulphate ions. (3 marks)

21. 2017 P2 Q3 (a)

A student used **Figure 2** to investigate the action of dilute sulphuric (VI) acid on some metals. Beaker **I** and **II** contained equal volumes of dilute sulphuric (VI) acid. To beaker **I**, a clean iron rod was dipped and to beaker **II**, a clean copper rod was dipped.

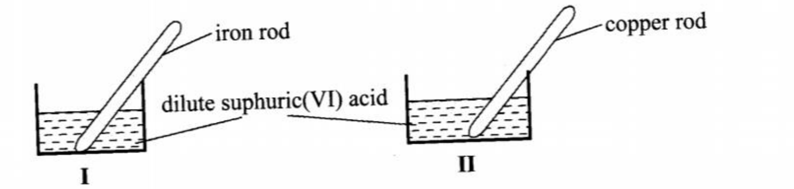


Figure 2

(i) Why was it necessary to clean the metal rods? (1 mark)

(ii) Describe the observations made in each beaker.

Beaker I: (1 mark)

Beaker II: (1 mark)

(iii) Explain the observations in (a) (ii). (2 marks)

22. 2018 P1 Q 4.

One of the allotropes of sulphur is rhombic sulphur.

(a) Name the other allotrope of sulphur. (1 mark)

(b) Draw a diagram to show the shape of the allotrope named in (a) above.

(1 mark)

(c) Write an equation for the reaction between concentrated sulphuric(VI) acid and sulphur.

(1 mark)

23. 2019 P1 Q9.

Sulphur(IV) oxide is prepared in the laboratory using the set-up in **Figure 3.** Study it and answer the questions that follow.

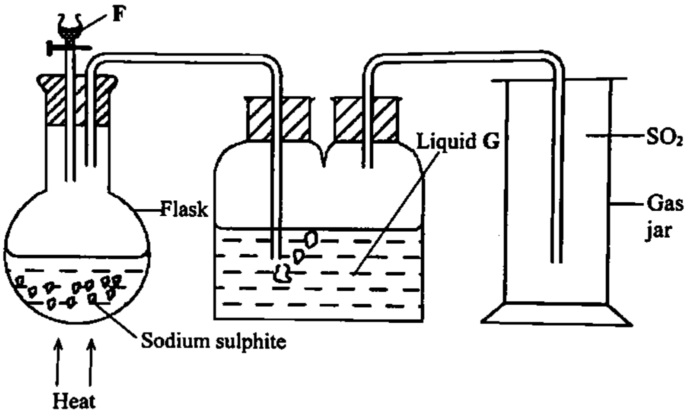


Figure 3

(a) Identify substance F. (1 mark)

(b) Write an equation for the reaction that takes place in the flask. (1 mark)

(c) State the purpose of liquid G. (1 mark)