

12.0.0 GAS LAWS (15 LESSONS)

(a) Gas laws

1. Matter is made up of small particles in accordance to Kinetic Theory of matter:

Naturally, there are basically **three** states of matter: **Solid, Liquid** and **gas**:

(i) A solid is made up of particles which are very closely packed with a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

(ii) A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

(iii) A gas is made up of particles free from each other. It thus has no definite /fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

2. Gases are affected by **physical conditions**. There are **two** physical conditions:

(i) **Temperature**

(ii) **Pressure**

3. The SI unit of temperature is **Kelvin(K)**.

Degrees Celsius/Centigrade($^{\circ}\text{C}$) are also used.

The two units can be interconverted from the relationship:

$$^{\circ}\text{C} + 273 = \text{K}$$

$$\text{K} - 273 = ^{\circ}\text{C}$$

Practice examples

1. Convert the following into Kelvin.

(i) 0°C

$$^{\circ}\text{C} + 273 = \text{K substituting} : 0^{\circ}\text{C} + 273 = \mathbf{273\text{ K}}$$

(ii) -273°C

$$^{\circ}\text{C} + 273 = \text{K substituting} : -273^{\circ}\text{C} + 273 = \mathbf{0\text{ K}}$$

(iii) 25°C

$$^{\circ}\text{C} + 273 = \text{K substituting} : 25^{\circ}\text{C} + 273 = \mathbf{298\text{ K}}$$

(iv) 100°C

$$^{\circ}\text{C} + 273 = \text{K substituting} : 100^{\circ}\text{C} + 273 = \mathbf{373\text{ K}}$$

2. Convert the following into degrees Celsius/Centigrade($^{\circ}\text{C}$).

(i) 10 K

$$\text{K} - 273 = ^{\circ}\text{C} \text{ substituting: } 10 - 273 = -263^{\circ}\text{C}$$

(ii) (i) 1 K

$$\text{K} - 273 = ^{\circ}\text{C} \text{ substituting: } 1 - 273 = -272^{\circ}\text{C}$$

(iii) 110 K

$$\text{K} - 273 = ^{\circ}\text{C} \text{ substituting: } 110 - 273 = -163^{\circ}\text{C}$$

(iv) -24 K

$$\text{K} - 273 = ^{\circ}\text{C} \text{ substituting: } -24 - 273 = -297^{\circ}\text{C}$$

The **standard** temperature is $273\text{K} = 0^{\circ}\text{C}$.

The **room** temperature is assumed to be $298\text{K} = 25^{\circ}\text{C}$

4. The SI unit of pressure is Pascal(**Pa**) / Newton per metre squared (**Nm^{-2}**) .
Millimeters' of mercury(**mmHg**) ,centimeters of mercury(**cmHg**) and
atmospheres are also commonly used.

The units are **not** interconvertible but Pascals(Pa) are equal to Newton per metre squared(Nm^{-2}).

The **standard** pressure is the **atmospheric** pressure.

Atmospheric pressure is **equal** to about:

- (i)101325 Pa
- (ii)101325 Nm^{-2}
- (iii)760 mmHg
- (iv)76 cmHg
- (v)one atmosphere.

5. Molecules of gases are always in continuous random motion at high speed. This motion is affected by the physical conditions of temperature and pressure.
Physical conditions change the volume occupied by gases in a **closed** system.
The effect of physical conditions of temperature and pressure was investigated and expressed in both Boyles and Charles laws.

6. Boyles law states that

“the **volume** of a **fixed** mass of a gas is **inversely** proportional to the **pressure** at constant/fixed **temperature** ”

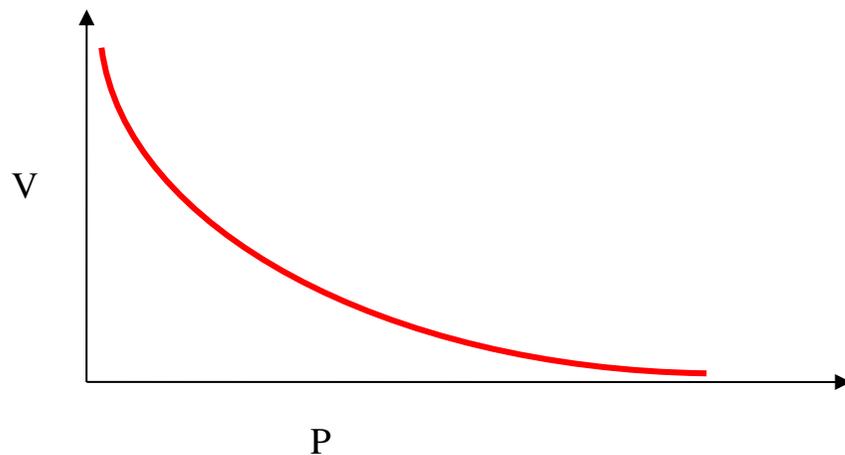
Mathematically:

$$\text{Volume} \propto \frac{1}{\text{Pressure}} \text{ (Fixed /constant Temperature)}$$

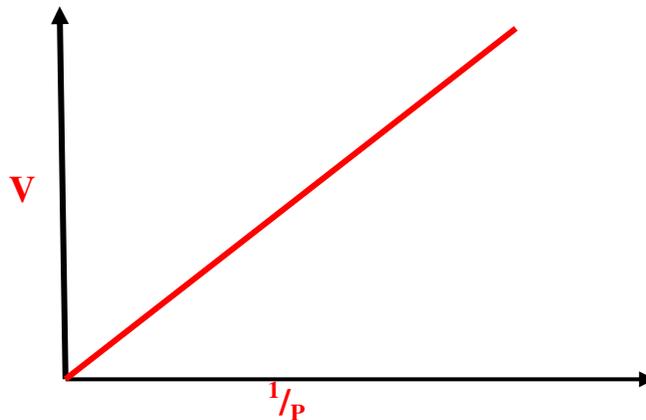
$$V \propto \frac{1}{P} \text{ (Fixed /constant T) i.e. } \mathbf{PV = Constant(k)}$$

From Boyles law , an **increase** in pressure of a gas cause a **decrease** in volume. i.e **doubling** the pressure cause the volume to be **halved**.

Graphically therefore a plot of volume(V) **against** pressure (P) produces a **curve**.



Graphically a plot of volume(V) **against** inverse/reciprocal of pressure ($1/p$) produces a **straight line**



For **two** gases then $\mathbf{P_1 V_1 = P_2 V_2}$

P_1 = Pressure of gas 1

V_1 = Volume of gas 1

P_2 = Pressure of gas 2

V_2 = Volume of gas 2

Practice examples:

1. A fixed mass of gas at 102300Pa pressure has a volume of 25cm³. Calculate its volume if the pressure is doubled.

Working

$$P_1 V_1 = P_2 V_2 \quad \text{Substituting : } 102300 \times 25 = (102300 \times 2) \times V_2$$
$$V_2 = \frac{102300 \times 25}{(102300 \times 2)} = \mathbf{12.5\text{cm}^3}$$

2. Calculate the pressure which must be applied to a fixed mass of 100cm³ of Oxygen for its volume to triple at 100000Nm⁻².

$$P_1 V_1 = P_2 V_2 \quad \text{Substituting : } 100000 \times 100 = P_2 \times (100 \times 3)$$
$$V_2 = \frac{100000 \times 100}{(100 \times 3)} = \mathbf{33333.3333 \text{ Nm}^{-2}}$$

3. A 60cm³ weather ballon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the ballon reach stratosphere where the pressure is 90000Pa?

$$P_1 V_1 = P_2 V_2 \quad \text{Substituting : } 101325 \times 60 = 90000 \times V_2$$
$$V_2 = \frac{101325 \times 60}{90000} = \mathbf{67.55 \text{ cm}^3}$$

The new volume at 67.55 cm³ **exceed** ballon capacity of 60.00 cm³. It will **burst** before reaching destination.

7. Charles law states that “the **volume** of a **fixed** mass of a gas is **directly proportional to the absolute temperature** at constant/**fixed** pressure”

Mathematically:

Volume \propto Pressure (Fixed /constant pressure)

$V \propto T$ (Fixed /constant P) i.e $\frac{V}{T} = \text{Constant}(k)$

From Charles law , an **increase** in temperature of a gas cause an **increase** in volume. i.e **doubling** the temperature cause the volume to be **doubled**.

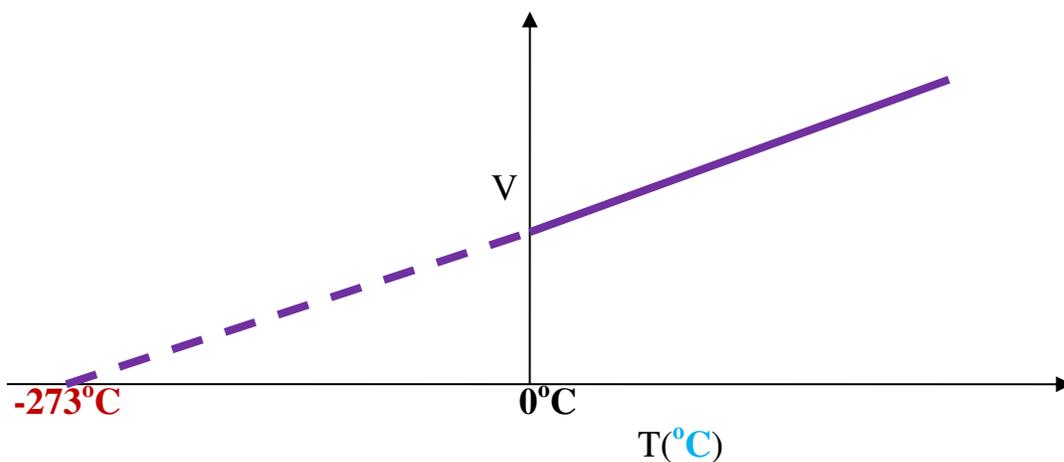
Gases expand/increase by $\frac{1}{273}$ by volume on heating. Gases contract/decrease by $\frac{1}{273}$ by volume on cooling at constant/fixed pressure.

The volume of a gas continues decreasing with decrease in temperature until at **-273°C / 0 K** the volume is **zero**. i.e. there is no gas.

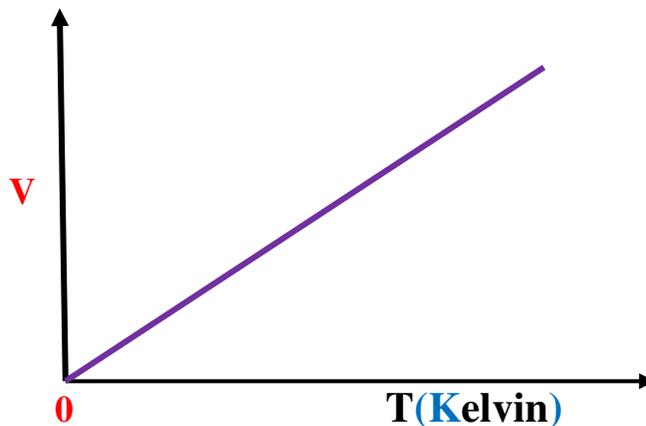
This temperature is called **absolute zero**. It is the **lowest** temperature at which a gas **can** exist.

Graphically therefore a plot of volume (**V**) **against** Temperature (**T**) in:

(i) $^{\circ}\text{C}$ produces a **straight line** that is **extrapolated** to the absolute zero of **-273°C** .



(ii) Kelvin/K produces a **straight line** from absolute zero of **0 Kelvin**



For two gases then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

T_1 = Temperature **in Kelvin** of gas 1

V_1 = Volume of gas 1

T_2 = Temperature **in Kelvin** of gas 2

V_2 = Volume of gas 2

Practice examples:

1. 500cm³ of carbon(IV)oxide at 0°C was transferred into a cylinder at -4°C. If the capacity of the cylinder is 450 cm³, explain what happened.

$$\begin{aligned}\frac{V_1}{T_1} &= \frac{V_2}{T_2} \text{ substituting } \frac{500}{(0+273)} = \frac{V_2}{(-4+273)} \\ &= \frac{500 \times (-4 \times 273)}{(0+273)} = \mathbf{492.674\text{cm}^3}\end{aligned}$$

The capacity of cylinder (500cm³) is **less** than new volume(492.674cm³).

7.326cm³(500-492.674cm³)of carbon(IV)oxide gas did not fit into the cylinder.

2. A mechanic was filling a deflated tyre with air in his closed garage using a hand pump. The capacity of the tyre was 40,000cm³ at room temperature. He rolled the tyre into the car outside. The temperature outside was 30°C.Explain what happens.

$$\begin{aligned}\frac{V_1}{T_1} &= \frac{V_2}{T_2} \text{ substituting } \frac{40000}{(25+273)} = \frac{V_2}{(30+273)} \\ &= \frac{40000 \times (30 \times 273)}{(25+273)} = \mathbf{40671.1409\text{cm}^3}\end{aligned}$$

The capacity of a tyre (40000cm³) is **less** than new volume(40671.1409cm³).

The tyre thus bursts.

3. A hydrogen gas balloon with 80cm³ was released from a research station at room temperature. If the temperature of the highest point it rose is -30°C , explain what happened.

$$\begin{aligned}\frac{V_1}{T_1} &= \frac{V_2}{T_2} \text{ substituting } \frac{80}{(25+273)} = \frac{V_2}{(-30+273)} \\ &= \frac{80 \times (-30 \times 273)}{(-30+273)} = \mathbf{65.2349\text{cm}^3}\end{aligned}$$

(25 + 273)

The capacity of balloon (80cm³) is **more** than new volume (65.2349cm³).
The balloon thus remained intact.

8. The continuous random motion of gases differ from gas to the other. The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion**.

The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas.**

The density of a gas depends on its molar mass/relative molecular mass. i.e. **The higher the density the higher the molar mass/relative atomic mass and thus the lower the rate of diffusion.**

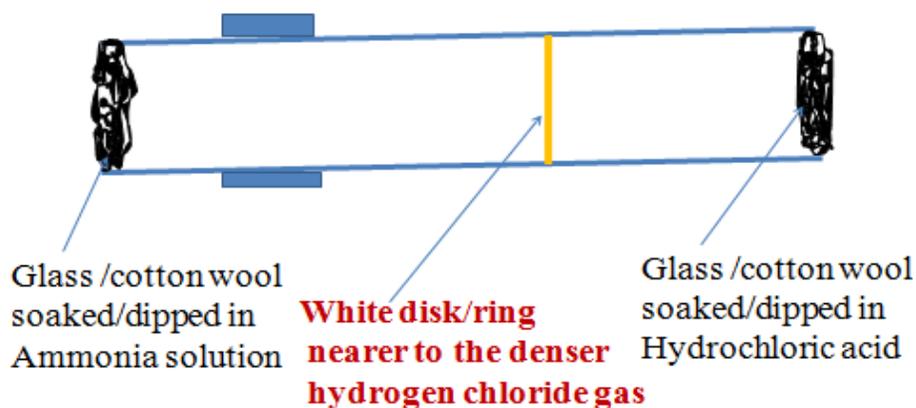
Examples

1. Carbon (IV) oxide(CO₂) has a molar mass of 44g. Nitrogen(N₂) has a molar mass of 28g. (N₂) is thus lighter/less dense than Carbon (IV) oxide(CO₂). N₂ diffuses faster than CO₂.

2. Ammonia(NH₃) has a molar mass of 17g. Nitrogen(N₂) has a molar mass of 28g. (N₂) is thus about **twice** lighter/less dense than Ammonia(NH₃). Ammonia(NH₃) diffuses twice faster than N₂.

3. Ammonia(NH₃) has a molar mass of 17g. Hydrogen chloride gas has a molar mass of 36.5g. Both gases on contact react to form **white fumes** of ammonium chloride. When a glass/cotton wool dipped in ammonia and another glass/cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to glass/cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower.

Diffusion of ammonia and hydrogen chloride



The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**.

Grahams law states that:

“the rate of diffusion of a gas is inversely proportional to the square root of its density, at the same/constant/fixed temperature and pressure”

Mathematically

$$R \propto \frac{1}{\sqrt{p}} \quad \text{and since density is proportional to mass then} \quad R \propto \frac{1}{\sqrt{m}}$$

For two gases then:

$$\frac{R_1}{\sqrt{M_2}} = \frac{R_2}{\sqrt{M_1}} \quad \text{where: } R_1 \text{ and } R_2 \text{ is the } \underline{\text{rate}} \text{ of diffusion of } 1^{\text{st}} \text{ and } 2^{\text{nd}} \text{ gas.}$$

$$M_1 \text{ and } M_2 \text{ is the } \underline{\text{molar mass}} \text{ of } 1^{\text{st}} \text{ and } 2^{\text{nd}} \text{ gas.}$$

Since rate is inverse of time. i.e. the higher the rate the less the time:

For two gases then:

$$\frac{T_1}{\sqrt{M_1}} = \frac{T_2}{\sqrt{M_2}} \quad \text{where: } T_1 \text{ and } T_2 \text{ is the } \underline{\text{time taken}} \text{ for } 1^{\text{st}} \text{ and } 2^{\text{nd}} \text{ gas to diffuse.}$$

$$M_1 \text{ and } M_2 \text{ is the } \underline{\text{molar mass}} \text{ of } 1^{\text{st}} \text{ and } 2^{\text{nd}} \text{ gas.}$$

Practice examples:

1. It takes 30 seconds for 100cm³ of carbon(IV)oxide to diffuse across a porous plate. How long will it take 150cm³ of nitrogen(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,N=14.0=16.0)

$$\text{Molar mass CO}_2=44.0 \quad \text{Molar mass NO}_2=46.0$$

Method 1

$$\begin{array}{l} 100\text{cm}^3 \text{ CO}_2 \text{ takes} \quad 30\text{seconds} \\ 150\text{cm}^3 \quad \text{takes} \quad \frac{150 \times 30}{100} = \underline{45\text{seconds}} \end{array}$$

$$\frac{T_{\text{CO}_2}}{T_{\text{NO}_2}} = \frac{\sqrt{\text{molar mass CO}_2}}{\sqrt{\text{molar mass NO}_2}} \Rightarrow \frac{45\text{seconds}}{T_{\text{NO}_2}} = \frac{\sqrt{44.0}}{\sqrt{46.0}}$$

$$T_{\text{NO}_2} = \frac{45\text{seconds} \times \sqrt{46.0}}{\sqrt{44.0}} = \underline{46.0114} \text{ seconds}$$

Method 2

$$\begin{array}{l}
 100\text{cm}^3 \text{ CO}_2 \text{ takes } 30\text{seconds} \\
 1\text{cm}^3 \text{ takes } \frac{100 \times 1}{30} = \underline{\underline{3.3333\text{cm}^3\text{sec}^{-1}}} \\
 \\
 \frac{R \text{ CO}_2}{R \text{ NO}_2} = \frac{\sqrt{\text{molar mass NO}_2}}{\sqrt{\text{molar mass CO}_2}} \Rightarrow \frac{3.3333\text{cm}^3\text{sec}^{-1}}{R \text{ NO}_2} = \frac{\sqrt{46.0}}{\sqrt{44.0}}
 \end{array}$$

$$R \text{ NO}_2 = \frac{3.3333\text{cm}^3\text{sec}^{-1} \times \sqrt{44.0}}{\sqrt{46.0}} = \underline{\underline{3.2601\text{cm}^3\text{sec}^{-1}}}$$

$$\begin{array}{l}
 3.2601\text{cm}^3 \text{ takes } 1\text{seconds} \\
 150\text{cm}^3 \text{ take } \frac{150\text{cm}^3}{3.2601\text{cm}^3} = \underline{\underline{46.0109\text{seconds}}}
 \end{array}$$

2. How long would 200cm³ of Hydrogen chloride take to diffuse through a porous plug if carbon(IV)oxide takes 200seconds to diffuse through.

$$\begin{array}{l}
 \text{Molar mass CO}_2 = 44\text{g} \quad \text{Molar mass HCl} = 36.5\text{g} \\
 \frac{T \text{ CO}_2}{T \text{ HCl}} = \frac{\sqrt{\text{molar mass CO}_2}}{\sqrt{\text{molar mass HCl}}} \Rightarrow \frac{200 \text{ seconds}}{T \text{ HCl}} = \frac{\sqrt{44.0}}{\sqrt{36.5}}
 \end{array}$$

$$T \text{ HCl} = \frac{200\text{seconds} \times \sqrt{36.5}}{\sqrt{44.0}} = \underline{\underline{182.1588 \text{ seconds}}}$$

3. Oxygen gas takes 250 seconds to diffuse through a porous diaphragm. Calculate the molar mass of gas Z which takes 227 second to diffuse.

$$\begin{array}{l}
 \text{Molar mass O}_2 = 32\text{g} \quad \text{Molar mass Z} = x \text{ g} \\
 \frac{T \text{ O}_2}{T \text{ Z}} = \frac{\sqrt{\text{molar mass O}_2}}{\sqrt{\text{molar mass Z}}} \Rightarrow \frac{250 \text{ seconds}}{227\text{seconds}} = \frac{\sqrt{32.0}}{\sqrt{x}}
 \end{array}$$

$$\sqrt{x} = \frac{227\text{seconds} \times \sqrt{32}}{250} = \underline{\underline{26.3828 \text{ grams}}}$$

4. 25cm³ of carbon(II)oxide diffuses across a porous plate in 25seconds. How long will it take 75cm³ of Carbon(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0, O=16.0)

$$\text{Molar mass CO}_2 = 44.0 \quad \text{Molar mass CO} = 28.0$$

Method 1

$$\begin{array}{l}
 25\text{cm}^3 \text{ CO takes } 25\text{seconds} \\
 75\text{cm}^3 \text{ takes } \frac{75 \times 25}{25} = \underline{\underline{75\text{seconds}}}
 \end{array}$$

$$\frac{T_{\text{CO}_2}}{T_{\text{CO}}} = \frac{\sqrt{\text{molar mass CO}_2}}{\sqrt{\text{molar mass CO}}} \Rightarrow \frac{T_{\text{CO}_2} \text{seconds}}{75} = \frac{\sqrt{44.0}}{\sqrt{28.0}}$$

$$T_{\text{CO}_2} = \frac{75 \text{seconds} \times \sqrt{44.0}}{\sqrt{28.0}} = \underline{\underline{94.0175}} \text{ seconds}$$

Method 2

25cm³ CO₂ takes 25seconds

$$1\text{cm}^3 \text{ takes } \frac{25 \times 1}{25} = \underline{\underline{1.0\text{cm}^3\text{sec}^{-1}}}$$

$$\frac{R_{\text{CO}_2}}{R_{\text{CO}}} = \frac{\sqrt{\text{molar mass CO}}}{\sqrt{\text{molar mass CO}_2}} \Rightarrow \frac{x \text{ cm}^3\text{sec}^{-1}}{1.0\text{cm}^3\text{sec}^{-1}} = \frac{\sqrt{28.0}}{\sqrt{44.0}}$$

$$R_{\text{CO}_2} = \frac{1.0\text{cm}^3\text{sec}^{-1} \times \sqrt{28.0}}{\sqrt{44.0}} = \underline{\underline{0.7977\text{cm}^3\text{sec}^{-1}}}$$

$$\frac{0.7977\text{cm}^3}{75\text{cm}^3} \text{ takes } \frac{1 \text{ seconds}}{0.7977\text{cm}^3} = \underline{\underline{94.0203\text{seconds}}}$$

13.0.0 THE MOLE-FORMULAE AND CHEMICAL EQUATIONS (40 LESSONS)

Introduction to the mole, molar masses and Relative atomic masses

1. The mole is the **SI** unit of the **amount** of substance.
2. The number of particles e.g. atoms, ions, molecules, electrons, cows, cars are all measured in terms of moles.
3. The number of particles in one mole is called the **Avogadros Constant**. It is denoted "**L**".

The Avogadros Constant contain **6.023 x10²³** particles. i.e.

$$\begin{array}{ll} 1 \text{ mole} = 6.023 \times 10^{23} \text{ particles} & = 6.023 \times 10^{23} \\ 2 \text{ moles} = 2 \times 6.023 \times 10^{23} \text{ particles} & = 1.205 \times 10^{24} \\ 0.2 \text{ moles} = 0.2 \times 6.023 \times 10^{23} \text{ particles} & = 1.205 \times 10^{22} \\ 0.0065 \text{ moles} = 0.0065 \times 6.023 \times 10^{23} \text{ particles} & = 3.914 \times 10^{21} \end{array}$$

3. The mass of one mole of a substance is called **molar mass**. The molar mass of:
(i) an **element** has mass equal to relative atomic mass /RAM(in grams) of the element e.g.

$$\begin{array}{l} \text{Molar mass of carbon(C)} = \text{relative atomic mass} = 12.0\text{g} \\ 6.023 \times 10^{23} \text{ particles of carbon} = 1 \text{ mole} = 12.0 \text{ g} \end{array}$$

$$\begin{array}{l} \text{Molar mass of sodium(Na)} = \text{relative atomic mass} = 23.0\text{g} \\ 6.023 \times 10^{23} \text{ particles of sodium} = 1 \text{ mole} = 23.0 \text{ g} \end{array}$$

$$\begin{array}{l} \text{Molar mass of Iron (Fe)} = \text{relative atomic mass} = 56.0\text{g} \\ 6.023 \times 10^{23} \text{ particles of iron} = 1 \text{ mole} = 56.0 \text{ g} \end{array}$$

- (ii) a **molecule** has mass equal to relative molecular mass /RMM (in grams) of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.

The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic** (e.g. **O₂, H₂, N₂, F₂, Cl₂, Br₂, I₂**) noble gases are **monoatomic** (e.g. **He, Ar, Ne, Xe**), Ozone gas (**O₃**) is **triatomic** e.g.

$$\begin{array}{l} \text{Molar mass Oxygen molecule(O}_2\text{)} = \text{relative molecular mass} = (16.0 \times 2)\text{g} = 32.0\text{g} \\ 6.023 \times 10^{23} \text{ particles of Oxygen molecule} = 1 \text{ mole} = 32.0 \text{ g} \end{array}$$

Molar mass **chlorine molecule**(Cl₂) = relative molecular mass = (35.5 x 2)g = 71.0g
 6.023 x 10²³ particles of chlorine molecule = 1 mole = 71.0 g

Molar mass **Nitrogen molecule**(N₂) = relative molecular mass = (14.0 x 2)g = 28.0g
 6.023 x 10²³ particles of Nitrogen molecule = 1 mole = 28.0 g

(ii) a **compound** has mass equal to relative formular mass /RFM (in grams) of the molecule. Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.

(i) Molar mass **Water**(H₂O) = relative formular mass = [(1.0 x 2) + 16.0]g = 18.0g

6.023 x 10²³ particles of Water molecule = 1 mole = 18.0 g

6.023 x 10²³ particles of Water molecule has:

- 2 x 6.023 x 10²³ particles of Hydrogen atoms

- 1 x 6.023 x 10²³ particles of Oxygen atoms

(ii) Molar mass **sulphuric(VI)acid**(H₂SO₄) = relative formular mass
 = [(1.0 x 2) + 32.0 + (16.0 x 4)]g = 98.0g

6.023 x 10²³ particles of sulphuric(VI)acid(H₂SO₄) = 1 mole = 98.0g

6.023 x 10²³ particles of sulphuric(VI)acid(H₂SO₄) has:

- 2 x 6.023 x 10²³ particles of Hydrogen atoms

- 1 x 6.023 x 10²³ particles of Sulphur atoms

- 4 x 6.023 x 10²³ particles of Oxygen atoms

(iii) Molar mass **sodium carbonate(IV)**(Na₂CO₃) = relative formular mass
 = [(23.0 x 2) + 12.0 + (16.0 x 3)]g = 106.0g

6.023 x 10²³ particles of sodium carbonate(IV)(Na₂CO₃) = 1 mole = 106.0g

6.023 x 10²³ particles of sodium carbonate(IV)(Na₂CO₃) has:

- 2 x 6.023 x 10²³ particles of Sodium atoms

- 1 x 6.023 x 10²³ particles of Carbon atoms

- 3 x 6.023 x 10²³ particles of Oxygen atoms

(iv) Molar mass **Calcium carbonate(IV)**(CaCO₃) = relative formular mass
 = [(40.0 + 12.0 + (16.0 x 3)]g = 100.0g.

6.023 x 10²³ particles of Calcium carbonate(IV)(CaCO₃) = 1 mole = 100.0g

6.023×10^{23} particles of Calcium carbonate(IV)(CaCO_3) has:

- $1 \times 6.023 \times 10^{23}$ particles of Calcium atoms
- $1 \times 6.023 \times 10^{23}$ particles of Carbon atoms
- $3 \times 6.023 \times 10^{23}$ particles of Oxygen atoms

(v) Molar mass **Water(H_2O)** = relative formula mass

$$= [(2 \times 1.0) + 16.0] \text{g} = 18.0 \text{g}$$

6.023×10^{23} particles of Water(H_2O) = 1 mole = 18.0g

6.023×10^{23} particles of Water(H_2O) has:

- $2 \times 6.023 \times 10^{23}$ particles of Hydrogen atoms
- $1 \times 6.023 \times 10^{23}$ particles of Oxygen atoms

Practice

1. Calculate the number of moles present in:

(i) 0.23 g of Sodium atoms

Molar mass of Sodium atoms = 23g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23 \text{g}}{23} = \mathbf{0.01 \text{moles}}$$

(ii) 0.23 g of Chlorine atoms

Molar mass of Chlorine atoms = 35.5 g

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23 \text{g}}{35.5} = \mathbf{0.0065 \text{moles} / 6.5 \times 10^{-3} \text{moles}}$$

(iii) 0.23 g of Chlorine molecules

Molar mass of Chlorine molecules = $(35.5 \times 2) = \mathbf{71.0 \text{g}}$

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23 \text{g}}{71} = \mathbf{0.0032 \text{moles} / 3.2 \times 10^{-3} \text{moles}}$$

(iv) 0.23 g of dilute sulphuric(VI) acid

Molar mass of $\text{H}_2\text{SO}_4 = [(2 \times 1) + 32 + (4 \times 14)] = \mathbf{98.0 \text{g}}$

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23 \text{g}}{98} = \mathbf{0.0023 \text{moles} / 2.3 \times 10^{-3} \text{moles}}$$

2. Calculate the number of atoms present in: (Avogadro's constant $L = 6.0 \times 10^{23}$)

(i) 0.23 g of dilute sulphuric (VI) acid

Method I

Molar mass of $\text{H}_2\text{SO}_4 = [(2 \times 1) + 32 + (4 \times 14)] = \mathbf{98.0 \text{g}}$

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23\text{g}}{98} = \mathbf{0.0023\text{moles}} / \mathbf{2.3 \times 10^{-3} \text{ moles}}$$

1 mole has 6.0×10^{23} atoms

$$2.3 \times 10^{-3} \text{ moles has } \frac{(2.3 \times 10^{-3} \times 6.0 \times 10^{23})}{1} = \mathbf{1.38 \times 10^{21} \text{ atoms}}$$

Method II

Molar mass of $\text{H}_2\text{SO}_4 = [(2 \times 1) + 32 + (4 \times 14)] = \mathbf{98.0\text{g}}$

$\mathbf{98.0\text{g}} = 1 \text{ mole has } 6.0 \times 10^{23} \text{ atoms}$

$$0.23 \text{ g therefore has } \frac{(0.23 \text{ g} \times 6.0 \times 10^{23})}{98} = \mathbf{1.38 \times 10^{21} \text{ atoms}}$$

(ii) 0.23 g of sodium carbonate(IV)decahydrate

Molar mass of $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} =$

$$[(2 \times 23) + 12 + (3 \times 16) + (10 \times 1.0) + (10 \times 16)] = \mathbf{276.0\text{g}}$$

Method I

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23\text{g}}{276} = \mathbf{0.00083\text{moles}} / \mathbf{8.3 \times 10^{-4} \text{ moles}}$$

1 mole has 6.0×10^{23} atoms

$$8.3 \times 10^{-4} \text{ moles has } \frac{(8.3 \times 10^{-4} \text{ moles} \times 6.0 \times 10^{23})}{1} = \mathbf{4.98 \times 10^{20} \text{ atoms}}$$

Method II

$276.0\text{g} = 1 \text{ mole has } 6.0 \times 10^{23} \text{ atoms}$

$$0.23 \text{ g therefore has } \frac{(0.23 \text{ g} \times 6.0 \times 10^{23})}{276.0} = \mathbf{4.98 \times 10^{20} \text{ atoms}}$$

(iii) 0.23 g of Oxygen gas

Molar mass of $\text{O}_2 = (2 \times 16) = \mathbf{32.0 \text{ g}}$

Method I

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} = > \frac{0.23\text{g}}{32} = \mathbf{0.00718\text{moles}} / \mathbf{7.18 \times 10^{-3} \text{ moles}}$$

1 mole has $\mathbf{2 \times 6.0 \times 10^{23} \text{ atoms}}$ in O_2

$$7.18 \times 10^{-3} \text{ moles has } \frac{(7.18 \times 10^{-3} \text{ moles} \times \mathbf{2 \times 6.0 \times 10^{23}})}{1} = \mathbf{8.616 \times 10^{21} \text{ atoms}}$$

Method II

$32.0\text{g} = 1 \text{ mole has } \mathbf{2 \times 6.0 \times 10^{23} \text{ atoms}}$ in O_2

$$0.23 \text{ g therefore has } \frac{(0.23 \text{ g} \times 2 \times 6.0 \times 10^{23})}{32.0} = 8.616 \times 10^{21} \text{ atoms}$$

(iv) 0.23 g of Carbon(IV)oxide gas

$$\text{Molar mass of CO}_2 = [12 + (2 \times 16)] = 44.0 \text{ g}$$

Method I

$$\text{Moles} = \frac{\text{mass in grams}}{\text{Molar mass}} \Rightarrow \frac{0.23 \text{ g}}{44} = 0.00522 \text{ moles} / 5.22 \times 10^{-3} \text{ moles}$$

1 mole has $3 \times 6.0 \times 10^{23}$ atoms in CO₂

$$7.18 \times 10^{-3} \text{ moles has } \frac{(5.22 \times 10^{-3} \text{ moles} \times 3 \times 6.0 \times 10^{23})}{1} = 9.396 \times 10^{21} \text{ atoms}$$

Method II

44.0g = 1 mole has $3 \times 6.0 \times 10^{23}$ atoms in CO₂

$$0.23 \text{ g therefore has } \frac{(0.23 \text{ g} \times 3 \times 6.0 \times 10^{23})}{44.0} = 9.409 \times 10^{21} \text{ atoms}$$

(c) Empirical and molecular formula

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

To determine the empirical formula of copper oxide**(a) Method 1: From copper to copper(II)oxide****Procedure.**

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

Sample results

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

Sample questions**1. Calculate the mass of copper powder used.**

$$\begin{array}{rcl}
 \text{Mass of crucible + copper before heating } (M_2) & = & 18.4 \\
 \text{Less Mass of crucible } (M_1) & & \underline{= - 15.6g} \\
 \text{Mass of copper} & & \mathbf{2.8 g}
 \end{array}$$

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

Method I

$$\text{Mass of crucible + copper after heating } (M_3) = 19.1g$$

$$\begin{aligned} \text{Mass of crucible + copper before heating (M}_2\text{)} &= - 18.4\text{g} \\ \text{Mass of Oxygen} &= \mathbf{0.7\text{ g}} \end{aligned}$$

Method II

$$\begin{aligned} \text{Mass of crucible + copper after heating (M}_3\text{)} &= 19.1\text{g} \\ \text{Mass of crucible} &= - 15.6\text{g} \\ \text{Mass of copper(II)Oxide} &= 3.5\text{ g} \\ \\ \text{Mass of copper(II)Oxide} &= 3.5\text{ g} \\ \text{Mass of copper} &= - 2.8\text{ g} \\ \text{Mass of Oxygen} &= \mathbf{0.7\text{ g}} \end{aligned}$$

3. Calculate the number of moles of:

(i) copper used (Cu = 63.5)

$$\text{number of moles of copper} = \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{2.8}{63.5} = \mathbf{0.0441\text{moles}}$$

(ii) Oxygen used (O = 16.0)

$$\text{number of moles of oxygen} = \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{0.7}{16.0} = \mathbf{0.0441\text{moles}}$$

4. Determine the mole ratio of the reactants

$$\begin{aligned} \text{Moles of copper} &= \frac{0.0441\text{moles}}{0.0441\text{moles}} = \frac{1}{1} \Rightarrow \text{Mole ratio Cu: O} = \mathbf{1:1} \end{aligned}$$

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

CuO (copper(II)oxide)

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

Observation

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

Explanation

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

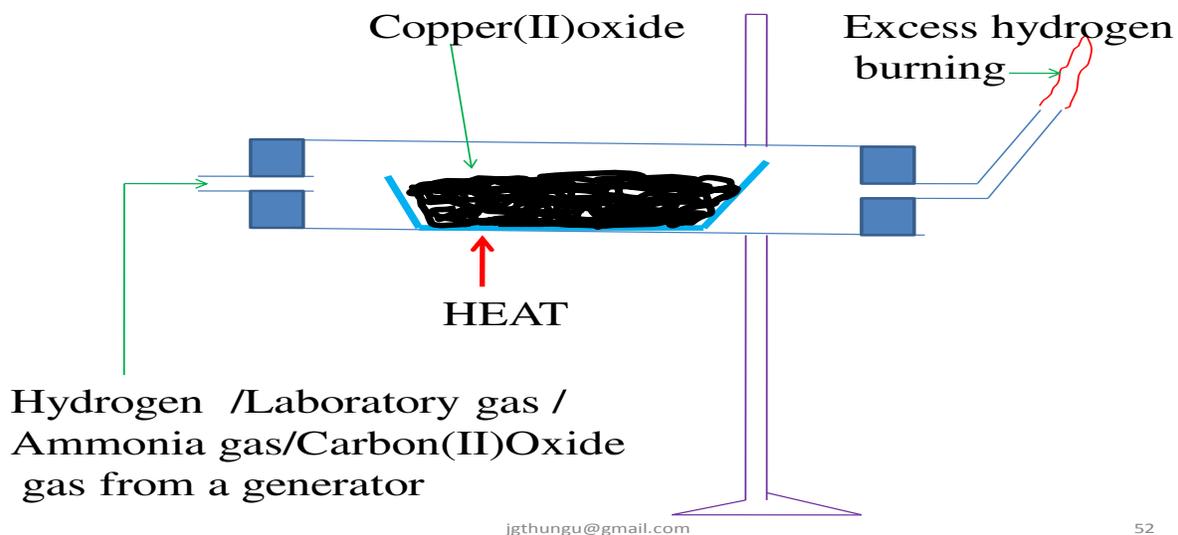
7. Explain why magnesium ribbon/shavings would be unsuitable in a similar experiment as the one above.

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

(b) Method 2: From copper(II)oxide to copper

Procedure.

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

Determining empirical formula from copper(II)oxide to copper

Pass slowly(to prevent copper(II)oxide from being blown away)a stream of either dry Hydrogen /ammonia/laboratory gas/ carbon(II)oxide gas for about two minutes from a suitable generator.

When all the in the apparatus set up is driven out ,heat the copper(II)oxide strongly for about five minutes until there is no further change. Stop heating.

Continue passing the gases until the glass tube is cool.

Turn off the gas generator.

Carefully remove the porcelain boat form the combustion tube.

Reweigh (M_3).

Sample results

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

Sample questions

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

$$\begin{array}{rcl}
 \text{Mass of boat before heating}(M_2) & = & 19.1 \\
 \text{Mass of empty boat}(M_1) & = & \underline{- 15.6\text{g}} \\
 \text{Mass of copper(II)Oxide} & & \mathbf{3.5\text{ g}}
 \end{array}$$

2. Calculate the mass of**(i) Oxygen.**

$$\begin{array}{rcl}
 \text{Mass of boat before heating}(M_2) & = & 19.1 \\
 \text{Mass of boat after heating } (M_3) & = & \underline{- 18.4\text{g}} \\
 \text{Mass of oxygen} & = & \mathbf{0.7\text{ g}}
 \end{array}$$

(ii)Copper

$$\begin{array}{rcl}
 \text{Mass of copper(II)Oxide} & = & 3.5\text{ g} \\
 \text{Mass of oxygen} & = & \underline{0.7\text{ g}} \\
 \text{Mass of oxygen} & = & \mathbf{2.8\text{ g}}
 \end{array}$$

3. Calculate the number of moles of:**(i) Copper used (Cu = 63.5)**

$$\text{number of moles of copper} = \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{2.8}{63.5} = \underline{\mathbf{0.0441\text{moles}}}$$

(ii) Oxygen used (O = 16.0)

$$\text{number of moles of oxygen} = \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{0.7}{16.0} = \underline{\mathbf{0.0441\text{moles}}}$$

4. Determine the mole ratio of the reactants

$$\begin{array}{rcl}
 \text{Moles of copper} & = & \frac{0.0441\text{moles}}{0.0441\text{moles}} = \frac{1}{1} \Rightarrow \text{Mole ratio Cu: O} = \mathbf{1:1} \\
 \text{Moles of oxygen} & &
 \end{array}$$

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

CuO (copper(II)oxide)

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

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

Explanation

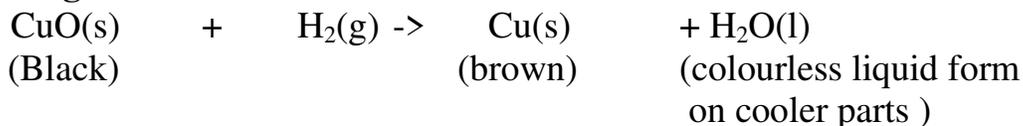
Copper(II)oxide powder is black. On heating it is reduced by a suitable reducing agent to brown copper metal.

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

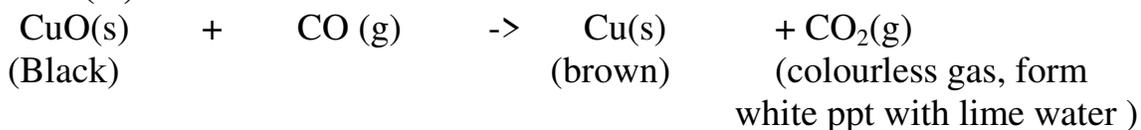
Magnesium is high in the reactivity series. None of the above reducing agents is strong enough to reduce the oxide to the metal.

8. Write the equation for the reaction that would take place when the reducing agent is:

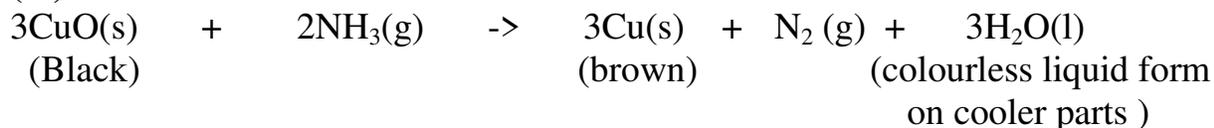
(i) Hydrogen



(ii) Carbon(II)oxide



(iii) Ammonia



9. Explain why the following is necessary during the above experiment;

(i) A stream of dry hydrogen gas should be passed before heating copper (II) Oxide.

Air combine with hydrogen in presence of heat causing an explosion

(ii) A stream of dry hydrogen gas should be passed after heating copper (II) Oxide has been stopped.

Hot metallic copper can be re-oxidized back to copper(II)oxide

(iii) A stream of excess carbon (II)oxide gas should be ignited to burn
Carbon (II)oxide is highly poisonous/toxic. On ignition it burns to form less toxic carbon (IV)oxide gas.

10. State two sources of error in this experiment.

(i) All copper(II)oxide may not be reduced to copper.

(ii) Some copper(II)oxide may be blown out the boat by the reducing agent.

4. Theoretically the empirical formula of a compound can be determined as in the following examples.

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

% of Oxygen = 100% - % of Copper => 100- 80 = **20%** of Oxygen

Element	Copper	Oxygen
Symbol	Cu	O
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{80}{63.5}$	$\frac{20}{16}$
Divide by the smallest value	$\frac{1.25}{1.25}$	$\frac{1.25}{1.25}$
Mole ratios	1	1

Empirical formula is **CuO**

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

Mass of Oxygen = 1.60 – 0.84 => **0.56 g** of Oxygen

Element	Magnesium	Oxygen
Symbol	Mg	O
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{0.84}{24}$	$\frac{0.56}{16}$
Divide by the smallest value	$\frac{0.35}{0.35}$	$\frac{0.35}{0.35}$
Mole ratios	1	1

Empirical formula is **MgO**

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

Mass of Oxygen = 100 – 47 => **53%** of Oxygen

Element	Silicon	Oxygen
Symbol	Si	O
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{47}{28}$	$\frac{53}{16}$
Divide by the smallest value	$\frac{1.68}{1.68}$	$\frac{3.31}{1.68}$
Mole ratios	1	1.94 = 2

Empirical formula is **SiO₂**

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

Mass of Oxygen = 100 - 47 => 53% of Oxygen

Element	Silicon	Oxygen
Symbol	Si	O
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{47}{28}$	$\frac{53}{16}$
Divide by the smallest value	$\frac{1.68}{1.68}$	$\frac{3.31}{1.68}$
Mole ratios	1	1.94 = 2

Empirical formula is SiO_2

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

Mass of evaporating dish = 300.0g

Mass of evaporating dish + hydrated salt = 305.0g

Mass of evaporating dish + anhydrous salt = 303.2g

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

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

Working

Mass of Hydrated salt = 305.0g - 300.0g = 5.0g

Mass of anhydrous salt = 303.2g - 300.0g = 3.2g

Mass of water in hydrated salt = 5.0g - 3.2g = 1.8g

Molar mass of water (H_2O) = **18.0g**

Molar mass of anhydrous copper (II) sulphate(VI) (CuSO_4) = **160.5g**

Element/compound	anhydrous copper (II) sulphate(VI)	Oxygen
Symbol	Si	O
Moles present = $\frac{\text{composition by mass}}{\text{Molar mass}}$	$\frac{3.2}{160.5}$	$\frac{1.8}{18}$
Divide by the smallest value	$\frac{0.0199}{0.0199}$	$\frac{0.1}{18}$
Mole ratios	1	5

The empirical formula of hydrated salt = $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

Hydrated salt has **five/5 molecules** of water of crystallizations

4. The molecular formula is the actual number of each kind of atoms present in a **molecule** of a compound.

The empirical formula of an ionic compound is the same as the chemical formula but for simple molecular structured compounds, the empirical formula may not be the same as the chemical formula.

The molecular formula is a multiple of empirical formula .It is determined from the relationship:

$$(i) \quad n = \frac{\text{Relative formular mass}}{\text{Relative empirical formula}}$$

where **n** is a whole number.

$$(ii) \text{ Relative empirical formula} \times n = \text{Relative formular mass}$$

where **n** is a whole number.

Practice sample examples

1. A hydrocarbon was found to contain 92.3% carbon and the remaining Hydrogen.

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

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

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

Empirical formula is **CH**

The molecular formular is thus determined :

$$n = \frac{\text{Relative formular mass}}{\text{Relative empirical formula}} = \frac{78}{13} = 6$$

Relative empirical formula

13

The molecular formula is $(\text{C H}) \times 6 = \underline{\text{C}_6\text{H}_6}$

2. A compound of carbon, hydrogen and oxygen contain 54.55% carbon, 9.09% and remaining 36.36% oxygen.

If its relative molecular mass is 88, determine its molecular formula(C=12.0, H =1.0, O= 16.0)

Element	Carbon	Hydrogen	Oxygen
Symbol	C	H	O
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{54.55}{12}$	$\frac{9.09}{1}$	$\frac{36.36}{16}$
Divide by the smallest value	$\frac{4.5458}{2.2725}$	$\frac{9.09}{2.2725}$	$\frac{2.2725}{2.2725}$
Mole ratios	2	4	1

Empirical formula is $\text{C}_2\text{H}_4\text{O}$

The molecular formula is thus determined :

$$n = \frac{\text{Relative formular mass}}{\text{Relative empirical formula}} = \frac{88}{44} = 2$$

The molecular formula is $(\text{C}_2\text{H}_4\text{O}) \times 2 = \text{C}_4\text{H}_8\text{O}_2$.

4.A hydrocarbon burns completely in excess air to form 5.28 g of carbon (IV) oxide and 2,16g of water.

If the molecular mass of the hydrocarbon is 84, draw and name its molecular structure.

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

$$\text{Mass of carbon in CO}_2 = \frac{\text{Mass of C in CO}_2}{\text{Molar mass of CO}_2} \times \text{mass of CO}_2 \Rightarrow \frac{12}{44} \times 5.28 = 1.44\text{g}\checkmark$$

$$\text{Mass of Hydrogen in H}_2\text{O} = \frac{\text{Mass of C in H}_2\text{O}}{\text{Molar mass of H}_2\text{O}} \times \text{mass of H}_2\text{O} \Rightarrow \frac{2}{18} \times 2.16 = 0.24\text{g}\checkmark$$

Element	Carbon	Hydrogen
---------	--------	----------

Symbol	C	H
Moles present = $\frac{\text{mass}}{\text{Molar mass}}$	$\frac{1.44\text{g}}{12}$	$\frac{0.24\text{g}}{1}$ ✓
Divide by the smallest value	$\frac{0.12}{0.12}$	$\frac{0.24}{0.12}$
Mole ratios	1	2 ✓

Empirical formula is **CH₂** ✓

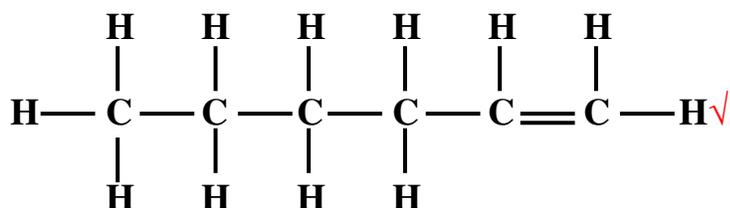
The molecular formula is thus determined :

$$n = \frac{\text{Relative formula mass}}{\text{Relative empirical formula}} = \frac{84}{14} = 6 \checkmark$$

The molecular formula is (CH₂) x 6 = **C₆H₁₂**. ✓

molecular name **Hexene** ✓ / **Hex-1-ene** (or any position isomer of Hexene)

Molecular structure



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

$$\text{Mass of N in A} = 5.2\% \times 0.085 = \underline{\underline{0.00442 \text{ g}}}$$

$$\text{Mass of C in A} = \frac{12}{44} \times 0.224 = \underline{\underline{0.0611\text{g}}}$$

$$\text{Mass of H in A} = \frac{2}{18} \times 0.0372 = \underline{\underline{0.0041\text{g}}}$$

$$\text{Mass of O in A} = 0.085\text{g} - 0.004442\text{g} = \underline{\underline{0.0806\text{g}}} \text{ (Mass of C,H,O)}$$

$$\Rightarrow 0.0611\text{g} + 0.0041\text{g} = \underline{\underline{0.0652\text{g}}} \text{ (Mass of C,H)}$$

$$0.0806\text{g} \text{ (Mass of C,H,O)} - 0.0652\text{g} \text{ (Mass of C,H)} = \underline{\underline{0.0154 \text{ g}}}$$

Element	Nitrogen	Carbon	Hydrogen	Oxygen
Symbol	N	C	H	O
Moles present = $\frac{\text{mass}}{\text{Molar mass}}$	$\frac{0.00442 \text{ g}}{14}$	$\frac{0.0611\text{g}}{12}$	$\frac{0.0041\text{g}}{1}$	$\frac{0.0154 \text{ g}}{16}$
Divide by the smallest value	$\frac{0.00032}{0.00032}$	$\frac{0.00509}{0.00032}$	$\frac{0.0041\text{g}}{0.00032}$	$\frac{0.00096}{0.00032}$

	0.00032	0.00032	0.00032	0.00032
Mole ratios	1	16	13	3

Empirical formula = $C_{16}H_{13}NO_3$

(d) Molar gas volume

The volume occupied by one mole of all gases at the same temperature and pressure is a constant. It is:

- (i) 24dm³/24litres/24000cm³ at room temperature(25°C/298K) and pressure(r.t.p).
i.e. 1mole of all gases =24dm³/24litres/24000cm³ at r.t.p

Examples

1mole of O₂ = 32g =6.0 x10²³ particles= 24dm³/24litres/24000cm³ at r.t.p

1mole of H₂ = 2g =6.0 x10²³ particles =24dm³/24litres/24000cm³ at r.t.p

1mole of CO₂ = 44g = 6.0 x10²³ particles =24dm³/24litres/24000cm³ at r.t.p

1mole of NH₃ = 17g =6.0 x10²³ particles = 24dm³/24litres/24000cm³ at r.t.p

1mole of CH₄ = 16g =6.0 x10²³ particles =24dm³/24litres/24000cm³ at r.t.p

- (ii)22.4dm³/22.4litres/22400cm³ at standard temperature(0°C/273K) and pressure(s.t.p)

i.e. 1mole of all gases =22.4dm³/22.4litres/22400cm³ at s.t.p

Examples

1mole of O₂ = 32g =6.0 x10²³ particles= 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of H₂ = 2g =6.0 x10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of CO₂ = 44g = 6.0 x10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of NH₃ = 17g =6.0 x10²³ particles= 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of CH₄ = 16g =6.0 x10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

The volume occupied by one mole of a gas at r.t.p or s.t.p is commonly called the **molar gas volume**. Whether the molar gas volume is at r.t.p or s.t.p must always be **specified**.

From the above therefore a less or more volume can be determined as in the examples below.

Practice examples

1. Calculate the number of particles present in:

(Avogadro's constant =6.0 x10²³ mole⁻¹)

(i) 2.24dm³ of Oxygen.

$$\begin{array}{l} 22.4\text{dm}^3 \quad \rightarrow \quad 6.0 \times 10^{23} \\ 2.24\text{dm}^3 \quad \rightarrow \quad \frac{2.24 \times 6.0 \times 10^{23}}{22.4} \end{array}$$

$$= \underline{6.0 \times 10^{22}} \text{ molecules} = 2 \times 6.0 \times 10^{22} = \underline{1.2 \times 10^{23}} \text{ atoms}$$

(ii) 2.24dm³ of Carbon(IV)oxide.

$$\begin{array}{l} 22.4\text{dm}^3 \quad \rightarrow \quad 6.0 \times 10^{23} \\ 2.24\text{dm}^3 \quad \rightarrow \quad \frac{2.24 \times 6.0 \times 10^{23}}{22.4} \end{array}$$

$$= \underline{6.0 \times 10^{22}} \text{ molecules} = (\text{CO}_2) = 3 \times 6.0 \times 10^{22} = \underline{1.8 \times 10^{23}} \text{ atoms}$$

2. 0.135 g of a gaseous hydrocarbon X on complete combustion produces 0.41g of carbon(IV)oxide and 0.209g of water.0.29g of X occupy 120cm³ at room temperature and 1 atmosphere pressure .Name X and draw its molecular structure.(C=12.0,O= 16.0,H=1.0,1 mole of gas occupies 24dm³ at r.t.p)

$$\begin{array}{l} \text{Molar mass CO}_2 = 44 \text{ gmole}^{-1} \checkmark \\ \text{Molar mass H}_2\text{O} = 18 \text{ gmole}^{-1} \checkmark \\ \text{Molar mass X} = \frac{0.29 \times (24 \times 1000)\text{cm}^3}{120\text{cm}^3} = 58 \text{ gmole}^{-1} \checkmark \end{array}$$

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

$$\begin{array}{l} \text{Mass of carbon in CO}_2 = \frac{\text{Mass of C in CO}_2}{\text{Molar mass of CO}_2} \times \text{mass of CO}_2 \Rightarrow \\ \frac{12}{44} \times 0.41 = \mathbf{0.1118g} \checkmark \end{array}$$

$$\begin{array}{l} \text{Mass of Hydrogen in H}_2\text{O} = \frac{\text{Mass of H in H}_2\text{O}}{\text{Molar mass of H}_2\text{O}} \times \text{mass of H}_2\text{O} \Rightarrow \\ \frac{2}{18} \times 0.209 = \mathbf{0.0232g} \checkmark \end{array}$$

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = $\frac{\% \text{ composition}}{\text{Molar mass}}$	$\frac{0.1118}{12}$	$\frac{0.0232g}{1} \checkmark$
Divide by the smallest value	$\frac{0.0093}{0.0093}$	$\frac{0.0232}{0.0093} \checkmark$
Mole ratios	1 x2	2.5x2

	2	5✓
--	---	----

Empirical formula is C_2H_5 ✓

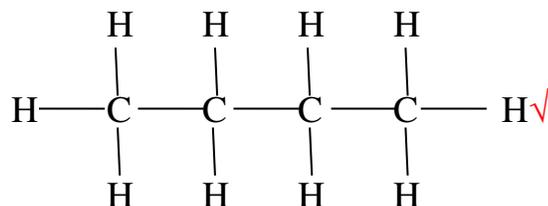
The molecular formula is thus determined :

$$n = \frac{\text{Relative formula mass}}{\text{Relative empirical formula}} = \frac{58}{29} = 2✓$$

The molecular formula is $(\text{C}_2\text{H}_5) \times 2 = \text{C}_4\text{H}_{10}$.✓

Molecule name **Butane**

Molecular structure

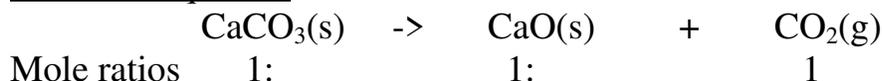


(e) Gravimetric analysis

Gravimetric analysis is the relationship between reacting masses and the volumes and /or masses of products. All reactants are in mole **ratios** to their products in accordance to their stoichiometric equation. Using the mole ratios of reactants and products any volume and/or mass can be determined as in the examples:

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

Chemical equation



Molar Mass $\text{CaCO}_3 = 100\text{g}$

Method 1

100g $\text{CaCO}_3(\text{s}) \rightarrow 24\text{dm}^3 \text{CO}_2(\text{g})$ at r.t.p

5.0 g $\text{CaCO}_3(\text{s}) \rightarrow \frac{5.0 \text{ g} \times 24\text{dm}^3}{100\text{g}} = \underline{\underline{1.2\text{dm}^3/1200\text{cm}^3}}$

Method 2

$$\text{Moles of } 5.0 \text{ g CaCO}_3(\text{s}) = \frac{5.0 \text{ g}}{100 \text{ g}} = \mathbf{0.05} \text{ moles}$$

Mole ratio 1:1

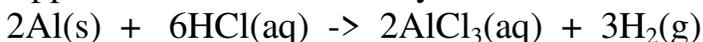
$$\text{Moles of CO}_2(\text{g}) = \mathbf{0.05} \text{ moles}$$

$$\text{Volume of CO}_2(\text{g}) = 0.05 \times 24000 \text{ cm}^3 = \mathbf{1200 \text{ cm}^3 / 1.2 \text{ dm}^3}$$

2. 1.0g of an alloy of aluminium and copper were reacted with excess hydrochloric acid. If 840cm³ of hydrogen at s.t.p was produced, calculate the % of copper in the alloy. (Al =27.0, one mole of a gas at s.t.p =22.4dm³)

Chemical equation

Copper does not react with hydrochloric acid



Method 1

$$3\text{H}_2(\text{g}) = 3 \text{ moles} \times (22.4 \times 1000) \text{ cm}^3 \Rightarrow 2 \times 27 \text{ g Al}$$

$$840 \text{ cm}^3 \Rightarrow \frac{840 \text{ cm}^3 \times 2 \times 27}{3 \times 22.4 \times 1000} = \mathbf{0.675 \text{ g}} \text{ of Aluminium}$$

$$\text{Total mass of alloy} - \text{mass of aluminium} = \text{mass of copper}$$

$$\Rightarrow 1.0 \text{ g} - 0.675 \text{ g} = \mathbf{0.325 \text{ g}} \text{ of copper}$$

$$\% \text{ copper} = \frac{\text{mass of copper} \times 100\%}{\text{Mass of alloy}} = \mathbf{32.5\%}$$

Method 2

Mole ratio $2\text{Al} : 3\text{H}_2 = 2:3$

$$\text{Moles of Hydrogen gas} = \frac{\text{volume of gas}}{\text{Molar gas volume}} \Rightarrow \frac{840 \text{ cm}^3}{22400 \text{ cm}^3} = \mathbf{0.0375 \text{ moles}}$$

$$\text{Moles of Al} = \frac{2}{3} \text{ moles of H}_2 \Rightarrow \frac{2}{3} \times 0.0375 \text{ moles} = \mathbf{0.025 \text{ moles}}$$

$$\text{Mass of Al} = \text{moles} \times \text{molar mass} \Rightarrow 0.025 \text{ moles} \times 27 = \mathbf{0.675 \text{ g}}$$

$$\text{Total mass of alloy} - \text{mass of aluminium} = \text{mass of copper}$$

$$\Rightarrow 1.0 \text{ g} - 0.675 \text{ g} = \mathbf{0.325 \text{ g}} \text{ of copper}$$

$$\% \text{ copper} = \frac{\text{mass of copper} \times 100\%}{\text{Mass of alloy}} = \mathbf{32.5\%}$$

(f) Gay Lussac's law

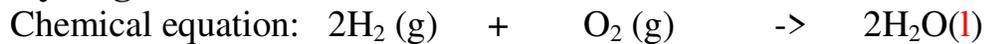
Gay Lussacs law states that **“when gases combine/react they do so in simple volume ratios to each other and to their gaseous products at constant/same temperature and pressure”**

Gay Lussacs law thus only apply to gases

Given the volume of one gas reactant, the other gaseous reactants can be deduced thus:

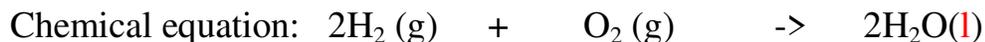
Examples

1. Calculate the volume of Oxygen required to completely react with 50cm³ of Hydrogen.



50cm³ of Oxygen is used

2. Calculate the volume of air required to completely reacts with 50cm³ of Hydrogen.(assume Oxygen is 21% by volume of air)



When the gaseous mixture was passed through, potassium hydroxide its volume decreased to 25cm³.

(a) What volume of Oxygen was used during the reaction. (1mk)

$$\text{Volume of Oxygen used} = 100 - 25 = \underline{75\text{cm}^3} \checkmark$$

(P was completely burnt)

(b) Determine the molecular formula of the hydrocarbon (2mk)



$$\underline{15\text{cm}^3} : \underline{75\text{cm}^3}$$

$$\underline{15} \quad \underline{15}$$

$$1 : 3 \checkmark$$

=> 1 atom of C react with 6 (3x2) atoms of Oxygen

Thus x = 1 and y = 2 => P has molecular formula CH₄ ✓

(g) Ionic equations

An ionic equation is a chemical statement showing the movement of ions (cations and anions) from reactants to products.

Solids, gases and liquids do not ionize/dissociate into free ions. **Only** ionic compounds in **aqueous/solution** or **molten** state ionize/dissociate into free cations and anions (**ions**)

An ionic equation is usually derived from a stoichiometric equation by using the following guidelines

Guidelines for writing ionic equations

1. Write the balanced stoichiometric equation
2. Indicate the state symbols of the reactants and products
3. **Split** into cations and anions all the reactants and products that exist in **aqueous** state.
4. **Cancel out** any cation and anion that appear on **both** the product and reactant side.
5. Rewrite the chemical equation. It is an ionic equation.

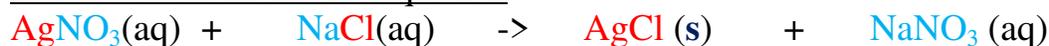
Practice

(a) Precipitation of an insoluble salt

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

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

Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side

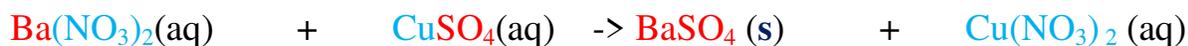


Rewrite the equation

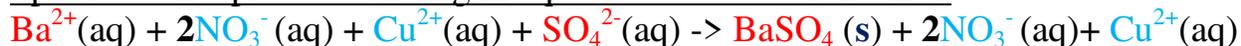


2. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

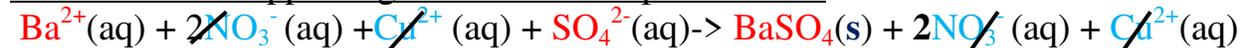
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side



Rewrite the equation

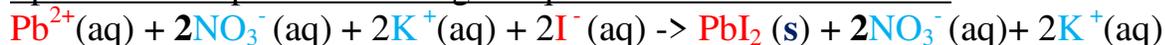


3.A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.

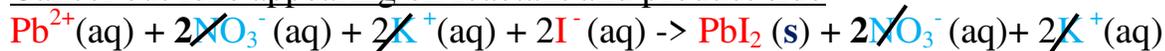
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side



Rewrite the equation



(b)Neutralization

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

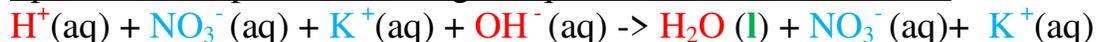
(i)Reaction of alkalis with acids

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

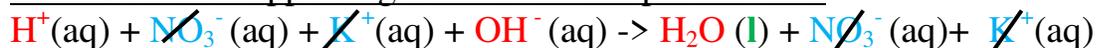
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side



Rewrite the equation



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

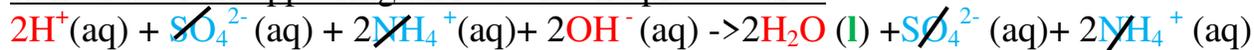
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side

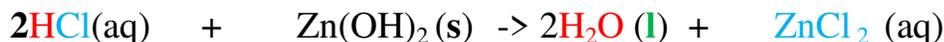


Rewrite the equation

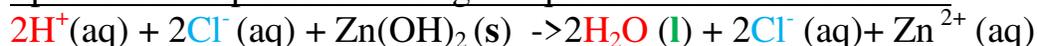


3. Reaction of hydrochloric acid with Zinc hydroxide

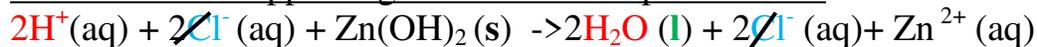
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side



Rewrite the equation



(h) Molar solutions

A molar solution is one whose concentration is known. The SI unit of concentration is **Molarity** denoted **M**.

Molarity may be defined as the number of moles of solute present in **one cubic decimeter** of solution.

One cubic decimeter is equal to **one litre** and also equal to **1000cm³**.

The higher the molarity the higher the concentration and the higher/more solute has been dissolved in the solvent to make one cubic decimeter/ litre/1000cm³

solution.

Examples

2M sodium hydroxide means 2 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm³ uniform **solution** mixture of sodium hydroxide and water.

0.02M sodium hydroxide means 0.02 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm³ uniform **solution** mixture of sodium hydroxide and water.

“2M” is **more concentrated** than “0.02M”.

Preparation of molar solution

Procedure

Weigh accurately 4.0 g of sodium hydroxide pellets into a 250cm³ volumetric flask.

Using a wash bottle add about 200cm³ of distilled water.

Stopper the flask.

Shake vigorously for three minutes.

Remove the stopper for a second then continue to shake for about another two minutes until **all** the solid has dissolved.

Add more water slowly upto **exactly** the 250 cm³ **mark**.

Sample questions

1. Calculate the number of moles of sodium hydroxide pellets present in:

(i) 4.0 g.

$$\text{Molar mass of NaOH} = (23 + 16 + 1) = 40\text{g}$$

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \quad \Rightarrow \quad \frac{4.0}{40} = \mathbf{0.1 / 1.0 \times 10^{-1} \text{ moles}}$$

(ii) 250 cm³ solution in the volumetric flask.

$$\text{Moles in 250 cm}^3 = \mathbf{0.1 / 1.0 \times 10^{-1} \text{ moles}}$$

(iii) one decimeter of solution

Method 1

$$\text{Moles in decimeters} = \mathbf{Molarity} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}}$$

$$\Rightarrow \frac{1.0 \times 10^{-1} \text{ moles} \times 1000\text{cm}^3}{250\text{cm}^3} =$$

$$= \mathbf{0.4 \text{ M} / 0.4 \text{ molesdm}^{-3}}$$

Method 2

250cm³ solution contain 1.0 x 10⁻¹ moles

$$1000\text{cm}^3 \text{ solution} = \text{Molarity} \text{ contain } \frac{1000 \times 1.0 \times 10^{-1} \text{ moles}}{250 \text{ cm}^3}$$

$$= \mathbf{0.4 \text{ M} / 0.4 \text{ molesdm}^{-3}}$$

Theoretical sample practice

1. Calculate the molarity of a solution containing:

(i) 4.0 g sodium hydroxide dissolved in 500cm³ solution

$$\text{Molar mass of NaOH} = (23 + 16 + 1) = 40\text{g}$$

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \quad \Rightarrow \quad \frac{4.0}{40} = \mathbf{0.1 / 1.0 \times 10^{-1} \text{ moles}}$$

Method 1

$$\begin{aligned} \text{Moles in decimeters} &= \mathbf{\text{Molarity}} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}} \\ \Rightarrow &\frac{1.0 \times 10^{-1} \text{ moles} \times 1000\text{cm}^3}{500\text{cm}^3} \\ &= \mathbf{\underline{0.2 \text{ M} / 0.2 \text{ molesdm}^{-3}}} \end{aligned}$$

Method 2

$$\begin{aligned} 500 \text{ cm}^3 \text{ solution} &\text{ contain } 1.0 \times 10^{-1} \text{ moles} \\ 1000\text{cm}^3 \text{ solution} &= \text{Molarity} \text{ contain } \frac{1000 \times 1.0 \times 10^{-1} \text{ moles}}{500 \text{ cm}^3} \\ &= \mathbf{\underline{0.2 \text{ M} / 0.2 \text{ molesdm}^{-3}}} \end{aligned}$$

(ii) 5.3 g anhydrous sodium carbonate dissolved in 50cm³ solution

$$\text{Molar mass of Na}_2\text{CO}_3 = (23 \times 2 + 12 + 16 \times 3) = 106 \text{ g}$$

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \quad \Rightarrow \quad \frac{5.3}{106} = \mathbf{0.05 / 5.0 \times 10^{-2} \text{ moles}}$$

Method 1

$$\begin{aligned} \text{Moles in decimeters} &= \mathbf{\text{Molarity}} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}} \\ \Rightarrow &\frac{1.0 \text{ moles} \times 1000\text{cm}^3}{50\text{cm}^3} \\ &= \mathbf{\underline{1.0 \text{ M}}} \end{aligned}$$

Method 2

$$\begin{aligned} 50 \text{ cm}^3 \text{ solution} &\text{ contain } 5.0 \times 10^{-2} \text{ moles} \\ 1000\text{cm}^3 \text{ solution} &= \text{Molarity} \text{ contain } \frac{1000 \times 5.0 \times 10^{-2} \text{ moles}}{50 \text{ cm}^3} \\ &= \mathbf{\underline{1.0\text{M} / 1.0 \text{ molesdm}^{-3}}} \end{aligned}$$

(iii) 5.3 g hydrated sodium carbonate decahydrate dissolved in 50cm³ solution

Molar mass of Na₂CO₃·10H₂O = (23 x 2 + 12 + 16 x 3 + 20 x 1 + 10 x 16) = 286g

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{5.3}{286} = 0.0185 / 1.85 \times 10^{-2} \text{ moles}$$

Method 1

$$\begin{aligned} \text{Moles in decimeters} &= \text{Molarity} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}} \\ &\Rightarrow \frac{1.85 \times 10^{-2} \text{ moles} \times 1000\text{cm}^3}{50\text{cm}^3} \\ &= \underline{\underline{0.37 \text{ M}/0.37 \text{ molesdm}^{-3}}} \end{aligned}$$

Method 2

50 cm³ solution contain 1.85 x 10⁻² moles

$$\begin{aligned} 1000\text{cm}^3 \text{ solution} &= \text{Molarity} \text{ contain } \frac{1000 \times 1.85 \times 10^{-2} \text{ moles}}{50 \text{ cm}^3} \\ &= \underline{\underline{3.7 \times 10^{-1} \text{ M} / 3.7 \times 10^{-1} \text{ molesdm}^{-3}}} \end{aligned}$$

(iv) 7.1 g of anhydrous sodium sulphate(VI) was dissolved in 20.0 cm³ solution. Calculate the molarity of the solution.

Method 1

20.0cm³ solution -> 7.1 g

$$1000\text{cm}^3 \text{ solution} \rightarrow \frac{1000}{20} \times 7.1 = \underline{\underline{3550 \text{ g dm}^{-3}}}$$

Molar mass Na₂SO₄ = 142 g

$$\text{Moles dm}^{-3} = \text{Molarity} = \frac{\text{Mass}}{\text{Molar mass}} \frac{3550}{142} = \underline{\underline{2.5 \text{ M}/ \text{molesdm}^{-3}}}$$

Method 2

Molar mass Na₂SO₄ = 142 g

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{7.1}{142} = 0.05 / 5.0 \times 10^{-2} \text{ moles}$$

Method 2(a)

$$\begin{aligned} \text{Moles in decimeters} &= \text{Molarity} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}} \\ &\Rightarrow \frac{5.0 \times 10^{-2} \text{ moles} \times 1000\text{cm}^3}{20\text{cm}^3} \\ &= \underline{\underline{2.5 \text{ M}/2.5 \text{ molesdm}^{-3}}} \end{aligned}$$

Method 2(b)

20 cm³ solution contain 5.0×10^{-2} moles

$$1000\text{cm}^3 \text{ solution} = \text{Molarity} \text{ contain } \frac{1000 \times 5.0 \times 10^{-2} \text{ moles}}{20 \text{ cm}^3}$$

$$= \underline{\underline{2.5 \text{ M}/2.5 \text{ molesdm}^{-3}}}$$

(iv) The density of sulphuric(VI) is 1.84gcm^{-3} Calculate the molarity of the acid.

Method 1

1.0cm³ solution \rightarrow 1.84 g

$$1000\text{cm}^3 \text{ solution} \rightarrow \frac{1000}{1} \times 1.84 = \underline{\underline{1840 \text{ g dm}^{-3}}}$$

Molar mass $\text{H}_2\text{SO}_4 = 98 \text{ g}$

$$\text{Moles dm}^{-3} = \text{Molarity} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{1840}{98}$$

$$= \underline{\underline{18.7755 \text{ M/ molesdm}^{-3}}}$$

Method 2

Molar mass $\text{H}_2\text{SO}_4 = 98 \text{ g}$

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{1.84}{98} = \underline{\underline{0.0188 / 1.88 \times 10^{-2} \text{ moles}}}$$

Method 2(a)

$$\text{Moles in decimeters} = \text{Molarity} = \frac{\text{Moles} \times 1000\text{cm}^3/1\text{dm}^3}{\text{Volume of solution}}$$

$$\Rightarrow \frac{1.88 \times 10^{-2} \text{ moles} \times 1000\text{cm}^3}{1.0\text{cm}^3}$$

$$= \underline{\underline{18.8\text{M}/18.8 \text{ molesdm}^{-3}}}$$

Method 2(b)

20 cm³ solution contain 1.88×10^{-2} moles

$$1000\text{cm}^3 \text{ solution} = \text{Molarity} \text{ contain } \frac{1000 \times 1.88 \times 10^{-2} \text{ moles}}{1.0 \text{ cm}^3}$$

$$= \underline{\underline{18.8\text{M}/18.8 \text{ molesdm}^{-3}}}$$

2. Calculate the mass of :

(i) 25 cm³ of 0.2M sodium hydroxide solution(Na =23.0.O =16.0, H=1.0)

Molar mass NaOH = **40g**

$$\text{Moles in } 25 \text{ cm}^3 = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.2 \times 25}{1000} = \underline{\underline{0.005/5.0 \times 10^{-3}} \text{ moles}}$$

$$\text{Mass of NaOH} = \text{Moles} \times \text{molar mass} = 5.0 \times 10^{-3} \times 40 = \underline{\underline{0.2 \text{ g}}}$$

(ii) 20 cm³ of 0.625 M sulphuric(VI)acid (S =32.0.O =16.0, H=1.0)

$$\text{Molar mass H}_2\text{SO}_4 = \underline{\underline{98\text{g}}}$$

$$\text{Moles in } 20 \text{ cm}^3 = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.625 \times 20}{1000} = \underline{\underline{0.0125/1.25.0 \times 10^{-3}} \text{ moles}}$$

$$\text{Mass of H}_2\text{SO}_4 = \text{Moles} \times \text{molar mass} \Rightarrow 5.0 \times 10^{-3} \times 40 = \underline{\underline{0.2 \text{ g}}}$$

(iii) 1.0 cm³ of 2.5 M Nitric(V)acid (N =14.0.O =16.0, H=1.0)

$$\text{Molar mass HNO}_3 = \underline{\underline{63 \text{ g}}}$$

$$\text{Moles in } 1 \text{ cm}^3 = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{2.5 \times 1}{1000} = \underline{\underline{0.0025 / 2.5. \times 10^{-3}} \text{ moles}}$$

$$\text{Mass of HNO}_3 = \text{Moles} \times \text{molar mass} \Rightarrow 2.5 \times 10^{-3} \times 40 = \underline{\underline{0.1 \text{ g}}}$$

3. Calculate the volume required to dissolve :

(a)(i) 0.25moles of sodium hydroxide solution to form a 0.8M solution

$$\text{Volume (in cm}^3\text{)} = \frac{\text{moles} \times 1000}{\text{Molarity}} \Rightarrow \frac{0.25 \times 1000}{0.8} = \underline{\underline{312.5\text{cm}^3}}$$

(ii) 100cm³ was added to the sodium hydroxide solution above. Calculate the concentration of the solution.

$$C_1 \times V_1 = C_2 \times V_2 \text{ where:}$$

$$C_1 = \text{molarity/concentration before diluting/adding water}$$

$$C_2 = \text{molarity/concentration after diluting/adding water}$$

$$V_1 = \text{volume before diluting/adding water}$$

$$V_2 = \text{volume after diluting/adding water}$$

$$\Rightarrow 0.8\text{M} \times 312.5\text{cm}^3 = C_2 \times (312.5 + 100)$$

$$C_2 = \underline{\underline{0.8\text{M} \times 312.5\text{cm}^3 = 0.6061\text{M}}}$$

412.5

(b)(ii) 0.01M solution containing 0.01moles of sodium hydroxide solution .

$$\text{Volume (in cm}^3\text{)} = \frac{\text{moles} \times 1000}{\text{Molarity}} \Rightarrow \frac{0.01 \times 1000}{0.01} = \underline{\underline{1000 \text{ cm}^3}}$$

(ii) Determine the quantity of water which must be added to the sodium hydroxide solution above to form a 0.008M solution.

$$C_1 \times V_1 = C_2 \times V_2 \text{ where:}$$

C_1 = molarity/concentration before diluting/adding water

C_2 = molarity/concentration after diluting/adding water

V_1 = volume before diluting/adding water

V_2 = volume after diluting/adding water

$$\Rightarrow 0.01\text{M} \times 1000 \text{ cm}^3 = 0.008 \times V_2$$

$$V_2 = \frac{0.01\text{M} \times 1000\text{cm}^3}{0.008} = \underline{\underline{1250\text{cm}^3}}$$

$$\text{Volume added} = 1250 - 1000 = \underline{\underline{250\text{cm}^3}}$$

(c) Volumetric analysis/Titration

Volumetric analysis/Titration is the process of determining unknown concentration of one reactant from a known concentration and volume of another.

Reactions take place in simple mole ratio of reactants and products.

Knowing the concentration/ volume of one reactant, the other can be determined from the relationship:

$$\frac{\underline{M_1 V_1}}{n_1} = \frac{\underline{M_2 V_2}}{n_2} \text{ where:}$$

M_1 = Molarity of 1st reactant

M_2 = Molarity of 2nd reactant

V_1 = Volume of 1st reactant

V_2 = Volume of 2nd reactant

n_1 = number of moles of 1st reactant from stoichiometric equation

n_2 = number of moles of 2nd reactant from stoichiometric equation

Examples

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

Stoichiometric equation: $\text{MCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Method 1

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2} \quad \rightarrow \quad \frac{M_1 \times 5.0 \text{ cm}^3}{1} = \frac{0.5 \text{ M} \times 25.0 \text{ cm}^3}{2}$$

$$\Rightarrow M_1 = \frac{0.5 \times 25.0 \times 1}{5.0 \times 2} = \underline{\underline{1.25 \text{ M} / 1.25 \text{ moledm}^{-3}}}$$

Method 2

Moles of HCl used = $\frac{\text{molarity} \times \text{volume}}{1000}$

$$\Rightarrow \frac{0.5 \times 25.0}{1000} = \underline{\underline{0.0125 / 1.25 \times 10^{-2} \text{ moles}}}$$

Mole ratio $\text{MCO}_3 : \text{HCl} = 1:2$

$$\text{Moles MCO}_3 = \frac{0.0125 / 1.25 \times 10^{-2} \text{ moles}}{2} = \underline{\underline{0.00625 / 6.25 \times 10^{-3} \text{ moles}}}$$

$$\text{Molarity MCO}_3 = \frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$$

$$= \underline{\underline{1.25 \text{ M} / 1.25 \text{ moledm}^{-3}}}$$

2. 2.0cm³ of 0.5M hydrochloric acid react with 0.1M of M₂CO₃. Calculate the volume of 0.1M M₂CO₃ used.

Stoichiometric equation: $\text{M}_2\text{CO}_3 (\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{MCl} (\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Method 1

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2} \quad \rightarrow \quad \frac{0.5 \times 2.0 \text{ cm}^3}{2} = \frac{0.1 \text{ M} \times V_2 \text{ cm}^3}{1}$$

$$\Rightarrow V_2 = \frac{0.5 \times 2.0 \times 1}{0.1 \times 2} = \underline{\underline{1.25 \text{ M} / 1.25 \text{ moledm}^{-3}}}$$

Method 2

Moles of HCl used = $\frac{\text{molarity} \times \text{volume}}{1000}$

$$\Rightarrow \frac{0.5 \times 2.0}{1000} = \underline{\underline{0.0125 / 1.25 \times 10^{-2} \text{ moles}}}$$

Mole ratio $\text{M}_2\text{CO}_3 : \text{HCl} = 1:2$

$$\text{Moles M}_2\text{CO}_3 = \frac{0.0125 / 1.25 \times 10^{-2} \text{ moles}}{2} = \underline{\underline{0.00625 / 6.25 \times 10^{-3} \text{ moles}}}$$

$$\text{Molarity M}_2\text{CO}_3 = \frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$$

$$= \underline{1.25\text{M} / 1.25 \text{ moledm}^{-3}}$$

3. 5.0cm³ of 0.1M sodium iodide react with 0.1M of Lead(II)nirate(V). Calculate(i) the volume of Lead(II)nirate(V) used.

(ii)the mass of Lead(II)Iodide formed
(Pb=207.0, I =127.0)

Stoichiometric equation: $2\text{NaI}(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{NaNO}_3(\text{aq}) + \text{PbI}_2(\text{s})$

(i)Volume of Lead(II)nirate(V) used

Method 1

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2} \quad \rightarrow \quad \frac{5 \times 0.1\text{cm}^3}{2} = \frac{0.1\text{M} \times V_2\text{cm}^3}{1}$$

$$\Rightarrow V_2 = \frac{0.1 \times 5.0 \times 1}{0.1 \times 2} = \underline{1.25\text{M} / 1.25 \text{ moledm}^{-3}}$$

Method 2

Moles of HCl used = $\frac{\text{molarity} \times \text{volume}}{1000}$

$$\Rightarrow \frac{0.1 \times 5.0}{1000} = \underline{0.0125 / 1.25 \times 10^{-2} \text{ moles}}$$

Mole ratio $\text{M}_2\text{CO}_3 : \text{HCl} = 1:2$

$$\text{Moles } \text{M}_2\text{CO}_3 = \frac{0.0125 / 1.25 \times 10^{-2} \text{ moles}}{2} = \underline{0.00625 / 6.25 \times 10^{-3} \text{ moles}}$$

Molarity $\text{M}_2\text{CO}_3 = \frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$

$$= \underline{1.25\text{M} / 1.25 \text{ moledm}^{-3}}$$

4. 0.388g of a monobasic organic acid B required 46.5 cm³ of 0.095M sodium hydroxide for complete neutralization. Name and draw the structural formula of B

Moles of NaOH used = $\frac{\text{molarity} \times \text{volume}}{1000}$

$$\Rightarrow \frac{0.095 \times 46.5}{1000} = \underline{0.0044175 / 4.4175 \times 10^{-3} \text{ moles}}$$

Mole ratio B : NaOH = 1:1

$$\text{Moles B} = \underline{\underline{0.0044175 / 4.4175 \times 10^{-3} \text{ moles}}}$$

$$\begin{aligned} \text{Molar mass B} &= \frac{\text{mass}}{\text{moles}} \Rightarrow \frac{0.388}{0.0044175 / 4.4175 \times 10^{-3} \text{ moles}} \\ &= \underline{\underline{87.8324 \text{ gmole}^{-1}}} \end{aligned}$$

X-COOH = 87.8324 where X is an alkyl group

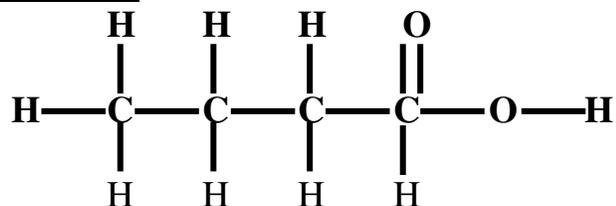
$$X = 87.8324 - 42 = 42.8324 = 43$$

By elimination: $\text{CH}_3 = 15$ $\text{CH}_3\text{CH}_2 = 29$ $\text{CH}_3\text{CH}_2\text{CH}_2 = 43$

Molecular formula : $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$

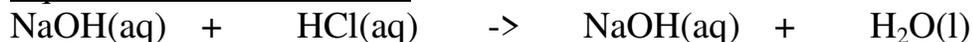
Molecule name : Butan-1-oic acid

Molecular structure



5. 10.5 g of an impure sample containing ammonium sulphate (VI) fertilizer was warmed with 250cm³ of 0.8M sodium hydroxide solution. The excess of the alkali was neutralized by 85cm³ of 0.5M hydrochloric acid. Calculate the % of impurities in the ammonium sulphate (VI) fertilizer. (N=14.0, S=32.0, O=16.0, H=1.0)

Equation for neutralization

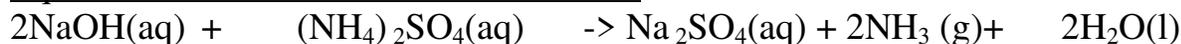


Mole ratio NaOH(aq):HCl(aq) = 1:1

$$\text{Moles of HCl} = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.5 \times 85}{1000} = \underline{\underline{0.0425 \text{ moles}}}$$

Excess moles of NaOH(aq) = **0.0425 moles**

Equation for reaction with ammonium salt



Mole ratio NaOH(aq): (NH₄)₂SO₄(aq) = 2:1

$$\text{Total moles of NaOH} = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.8 \times 250}{1000} = \underline{\underline{0.2 \text{ moles}}}$$

$$\text{Moles of NaOH that reacted with } (\text{NH}_4)_2\text{SO}_4 = 0.2 - 0.0425 = \underline{\underline{0.1575 \text{ moles}}}$$

$$\text{Moles } (\text{NH}_4)_2\text{SO}_4 = \frac{1}{2} \times 0.1575 \text{ moles} = \underline{\underline{0.07875 \text{ moles}}}$$

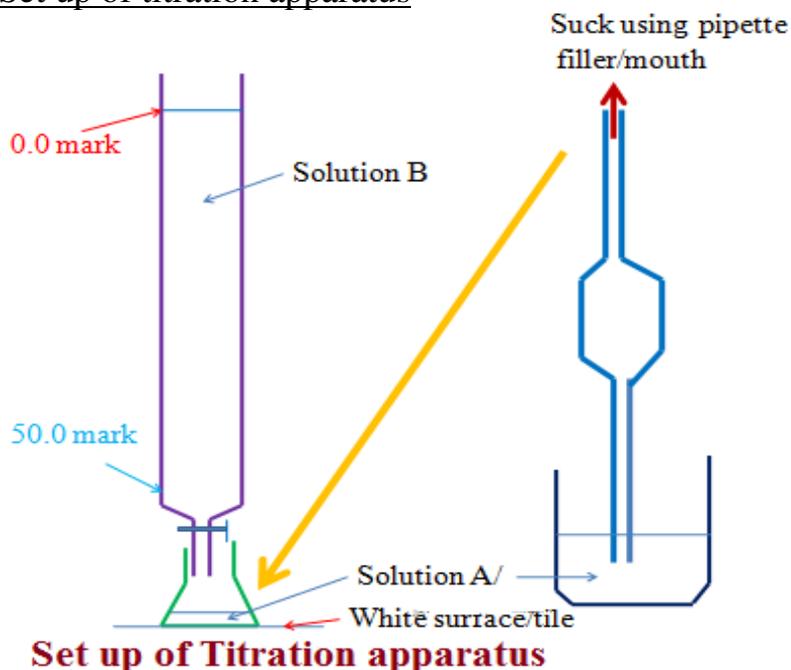
$$\text{Molar mass } (\text{NH}_4)_2\text{SO}_4 = \underline{\underline{132 \text{ gmole}^{-1}}}$$

$$\begin{aligned} \text{Mass of in impure sample} &= \text{moles} \times \text{molar mass} \Rightarrow 0.07875 \times 132 = \mathbf{10.395 \text{ g}} \\ \text{Mass of impurities} &= 10.5 - 10.395 = \mathbf{0.105 \text{ g}} \\ \% \text{ impurities} &= \frac{0.105 \times 100}{10.5} = \mathbf{1.0 \%} \end{aligned}$$

Practically volumetric analysis involves **titration**.

Titration generally involves filling a burette with known/unknown concentration of a solution then adding the solution to unknown/known concentration of another solution in a conical flask until there is complete reaction. If the solutions used are both colourless, an **indicator** is added to the conical flask. When the reaction is over, a **slight/little** excess of burette contents **change** the colour of the indicator. This is called the **end point**.

Set up of titration apparatus



The titration process involves determination of **titre**. The titre is the volume of burette contents/reading **before** and **after** the end point. Burette contents/reading **before** titration is usually called the **Initial** burette reading. Burette contents/reading **after** titration is usually called the **Final** burette reading. The titre value is thus a sum of the **Final less Initial** burette readings.

To reduce errors, titration process should be repeated at least once more. The results of titration are recorded in a **titration table** as below

Sample titration table

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

As **evidence** of a titration **actually** done examining body requires the candidate to record their burette readings before and after the titration.

For KCSE candidates burette readings **must** be recorded **in** a titration table in the **format provided** by the Kenya National Examination Council.

As **evidence** of all titration **actually** done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table **in the format provided**.

Calculate the average volume of solution used

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

As **evidence** of understanding the degree of accuracy of burettes , all readings must be recorded to **a** decimal point.

As **evidence** of accuracy in carrying the out the titration , candidates value should be **within 0.2** of the **school value** .

The school value is the **teachers** readings **presented** to the examining body/council based on the **concentrations** of the solutions s/he presented to her/his candidates.

Bonus mark is awarded for averaged reading **within 0.1** school value as Final answer.

Calculations involved after the titration require candidates **thorough** practical and theoretical **practice mastery** on the:

- (i)relationship **among** the mole, molar mass, mole ratios, concentration, molarity.
- (ii) mathematical application of 1st principles.

Very useful information which candidates forget appears usually in the beginning of the question paper as:

“You are provided with...”

All calculation must be to the **4th decimal point** unless they divide fully to a lesser decimal point.

Candidates are expected to use a non programmable scientific calculator.

(a) Sample Titration Practice 1 (Simple Titration)**You are provided with:**

0.1M sodium hydroxide solution A

Hydrochloric acid solution B

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

Procedure

Fill the burette with solution B. Pipette 25.0cm³ of solution A into a conical flask. Titrate solution A with solution B using phenolphthalein indicator to complete the titration table 1

Sample results Titration table 1

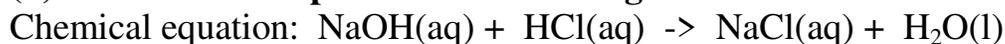
Titration number	1	2	3
Final burette reading (cm ³)	20.0	20.0	20.0
Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of solution B used(cm ³)	20.0	20.0	20.0

Sample worked questions**1. Calculate the average volume of solution B used**

$$\text{Average titre} = \frac{\text{Titre 1} + \text{Titre 2} + \text{Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3} \right) = \underline{\underline{20.0\text{cm}^3}}$$

2. How many moles of:**(i) solution A were present in 25cm³ solution.**

$$\text{Moles of solution A} = \frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.1 \times 25}{1000} = \underline{\underline{2.5 \times 10^{-3}}} \text{ moles}$$

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

$$\text{Mole ratio 1:1} \Rightarrow \text{Moles of A} = \text{Moles of B} = \underline{\underline{2.5 \times 10^{-3}}} \text{ moles}$$

(iii) solution B in moles per litre.

$$\text{Moles of B per litre} = \frac{\text{moles} \times 1000}{\text{Volume}} = \frac{2.5 \times 10^{-3} \times 1000}{20} = \underline{\underline{0.1\text{M}}}$$

(b) Sample Titration Practice 2 (Redox Titration)**You are provided with:**

Acidified Potassium manganate(VII) solution A

0.1M of an iron (II)salt solution B

8.5g of ammonium iron(II)sulphate(VI) crystals $(\text{NH}_4)_2 \text{SO}_4 \text{FeSO}_4 \cdot x\text{H}_2\text{O}$
solid C

You are required to

(i) standardize acidified potassium manganate(VII)

(ii) determine the value of x in the formula $(\text{NH}_4)_2 \text{SO}_4 \text{FeSO}_4 \cdot x\text{H}_2\text{O}$.**Procedure 1**Fill the burette with solution A. Pipette 25.0cm³ of solution B into a conical flask.

Titrate solution A with solution B until a pink colour just appears.

Record your results to complete table 1.

Table 1: **Sample results**

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

Sample worked questions

1. Calculate the average volume of solution A used

$$\text{Average titre} = \frac{\text{Titre 1} + \text{Titre 2} + \text{Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3} \right) = \underline{\underline{20.0\text{cm}^3}}$$

2. How many moles of:(i) solution B were present in 25cm³ solution.

$$\text{Moles of solution A} = \frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.1 \times 25}{1000} = \underline{\underline{2.5 \times 10^{-3}}} \text{ moles}$$

(ii) solution A were present in the average volume. Assume one mole of B react with five moles of B

Mole ratio A : B = 1:5

$$\Rightarrow \text{Moles of A} = \frac{\text{Moles of B}}{5} = \frac{2.5 \times 10^{-3}}{5} \text{ moles} = \underline{\underline{5.0 \times 10^{-4}}} \text{ moles}$$

(iii) solution B in moles per litre.

$$\begin{aligned} \text{Moles of B per litre} &= \frac{\text{moles} \times 1000}{\text{Volume}} = \frac{2.5 \times 10^{-3} \times 1000}{20} \\ &= \mathbf{0.025 \text{ M / moles per litre / moles l}^{-1}} \end{aligned}$$

Procedure 2

Place all the solid C into the 250cm³ volumetric flask carefully. Add about 200cm³ of distilled water. Shake to dissolve. Make up to the 250cm³ of solution by adding more distilled water. Label this solution C. Pipette 25cm³ of solution C into a conical flask, Titrate solution C with solution A until a permanent pink colour just appears. Complete table 2.

Table 2: **Sample results**

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

Sample worked questions

1. Calculate the average volume of solution A used

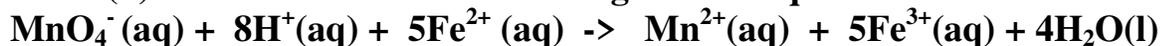
$$\text{Average titre} = \frac{\text{Titre 1} + \text{Titre 2} + \text{Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3} \right) = \mathbf{20.0\text{cm}^3}$$

2. How many moles of:

(i) solution A were present in the average titre.

$$\text{Moles of solution A} = \frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.025 \times 20}{1000} = \mathbf{5.0 \times 10^{-4} \text{ moles}}$$

(ii) solution C in 25cm³ solution given the equation for the reaction:



Mole ratio $\text{MnO}_4^- (\text{aq}) : 5\text{Fe}^{2+} (\text{aq}) = 1:5 \Rightarrow$

$$\text{Moles of } 5\text{Fe}^{2+} (\text{aq}) = \frac{\text{Moles of } \text{MnO}_4^- (\text{aq})}{5} = \frac{5.0 \times 10^{-4}}{5} \text{ moles} = \mathbf{1.0 \times 10^{-4} \text{ moles}}$$

(iii) solution B in 250cm³.

$$\text{Moles of B per litre} = \frac{\text{moles} \times 250}{\text{Volume}} = \frac{1.0 \times 10^{-4} \times 250}{25} = \mathbf{1.0 \times 10^{-3} \text{ moles}}$$

3. Calculate the molar mass of solid C and hence the value of x in the chemical formula $(\text{NH}_4)_2\text{SO}_4\text{FeSO}_4 \cdot x\text{H}_2\text{O}$.

(N=14.0, S=32.0, Fe=56.0, H=1.0 O=16.0)

$$\text{Molar mass} = \frac{\text{mass per litre}}{\text{Moles per litre}} = \frac{8.5}{1.0 \times 10^{-3} \text{ moles}} = \underline{\underline{8500 \text{ g}}}$$

$$\begin{aligned} \text{NH}_4)_2\text{SO}_4\text{FeSO}_4 \cdot x\text{H}_2\text{O} &= 8500 \\ 284 + 18x &= 8500 \\ 8500 - 284 &= \frac{8216}{18} = \frac{18x}{18} = \underline{\underline{454.4444}} \end{aligned}$$

$$x = \underline{\underline{454}} \text{ (whole number)}$$

(c) Sample Titration Practice 3 (**Back titration**)

You are provided with:

- (i) an impure calcium carbonate labeled M
- (ii) Hydrochloric acid labeled solution N
- (iii) solution L containing 20g per litre sodium hydroxide.

You are required to determine the concentration of N in moles per litre and the % of calcium carbonate in mixture M.

Procedure 1

Pipette 25.0cm³ of solution L into a conical flask. Add 2-3 drops of phenolphthalein indicator. Titrate with dilute hydrochloric acid solution N and record your results in table 1(4mark)

Sample Table 1

	1	2	3
Final burette reading (cm ³)	6.5	6.5	6.5
Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of N used (cm ³)	6.5	6.5	6.5

Sample questions

(a) Calculate the average volume of solution N used

$$\frac{6.5 + 6.5 + 6.5}{3} = \underline{\underline{6.5}} \text{ cm}^3$$

(b) How many moles of sodium hydroxide are contained in 25cm³ of solution L

Molar mass NaOH = **40g**

$$\text{Molarity of L} = \frac{\text{mass per litre}}{\text{molar mass}} \Rightarrow \frac{20}{40} = \underline{\underline{0.5M}}$$

Molar mass NaOH 40

$$\text{Moles NaOH in } 25\text{cm}^3 = \frac{\text{molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.5\text{M} \times 25\text{cm}^3}{1000} = \underline{0.0125 \text{ moles}}$$

(c) Calculate:

(i) the number of moles of hydrochloric acid that react with sodium hydroxide in (b) above.

Mole ratio NaOH : HCl from stoichiometric equation = **1:1**

Moles HCl = Moles NaOH \Rightarrow **0.0125 moles**

(ii) the molarity of hydrochloric acid solution N.

$$\text{Molarity} = \frac{\text{moles} \times 1000}{6.5} \Rightarrow \frac{0.0125 \text{ moles} \times 1000}{6.5} = \underline{1.9231\text{M}/\text{mole dm}^{-3}}$$

Procedure 2

Place the 4.0 g of M provided into a conical flask and add 25.0cm³ of the dilute hydrochloric acid to it using a clean pipette. Swirl the contents of the flask vigorously until effervescence stop. Using a 100ml measuring cylinder add 175cm³ distilled water to make up the solution up to 200cm³. Label this solution K. Using a clean pipette transfer 25.0cm³ of the solution into a clean conical flask and titrate with solution L from the burette using 2-3 drops of methyl orange indicator. Record your observations in table 2.

Sample Table 2

	1	2	3
Final burette reading (cm ³)	24.5	24.5	24.5
Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of N used (cm ³)	24.5	24.5	24.5

Sample calculations

(a) Calculate the average volume of solution L used (1mk)

$$\frac{24.5 + 24.5 + 24.5}{3} = \underline{24.5\text{cm}^3}$$

(b) How many moles of sodium hydroxide are present in the average volume of solution L used?

$$\text{Moles} = \frac{\text{molarity} \times \text{average burette volume}}{1000} \Rightarrow \frac{0.5 \times 24.5}{1000}$$

$$= \underline{0.01225 / 1.225 \times 10^{-2} \text{ moles}}$$

(c) How many moles of hydrochloric acid are present in the original 200cm³ of solution K?

Mole ratio NaOH: HCl = 1:1 => moles of HCl = $0.01225 / 1.225 \times 10^{-2}$ moles

Moles in 200cm³ = $\frac{200\text{cm}^3 \times 0.01225 / 1.225 \times 10^{-2} \text{ moles}}{25\text{cm}^3(\text{volume pipetted})}$

= $0.49 / 4.9 \times 10^{-1}$ moles

(d) How many moles of hydrochloric acid were contained in original 25 cm³ solution N used

Original moles = $\frac{\text{Original molarity} \times \text{pipetted volume}}{1000\text{cm}^3}$ =>

$\frac{1.9231\text{M}/\text{mole dm}^{-3} \times 25}{1000} = \underline{0.04807/4.807 \times 10^{-2}}$ moles

(e) How many moles of hydrochloric acid were used to react with calcium carbonate present?

Moles that reacted = original moles – moles in average titre =>

$0.04807/4.807 \times 10^{-2} \text{ moles} - 0.01225 / 1.225 \times 10^{-2} \text{ moles}$

= $0.03582/3.582 \times 10^{-2}$ moles

(f) Write the equation for the reaction between calcium carbonate and hydrochloric acid.



(g) Calculate the number of moles of calcium carbonate that reacted with hydrochloric acid.

From the equation $\text{CaCO}_3(\text{s}):2\text{HCl}(\text{aq}) = 1:2$

=> Moles $\text{CaCO}_3(\text{s}) = \frac{1}{2} \text{ moles HCl}$
 = $\frac{1}{2} \times 0.03582/3.582 \times 10^{-2} \text{ moles}$
 = $0.01791 / 1.791 \times 10^{-2}$ moles

(h) Calculate the mass of calcium carbonate in 4.0g of mixture M (Ca=40.0, O = 16.0, C=12.0)

Molar mass $\text{CaCO}_3 = 100\text{g}$

Mass $\text{CaCO}_3 = \text{moles} \times \text{molar mass} \Rightarrow 0.01791 / 1.791 \times 10^{-2} \text{ moles} \times 100\text{g}$
 $= 1.791\text{g}$

(i) Determine the % of calcium carbonate present in the mixture

$$\% \text{CaCO}_3 = \frac{\text{mass of pure} \times 100\%}{\text{Mass of impure}} \Rightarrow \frac{1.791\text{g} \times 100\%}{4.0} = \underline{\underline{44.775\%}}$$

(d) Sample titration practice 4 (Multiple titration)

You are provided with:

- (i) sodium L containing 5.0g per litre of a dibasic organic acid $\text{H}_2\text{X} \cdot 2\text{H}_2\text{O}$.
- (ii) solution M which is acidified potassium manganate(VII)
- (iii) solution N a mixture of sodium ethanedioate and ethanedioic acid
- (iv) 0.1M sodium hydroxide solution P
- (v) 1.0M sulphuric(VI)

You are required to:

- (i) standardize solution M using solution L
- (ii) use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

Procedure 1

Fill the burette with solution M. Pipette 25.0cm³ of solution L into a conical flask. Heat this solution to about 70°C (**but not to boil**). Titrate the hot solution L with solution M until a permanent pink colour just appears. Shake thoroughly during the titration. Repeat this procedure to complete table 1.

Sample Table 1

	1	2	3
Final burette reading (cm ³)	24.0	24.0	24.0
Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of N used (cm ³)	24.0	24.0	24.0

Sample calculations

(a) Calculate the average volume of solution L used (1mk)

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

(b) Given that the concentration of the dibasic acid is $0.05 \text{ moles dm}^{-3}$. determine the value of x in the formula $\text{H}_2\text{X} \cdot 2\text{H}_2\text{O}$ (H=1.0, O=16.0)

$$\text{Molar mass } \text{H}_2\text{X} \cdot 2\text{H}_2\text{O} = \frac{\text{mass per litre}}{\text{Moles/litre}} \Rightarrow \frac{5.0 \text{g/litre}}{0.05 \text{ moles dm}^{-3}} = \underline{100 \text{g}}$$

$$\text{H}_2\text{X} \cdot 2\text{H}_2\text{O} = 100$$

$$X = 100 - ((2 \times 1) + 2 \times (2 \times 1) + (2 \times 16)) \Rightarrow 100 - 34 = \underline{66}$$

(c) Calculate the number of moles of the dibasic acid $\text{H}_2\text{X} \cdot 2\text{H}_2\text{O}$.

$$\text{Moles} = \frac{\text{molarity} \times \text{pipette volume}}{1000} \Rightarrow \frac{0.5 \times 25}{1000} = \underline{0.0125/1.25 \times 10^{-2}} \text{ moles}$$

(d) Given the mole ratio manganate(VII) (MnO_4^-): acid H_2X is 2:5, calculate the number of moles of manganate(VII) (MnO_4^-) in the average titre.

$$\begin{aligned} \text{Moles } \text{H}_2\text{X} &= \frac{2}{5} \text{ moles of } \text{MnO}_4^- \\ &\Rightarrow \frac{2}{5} \times 0.0125/1.25 \times 10^{-2} \text{ moles} \\ &= \underline{0.005/5.0 \times 10^{-3}} \text{ moles} \end{aligned}$$

(e) Calculate the concentration of the manganate(VII) (MnO_4^-) in moles per litre.

$$\begin{aligned} \text{Moles per litre/molarity} &= \frac{\text{moles} \times 1000}{\text{average burette volume}} \\ &\Rightarrow \frac{0.005/5.0 \times 10^{-3} \text{ moles} \times 1000}{24.0} = \underline{0.2083} \text{ moles l}^{-1}/\text{M} \end{aligned}$$

Procedure 2

With solution M still in the burette, pipette 25.0 cm³ of solution N into a conical flask. Heat the conical flask containing solution N to about 70°C. Titrate while hot with solution M. Repeat the experiment to complete table 2.

Sample Table 2

	1	2	3
Final burette reading (cm ³)	12.5	12.5	12.5

Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of N used (cm ³)	12.5	12.5	12.5

Sample calculations

(a) Calculate the average volume of solution L used (1mk)

$$\frac{12.5 + 12.5 + 12.5}{3} = 12.5 \text{ cm}^3$$

(b) Calculations:

(i) How many moles of manganate(VII) ions are contained in the average volume of solution M used?

$$\begin{aligned} \text{Moles} &= \frac{\text{molarity of solution M} \times \text{average burette volume}}{1000} \\ \Rightarrow \frac{0.2083 \text{ moles l}^{-1} / \text{M} \times 12.5}{1000} &= \underline{0.0026 / 2.5 \times 10^{-3}} \text{ moles} \end{aligned}$$

(ii) The reaction between manganate(VII) ions and ethanedioate ions that reacted with is as in the equation:



Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M.

From the stoichiometric equation, mole ratio $\text{MnO}_4^- (\text{aq}) : \text{C}_2\text{O}_4^{2-} (\text{aq}) = 2:5$

$$\begin{aligned} \Rightarrow \text{moles } \text{C}_2\text{O}_4^{2-} &= \frac{5}{2} \text{ moles } \text{MnO}_4^- \Rightarrow \frac{5}{2} \times 0.0026 / 2.5 \times 10^{-3} \text{ moles} \\ &= \underline{0.0065 / 6.5 \times 10^{-3}} \text{ moles} \end{aligned}$$

(iii) Calculate the number of moles of ethanedioate ions contained in 250cm³ solution N.

$$\begin{aligned} 25 \text{ cm}^3 \text{ pipette volume} &\rightarrow 0.0065 / 6.5 \times 10^{-3} \text{ moles} \\ 250 \text{ cm}^3 &\rightarrow \end{aligned}$$

$$\frac{0.0065 / 6.5 \times 10^{-3} \text{ moles} \times 250}{25} = \underline{0.065 / 6.5 \times 10^{-2}} \text{ moles}$$

Procedure 3

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25cm³ of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to complete table 3.

Sample Table 2

	1	2	3
Final burette reading (cm ³)	24.9	24.9	24.9
Initial burette reading (cm ³)	0.0	0.0	0.0
Volume of N used (cm ³)	24.9	24.9	24.9

Sample calculations

(a) Calculate the average volume of solution L used (1mk)

$$\frac{24.9 + 24.9 + 24.9}{3} = 24.9 \text{ cm}^3$$

(b) Calculations:

(i) How many moles of sodium hydroxide solution P were contained in the average volume?

$$\begin{aligned} \text{Moles} &= \frac{\text{molarity of solution P} \times \text{average burette volume}}{1000} \\ \Rightarrow \frac{0.1 \text{ moles l}^{-1} \times 24.9}{1000} &= \underline{0.00249 / 2.49 \times 10^{-3}} \text{ moles} \end{aligned}$$

(ii) Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:



Calculate the number of moles of ethanedioic acid that were used in the reaction

From the stoichiometric equation, mole ratio NaOH(aq): H₂C₂O₄(aq) = 2:1

$$\begin{aligned} \Rightarrow \text{moles H}_2\text{C}_2\text{O}_4 &= \frac{1}{2} \text{ moles NaOH} \Rightarrow \frac{1}{2} \times 0.00249 / 2.49 \times 10^{-3} \text{ moles} \\ &= \underline{0.001245 / 1.245 \times 10^{-3}} \text{ moles.} \end{aligned}$$

(iii) How many moles of ethanedioic acid were contained in 250cm³ of solution N?

$$\begin{aligned}
 &25\text{cm}^3 \text{ pipette volume} \rightarrow 0.001245/1.245 \times 10^{-3} \text{ moles} \\
 &250\text{cm}^3 \quad \quad \quad \rightarrow \\
 &\quad \quad \quad \frac{0.001245/1.245 \times 10^{-3} \text{ moles} \times 250}{25} = \underline{\underline{0.01245/1.245 \times 10^{-2}}} \text{ moles}
 \end{aligned}$$

(iii) Determine the % by mass of sodium ethanedioate in the mixture (H= 1.0,O=16.0,C=12.0 and total mass of mixture =2.0 g in 250cm³ solution)

$$\begin{aligned}
 \text{Molar mass H}_2\text{C}_2\text{O}_4 &= \mathbf{90.0\text{g}} \\
 \text{Mass of H}_2\text{C}_2\text{O}_4 \text{ in } 250\text{cm}^3 &= \text{moles in } 250\text{cm}^3 \times \text{molar mass H}_2\text{C}_2\text{O}_4 \\
 &=> 0.01245/1.245 \times 10^{-2} \text{ moles} \times 90.0 \\
 &= \underline{\underline{1.1205\text{g}}}
 \end{aligned}$$

$$\begin{aligned}
 &\% \text{ by mass of sodium ethanedioate} \\
 &= \frac{(\text{Mass of mixture} - \text{mass of H}_2\text{C}_2\text{O}_4)}{\text{Mass of mixture}} \times 100\% \\
 &=> \frac{2.0 - 1.1205 \text{ g}}{2.0} = \underline{\underline{43.975\%}}
 \end{aligned}$$

Note

- (i) L is 0.05M Oxalic acid
- (ii) M is 0.01M KMnO₄
- (iii) N is 0.03M oxalic acid(without sodium oxalate)

Practice example 5.(Determining equation for a reaction)

You are provided with

- 0.1M hydrochloric acid solution A
- 0.5M sodium hydroxide solution B

You are to determine the equation for the reaction between solution A and B

Procedure

Fill the burette with solution A. Using a pipette and pipette filler transfer 25.0cm³ of solution B into a conical flask. Add 2-3 drops of phenolphthalein indicator. Run solution A into solution B until a permanent pink colour just appears. Record your results in Table 1. Repeat the experiment to obtain three concordant results to complete Table 1

Table 1(Sample results)

Titration	1	2	3
Final volume(cm ³)	12.5	25.0	37.5
Initial volume(cm ³)	0.0	12.5	25.0
Volume of solution A used(cm ³)	12.5	12.5	12.5

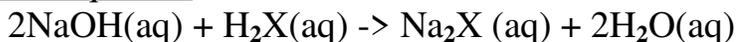
Sample questions

Calculate the average volume of solution A used.

$$\frac{12.5+12.5+12.5}{3} = \mathbf{12.5\text{cm}^3}$$

Theoretical Practice examples

1. 1.0g of dibasic acid $\text{HOOC}(\text{CH}_2)_x\text{COOH}$ was dissolved in 250cm³ solution. 25.0 cm³ of this solution reacted with 30.0cm³ of 0.06M sodium hydroxide solution. Calculate the value of x in $\text{HOOC}(\text{CH}_2)_x\text{COOH}$. (C=12.0,H=1.0,O=16.)

Chemical equation

Mole ratio $\text{NaOH}(\text{aq}) : \text{H}_2\text{X}(\text{aq}) = 2:1$

Method 1

$$\frac{M_a V_a = n_a}{M_b V_b = n_b} \Rightarrow \frac{M_a \times 25.0}{0.06 \times 30.0} = \frac{1}{2} \Rightarrow M_a = \frac{0.06 \times 30.0 \times 1}{25.0 \times 2}$$

Molarity of acid = $\mathbf{0.036\text{M/Mole l}^{-1}}$

$$\text{Mass of acid per litre} = \frac{1.0 \times 1000}{250} = \mathbf{4.0\text{ g/l}}$$

$$0.036\text{M/ Mole l}^{-1} \rightarrow 4.0\text{ g/l}$$

$$1\text{ mole} = \text{molar mass of } \text{HOOC}(\text{CH}_2)_x\text{COOH} = \frac{4.0 \times 1}{0.036} = \mathbf{111.1111\text{ g}}$$

$$\text{Molar mass } (\text{CH}_2)_x = 111.1111 - (\text{HOOC}(\text{COOH}) = 90.0) = \mathbf{21.1111}$$

$$(\text{CH}_2)_x = 14x = \frac{21.1111}{14} = \mathbf{1.5 = 1\text{ (whole number)}}$$

Method 2

$$\text{Moles of sodium hydroxide} = \frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.06 \times 30}{1000} = \underline{1.8 \times 10^{-3} \text{ moles}}$$

$$\text{Moles of Hydrochloric acid} = \frac{1}{2} \times 1.8 \times 10^{-3} \text{ moles} = 9.0 \times 10^{-4} \text{ moles}$$

$$\text{Molarity of Hydrochloric acid} = \frac{\text{moles} \times 1000}{\text{Volume}} = \frac{9.0 \times 10^{-4} \text{ moles} \times 1000}{25}$$

$$\text{Molarity of acid} = \underline{0.036 \text{ M/Mole l}^{-1}}$$

$$\text{Mass of acid per litre} = \frac{1.0 \times 1000}{250} = \underline{4.0 \text{ g/l}}$$

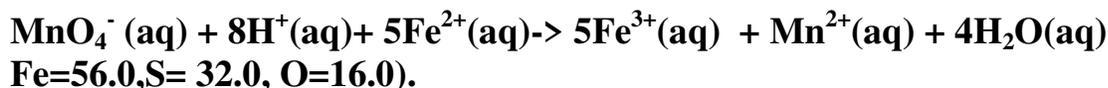
$$0.036 \text{ M/Mole l}^{-1} \rightarrow 4.0 \text{ g/l}$$

$$1 \text{ mole} = \text{molar mass of HOOC(CH}_2)_x\text{COOH} = \frac{4.0 \times 1}{0.036} = \underline{111.1111 \text{ g}}$$

$$\text{Molar mass (CH}_2)_x = 111.1111 - (\text{HOOC-COOH} = 90.0) = \underline{21.1111}$$

$$(\text{CH}_2)_x = 14x = \frac{21.1111}{14} = \underline{1.5} = \underline{1 \text{ (whole number)}}$$

2. 20.0cm³ of 0.05 M acidified potassium manganate(VII)solution oxidized 25.0cm³ of Fe²⁺(aq) ions in 40.0g/l of impure Iron (II)sulphate(VI) to Fe³⁺(aq) ions. Calculate the percentage impurities in the Iron (II)sulphate(VI).



$$\text{Moles of MnO}_4^- (\text{aq}) = \frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.05 \times 20.0}{1000} = \underline{0.001 \text{ Moles}}$$

$$\text{Mole ratio MnO}_4^- (\text{aq}): 5\text{Fe}^{2+}(\text{aq}) = 1:5$$

$$\text{Moles } 5\text{Fe}^{2+}(\text{aq}) = 5 \times 0.001 = \underline{0.005 \text{ Moles}}$$

$$\begin{aligned} \text{Moles of } 5\text{Fe}^{2+}(\text{aq}) \text{ per litre/molarity} &= \frac{\text{Moles} \times 1000}{\text{Volume}} = \frac{0.005 \times 1000}{25.0} \\ &= \underline{0.2 \text{ M/ Moles/litre}} \end{aligned}$$

$$\text{Molar mass} = \text{FeSO}_4 = \underline{152 \text{ g}}$$

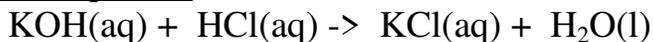
$$\text{Mass of in the mixture} = \text{Moles} \times \text{molar mass} \Rightarrow 0.2 \times 152 = \underline{30.4 \text{ g}}$$

$$\text{Mass of impurity} = 40.0 - 30.4 = \underline{9.6 \text{ g}}$$

$$\% \text{ impurity} = \frac{9.6 \text{ g} \times 100}{40.0} = \underline{\underline{24.0 \%}} \text{ impurity}$$

3.9.7 g of a mixture of Potassium hydroxide and Potassium chloride was dissolved to make one litre solution. 20.0cm³ of this solution required 25.0cm³ of 0.12M hydrochloric acid for completed neutralization. Calculate the percentage by mass of Potassium chloride. (K=39.0, Cl= 35.5)

Chemical equation



$$\text{Moles of HCl} = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.12 \times 25.0}{1000} = \underline{\underline{0.003/3.0 \times 10^{-3}}} \text{ moles}$$

Mole ratio KOH(aq) : HCl(aq) = 1:1

$$\text{Moles KOH} = \underline{\underline{0.003/3.0 \times 10^{-3}}} \text{ moles}$$

Method 1

Molar mass KOH = **56.0g**

$$\text{Mass KOH in 25cm}^3 = \underline{\underline{0.003/3.0 \times 10^{-3}}} \text{ moles} \times 56.0 = \underline{\underline{0.168g}}$$

$$\text{Mass KOH in 1000cm}^3/1 \text{ litre} = \frac{0.168 \times 1000}{20} = \underline{\underline{8.4 \text{ g/l}}}$$

$$\text{Mass of KCl} = 9.7\text{g} - 8.4\text{g} = \underline{\underline{1.3 \text{ g}}}$$

$$\% \text{ of KCl} = \frac{1.3 \times 100}{9.7} = \underline{\underline{13.4021 \%}}$$

Method 2

$$\text{Moles KOH in 1000cm}^3 /1 \text{ litre} = \frac{\text{Moles in 20cm}^3 \times 1000}{20} \Rightarrow \frac{0.003 \times 1000}{20} = \underline{\underline{0.15\text{M/Moles /litre}}}$$

Molar mass KOH = **56.0g**

$$\text{Mass KOH in 1000/1 litre} = 0.15\text{M/Moles /litre} \times 56.0 = \underline{\underline{8.4\text{g/l}}}$$

$$\text{Mass of KCl} = 9.7\text{g} - 8.4\text{g} = \underline{\underline{1.3 \text{ g}}}$$

$$\% \text{ of KCl} = \frac{1.3 \times 100}{9.7} = \underline{\underline{13.4021 \%}}$$

4. A certain carbonate, GCO₃, reacts with dilute hydrochloric acid according to the equation given below:



If 1 g of the carbonate reacts completely with 20 cm³ of 1 M hydrochloric acid, calculate the relative atomic mass of G (C = 12.0 = 16.0)

$$\text{Moles of HCl} = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{1 \times 20}{1000} = \underline{\underline{0.02 \text{ moles}}}$$

Mole ratio HCl; $\text{GCO}_3 = 2:1$

$$\text{Moles of GCO}_3 = \frac{0.02 \text{ moles}}{2} = \underline{\underline{0.01 \text{ moles}}}$$

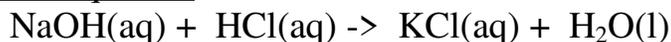
$$\text{Molar mass of GCO}_3 = \frac{\text{mass}}{\text{moles}} \Rightarrow \frac{1}{0.01 \text{ moles}} = \underline{\underline{100 \text{ g}}}$$

$$G = \text{GCO}_3 - \text{CO}_3 \Rightarrow 100\text{g} - (12 + 16 \times 3 = 60) = \underline{\underline{40(\text{no units})}}$$

5. 46.0g of a metal carbonate MCO_3 was dissolved 160cm³ of 0.1M excess hydrochloric acid and the resultant solution diluted to one litre. 25.0cm³ of this solution required 20.0cm³ of 0.1M sodium hydroxide solution for complete neutralization. Calculate the atomic mass of 'M'

Equation

Chemical equation



$$\text{Moles of NaOH} = \frac{\text{Molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.1 \times 20}{1000} = \underline{\underline{0.002 \text{ moles}}}$$

Mole ratio HCl; NaOH = 1:1

$$\text{Excess moles of HCl} = \underline{\underline{0.002 \text{ moles}}}$$

$$\begin{array}{l} 25\text{cm}^3 \quad \rightarrow \quad 0.002 \text{ moles} \\ 1000\text{cm}^3 \quad \rightarrow \quad \frac{1000 \times 0.002}{25\text{cm}^3} = \underline{\underline{0.08 \text{ moles}}} \end{array}$$

$$\text{Original moles of HCl} = \text{Molarity} \times \text{volume} \Rightarrow 1\text{M} \times 1\text{litre} = \underline{\underline{1.0 \text{ moles}}}$$

$$\text{Moles of HCl reacted with MCO}_3 = 1.0 - 0.08 \text{ moles} = \underline{\underline{0.92 \text{ moles}}}$$

Chemical equation



Mole ratio $\text{MCO}_{3(s)} : \text{HCl}_{(aq)} = 1:2$

$$\text{Moles of MCO}_3 = \frac{0.92 \text{ moles}}{2} \Rightarrow \underline{\underline{0.46 \text{ moles}}}$$

$$\text{Molar mass of MCO}_3 = \frac{\text{mass}}{\text{moles}} \Rightarrow \frac{46\text{g}}{0.46 \text{ moles}} = \underline{\underline{100 \text{ g}}}$$

$$M = \text{MCO}_3 - \text{CO}_3 \Rightarrow 100\text{g} - (12 + 16 \times 3 = 60) = \underline{\underline{40}}$$

6. 25.0cm³ of a mixture of Fe²⁺ and Fe³⁺ ions in an aqueous salt was acidified with sulphuric(VI) acid then titrated against potassium manganate(VI). The salt required 15cm³ of 0.02M potassium manganate(VI) for complete reaction.

A second 25cm³ portion of the Fe²⁺ and Fe³⁺ ion salt was reduced by Zinc then titrated against the same concentration of potassium manganate(VI). 19.0cm³ of potassium manganate(VI) solution was used for complete reaction.

Calculate the concentration of Fe²⁺ and Fe³⁺ ion in the solution on moles per litre.

Mole ratio Fe²⁺ : MnO₄⁻ = 5:1

$$\text{Moles MnO}_4^- \text{ used} = \frac{0.02 \times 15}{1000} = 3.0 \times 10^{-4} \text{ moles}$$

$$\text{Moles Fe}^{2+} = \frac{3.0 \times 10^{-4} \text{ moles}}{5} = 6.0 \times 10^{-5} \text{ moles}$$

$$\text{Molarity of Fe}^{2+} = \frac{6.0 \times 10^{-4} \text{ moles} \times 1000}{25} = \underline{2.4 \times 10^{-3} \text{ moles l}^{-1}}$$

Since Zinc reduces Fe³⁺ to Fe²⁺ in the mixture:

$$\text{Moles MnO}_4^- \text{ that reacted with all Fe}^{2+} = \frac{0.02 \times 19}{1000} = 3.8 \times 10^{-4} \text{ moles}$$

$$\text{Moles of all Fe}^{2+} = \frac{3.8 \times 10^{-4} \text{ moles}}{5} = 7.6 \times 10^{-5} \text{ moles}$$

$$\text{Moles of Fe}^{3+} = 3.8 \times 10^{-4} - 6.0 \times 10^{-5} = 1.6 \times 10^{-4} \text{ moles}$$

$$\text{Molarity of Fe}^{3+} = \frac{1.6 \times 10^{-4} \text{ moles} \times 1000}{25} = \underline{4.0 \times 10^{-4} \text{ moles l}^{-1}}$$

14.0.0 ORGANIC CHEMISTRY I (HYDROCARBONS) (25 LESSONS)

Introduction to Organic chemistry

Organic chemistry is the branch of chemistry that studies carbon compounds present in living things, once living things or synthetic/man-made.

Compounds that makes up living things whether alive or dead mainly contain carbon. Carbon is tetravalent.

It is able to form stable covalent bonds with itself and many non-metals like hydrogen, nitrogen ,oxygen and halogens to form a variety of compounds. This is because:

(i) carbon uses all the four valence electrons to form four strong covalent bond.

(ii)carbon can covalently bond to form a single, double or triple covalent bond with itself.

(iii)carbon atoms can covalently bond to form a very long chain or ring.

When carbon covalently bond with Hydrogen, it forms a group of organic compounds called **Hydrocarbons**

A.HYDROCARBONS (HCs)

Hydrocarbons are a group of organic compounds containing /made up of hydrogen and carbon atoms only.

Depending on the type of bond that exist between the individual carbon atoms, hydrocarbon are classified as:

(i) Alk**a**nes

(ii) Alk**e**nes

(iii) Alk**y**nes

(i) Alk**a**nes

(a)Nomenclature/Naming

These are hydrocarbons with a general formula C_nH_{2n+2} where **n** is the number of **Carbon** atoms in a molecule.

The carbon atoms are linked by single bond to each other and to hydrogen atoms.

They include:

n	General/ Molecular formula	Structural formula	Name
1	CH ₄	$ \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array} $	Methane
2	C ₂ H ₆	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} $	Ethane
3	C ₃ H ₈	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array} $	Propane
4	C ₄ H ₁₀	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $	Butane
5	C ₅ H ₁₂	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $ $ \text{CH}_3 (\text{CH}_2)_3 \text{CH}_3 $	Pentane
6	C ₆ H ₁₄	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $	Hexane

		$\begin{array}{ccccccc} \text{H} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \end{array}$ $\text{CH}_3 (\text{CH}_2)_6 \text{CH}_3$	
7	C_7H_{16}	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \end{array}$	Heptane
8	C_8H_{18}	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \end{array}$	Octane
9	C_9H_{20}	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \end{array}$	Nonane
10	$\text{C}_{10}\text{H}_{22}$	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \end{array}$	decane

Note

1. The **general formula/molecular formula** of a compound shows the number of each atoms of elements making the compound e.g.

Decane has a general/molecular formula $\text{C}_{10}\text{H}_{22}$; this means there are 10 carbon atoms and 22 hydrogen atoms in a molecule of decane.

2. The **structural formula** shows the arrangement/bonding of atoms of each element making the compound e.g

Decane has the structural formula as in the table above; this means the 1st carbon from left to right is bonded to three hydrogen atoms and one carbon atom.

The 2nd carbon atom is joined/bonded to two other carbon atoms and two Hydrogen atoms.

3. Since carbon is **tetravalent**, each atom of carbon in the alkane **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons.

4. Since Hydrogen is **monovalent**, each atom of hydrogen in the alkane **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

5. One member of the alkane differ from the next/previous by a CH₂ group.

e.g

Propane differ from ethane by one carbon and two Hydrogen atoms form ethane.

Ethane differ from methane also by one carbon and two Hydrogen atoms

6. A group of compounds that differ by a CH₂ group from the next /previous **consecutively** is called a **homologous series**.

7. A homologous series:

(i) differ by a CH₂ group from the next /previous consecutively

(ii) have similar chemical properties

(iii) have similar chemical formula that can be represented by a general formula e.g alkanes have the general formula C_nH_{2n+2}.

(iv) the physical properties (e.g. melting/boiling points) show steady gradual change)

8. The 1st four alkanes have the prefix **meth_**, **eth_**, **prop_** and **but_** to represent 1, 2, 3 and 4 carbons in the compound. All other use the numeral prefix **pent_**, **Hex_**, **hept_**, etc to show also the number of carbon atoms.

9. If one hydrogen atom in an alkane is removed, an alkyl group is formed. e.g

Alkane name	molecular structure C _n H _{2n+2}	Alkyl name	Molecular structure C _n H _{2n+1}
methane	CH ₄	methyl	CH₃
ethane	CH ₃ CH ₃	ethyl	CH₃ CH₂
propane	CH ₃ CH ₂ CH ₃	propyl	CH₃ CH₂ CH₂
butane	CH ₃ CH ₂ CH ₂ CH ₃	butyl	CH₃ CH₂ CH₂ CH₂

(b) Isomers of alkanes

Isomers are compounds with the same molecular **general formula** but different molecular **structural formula**.

Isomerism is the existence of a compounds having the same general/molecular formula but different structural formula.

The 1st three alkanes do not form isomers. Isomers are named by using the IUPAC (International Union of Pure and Applied Chemistry) system of **nomenclature/naming**.

The IUPAC system of nomenclature uses the following basic rules/guidelines:

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

H

(ii) 2,2-dimethylpentane

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **CH₃ CH₂ CH₂ CH₃**

2. Number the longest chain from the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon "2"

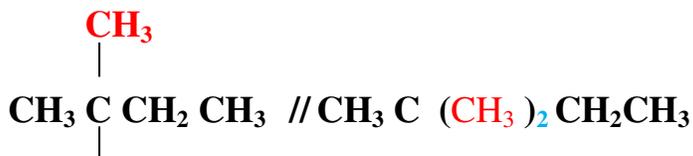
3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon "2"

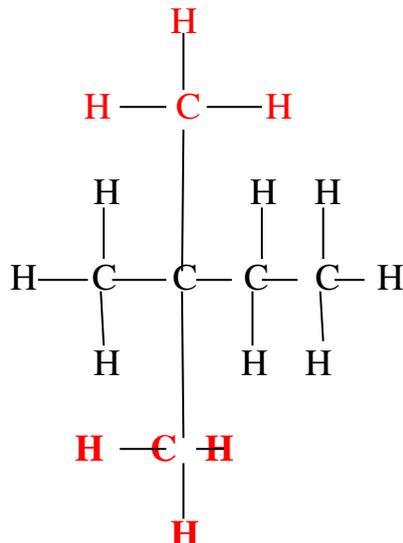
Number of branches at carbon "2"

Type of the branch two "methyl" hence

Molecular formular



Structural formula



(iii) 2,2,3-trimethylbutane

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name $\text{CH}_3 \text{CH}_2 \text{CH}_2 \text{CH}_3$

2. Number the longest chain from the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon "2 and 3"

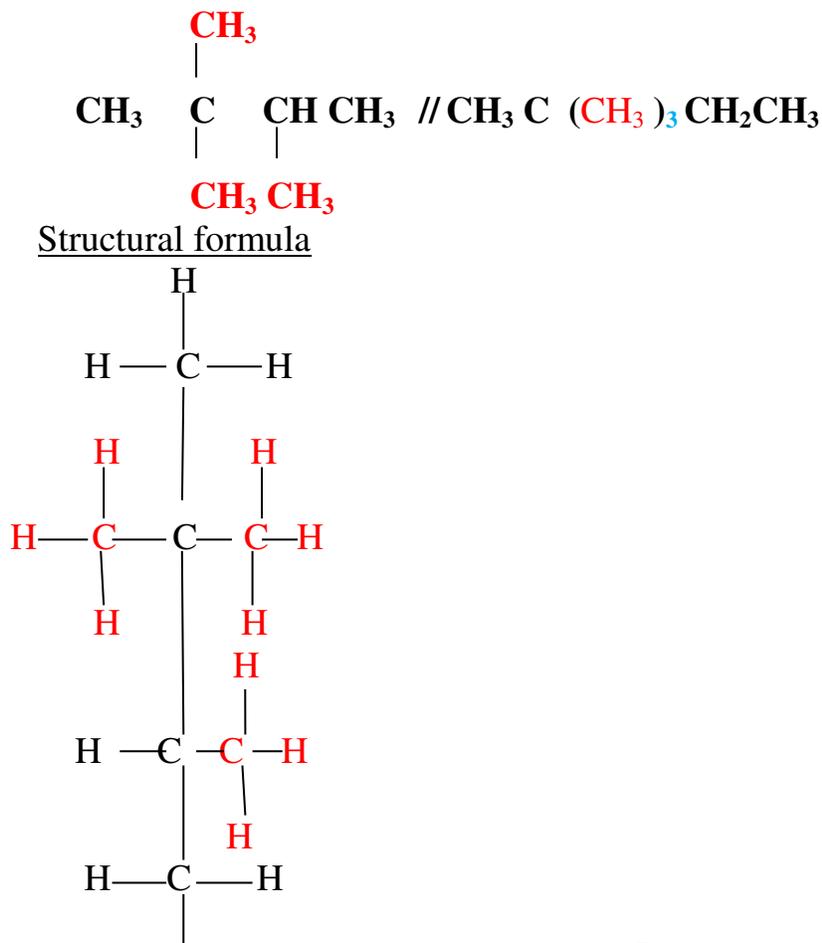
3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon "2 and 3"

Number of branches at carbon "3"

Type of the branch three "methyl" hence

Molecular formular

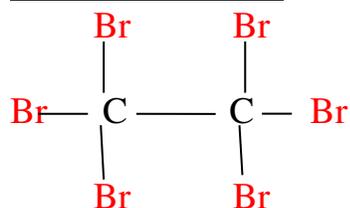


H
(iv) 1,1,1,2,2,2-hexabromoethane

Molecular formula



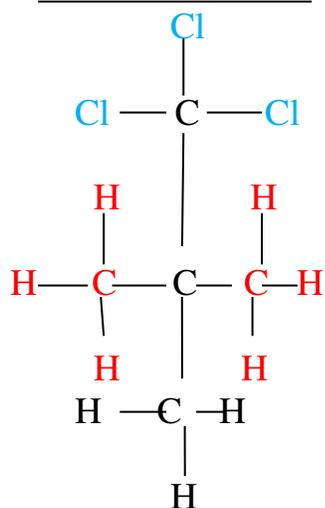
Structural formula



(v) 1,1,1-tetrachloro-2,2-dimethylbutane



Structural formula



(c) Occurrence and extraction

Crude oil, **natural gas** and **biogas** are the main sources of alkanes:

(i) Natural gas is found on top of crude oil deposits and consists mainly of methane.

(ii) Biogas is formed from the decay of waste organic products like animal dung and cellulose. When the decay takes place in absence of oxygen, 60-75% by volume of the gaseous mixture of methane gas is produced.

(iii) Crude oil is a mixture of many flammable hydrocarbons/substances. Using fractional distillation, each hydrocarbon fraction can be separated from the other. The hydrocarbon with lower /smaller number of carbon atoms in the chain have lower boiling point and thus collected first.

As the carbon **chain increase**, the **boiling** point, **viscosity** (ease of flow) and colour **intensity increase** as **flammability decrease**. Hydrocarbons in crude oil are not pure. They thus have no sharp fixed boiling point.

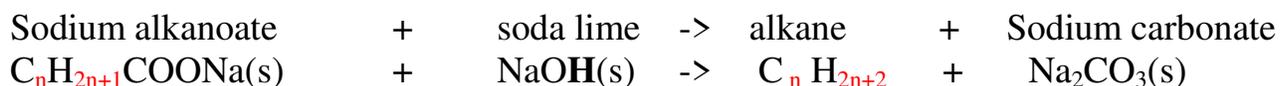
Uses of different crude oil fractions

Carbon atoms in a molecule	Common name of fraction	Uses of fraction
1-4	Gas	L.P.G gas for domestic use
5-12	Petrol	Fuel for petrol engines
9-16	Kerosene/Paraffin	Jet fuel and domestic lighting/cooking
15-18	Light diesel	Heavy diesel engine fuel
18-25	Diesel oil	Light diesel engine fuel
20-70	Lubricating oil	Lubricating oil to reduce friction.
Over 70	Bitumen/Asphalt	Tarmacking roads

(d) School laboratory preparation of alkanes

In a school laboratory, alkanes may be prepared from the reaction of a sodium alkanoate with solid sodium hydroxide/soda lime.

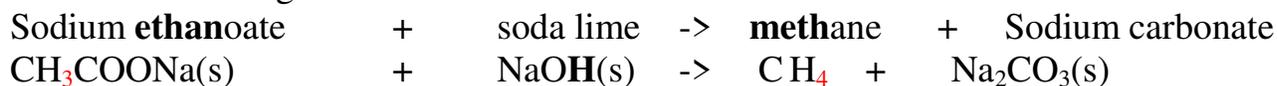
Chemical equation:



The "H" in NaOH is transferred/moves to the $\text{C}_n\text{H}_{2n+1}$ in $\text{C}_n\text{H}_{2n+1}\text{COONa(s)}$ to form $\text{C}_n\text{H}_{2n+2}$.

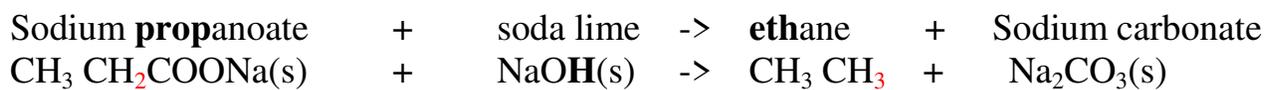
Examples

1. **Methane** is prepared from the heating of a mixture of sodium **ethanoate** and soda lime and collecting over water



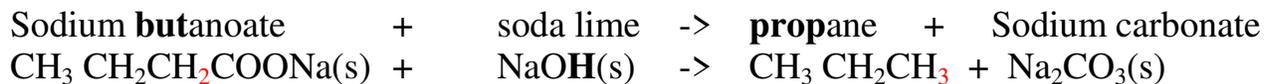
The “H” in NaOH is transferred/moves to the CH₃ in CH₃COONa(s) to form CH₄.

2. **Ethane** is prepared from the heating of a mixture of sodium **propanoate** and soda lime and collecting over water



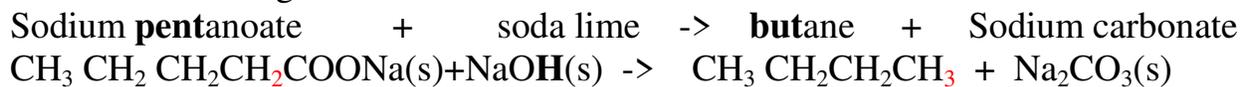
The “H” in NaOH is transferred/moves to the CH₃CH₂ in CH₃CH₂COONa (s) to form CH₃CH₃

3. **Propane** is prepared from the heating of a mixture of sodium **butanoate** and soda lime and collecting over water



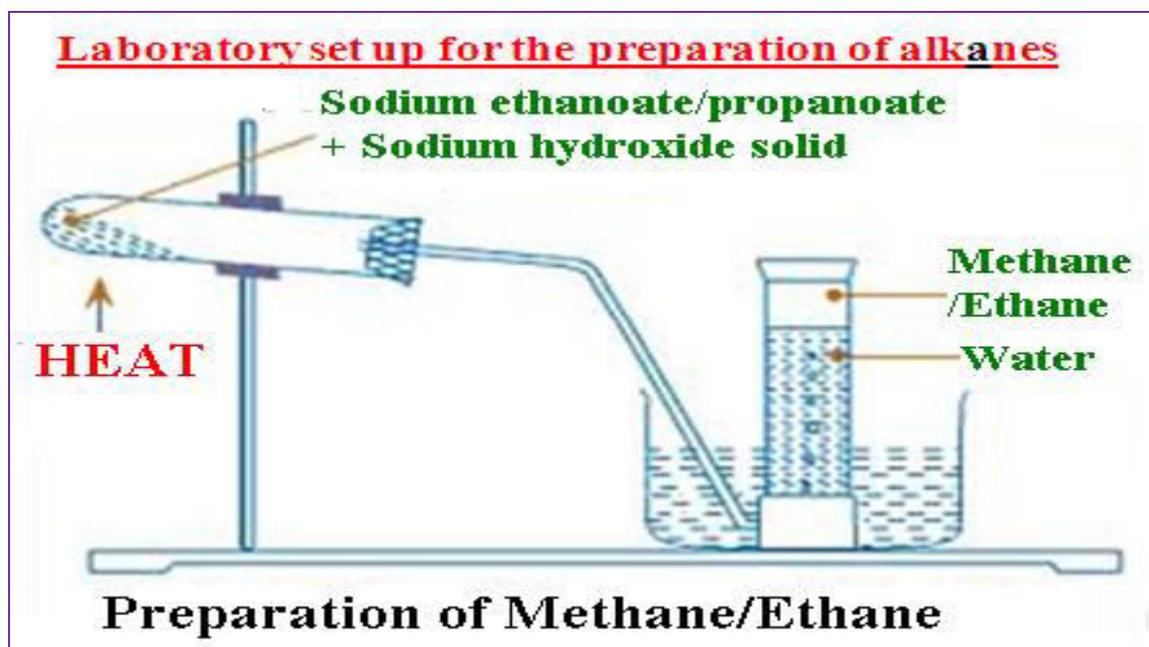
The “H” in NaOH is transferred/moves to the CH₃CH₂CH₂ in CH₃CH₂CH₂COONa (s) to form CH₃CH₂CH₃

4. **Butane** is prepared from the heating of a mixture of sodium **pentanoate** and soda lime and collecting over water



The “H” in NaOH is transferred/moves to the CH₃CH₂CH₂CH₂ in CH₃CH₂CH₂CH₂COONa (s) to form CH₃CH₂CH₂CH₃

Laboratory set up for the preparation of alkanes



(d) Properties of alkanes

I. Physical properties

Alkanes are colourless gases, solids and liquids that are not poisonous.

They are slightly soluble in water.

The solubility decrease as the carbon chain and thus the molar mass increase

The melting and boiling point increase as the carbon chain increase.

This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.

The 1st four straight chain alkanes (methane, ethane, propane and butane) are therefore gases, the next six (pentane, hexane, heptane, octane, nonane, and decane) are liquids while the rest from undecane (11 carbon atoms) are solids.

The density of straight chain alkanes increase with increasing carbon chain as the intermolecular forces increases.

This reduces the volume occupied by a given mass of the compound.

Summary of physical properties of alkanes

Alkane	General formula	Melting point(K)	Boiling point(K)	Density gcm ⁻³	State at room(298K) temperature and pressure atmosphere (101300Pa)
Methane	CH ₄	90	112	0.424	gas

Ethane	CH ₃ CH ₃	91	184	0.546	gas
Propane	CH ₃ CH ₂ CH ₃	105	231	0.501	gas
Butane	CH ₃ (CH ₂) ₂ CH ₃	138	275	0.579	gas
Pentane	CH ₃ (CH ₂) ₃ CH ₃	143	309	0.626	liquid
Hexane	CH ₃ (CH ₂) ₄ CH ₃	178	342	0.657	liquid
Heptane	CH ₃ (CH ₂) ₅ CH ₃	182	372	0.684	liquid
Octane	CH ₃ (CH ₂) ₆ CH ₃	216	399	0.703	liquid
Nonane	CH ₃ (CH ₂) ₇ CH ₃	219	424	0.708	liquid
Octane	CH ₃ (CH ₂) ₈ CH ₃	243	447	0.730	liquid

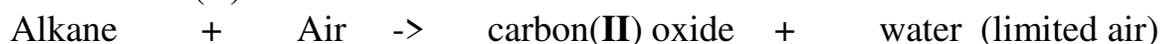
II. Chemical properties

(i) Burning.

Alkanes burn with a **blue**/non-luminous **non-sooty**/non-smoky flame in **excess** air to form carbon(IV) oxide and water.

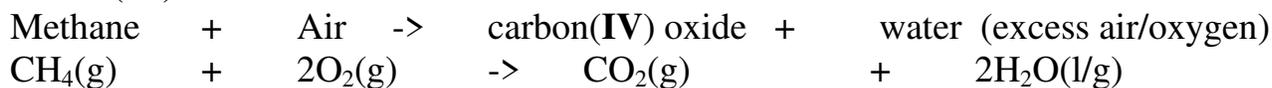


Alkanes burn with a **blue**/non-luminous **no-sooty**/non-smoky flame in **limited** air to form carbon(II) oxide and water.

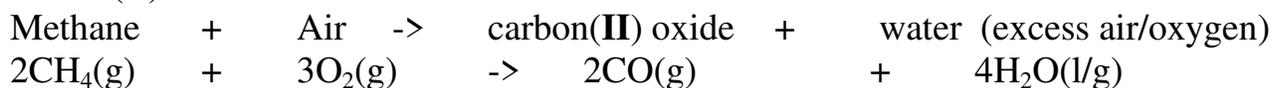


Examples

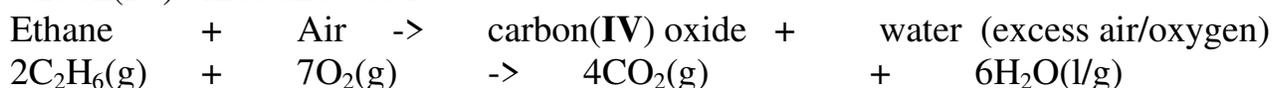
1.(a) Methane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.



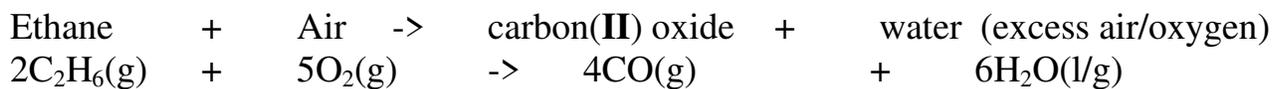
(b) Methane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.



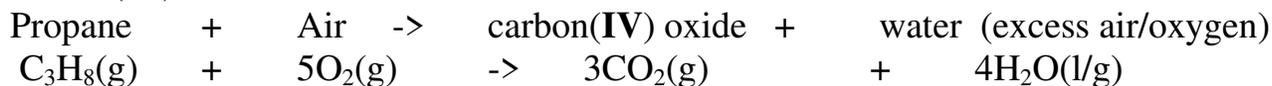
2.(a) Ethane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.



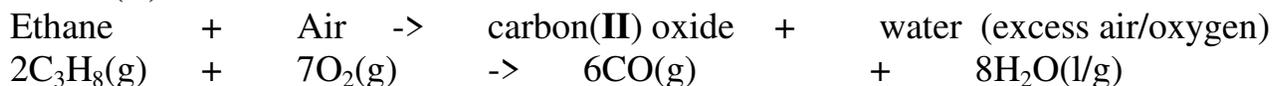
(b) Ethane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.



3.(a) Propane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.



(b) Ethane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.



ii)Substitution

Substitution reaction is one in which a hydrogen atom is replaced by a halogen in presence of ultraviolet light.

Alkanes react with halogens in presence of ultraviolet light to form halogenoalkanes.

During substitution:

- (i) the halogen molecule is split into free atom/radicals.
- (ii) one free halogen radical/atoms knock /remove one hydrogen from the alkane leaving an alkyl radical.
- (iii) the alkyl radical combine with the other free halogen atom/radical to form halogenoalkane.
- (iv) the chlorine atoms substitute repeatedly in the alkane. Each substitution removes a hydrogen atom from the alkane and form hydrogen halide.
- (v) substitution stops when all the hydrogen in alkanes are replaced with halogens.

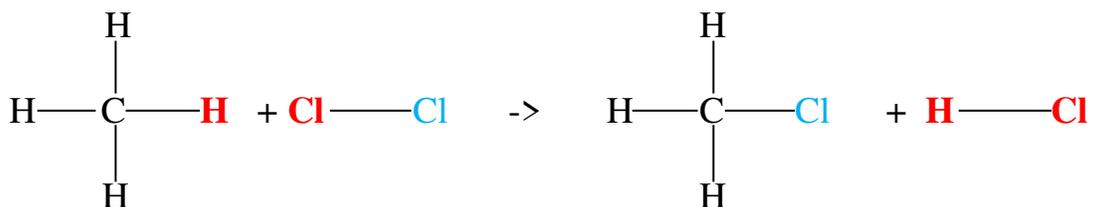
Substitution reaction is a highly **explosive** reaction in presence of **sunlight / ultraviolet** light that act as **catalyst**.

Examples of substitution reactions

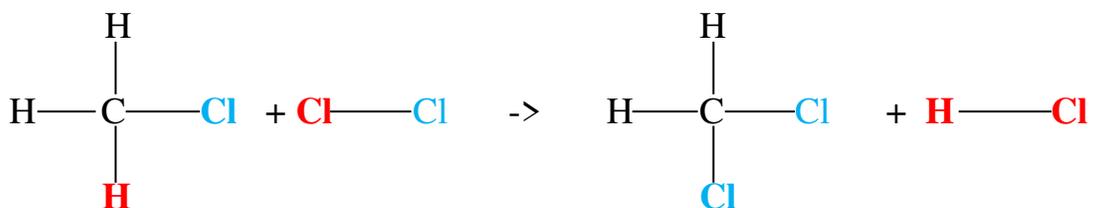
Methane has no effect on bromine or chlorine in diffused light/dark. In sunlight , a mixture of chlorine and methane explode to form colourless mixture of chloromethane and hydrogen chloride gas. The pale green colour of chlorine gas fades.

Chemical equation

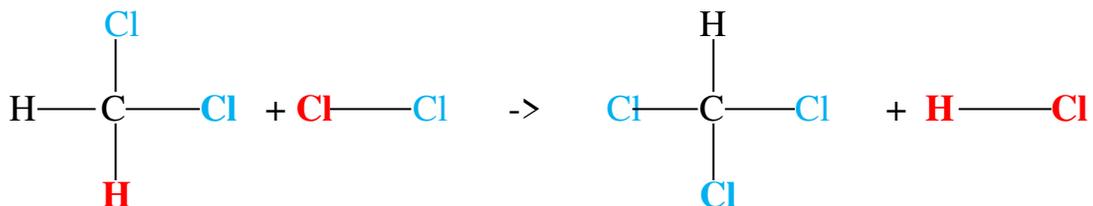




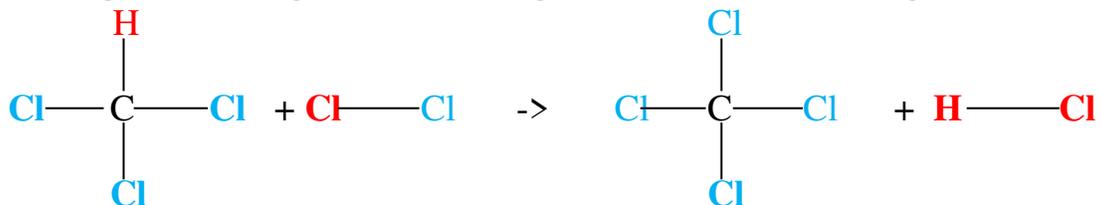
(b) Chloromethane + chlorine \rightarrow dichloromethane + Hydrogen chloride



(c) dichloromethane + chlorine \rightarrow trichloromethane + Hydrogen chloride



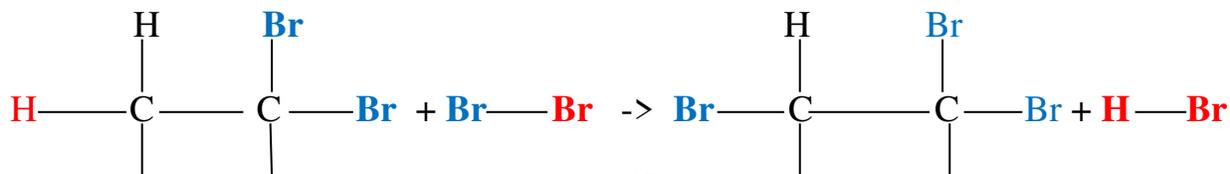
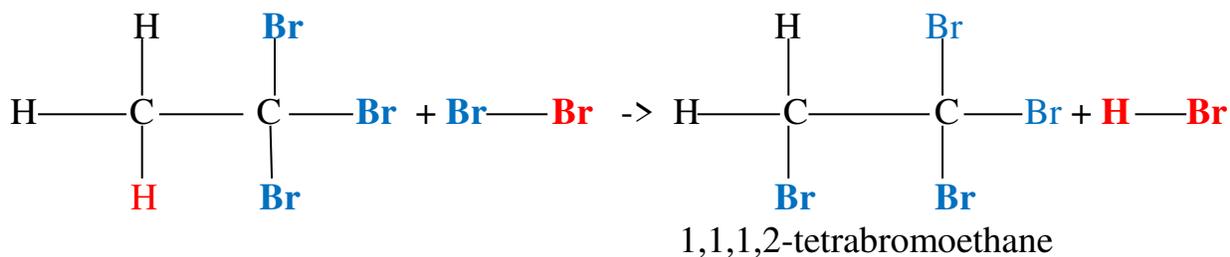
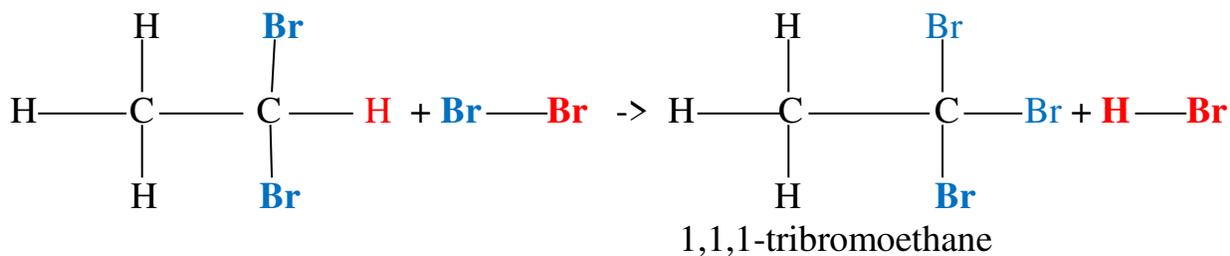
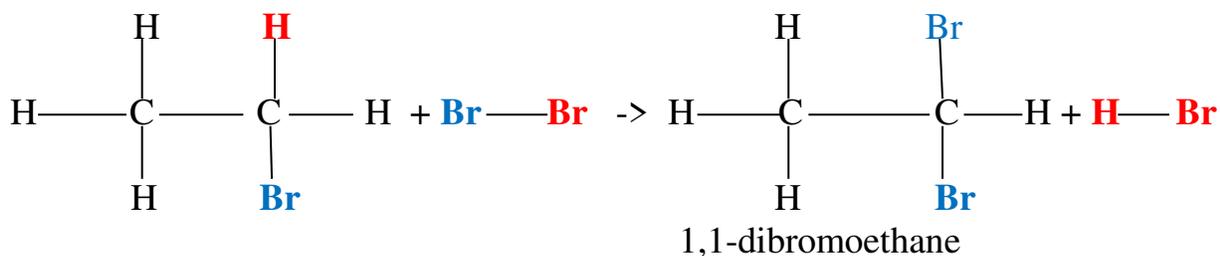
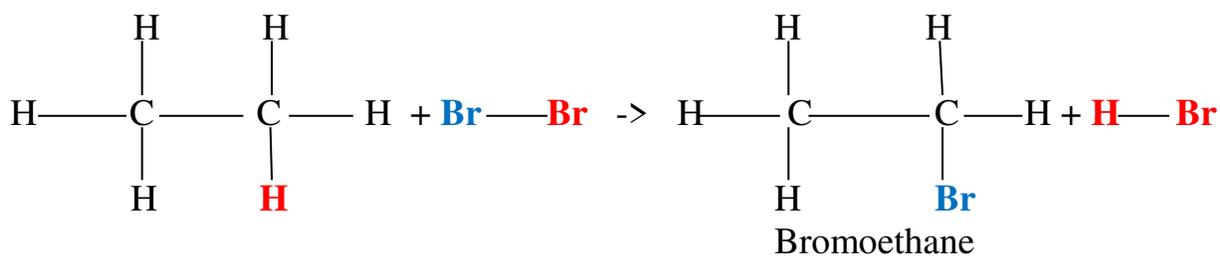
(c) trichloromethane + chlorine \rightarrow tetrachloromethane + Hydrogen chloride

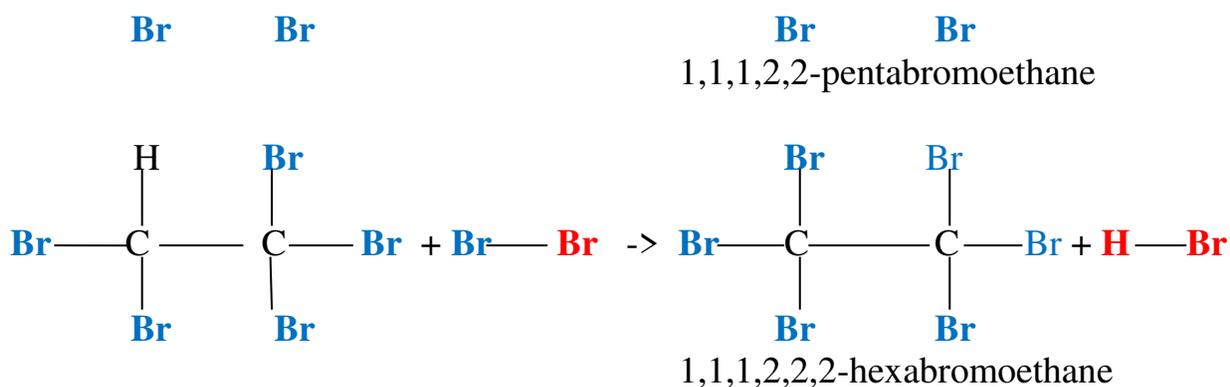


Ethane has no effect on bromine or chlorine in diffused light/dark. In sunlight, a mixture of bromine and ethane explodes to form a colourless mixture of bromoethane and hydrogen chloride gas. The red/brown colour of bromine gas fades.

Chemical equation

(a) Ethane + chlorine → Chloroethane + Hydrogen chloride





Uses of alkanes

1. Most alkanes are used as fuel e.g. Methane is used as biogas in homes. Butane is used as the Laboratory gas.
2. On cracking, alkanes are a major source of Hydrogen for the manufacture of ammonia/Haber process.
3. In manufacture of Carbon black which is a component in printers ink.
4. In manufacture of useful industrial chemicals like methanol, methanol, and chloromethane.

(ii) Alkenes

(a) Nomenclature/Naming

These are hydrocarbons with a general formula C_nH_{2n} and $\begin{array}{c} | \quad | \\ -\text{C}=\text{C}- \end{array}$ double bond as the functional group. n is the number of Carbon atoms in the molecule.

The carbon atoms are linked by at least one **double** bond to each other and single bonds to hydrogen atoms.

They include:

n	General/ Molecular formula	Structural formula	Name
1		Does not exist	
2	C_2H_4	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}=\text{C}-\text{H} \end{array} $	Ethene

		CH_2CH_2	
3	C_3H_8	$\begin{array}{ccccccc} & \text{H} & & \text{H} & & \text{H} & \\ & & & & & & \\ \text{H} & -\text{C} & = & \text{C} & - & \text{C} & -\text{H} \\ & & & & & & \\ & & & & & \text{H} & \end{array}$ <p style="text-align: center;">CH_2CHCH_3</p>	Propene
4	C_4H_{10}	$\begin{array}{ccccccc} & \text{H} & & \text{H} & & \text{H} & & \text{H} \\ & & & & & & & \\ \text{H} & -\text{C} & = & \text{C} & - & \text{C} & - & \text{C} & -\text{H} \\ & & & & & & & & \\ & & & & & \text{H} & & \text{H} & \end{array}$ <p style="text-align: center;">$\text{CH}_2\text{CHCH}_2\text{CH}_3$</p>	Butene
5	C_5H_{12}	$\begin{array}{ccccccc} & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \text{H} \\ & & & & & & & & & \\ \text{H} & -\text{C} & = & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & -\text{H} \\ & & & & & & & & & & \\ & & & & & \text{H} & & \text{H} & & \text{H} & \end{array}$ <p style="text-align: center;">$\text{CH}_2\text{CH}(\text{CH}_2)_2\text{CH}_3$</p>	Pentene
6	C_6H_{14}	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & \\ \text{H} & -\text{C} & = & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & -\text{H} \\ & & & & & & & & & & & & \\ & & & & & \text{H} & & \text{H} & & \text{H} & & \text{H} & \end{array}$ <p style="text-align: center;">$\text{CH}_2\text{CH}(\text{CH}_2)_3\text{CH}_3$</p>	Hexene
7	C_7H_{16}	$\begin{array}{ccccccc} & \text{H} & & \text{H} \\ & & & & & & & & & & & & & \\ \text{H} & -\text{C} & - & \text{C} & -\text{H} \\ & & & & & & & & & & & & & & \\ & & & \text{H} & \end{array}$	Heptene

		$\text{CH}_2 \text{CH} (\text{CH}_2)_4 \text{CH}_3$	
8	C_8H_{18}	$ \begin{array}{cccccccc} & \text{H} \\ & & & & & & & \\ \text{H} & - \text{C} = & \text{C} - \text{H} \\ & & & & & & & \\ & & & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \end{array} $ <p style="text-align: center;">$\text{CH}_2 \text{CH} (\text{CH}_2)_5 \text{CH}_3$</p>	Octene
9	C_9H_{20}	$ \begin{array}{ccccccccc} & \text{H} \\ & & & & & & & & \\ \text{H} & - \text{C} = & \text{C} - \text{H} \\ & & & & & & & & \\ & & & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \end{array} $ <p style="text-align: center;">$\text{CH}_2 \text{CH} (\text{CH}_2)_6 \text{CH}_3$</p>	Nonene
10	$\text{C}_{10}\text{H}_{22}$	$ \begin{array}{cccccccccc} & \text{H} \\ & & & & & & & & & \\ \text{H} & - \text{C} = & \text{C} - \text{H} \\ & & & & & & & & & \\ & & & \text{H} \end{array} $ <p style="text-align: center;">$\text{CH}_2 \text{CH} (\text{CH}_2)_7 \text{CH}_3$</p>	decene

Note

1. Since carbon is **tetravalent**, each atom of carbon in the alkene **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the double bond.

2. Since Hydrogen is **monovalent**, each atom of hydrogen in the alkene **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3. One member of the alkene, like alkanes, differ from the next/previous by a CH_2 group. They also form a homologous series.

e.g

Propene differ from ethene by one carbon and two Hydrogen atoms from ethene.

4. A homologous series of alkenes like that of alkanes:

- (i) differ by a CH_2 group from the next /previous consecutively
- (ii) have similar chemical properties

(iii) have similar chemical formula represented by the general formula C_nH_{2n}
(iv) the physical properties also show steady gradual change

5. The $C=C$ double bond in alkene is the functional group. A functional group is the **reacting site** of a molecule/compound.

6. The $C=C$ double bond in alkene can easily be broken to accommodate more two more monovalent atoms. The $C=C$ double bond in alkenes make it thus **unsaturated**.

7. An unsaturated hydrocarbon is one with a double $C=C$ or triple $C\equiv C$ carbon bonds in their molecular structure. Unsaturated hydrocarbon easily reacts to be **saturated**.

8. A saturated hydrocarbon is one without a double $C=C$ or triple $C\equiv C$ carbon bonds in their molecular structure.

Most of the reactions of alkenes take place at the $C=C$ bond.

(b) Isomers of alkenes

Isomers of alkenes like alkanes have the same molecular **general formula** but **different** molecular **structural formula**.

Ethene and propene do not form isomers. Isomers of alkenes are also named by using the IUPAC (International Union of Pure and Applied Chemistry) system of **nomenclature/naming**.

The IUPAC system of nomenclature of naming alkenes uses the following basic rules/guidelines:

1. Identify the longest continuous/straight carbon chain which contains the $C=C$ **double** bond get/determine the **parent** alkene.

2. Number the longest chain from the end of the chain which contains the $C=C$ **double** bond so the $C=C$ **double** bond lowest number possible.

3. Indicate the positions by splitting "**alk-positions-ene**" e.g. but-2-ene, pent-1,3-diene.

4. The position **indicated** must be for the carbon atom at the **lower** position in the $C=C$ **double bond**. i.e

But-2-ene means the double $C=C$ is between Carbon "2" and "3"

Pent-1,3-diene means there are two double bond one between carbon "1" and "2" and another between carbon "3" and "4"

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alkene. Name them fluoro-, chloro-, bromo-, iodo- if they are halogens

6. Use prefix di-, tri-, tetra-, penta-, hexa- to show the number of **double** $C=C$ bonds and **branches** attached to the alkene.

7. Position isomers can be formed when the $=C=C=$ double bond is shifted between carbon atoms e.g.

But-2-ene means the double $=C=C=$ is between Carbon "2" and "3"

But-1-ene means the double $=C=C=$ is between Carbon "1" and "2"

Both But-1-ene and But-2-ene are position isomers of Butene

8. Position isomers are molecules/compounds having the same general formula but different position of the functional group. i.e.

Butene has the molecular/general formula C_4H_8 position but can form both But-1-ene and But-2-ene as position isomers.

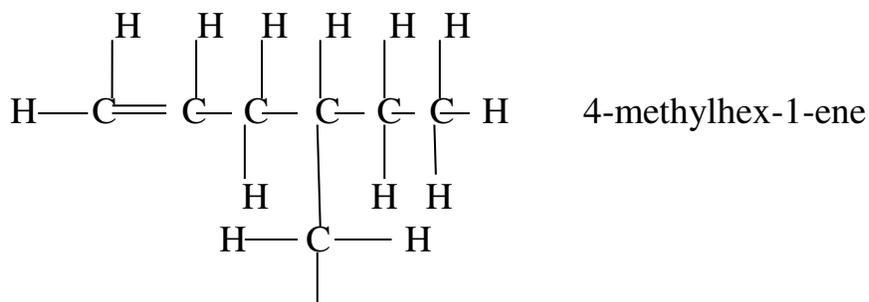
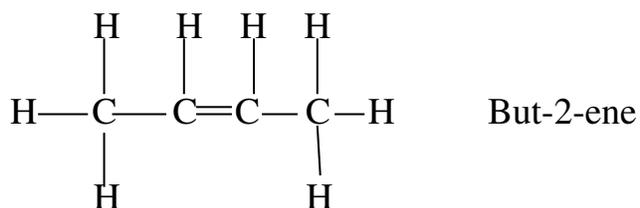
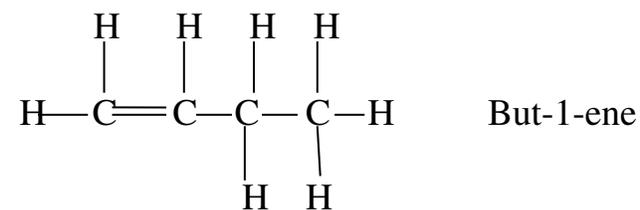
9. Like alkanes, an alkyl group can be attached to the alkene. Chain/branch isomers are thus formed.

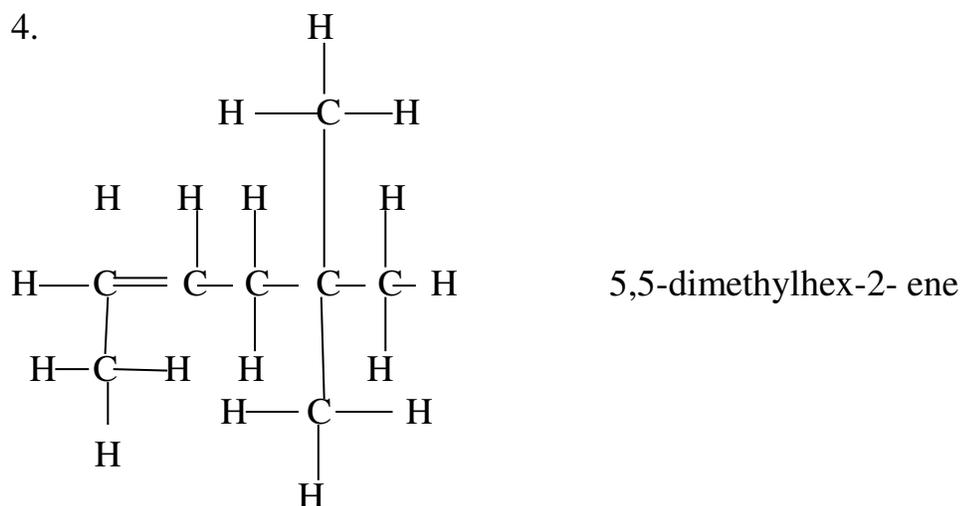
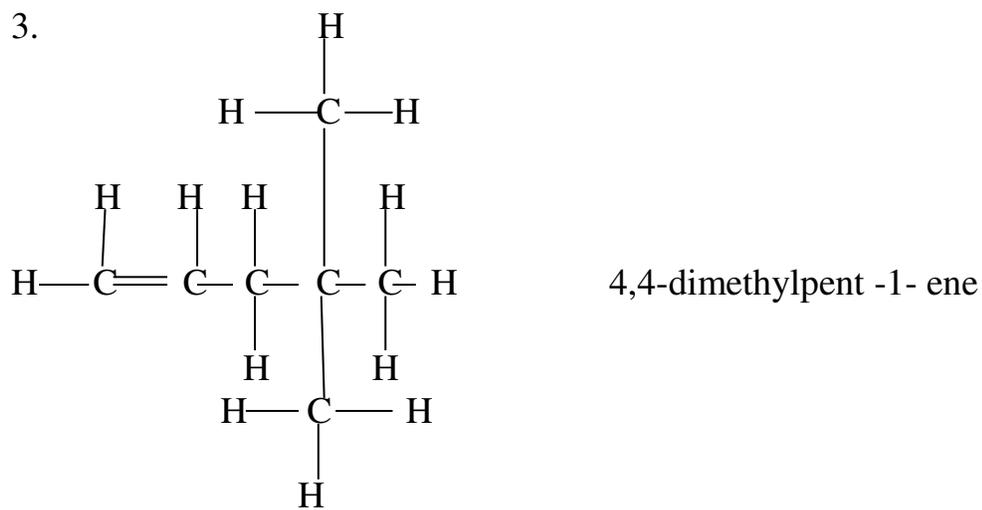
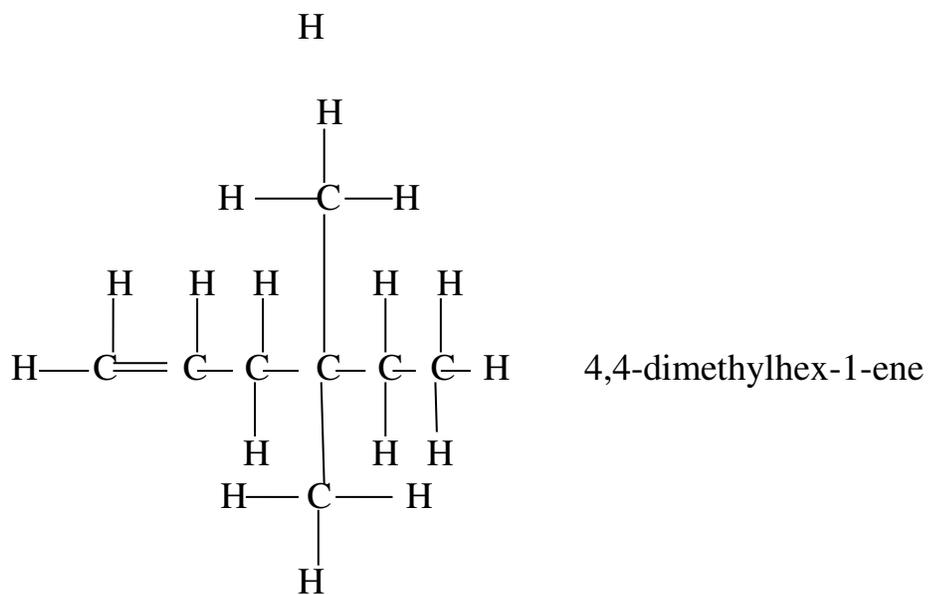
10. Chain/branch isomers are molecules/compounds having the same general formula but different structural formula e.g

Butene and 2-methyl propene both have the same general formula but different branching chain.

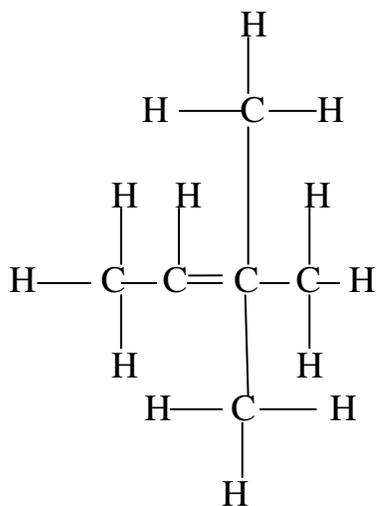
Practice on IUPAC nomenclature of alkenes

Name the following isomers of alkene





5.



2,2-dimethylbut -2- ene

8. $\text{H}_2\text{C CHCH}_2 \text{CH}_2 \text{CH}_3$

pent -1- ene

9. $\text{H}_2\text{C C}(\text{CH}_3)\text{CH}_2 \text{CH}_2 \text{CH}_3$

2-methylpent -1- ene

10. $\text{H}_2\text{C C}(\text{CH}_3)\text{C}(\text{CH}_3)_2 \text{CH}_2 \text{CH}_3$

2,3,3-trimethylpent -1- ene

11. $\text{H}_2\text{C C}(\text{CH}_3)\text{C}(\text{CH}_3)_2 \text{C}(\text{CH}_3)_2 \text{CH}_3$

2,3,3,4,4-pentamethylpent -1- ene

12. $\text{H}_3\text{C C}(\text{CH}_3)\text{C}(\text{CH}_3) \text{C}(\text{CH}_3)_2 \text{CH}_3$

2,3,4,4-tetramethylpent -2- ene

13. $\text{H}_2\text{C C}(\text{CH}_3)\text{C}(\text{CH}_3) \text{C}(\text{CH}_3) \text{CH}_3$

2,3,4-trimethylpent -1,3- diene

14. $\text{H}_2\text{C CBrCBr CBr CH}_3$

2,3,4-tribromopent -1,3- diene

15. $\text{H}_2\text{C CHCH CH}_2$

But -1,3- diene

16. $\text{Br}_2\text{C CBrCBr CBr}_2$

1,1,2,3,4,4-hexabromobut -1,3- diene

17. $\text{I}_2\text{C ClCl Cl}_2$

1,1,2,3,4,4-hexaiodobut -1,3- diene

18. $\text{H}_2\text{C C}(\text{CH}_3)\text{C}(\text{CH}_3) \text{CH}_2$

2,3-dimethylbut -1,3- diene

(c) Occurrence and extraction

At industrial level, alkenes are obtained from the cracking of alkanes. Cracking is the process of breaking long chain alkanes to smaller/shorter alkanes, an alkene and hydrogen gas at high temperatures.

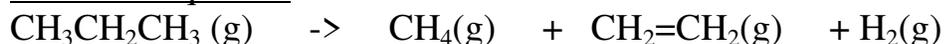
Cracking is a major source of useful hydrogen gas for manufacture of ammonia/nitric(V) acid/HCl i.e.

Long chain alkane \rightarrow smaller/shorter alkane + Alkene + Hydrogen gas

Examples

1. When irradiated with high energy radiation, Propane undergoes cracking to form methane gas, ethene and hydrogen gas.

Chemical equation



2. Octane undergoes cracking to form hydrogen gas, butene and butane gases

Chemical equation

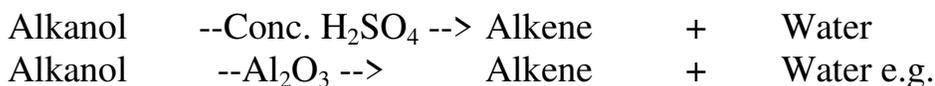


(d) School laboratory preparation of alkenes

In a school laboratory, alkenes may be prepared from dehydration of alkanols using:

(i) concentrated sulphuric(VI) acid (H_2SO_4).

(a) aluminium(III) oxide (Al_2O_3) i.e.

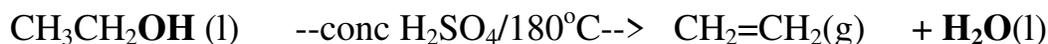


1.(a) At about 180°C , concentrated sulphuric(VI) acid dehydrates/removes water from ethanol to form ethene.

The gas produced contains traces of carbon(IV) oxide and sulphur(IV) oxide gas as impurities.

It is thus passed through concentrated sodium/potassium hydroxide solution to remove the impurities.

Chemical equation



(b) On heating strongly aluminium(III) oxide (Al_2O_3), it dehydrates/removes water from ethanol to form ethene.

Ethanol vapour passes through the hot aluminium (III) oxide which catalyses the dehydration.

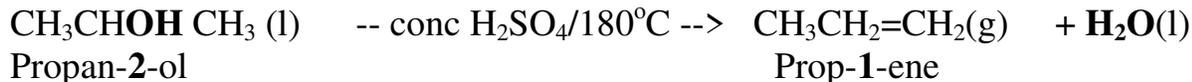
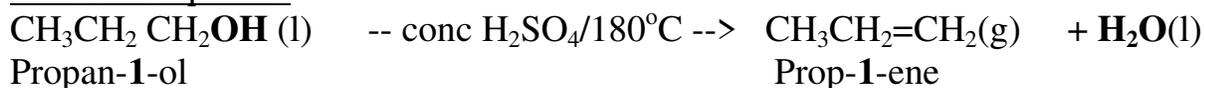
Activated aluminium(III)oxide has a very high affinity for water molecules/elements of water and thus dehydrates/ removes water from ethanol to form ethene.

Chemical equation



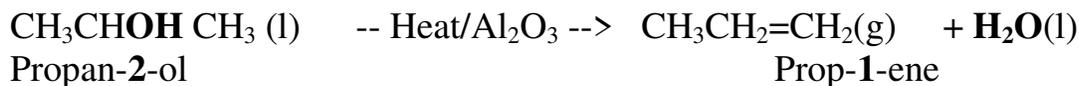
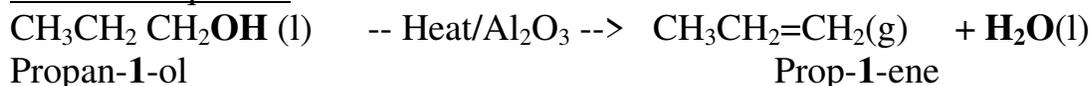
2(a) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by conc H_2SO_4 at about 180°C to propene(propene has no position isomers).

Chemical equation



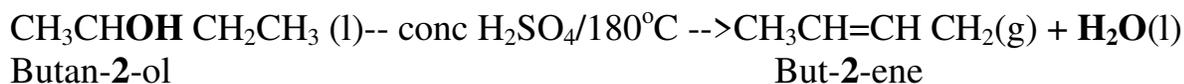
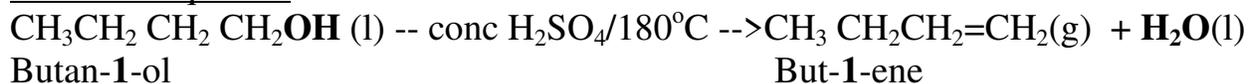
(b) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by heating strongly aluminium(III)oxide(Al_2O_3) form propene

Chemical equation



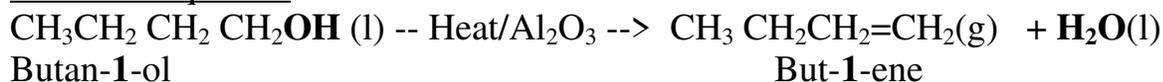
3(a) Butan-1-ol and Butan-2-ol(position isomers of butanol) are dehydrated by conc H_2SO_4 at about 180°C to But-1-ene and But-2-ene respectively

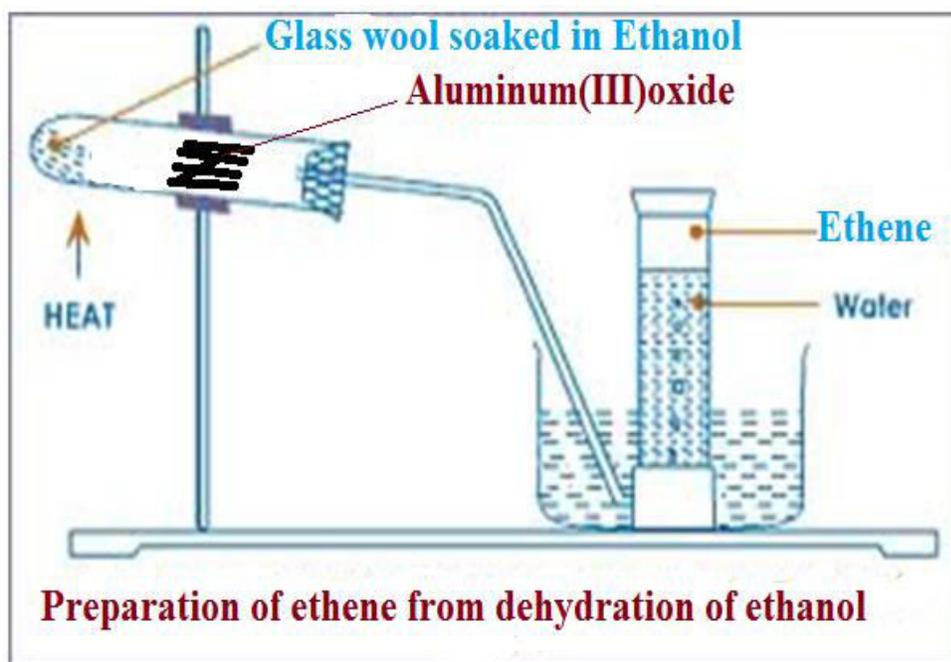
Chemical equation



(b) Butan-1-ol and Butan-2-ol are dehydrated by heating strongly aluminium (III) oxide (Al_2O_3) form But-1-ene and But-2-ene respectively.

Chemical equation





(e) Properties of alkenes

I. Physical properties

Like alkanes, alkenes are colourless gases, solids and liquids that are not poisonous. They are slightly soluble in water.

The solubility in water decreases as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene.

The melting and boiling point increase as the carbon chain increases.

This is because of the increase in van-der-Waals /intermolecular forces as the carbon chain increases.

The 1st four straight chain alkenes (ethene, propene, but-1-ene and pent-1-ene) are gases at room temperature and pressure.

The density of straight chain alkenes, like alkanes, increases with increasing carbon chain as the intermolecular forces increase, reducing the volume occupied by a given mass of the alkene.

Summary of physical properties of the 1st five alkenes

Alkene	General formula	Melting point(°C)	Boiling point(K)	State at room(298K) temperature and pressure atmosphere (101300Pa)

Ethene	CH ₂ CH ₂	-169	-104	gas
Propene	CH ₃ CHCH ₂	-145	-47	gas
Butene	CH ₃ CH ₂ CHCH ₂	-141	-26	gas
Pent-1-ene	CH ₃ (CH ₂ CHCH ₂	-138	30	liquid
Hex-1-ene	CH ₃ (CH ₂) CHCH ₂	-98	64	liquid

II. Chemical properties

(a) Burning/combustion

Alkenes burn with a **yellow/** luminous **sooty/** smoky flame in **excess** air to form carbon(IV) oxide and water.

Alkene + Air → carbon(IV) oxide + water (excess air/oxygen)

Alkenes burn with a **yellow/** luminous **sooty/** smoky flame in **limited** air to form carbon(II) oxide and water.

Alkene + Air → carbon(II) oxide + water (limited air)

Burning of alkenes with a **yellow/** luminous **sooty/** smoky flame is a confirmatory test for the **presence** of the =C=C= double bond because they have **higher C:H ratio**.

A homologous series with $\begin{array}{c} | \\ -C=C- \\ | \end{array}$ double or $-C\equiv C-$ triple bond is said to be **unsaturated**.

A homologous series with $\begin{array}{c} | \\ -C-C- \\ | \end{array}$ single bond is said to be **saturated**. Most of the reactions of the unsaturated compound involve trying to be saturated to form a

$\begin{array}{c} | \\ -C-C- \\ | \end{array}$ single bond.

Examples of burning alkenes

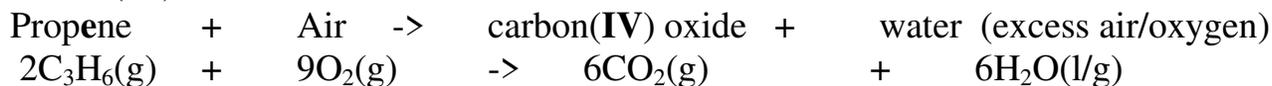
1.(a) Ethene when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.

Ethene + Air → carbon(IV) oxide + water (excess air/oxygen)
 $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l/g)$

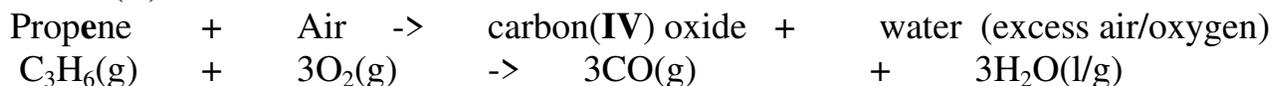
(b) Ethene when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.



2.(a) Propene when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.



(a) Propene when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.



(b) Addition reactions

An addition reaction is one which an unsaturated compound reacts to form a saturated compound. Addition reactions of alkenes are named from the reagent used to cause the addition/convert the double =C=C= to single C-C bond.

(i) Hydrogenation

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at high temperatures react with alkenes to form alkanes.

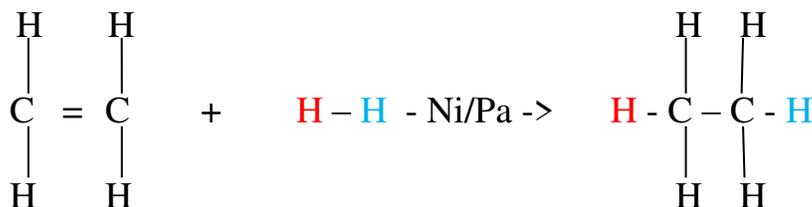
Examples

1. When Hydrogen gas is passed through liquid vegetable and animal **oil** at about 180°C in presence of Nickel catalyst, solid fat is formed.

Hydrogenation is thus used to **harden** oils to solid fat especially margarine.

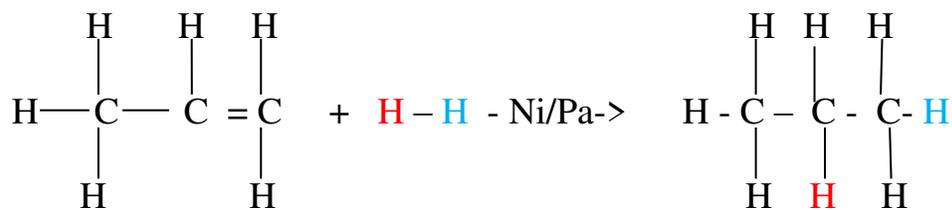
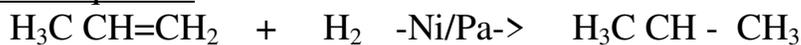
During hydrogenation, one hydrogen atom in the hydrogen molecule attach itself to one carbon and the other hydrogen to the second carbon breaking the double bond to single bond.

Chemical equation



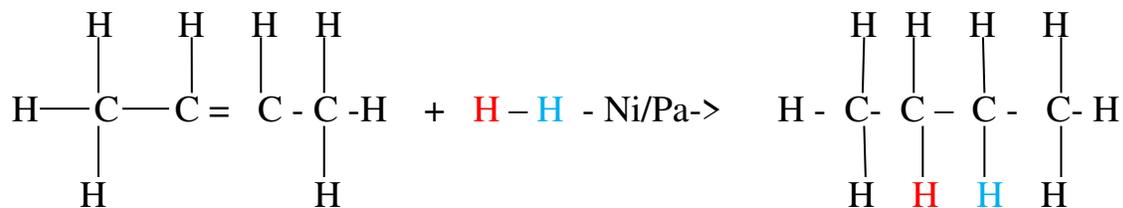
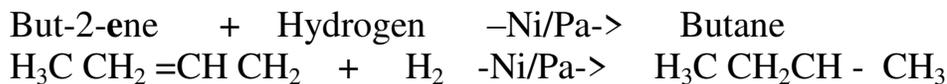
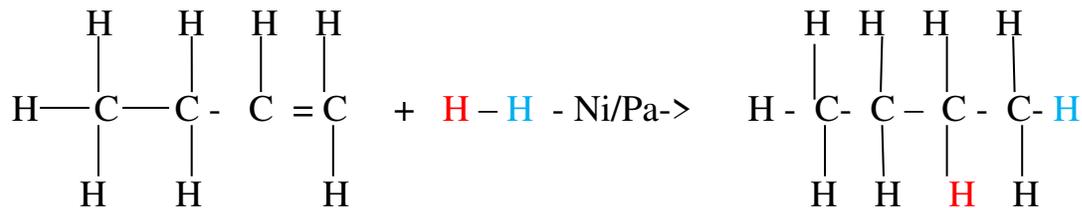
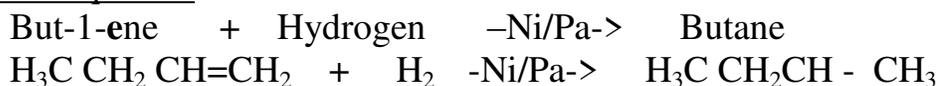
2. Propene undergo hydrogenation to form Propane

Chemical equation

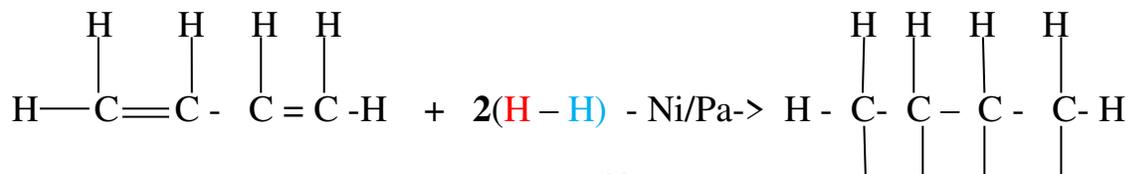
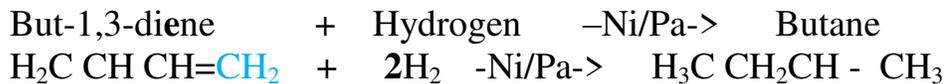


3. Both But-1-ene and But-2-ene undergo hydrogenation to form Butane

Chemical equation



4. But-1,3-diene should undergo hydrogenation to form Butane. The reaction uses **two moles** of hydrogen molecules/**four** hydrogen atoms to break the two double bonds.



H H H H

(ii) Halogenation.

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alkene to form an alkane.

The double bond in the alkene breaks and forms a single bond.

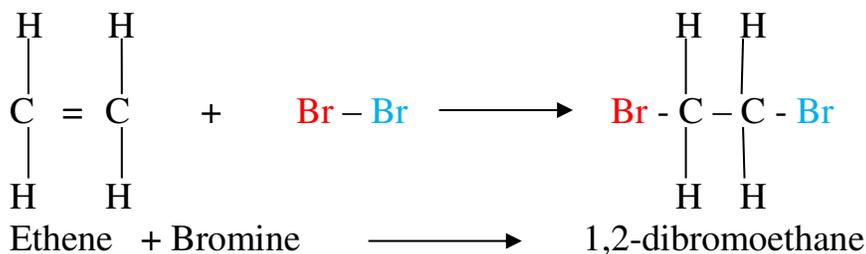
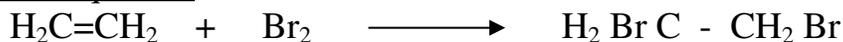
The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases/reduces.

One bromine atom bonds at the 1st carbon in the double bond while the other goes to the 2nd carbon.

Examples

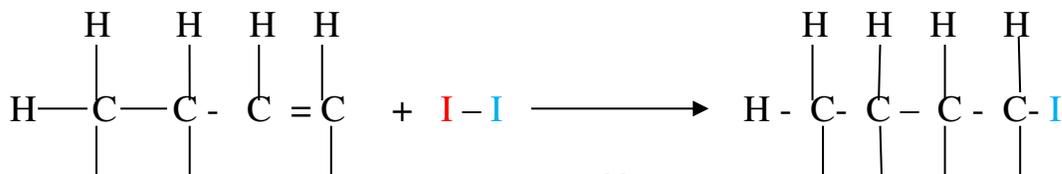
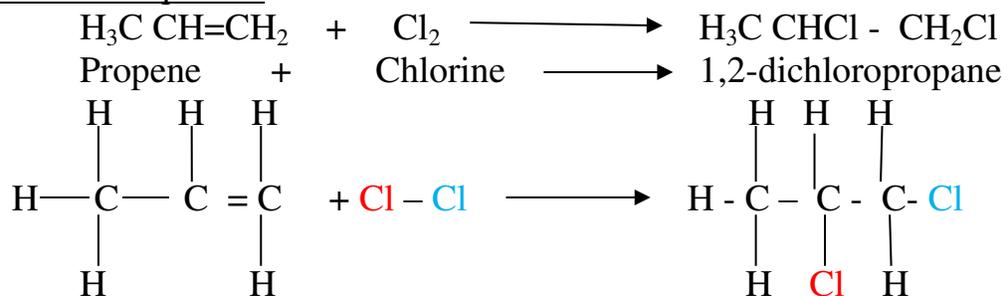
1. Ethene reacts with bromine to form 1,2-dibromoethane.

Chemical equation



2. Propene reacts with chlorine to form 1,2-dichloropropane.

Chemical equation

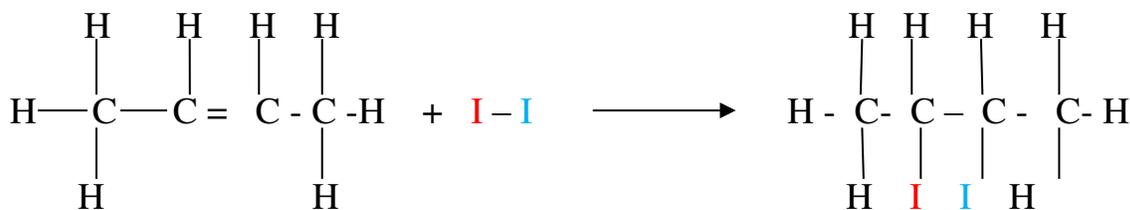
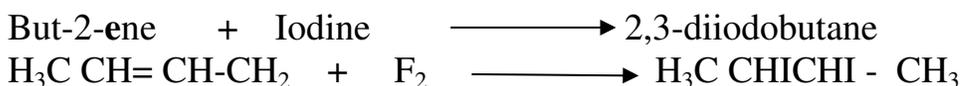
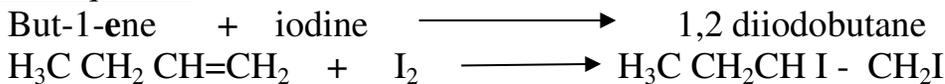


H H H H

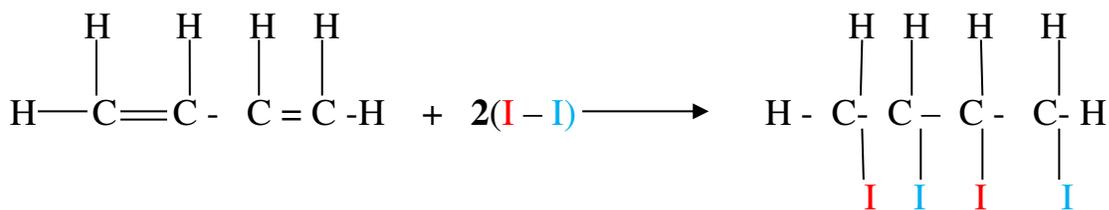
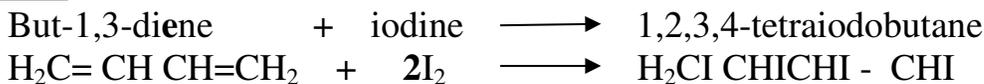
H H H H

3. Both But-1-ene and But-2-ene undergo halogenation with iodine to form 1,2-diiodobutane and 2,3-diiodobutane

Chemical equation



4. But-1,3-diene should undergo halogenation to form Butane. The reaction uses **two moles** of iodine molecules/**four** iodine atoms to break the two double bonds.



(iii) Reaction with hydrogen halides.

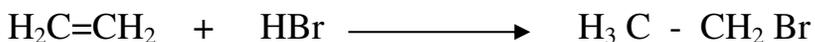
Hydrogen halides reacts with alkene to form a halogenoalkane. The double bond in the alkene break and form a single bond.

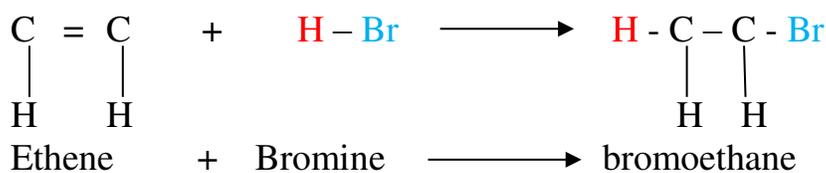
The main compound is one which the **hydrogen** atom bond at the carbon with **more hydrogen**.

Examples

1. Ethene reacts with hydrogen bromide to form bromoethane.

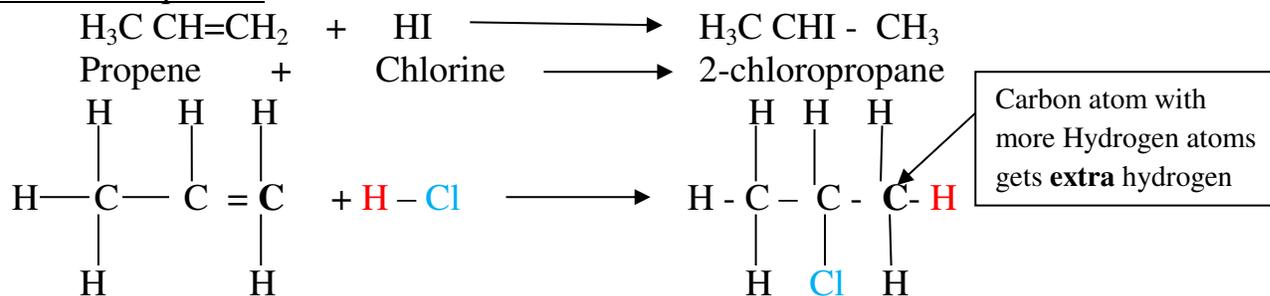
Chemical equation





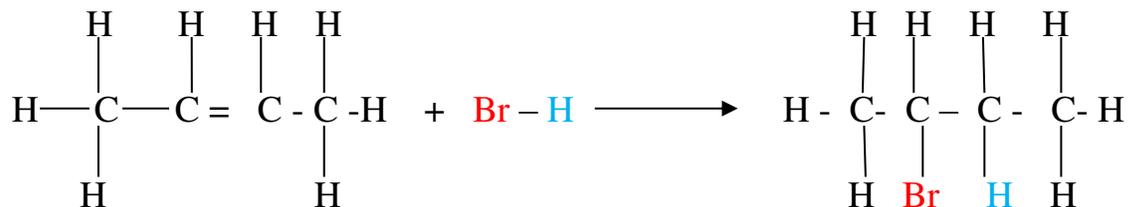
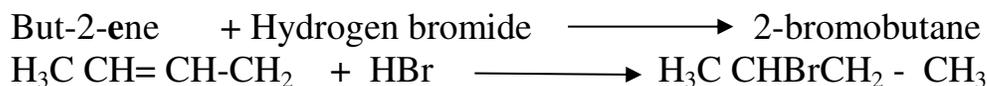
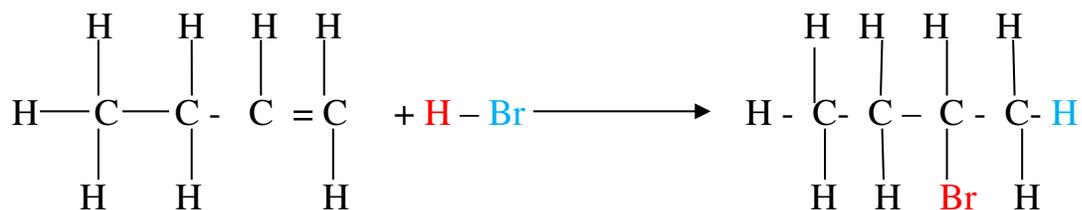
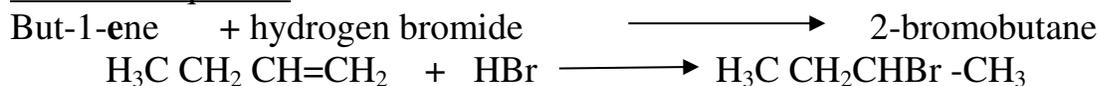
2. Propene reacts with hydrogen iodide to form 2-iodopropane.

Chemical equation



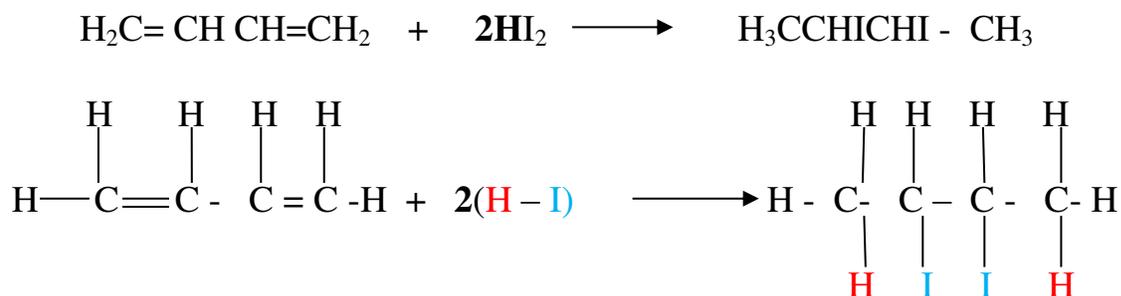
3. Both But-1-ene and But-2-ene reacts with hydrogen bromide to form 2-bromobutane

Chemical equation



4. But-1,3-diene react with hydrogen iodide to form 2,3- diiodobutane. The reaction uses **two moles** of hydrogen iodide molecules/**two** iodine atoms and two hydrogen atoms to break the two double bonds.





(iv) Reaction with bromine/chlorine water.

Chlorine and bromine water is formed when the halogen is dissolved in distilled water. Chlorine water has the formula HOCl (hypochlorous/chloric(I) acid). Bromine water has the formula HOBr (hydrobromic(I) acid).

During the addition reaction, the halogen moves to one carbon and the OH to the other carbon in the alkene at the =C=C= double bond to form a **halogenoalkanol**.

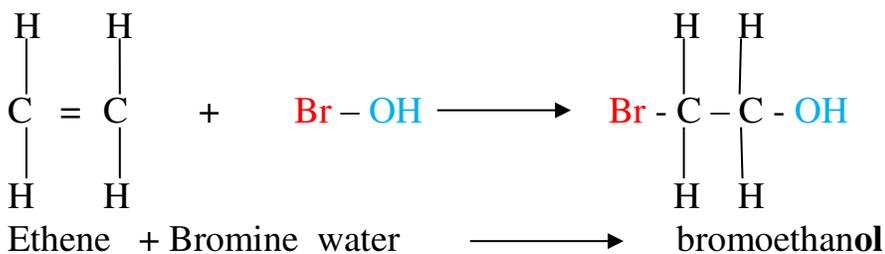
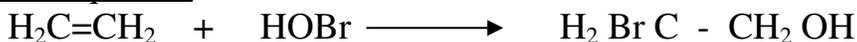
Bromine water + Alkene → bromoalkanol

Chlorine water + Alkene → chloroalkanol

Examples

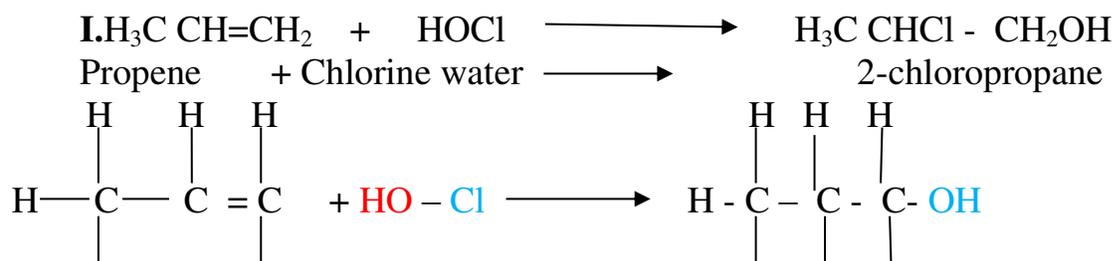
1. Ethene reacts with bromine water to form bromoethanol.

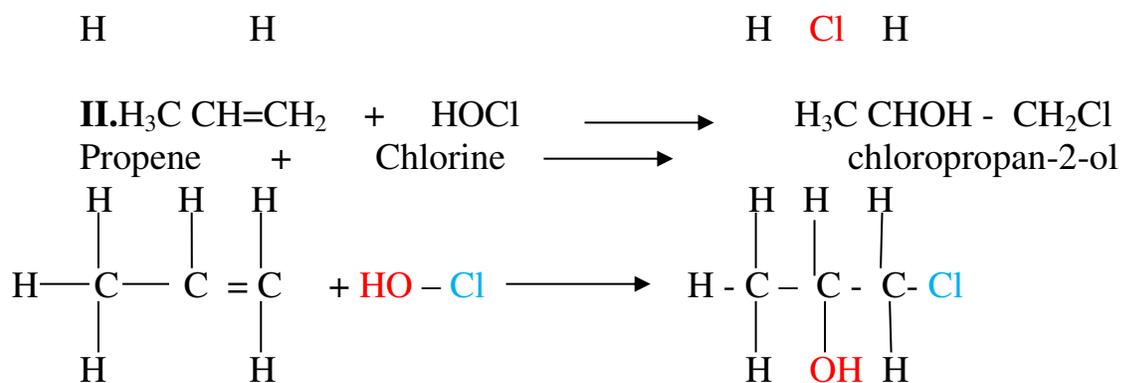
Chemical equation



2. Propene reacts with chlorine water to form chloropropan-2-ol / 2-chloropropan-1-ol.

Chemical equation

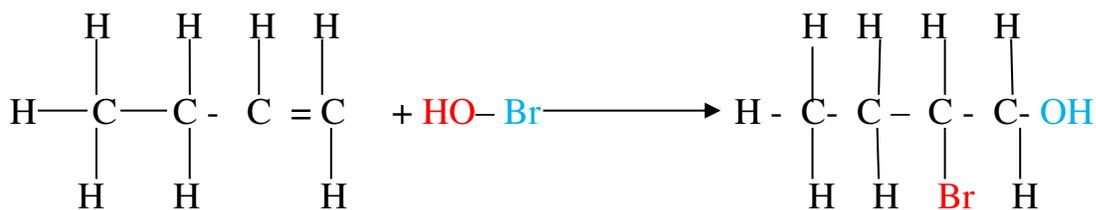
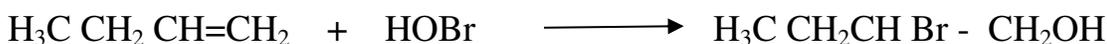




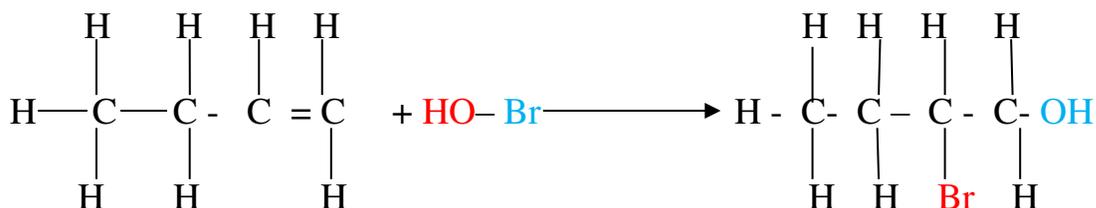
3. Both But-1-ene and But-2-ene react with bromine water to form 2-bromobutan-1-ol / 3-bromobutan-2-ol respectively

Chemical equation

I. But-1-ene + bromine water \longrightarrow 2-bromobutan-1-ol



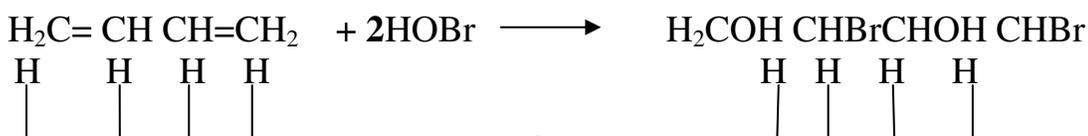
II. But-2-ene + bromine water \longrightarrow 3-bromobutan-2-ol

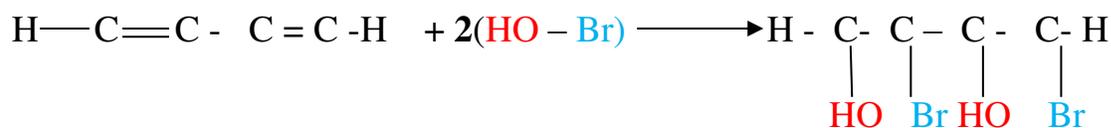
$$\text{H}_3\text{C}-\text{CH}=\text{CH}-\text{CH}_3 + \text{HOBr} \longrightarrow \text{H}_3\text{C}-\text{CH}_2\text{OH}-\text{CH}(\text{Br})-\text{CH}_3$$


4. But-1,3-diene reacts with bromine water to form Butan-1,3-diol.

The reaction uses **two moles** of bromine water molecules to break the two double bonds.

But-1,3-diene + bromine water \longrightarrow 2,4-dibromobutan-1,3-diol





(v) Oxidation.

Alkenes are oxidized to alkanols with **duo/double** functional groups by oxidizing agents.

When an alkene is bubbled into orange acidified potassium/sodium dichromate (VI) solution, the colour of the oxidizing agent changes to green.

When an alkene is bubbled into purple acidified potassium/sodium manganate(VII) solution, the oxidizing agent is decolorized.

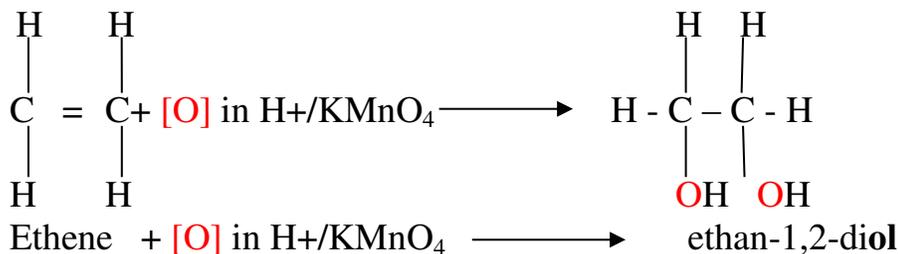
Examples

1 Ethene is oxidized to ethan-1,2-diol by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.

The purple acidified potassium/sodium manganate(VII) solution is decolorized.

The orange acidified potassium/sodium dichromate(VI) solution turns to green.

Chemical equation

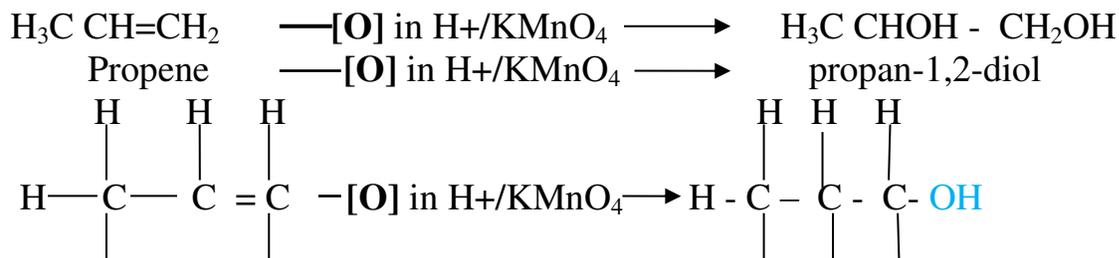


2. Propene is oxidized to propan-1,2-diol by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.

The purple acidified potassium/sodium manganate(VII) solution is decolorized.

The orange acidified potassium/sodium dichromate(VI) solution turns to green.

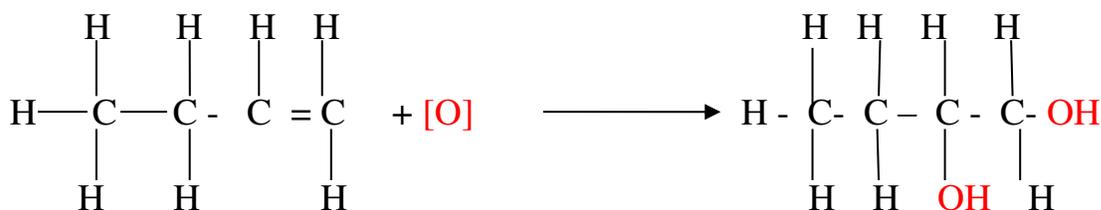
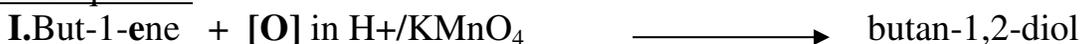
Chemical equation





3. Both But-1-ene and But-2-ene react with bromine water to form butan-1,2-diol and butan-2,3-diol

Chemical equation



(v) Hydrolysis.

Hydrolysis is the reaction of a compound with water/addition of H-OH to a compound.

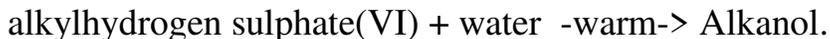
Alkenes undergo hydrolysis to form alkanols .

This takes place in two steps:

(i) Alkenes react with **concentrated sulphuric(VI) acid** at room temperature and pressure to form **alkylhydrogen sulphate(VI)**.



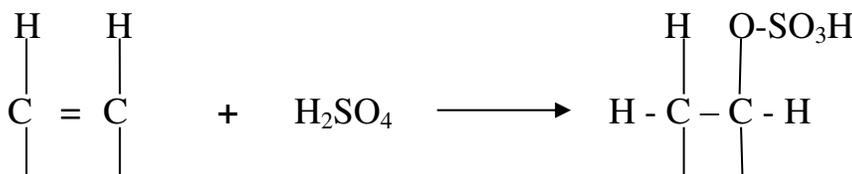
(ii) On adding **water** to alkylhydrogen sulphate(VI) then warming, an alkanol is formed.



Examples

(i) Ethene reacts with cold concentrated sulphuric(VI) acid to form ethyl hydrogen sulphate(VII)

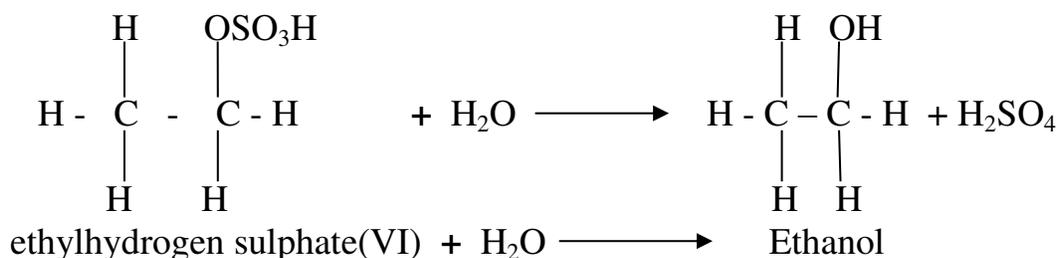
Chemical equation





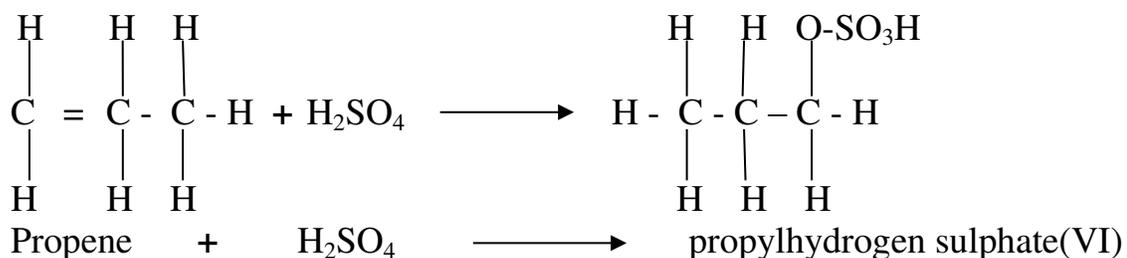
(ii) Ethylhydrogen sulphate(VI) is hydrolysed by water to ethanol

Chemical equation



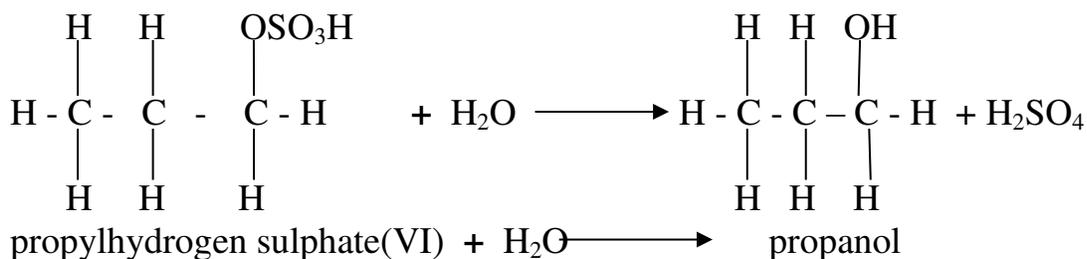
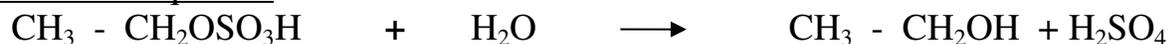
2. Propene reacts with cold concentrated sulphuric(VI) acid to form propyl hydrogen sulphate(VII)

Chemical equation



(ii) Propylhydrogen sulphate(VI) is hydrolysed by water to propanol

Chemical equation



(vi) Polymerization/self addition

Addition polymerization is the process where a small unsaturated monomer (alkene) molecule join together to form a large saturated molecule.

Only alkenes undergo addition polymerization.

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix “**poly**” before the name of monomer to form a **polyalkene**

During addition polymerization

(i)the double bond in alkenes break

(ii)free radicals are formed

(iii)the free radicals collide with each other and join to form a larger molecule.

The more collisions the larger the molecule.

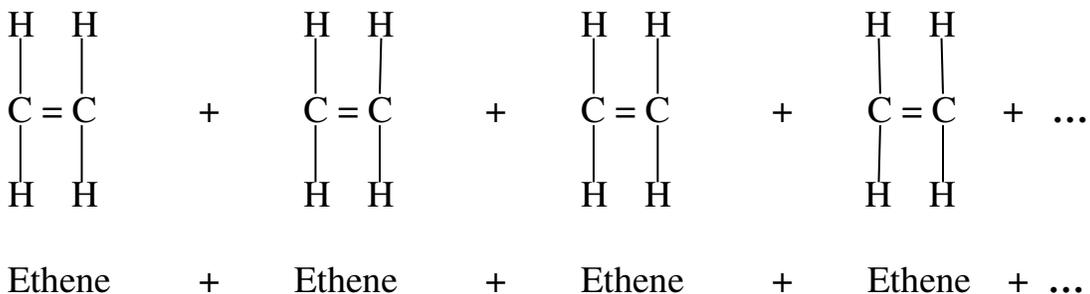
Examples of addition polymerization

1. Formation of Polyethene

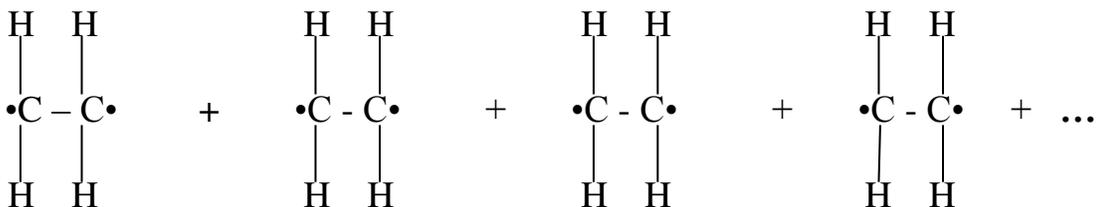
Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

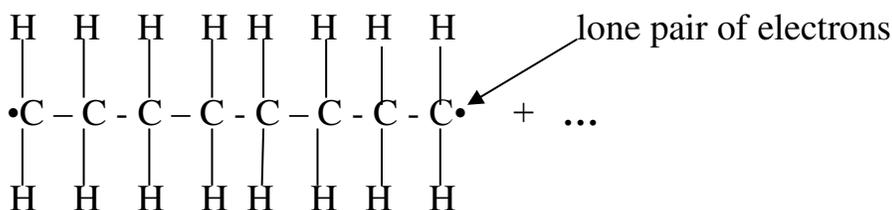


(ii)the double bond joining the ethane molecule break to free readicals



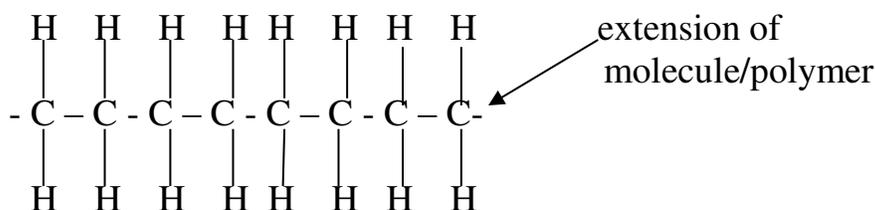
Ethene radical + Ethene radical + Ethene radical + Ethene radical + ...

(iii)the free radicals collide with each other and join to form a larger molecule

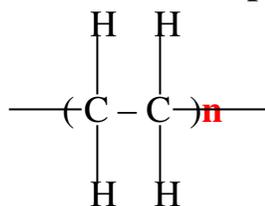


Lone pair of electrons can be used to join more monomers to form longer polyethene.

Polyethene molecule can be represented as:



Since the molecule is a **repetition** of one monomer, then the polymer is:



Where **n** is the number of monomers in the polymer. The number of monomers in the polymer can be determined from the molar mass of the polymer and monomer from the relationship:

$$\text{Number of monomers/repeating units in monomer} = \frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$$

Examples

Polythene has a molar mass of 4760. Calculate the number of ethene molecules in the polymer (C=12.0, H=1.0)

$$\text{Number of monomers/repeating units in polyomer} = \frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$$

$$\Rightarrow \text{Molar mass ethene (C}_2\text{H}_4) = 28 \text{ Molar mass polyethene} = 4760$$

$$\text{Substituting} \quad \frac{4760}{28} = \underline{170 \text{ ethene molecules}}$$

The **commercial** name of polyethene is **polythene**.

It is an elastic, tough, transparent and durable plastic.

Polythene is used:

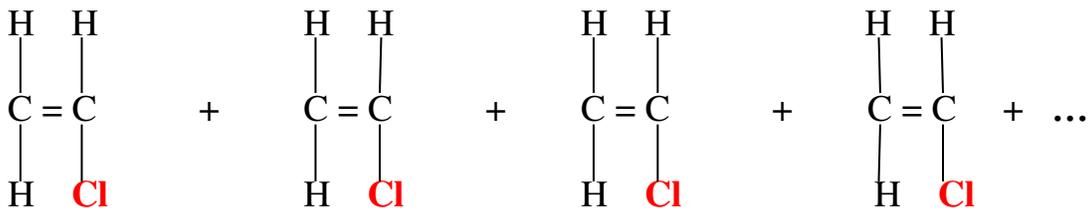
- (i) in making plastic bag
- (ii) bowls and plastic bags
- (iii) packaging materials

2. Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

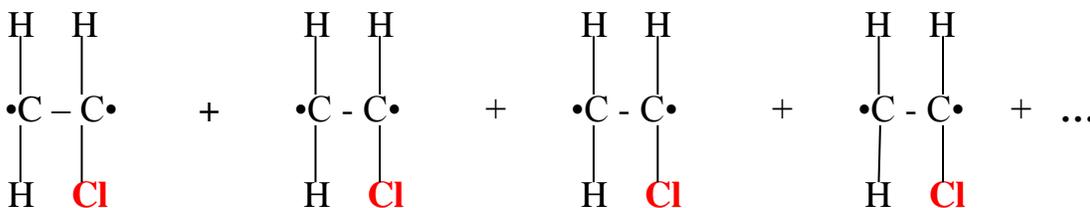
During polymerization:

(i) many molecules are brought nearer to each other by the high pressure (which reduces the volume occupied by reacting particles)

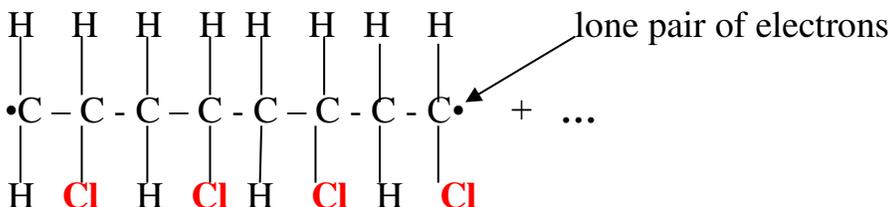


chloroethene + chloroethene + chloroethene + chloroethene + ...

(ii) the double bond joining the chloroethene molecule break to free radicals

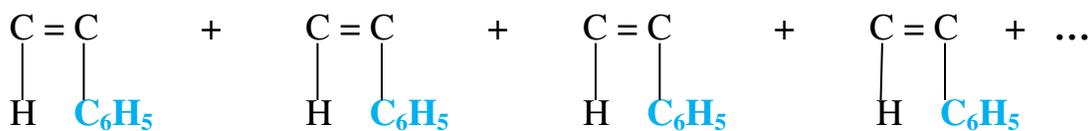


(iii) the free radicals collide with each other and join to form a larger molecule



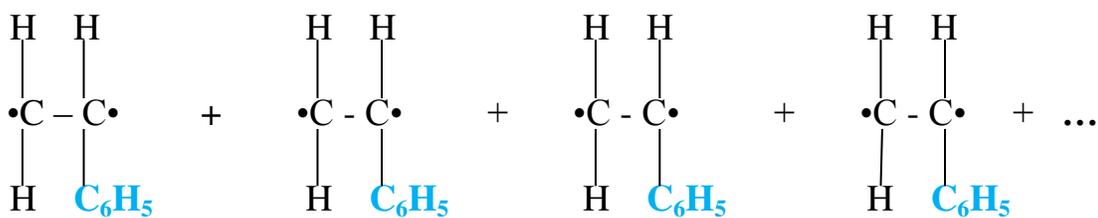
Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:

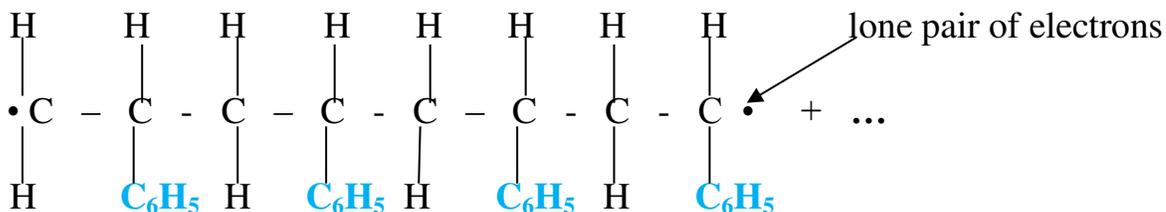


phenylethene + phenylethene + phenylethene + phenylethene + ...

(ii) the double bond joining the phenylethene molecule break to free radicals

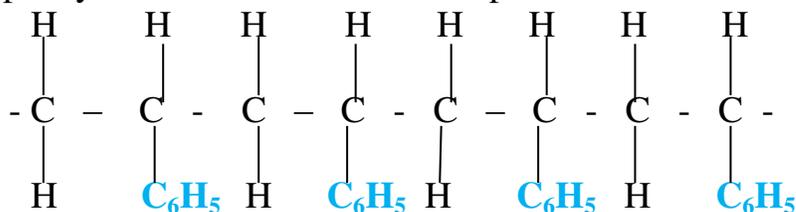


(iii) the free radicals collide with each other and join to form a larger molecule

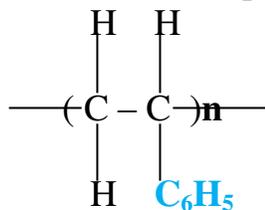


Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polyphenylthene has a molar mass of 4760. Calculate the number of phenylethene molecules in the polymer (C=12.0, H=1.0,)

$$\text{Number of monomers/repeating units in monomer} = \frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$$

$$\Rightarrow \text{Molar mass ethene (C}_8\text{H}_8) = 104 \text{ Molar mass polyethene} = 4760$$

$$\text{Substituting } \frac{4760}{104} = 45.7692 \Rightarrow 45 \text{ polyphenylethene molecules (whole number)}$$

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i) in making packaging material for carrying delicate items like computers, radion, calculators.

(ii) ceiling tiles

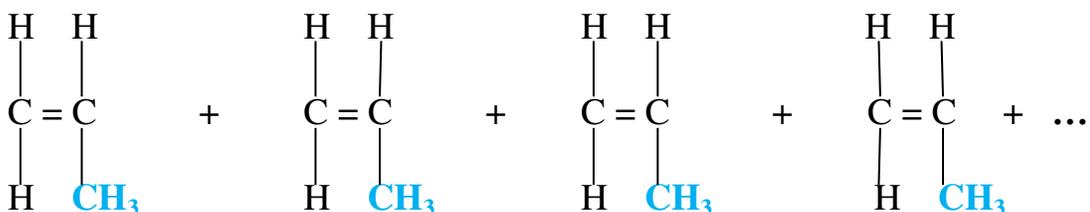
(iii) clothe linings

4. Formation of Polypropene

Polypropene is an addition polymer formed when propene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

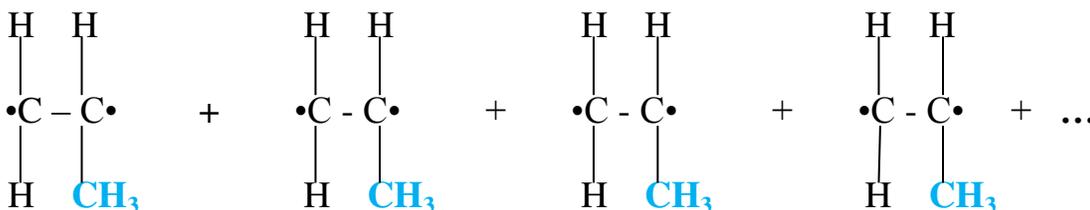
During polymerization:

(i) many molecules are brought nearer to each other by the high pressure (which reduces the volume occupied by reacting particles)

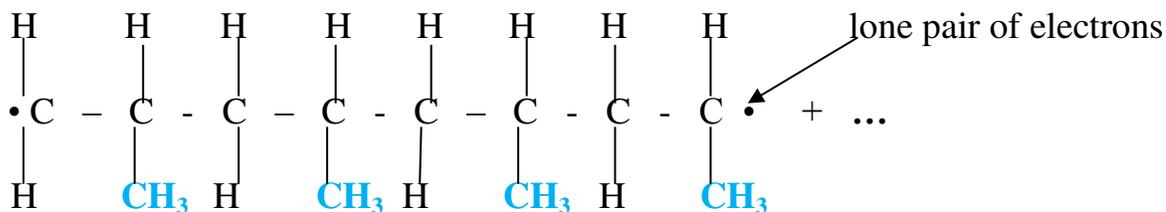


propene + propene + propene + propene + ...

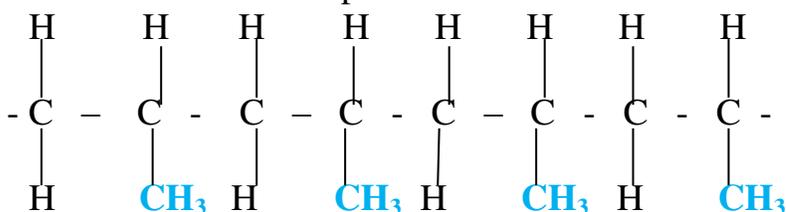
(ii) the double bond joining the phenylethene molecule break to free radicals



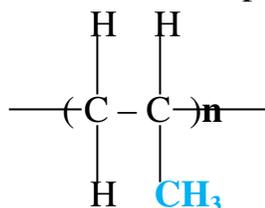
(iii) the free radicals collide with each other and join to form a larger molecule



Lone pair of electrons can be used to join more monomers to form longer propene. propene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polypropene has a molar mass of 4760. Calculate the number of propene molecules in the polymer (C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = $\frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$

=> Molar mass propene (C₃H₆) = 44 Molar mass polyethene = 4760

Substituting $\frac{4760}{44} = 108.1818 \Rightarrow 108$ propene molecules (whole number)

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

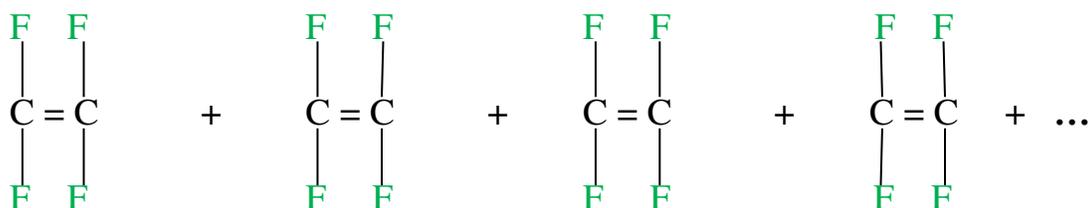
- (i) in making packaging material for carrying delicate items like computers, radion, calculators.
- (ii) ceiling tiles
- (iii) clothe linings

5. Formation of Polytetrafluoroethene

Polytetrafluoroethene is an addition polymer formed when tetrafluoroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

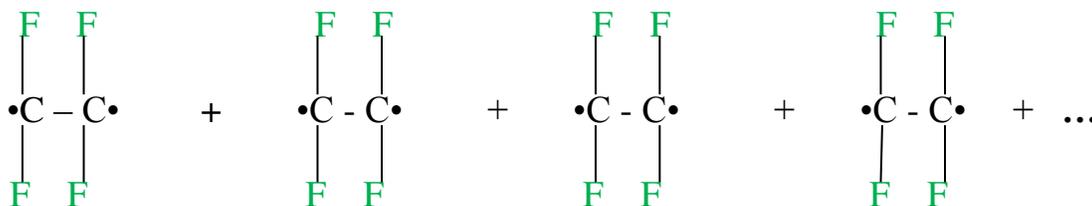
During polymerization:

(i) many molecules are brought nearer to each other by the high pressure (which reduces the volume occupied by reacting particles)

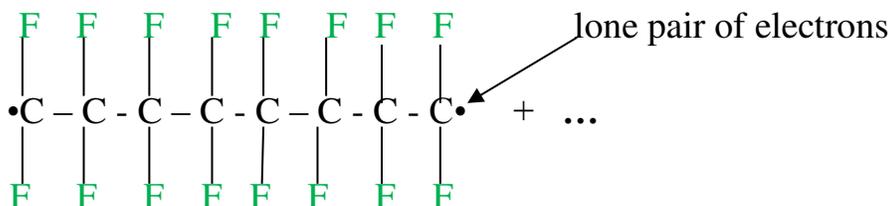


tetrafluoroethene + tetrafluoroethene + tetrafluoroethene + tetrafluoroethene + ...

(ii) the double bond joining the tetrafluoroethene molecule breaks to form free radicals



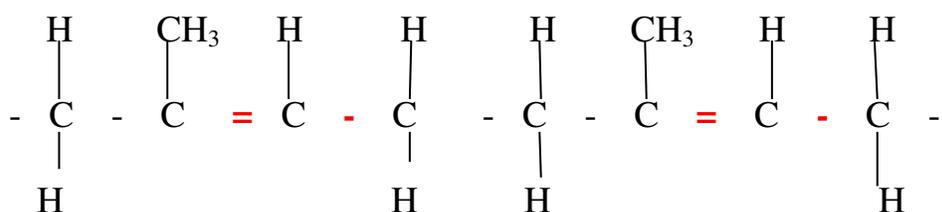
(iii) the free radicals collide with each other and join to form a larger molecule



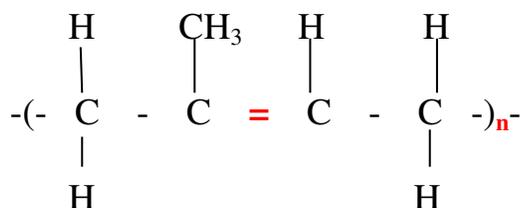
Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:



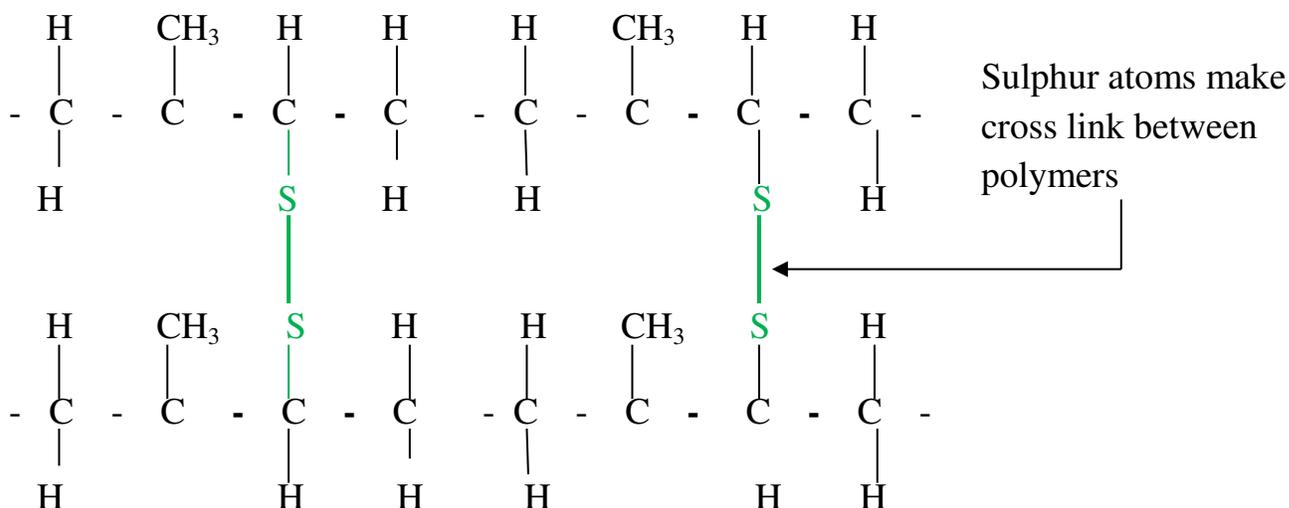


Generally the structure of rubber is thus;



Pure rubber is soft and sticky. It is used to make erasers, car tyres. Most of it is vulcanized. Vulcanization is the process of heating rubber with sulphur to make it harder/tougher.

During vulcanization the sulphur atoms form a cross link between chains of rubber molecules/polymers. This decreases the number of C=C double bonds in the polymer.



Vulcanized rubber is used to make **tyres, shoes** and **valves**.

7. Formation of synthetic rubber

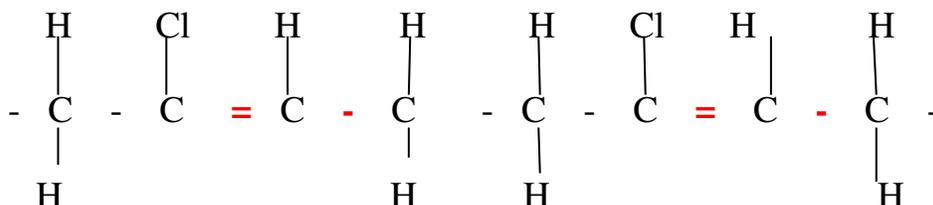
Synthetic rubber is able to resist action of oil, abrasion and organic solvents which rubber cannot.

Common synthetic rubber is a polymer of 2-chlorobut-1,3-diene ;

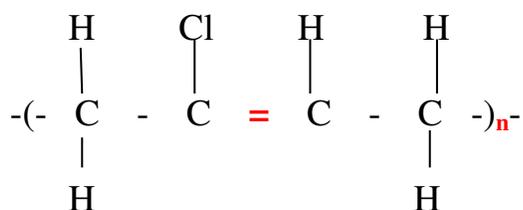




During polymerization to synthetic rubber, one double C=C bond is broken to self add to another molecule. The double bond remaining move to carbon "2" thus;



Generally the structure of rubber is thus;



Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

(c) Test for the presence of $-\overset{|}{\text{C}}=\overset{|}{\text{C}}-$ double bond.

(i) Burning/combustion

All unsaturated hydrocarbons with a $-\text{C}=\text{C}-$ or $-\text{C}\equiv\text{C}-$ bond burn with a yellow sooty flame.

Experiment

Scoop a sample of the substance provided in a clean metallic spatula. Introduce it on a Bunsen burner.

Observation	Inference
Solid melt then burns with a yellow sooty flame	$-\overset{ }{\text{C}}=\overset{ }{\text{C}}-$, $-\text{C}\equiv\text{C}-$ bond

(ii) Oxidation by acidified $\text{KMnO}_4/\text{K}_2\text{Cr}_2\text{O}_7$

Bromine water, Chlorine water and Oxidizing agents acidified $\text{KMnO}_4/\text{K}_2\text{Cr}_2\text{O}_7$ change to **unique** colour in presence of $-\overset{|}{\text{C}}=\overset{|}{\text{C}}-$

or $-\text{C} \equiv \text{C}-$ bond.

Experiment

Scoop a sample of the substance provided into a clean test tube. Add 10cm³ of distilled water. Shake. Take a portion of the solution mixture. Add three drops of acidified $\text{KMnO}_4/\text{K}_2\text{Cr}_2\text{O}_7$.

Observation	Inference
Acidified KMnO_4 decolorized	$-\overset{\text{I}}{\text{C}} = \overset{\text{I}}{\text{C}}-$
Orange colour of acidified $\text{K}_2\text{Cr}_2\text{O}_7$ turns green	$-\text{C} \equiv \text{C}-$ bond
Bromine water is decolorized	
Chlorine water is decolorized	

(d) Some uses of Alkenes

1. In the manufacture of plastic
2. Hydrolysis of ethene is used in industrial manufacture of ethanol.
3. In ripening of fruits.
4. In the manufacture of detergents.

(iii) Alkynes

(a) Nomenclature/Naming

These are hydrocarbons with a general formula C_nH_{2n-2} and $-C\equiv C-$ double bond as the functional group. n is the number of Carbon atoms in the molecule.

The carbon atoms are linked by at least one **triple** bond to each other and single bonds to hydrogen atoms.

They include:

n	General/ Molecular formula	Structural formula	Name
1		Does not exist	-
2	C_2H_2	$H-C\equiv C-H$ <p style="text-align: center;">CH CH</p>	Eth y ne
3	C_3H_4	$H-C\equiv C-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-H$ <p style="text-align: center;">CH C CH₃</p>	Prop y ne
4	C_4H_6	$H-C\equiv C-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-H$ <p style="text-align: center;">CH C CH₂CH₃</p>	But y ne
5	C_5H_8	$H-C\equiv C-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-H$ <p style="text-align: center;">CH C (CH₂)₂CH₃</p>	Pent y ne
6	C_6H_{10}	$H-C\equiv C-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-\overset{\overset{H}{ }}{\underset{\underset{H}{ }}{C}}-H$	Hex y ne

		$\begin{array}{ccccccc} \text{H} & - & \text{C} & \equiv & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & & & & & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \\ & & & & & & \text{CH} & & \text{C} & & (\text{CH}_2)_3 & & \text{CH}_3 & & \end{array}$	
--	--	---	--

7	C_7H_{12}	$\begin{array}{ccccccc} & & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ & & & & & & \\ \text{H} & - & \text{C} & \equiv & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & \\ & & \text{H} & & \\ & & \text{CH} & & \text{C} & & (\text{CH}_2)_4 & & \text{CH}_3 & & & & & & \end{array}$	Heptyne
8	C_8H_{14}	$\begin{array}{ccccccc} & & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ & & & & & & & \\ \text{H} & - & \text{C} & \equiv & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & & & \\ & & \text{H} & & \\ & & \text{CH} & & \text{C} & & (\text{CH}_2)_5 & & \text{CH}_3 & & & & & & & & \end{array}$	Octyne
9	C_9H_{16}	$\begin{array}{ccccccc} & & \text{H} \\ & & & & & & & & \\ \text{H} & - & \text{C} & \equiv & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & & & & & \\ & & \text{H} & & \\ & & \text{CH} & & \text{C} & & (\text{CH}_2)_6 & & \text{CH}_3 & & & & & & & & & & \end{array}$	Nonyne
10	$\text{C}_{10}\text{H}_{18}$	$\begin{array}{ccccccc} & & \text{H} \\ & & & & & & & & & & \\ \text{H} & - & \text{C} & \equiv & \text{C} & - & \text{H} \\ & & & & & & & & & & & & & & & & & & & & \\ & & \text{H} & & \\ & & \text{CH} & & \text{C} & & (\text{CH}_2)_7 & & \text{CH}_3 & & & & & & & & & & & & \end{array}$	Decyne

Note

1. Since carbon is **tetravalent**, each atom of carbon in the alkyne **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the triple bond.

2. Since Hydrogen is **monovalent**, each atom of hydrogen in the alkyne **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3. One member of the alkyne, like alkenes and alkanes, differ from the next/previous by a CH_2 group (molar mass of 14 atomic mass units). They thus form a homologous series.

e.g

Propyne differ from ethyne by (14 a.m.u) one carbon and two Hydrogen atoms from ethyne.

4. A homologous series of alkenes like that of alkanes:

- (i) differ by a CH_2 group from the next /previous consecutively
- (ii) have similar chemical properties
- (iii) have similar chemical formula with general formula $\text{C}_n\text{H}_{2n-2}$
- (iv) the physical properties also show steady gradual change

5. The $-\text{C}\equiv\text{C}-$ triple bond in alkyne is the functional group. The functional group is the **reacting site** of the alkynes.

6. The $-\text{C}=\text{C}-$ double bond in alkyne can easily be broken to accommodate more /four more monovalent atoms. The $-\text{C}\equiv\text{C}-$ triple bond in alkynes make it thus **unsaturated** like alkenes.

7. Most of the reactions of alkynes like alkenes take place at the $-\text{C}\equiv\text{C}-$ triple bond.

(b) Isomers of alkynes

Isomers of alkynes have the same molecular **general formula** but different molecular **structural formula**.

Isomers of alkynes are also named by using the IUPAC (International Union of Pure and Applied Chemistry) system of **nomenclature/naming**.

The IUPAC system of nomenclature of naming alkynes uses the following basic rules/guidelines:

1. Identify the longest continuous/straight carbon chain which contains the $-\text{C}\equiv\text{C}-$ **triple bond** to get/determine the **parent** alkene.

2. Number the longest chain from the end of the chain which contains the $-\text{C}\equiv\text{C}-$ **triple bond** so as $-\text{C}\equiv\text{C}-$ **triple bond** get lowest number possible.

3. Indicate the positions by splitting “**alk-positions-yne**” e.g. but-2-yne, pent-1,3-diyne.

4. The position **indicated** must be for the carbon atom at the **lower** position in the $-\text{C}\equiv\text{C}-$ **triple bond**. i.e

But-2-yne means the triple $-\text{C}\equiv\text{C}-$ is between Carbon “2” and “3”

Pent-1,3-diyne means there are two triple bonds; one between carbon “1” and “2” and another between carbon “3” and “4”

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alkyne. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

6. Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of **triple** - $C \equiv C$ - bonds and **branches** attached to the alkyne.

7. Position isomers can be formed when the - $C \equiv C$ - triple bond is shifted between carbon atoms e.g.

But-2-yne means the double - $C = C$ - is between Carbon "2" and "3"

But-1-yne means the double - $C = C$ - is between Carbon "1" and "2"

Both But-1-yne and But-2-yne are position isomers of Butyne.

9. Like alkanes and alkynes, an alkyl group can be attached to the alkyne.

Chain/branch isomers are thus formed.

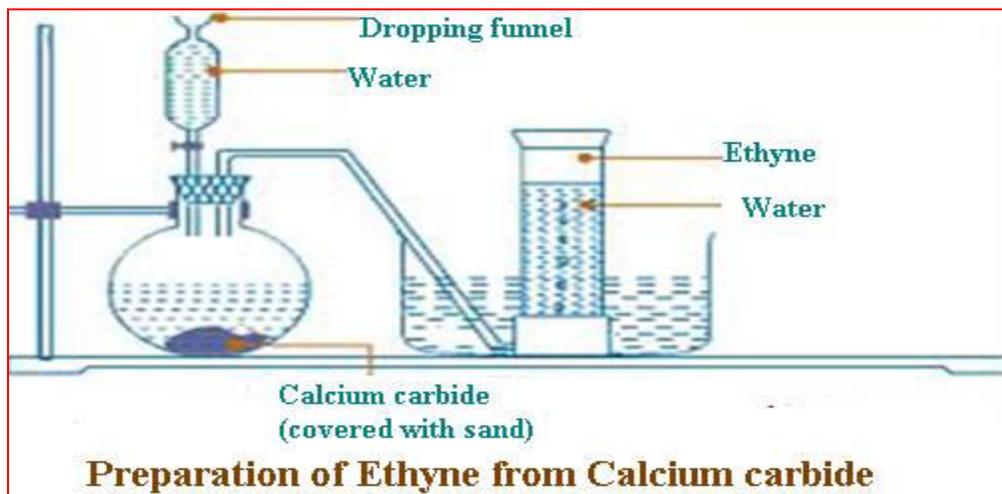
Butyne and 2-methyl propyne both have the same general formula but different branching chain.

(More on powerpoint)

(c) Preparation of Alkynes.

Ethyne is prepared from the reaction of water on calcium carbide. The reaction is highly exothermic and thus a layer of sand should be put above the calcium carbide to absorb excess heat to prevent the reaction flask from breaking.

Copper(II)sulphate(VI) is used to catalyze the reaction



Chemical equation



(d) Properties of alkynes

I. Physical properties

Like alkanes and alkenes, alkynes are colourless gases, solids and liquids that are not poisonous.

They are slightly soluble in water. The solubility in water decreases as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene. Ethyne has a pleasant taste when pure.

The melting and boiling point increase as the carbon chain increases.

This is because of the increase in van-der-Waals / intermolecular forces as the carbon chain increases. The first three straight chain alkynes (ethyne, propyne and but-1-yne) are gases at room temperature and pressure.

The density of straight chain alkynes increases with increasing carbon chain as the intermolecular forces increase reducing the volume occupied by a given mass of the alkyne.

Summary of physical properties of the first five alkynes

Alkyne	General formula	Melting point(°C)	Boiling point(°C)	State at room(298K) temperature and pressure atmosphere (101300Pa)
Ethyne	$\text{CH} \equiv \text{CH}$	-82	-84	gas
Propyne	$\text{CH}_3 \text{C} \equiv \text{CH}$	-103	-23	gas
Butyne	$\text{CH}_3\text{CH}_2 \text{C} \equiv \text{CH}$	-122	8	gas
Pent-1-yne	$\text{CH}_3(\text{CH}_2)_2 \text{C} \equiv \text{CH}$	-119	39	liquid
Hex-1-yne	$\text{CH}_3(\text{CH}_2)_3 \text{C} \equiv \text{CH}$	-132	71	liquid

II. Chemical properties

(a) Burning/combustion

Alkynes burn with a **yellow**/ luminous very **sooty**/ smoky flame in **excess** air to form carbon(IV) oxide and water.

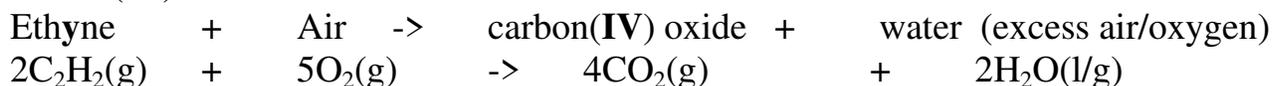


Alkenes burn with a **yellow**/ luminous very **sooty**/ smoky flame in **limited** air to form carbon(II) oxide/carbon and water.

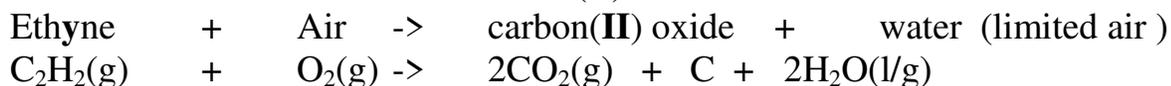
Alkyne + Air \rightarrow carbon(II) oxide /carbon + water (limited air)
 Burning of alkynes with a **yellow/ luminous sooty/ smoky** flame is a confirmatory test for the **presence** of the - $C \equiv C$ - triple bond because they have very **high C:H ratio**.

Examples of burning alkynes

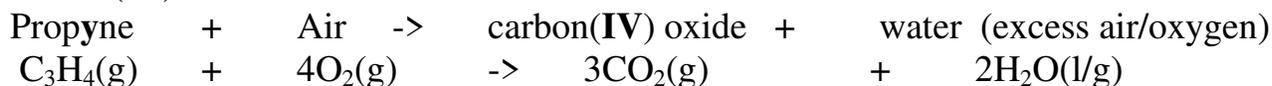
1.(a) Ethyne when ignited burns with a **yellow** very **sooty** flame in **excess** air to form carbon(IV) oxide and water.



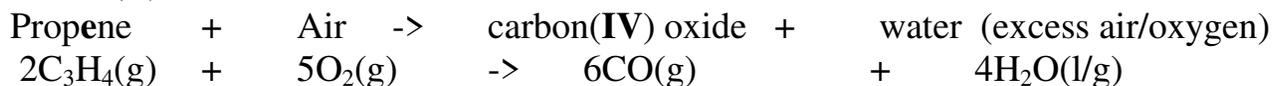
(b) Ethyne when ignited burns with a **yellow sooty** flame in **limited** air to form a mixture of unburnt carbon and carbon(II) oxide and water.



2.(a) Propyne when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.



(a) Propyne when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.



(b) Addition reactions

An addition reaction is one which an unsaturated compound reacts to form a saturated compound. Addition reactions of alkynes are also named from the reagent used to cause the addition/convert the triple - $C \equiv C$ - to single C- C bond.

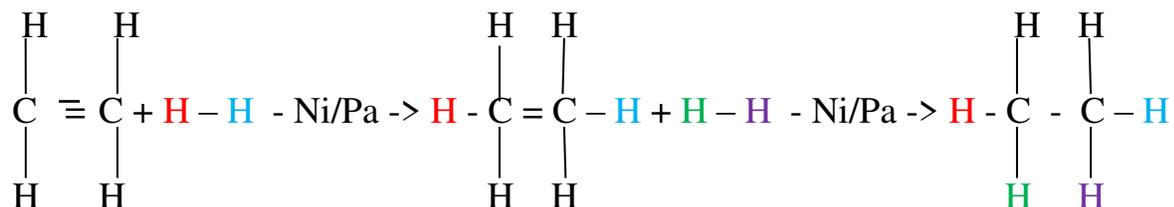
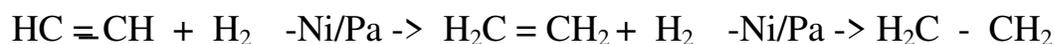
(i) Hydrogenation

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at 150°C temperatures react with alkynes to form alkenes then alkanes.

Examples

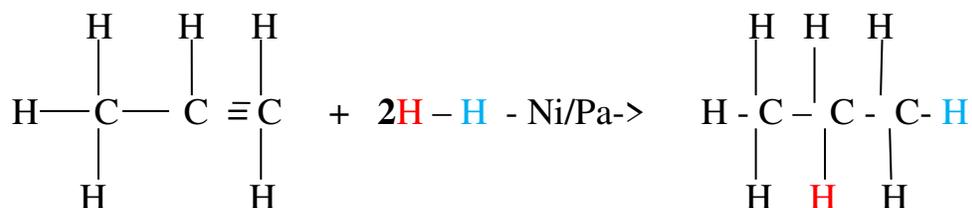
1. During hydrogenation, **two** hydrogen atom in the hydrogen molecule attach itself to one carbon and the other **two** hydrogen to the second carbon breaking the **triple** bond to **double** the **single**.

Chemical equation



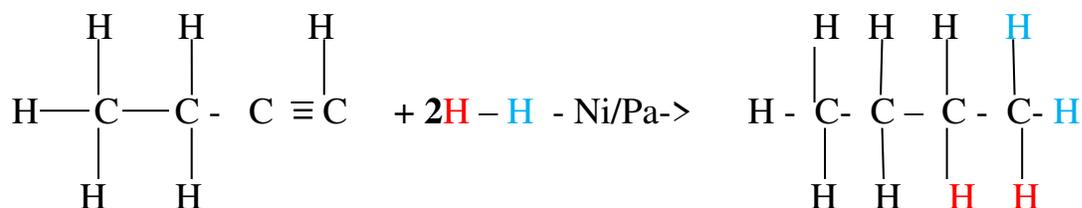
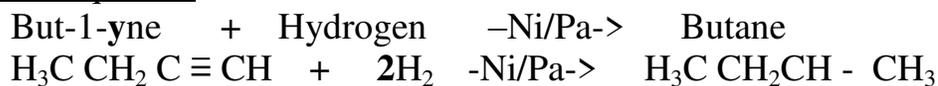
2. Propyne undergo hydrogenation to form Propane

Chemical equation



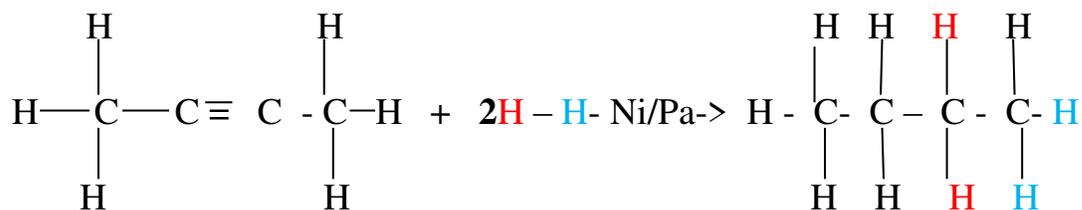
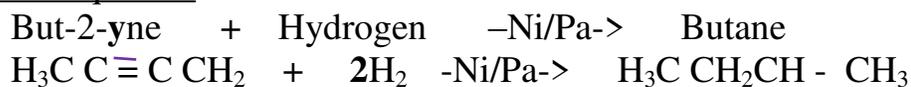
3(a) But-1-yne undergo hydrogenation to form Butane

Chemical equation



(b) But-2-yne undergo hydrogenation to form Butane

Chemical equation



(ii) Halogenation.

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alkyne to form an alkene then alkane.

The reaction of alkynes with halogens with alkynes is **faster** than with alkenes. The triple bond in the alkyne break and form a double then single bond.

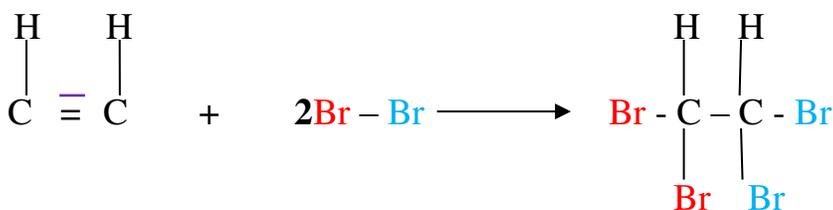
The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases.

Two bromine atoms bond at the 1st carbon in the triple bond while the other two goes to the 2nd carbon.

Examples

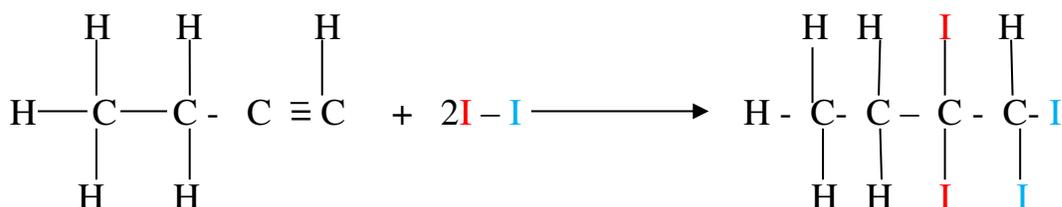
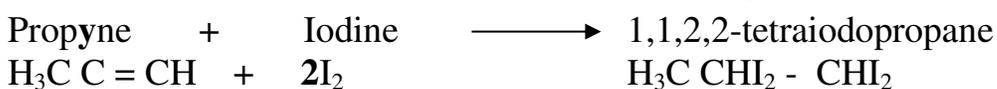
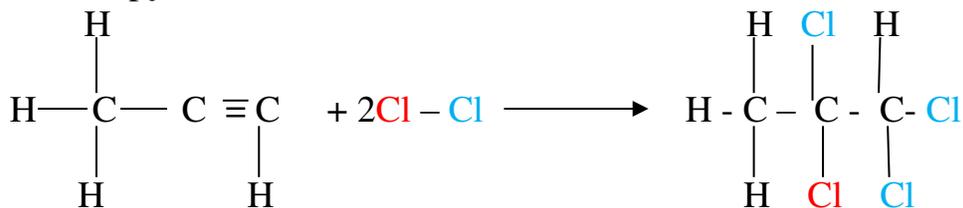
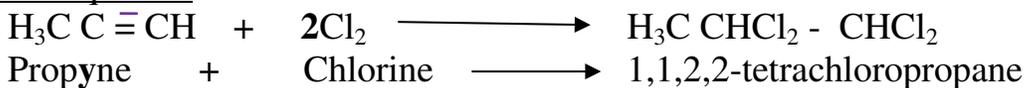
1.Ethyne reacts with brown bromine vapour to form 1,1,2,2-tetrabromoethane.

Chemical equation



2.Propyne reacts with chlorine to form 1,1,2,2-tetrachloropropane.

Chemical equation

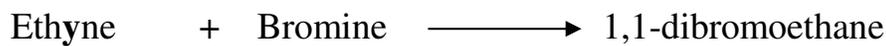
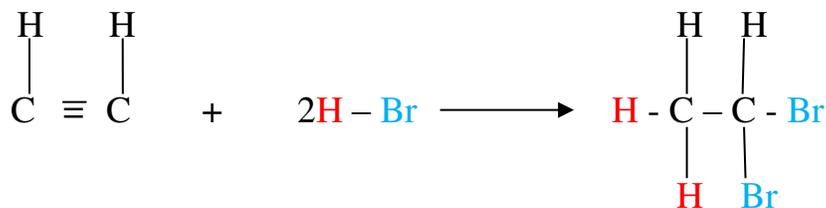


The main compound is one which the **hydrogen** atom bond at the carbon with **more hydrogen** .

Examples

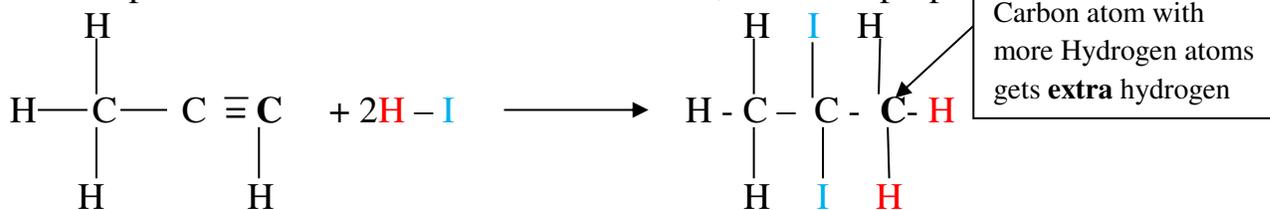
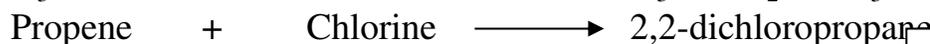
1. Ethyne reacts with hydrogen bromide to form bromoethane.

Chemical equation



2. Propyne reacts with hydrogen iodide to form 2,2-diiodopropane (as the main product)

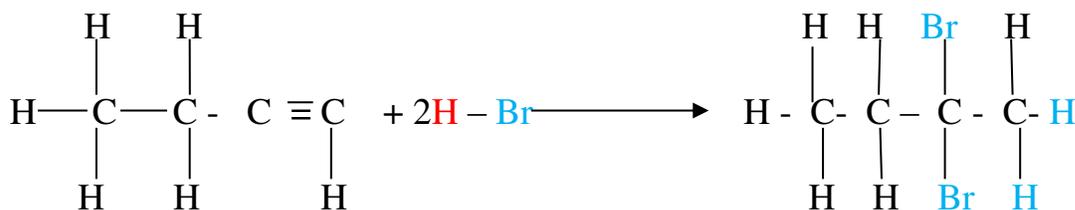
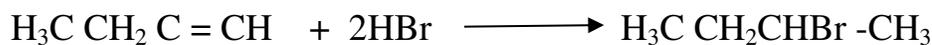
Chemical equation



Carbon atom with more Hydrogen atoms gets **extra** hydrogen

3. Both But-1-yne and But-2-yne reacts with hydrogen bromide to form 2,2-dibromobutane

Chemical equation



(A) INTRODUCTION.

Alkanols belong to a homologous series of organic compounds with a general formula $C_nH_{2n+1}OH$ and thus $-OH$ as the functional group. The 1st ten alkanols include

n	General / molecular formula	Structural formula	IUPAC name
1	CH_3OH	$\begin{array}{c} H - C - O - H \\ \\ H \end{array}$	Methanol
2	CH_3CH_2OH C_2H_5OH	$\begin{array}{c} H \quad H \\ \quad \\ H - C - C - O - H \\ \quad \\ H \quad H \end{array}$	Ethanol
3	$CH_3(CH_2)_2OH$ C_3H_7OH	$\begin{array}{c} H \quad H \quad H \\ \quad \quad \\ H - C - C - C - O - H \\ \quad \quad \\ H \quad H \quad H \end{array}$	Propanol
4	$CH_3(CH_2)_3OH$ C_4H_9OH	$\begin{array}{c} H \quad H \quad H \quad H \\ \quad \quad \quad \\ H - C - C - C - C - O - H \\ \quad \quad \quad \\ H \quad H \quad H \quad H \end{array}$	Butanol
5	$CH_3(CH_2)_4OH$ $C_5H_{11}OH$	$\begin{array}{c} H \quad H \quad H \quad H \quad H \\ \quad \quad \quad \quad \\ H - C - C - C - C - C - O - H \\ \quad \quad \quad \quad \\ H \quad H \quad H \quad H \quad H \end{array}$	Pentanol
6	$CH_3(CH_2)_5OH$ $C_6H_{13}OH$	$\begin{array}{c} H \quad H \quad H \quad H \quad H \quad H \\ \quad \quad \quad \quad \quad \\ H - C - C - C - C - C - C - O - H \\ \quad \quad \quad \quad \quad \\ H \quad H \quad H \quad H \quad H \quad H \end{array}$	Hexanol

		H H H H H H	
7	CH ₃ (CH ₂) ₆ OH C ₇ H ₁₅ OH	$ \begin{array}{ccccccc} \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ & & & & & & \\ \text{H}-\text{C} & -\text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}-\text{O}-\text{H} \\ & & & & & & \\ \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \end{array} $	Heptanol
8	CH ₃ (CH ₂) ₇ OH C ₈ H ₁₇ OH	$ \begin{array}{cccccccc} \text{H} & \text{H} \\ & & & & & & & \\ \text{H}-\text{C} & -\text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}-\text{O}-\text{H} \\ & & & & & & & \\ \text{H} & \text{H} \end{array} $	Octanol
9	CH ₃ (CH ₂) ₈ OH C ₉ H ₁₉ OH	$ \begin{array}{ccccccccc} \text{H} & \text{H} \\ & & & & & & & & \\ \text{H}-\text{C} & -\text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}- & \text{C}-\text{O}-\text{H} \\ & & & & & & & & \\ \text{H} & \text{H} \end{array} $	Nonanol
10	CH ₃ (CH ₂) ₉ OH C ₁₀ H ₂₁ OH	$ \begin{array}{cccccccccc} \text{H} & \text{H} \\ & & & & & & & & & \\ \text{H}-\text{C} & -\text{C}- & \text{C}- & \text{C}-\text{O}-\text{H} \\ & & & & & & & & & \\ \text{H} & \text{H} \end{array} $	Decanol

Alkanols like Hydrocarbons(alkanes/alkenes/alkynes) form a homologous series where:

- (i)general name is derived from the alkane name then ending with “-ol”
- (ii)the members have –OH as the functional group
- (iii)they have the same general formula represented by R-OH where R is an alkyl group.
- (iv) each member differ by –CH₂ group from the next/previous.
- (v)they show a similar and gradual change in their physical properties e.g. boiling and melting points.
- (vi)they show similar and gradual change in their chemical properties.

B. ISOMERS OF ALKANOLS.

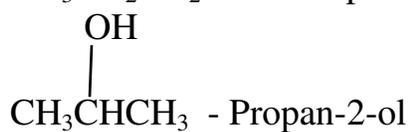
Alkanols exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines:

- (i) Like alkanes, identify the **longest** carbon chain to be the parent name.
- (ii) Identify the position of the **-OH** functional group to give it the **smallest /lowest** position.
- (iii) Identify the type and position of the **side** branches.

Practice examples of isomers of alkanols

(i) Isomers of propanol C_3H_7OH

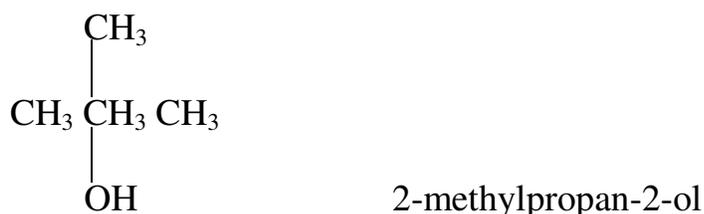
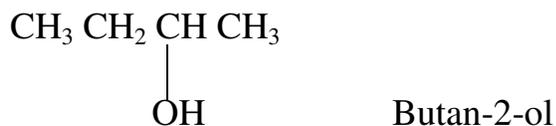
$CH_3CH_2CH_2OH$ - Propan-1-ol



Propan-2-ol and Propan-1-ol are position isomers because only the position of the –OH functional group changes.

(ii) Isomers of Butanol C_4H_9OH

$CH_3CH_2CH_2CH_2OH$ Butan-1-ol

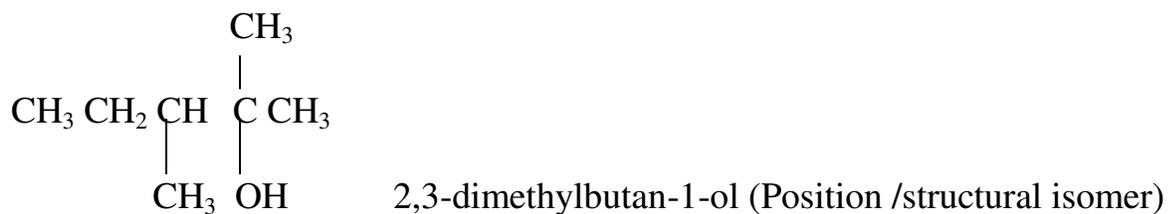
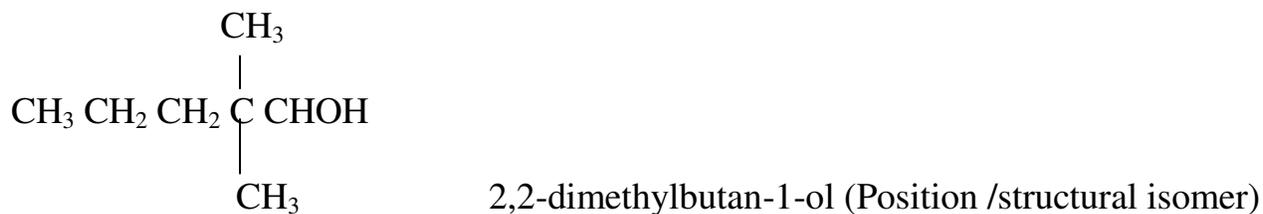
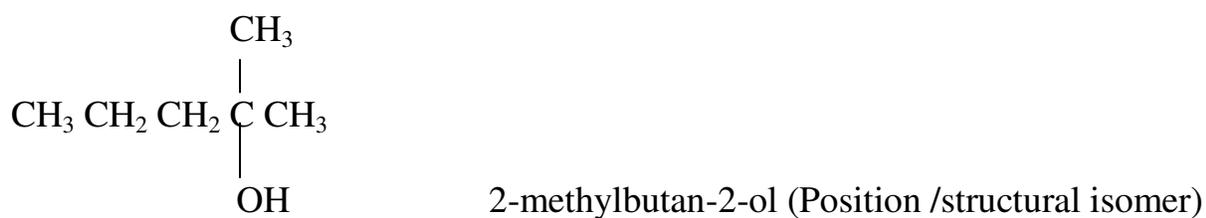
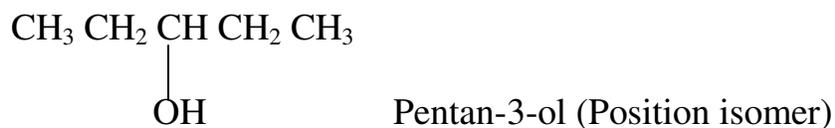
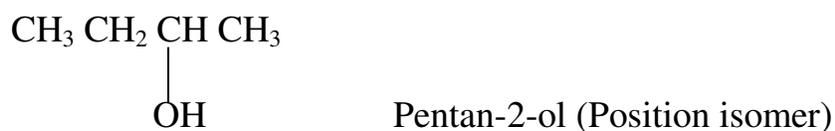


Butan-2-ol and Butan-1-ol are position isomers because only the position of the –OH functional group changes.

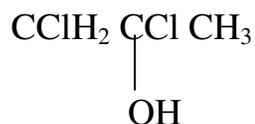
2-methylpropan-2-ol is both a structural and position isomers because both the position of the functional group and the arrangement of the atoms in the molecule changes.

(iii) Isomers of Pentanol $C_5H_{11}OH$

$CH_3CH_2CH_2CH_2CH_2OH$ Pentan-1-ol (Position isomer)



(iv) 1,2-dichloropropan-2-ol



(v) 1,2-dichloropropan-1-ol



To increase the concentration, fractional distillation is done to produce spirits (e.g. Brandy=40% ethanol).

Methanol is much more poisonous /toxic than ethanol.

Taken large quantity in small quantity it causes instant blindness and liver, killing the consumer victim within hours.

School laboratory preparation of ethanol from fermentation of glucose

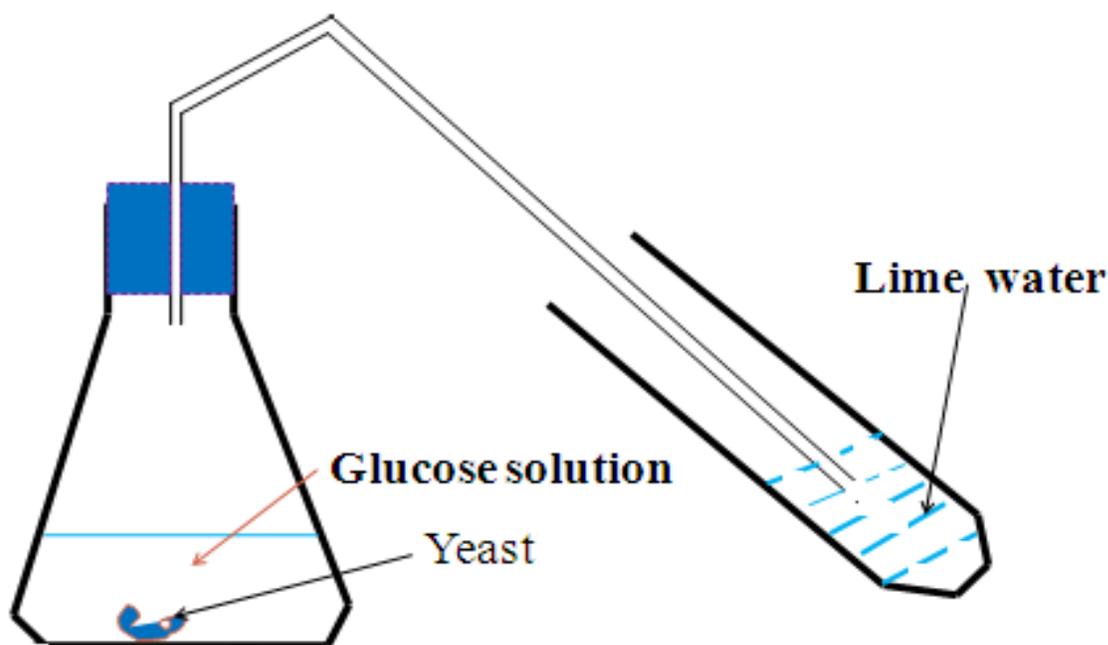
Measure 100cm³ of pure water into a conical flask.

Add about five spatula end full of glucose.

Stir the mixture to dissolve.

Add about one spatula end full of yeast.

Set up the apparatus as below.



Preserve the mixture for about **three** days.

D. PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOLS

Use the prepared sample above for the following experiments that shows the characteristic properties of alkanols

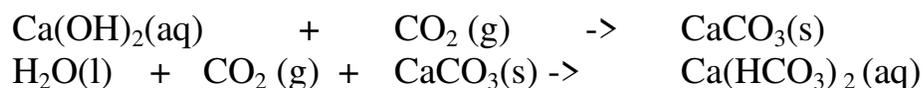
(a) Role of yeast

Yeast is a single cell fungus which contains the enzyme maltase and zymase that catalyse the fermentation process.

(b) Observations in lime water.

A white precipitate is formed that dissolve to a colourless solution later. Lime water/Calcium hydroxide reacts with carbon(IV)Oxide produced during the fermentation to form insoluble calcium carbonate and water.

More carbon (IV)Oxide produced during fermentation react with the insoluble calcium carbonate and water to form soluble calcium hydrogen carbonate.



(c) Effects on litmus paper

Experiment

Take the prepared sample and test with both blue and red litmus papers.

Repeat the same with pure ethanol and methylated spirit.

Sample Observation table

Substance/alkanol	Effect on litmus paper
Prepared sample	Blue litmus paper remain blue Red litmus paper remain red
Absolute ethanol	Blue litmus paper remain blue Red litmus paper remain red
Methylated spirit	Blue litmus paper remain blue Red litmus paper remain red

Explanation

Alkanols are neutral compounds/solution that have characteristic sweet smell and taste.

They have no effect on both blue and red litmus papers.

(d) Solubility in water.

Experiment

Place about 5cm³ of prepared sample into a clean test tube Add equal amount of distilled water.

Repeat the same with pure ethanol and methylated spirit.

Observation

No layers formed between the two liquids.

Explanation

Ethanol is **miscible** in water. Both ethanol and water are polar compounds .

The solubility of alkanols decrease with increase in the alkyl chain/molecular mass. The alkyl group is insoluble in water while –OH functional group is soluble in water.

As the molecular chain becomes **longer** ,the effect of the **alkyl** group **increases** as the effect of the functional group **decreases**.

e)Melting/boiling point.

Experiment

Place pure ethanol in a long boiling tube .Determine its boiling point.

Observation

Pure ethanol has a boiling point of 78°C at sea level/one atmosphere pressure.

Explanation

The melting and boiling point of alkanols increase with increase in molecular chain/mass .

This is because the intermolecular/van-der-waals forces of attraction between the molecules increase.

More heat energy is thus required to weaken the longer chain during melting and break during boiling.

f)Density

Density of alkanols increase with increase in the intermolecular/van-der-waals forces of attraction between the molecule, making it very close to each other.

This reduces the volume occupied by the molecule and thus increase the their mass per unit volume (density).

Summary table showing the trend in physical properties of alkanols

Alkanol	Melting point (°C)	Boiling point (°C)	Density gcm ⁻³	Solubility in water
Methanol	-98	65	0.791	soluble
Ethanol	-117	78	0.789	soluble
Propanol	-103	97	0.803	soluble
Butanol	-89	117	0.810	Slightly soluble
Pentanol	-78	138	0.814	Slightly soluble
Hexanol	-52	157	0.815	Slightly soluble
Heptanol	-34	176	0.822	Slightly soluble
Octanol	-15	195	0.824	Slightly soluble
Nonanol	-7	212	0.827	Slightly soluble
Decanol	6	228	0.827	Slightly soluble

g) Burning

Experiment

Place the prepared sample in a watch glass. Ignite. Repeat with pure ethanol and methylated spirit.

Observation/Explanation

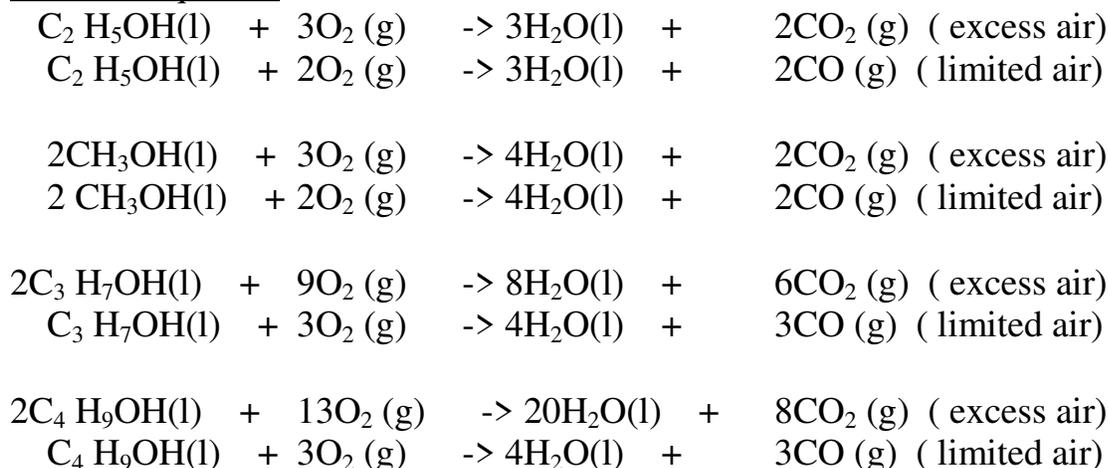
Fermentation produce ethanol with a lot of water(about a ratio of 1:3)which prevent the alcohol from igniting.

Pure ethanol and methylated spirit easily catch fire / highly flammable.

They burn with an almost colourless non-sooty/non-smoky **blue** flame to form **carbon(IV) oxide** (in excess air/oxygen)or **carbon(II) oxide** (limited air) and **water**.

Ethanol is thus a **saturated** compound like alkanes.

Chemica equation



Due to its flammability, ethanol is used;

- (i) as a fuel in spirit lamps
- (ii) as gasohol when blended with gasoline

(h)Formation of alkoxides

Experiment

Cut a very small piece of sodium. Put it in a beaker containing about 20cm³ of the prepared sample in a beaker.

Test the products with litmus papers. Repeat with absolute ethanol and methylated spirit.

Sample observations

Substance/alkanol	Effect of adding sodium
Fermentation prepared sample	(i)effervescence/fizzing/bubbles

	(ii) colourless gas produced that extinguish burning splint with explosion/ "Pop" sound (iii) colourless solution formed (iv) blue litmus papers remain blue (v) red litmus papers turn blue
Pure/absolute ethanol/methylated spirit	(i) slow effervescence/fizzing/bubbles (ii) colourless gas slowly produced that extinguish burning splint with explosion/ "Pop" sound (iii) colourless solution formed (iv) blue litmus papers remain blue (v) red litmus papers turn blue

Explanations

Sodium/potassium reacts slowly with alkanols to form basic solution called **alkoxides** and producing **hydrogen** gas.

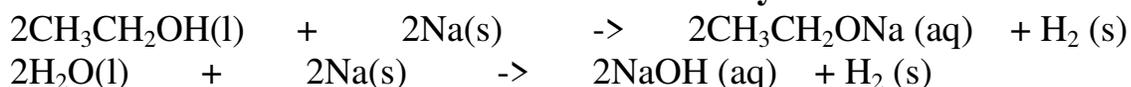
If the alkanol has some water the metals react faster with the water to form **soluble hydroxides/alkalis** i.e.



Examples

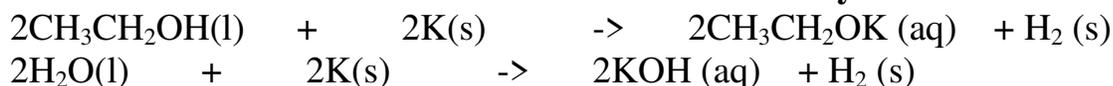
1. Sodium metal reacts with ethanol to form sodium **ethoxide**

Sodium metal reacts with water to form sodium **Hydroxide**



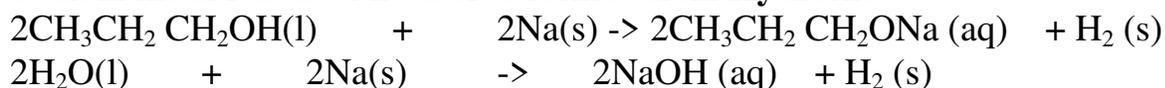
2. Potassium metal reacts with ethanol to form Potassium **ethoxide**

Potassium metal reacts with water to form Potassium **Hydroxide**



3. Sodium metal reacts with propanol to form sodium **propoxide**

Sodium metal reacts with water to form sodium **Hydroxide**



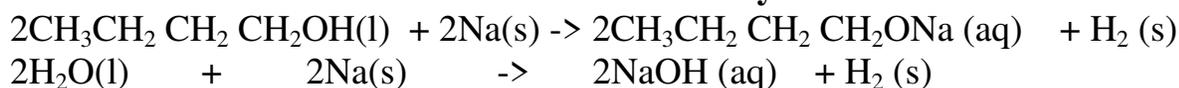
4. Potassium metal reacts with propanol to form Potassium **propoxide**

Potassium metal reacts with water to form Potassium **Hydroxide**



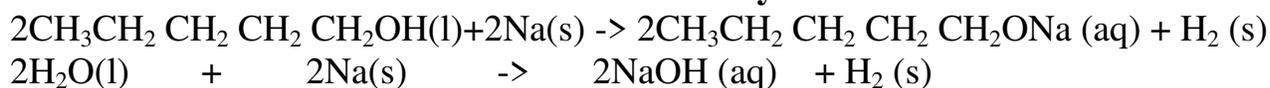
5. Sodium metal reacts with butanol to form sodium **butoxide**

Sodium metal reacts with water to form sodium **Hydroxide**



6. Sodium metal reacts with pentanol to form sodium **pentoxide**

Sodium metal reacts with water to form sodium **Hydroxide**



(i) Formation of Esters/Esterification

Experiment

Place 2cm³ of ethanol in a boiling tube.

Add equal amount of ethanoic acid. To the mixture add carefully 2 drops of concentrated sulphuric(VI) acid.

Warm/Heat gently.

Pour the mixture into a beaker containing about 50cm³ of cold water.

Smell the products.

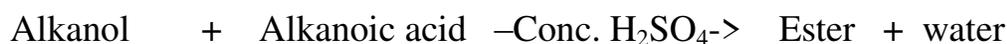
Repeat with methanol

Sample observations

Substance/alkanol	Effect on adding equal amount of ethanol/concentrated sulphuric(VI) acid
Absolute ethanol	Sweet fruity smell
Methanol	Sweet fruity smell

Explanation

Alkanols react with alkanolic acids to form a group of homologous series of sweet smelling compounds called esters and water. This reaction is catalyzed by concentrated sulphuric(VI) acid in the laboratory.



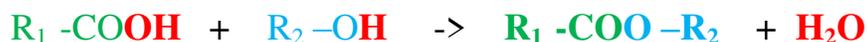
Naturally esterification is catalyzed by sunlight. Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanolic acids that

create a variety of known natural (mostly in fruits) and synthetic (mostly in juices) esters .

Esters derive their names from the alkanol first then alkanolic acids. The alkanol “becomes” an **alkyl** group and the alkanolic acid “becomes” **alkanoate** hence **alkylalkanoate**. e.g.

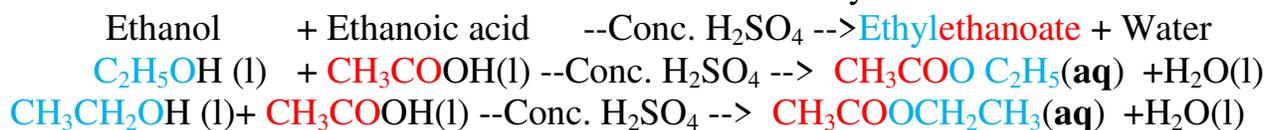
Ethanol	+	Ethanoic acid	->	Ethylethanoate	+	Water
Ethanol	+	Propanoic acid	->	Ethylpropanoate	+	Water
Ethanol	+	Methanoic acid	->	Ethylmethanoate	+	Water
Ethanol	+	butanoic acid	->	Ethylbutanoate	+	Water
Propanol	+	Ethanoic acid	->	Propylethanoate	+	Water
Methanol	+	Ethanoic acid	->	Methylethanoate	+	Water
Methanol	+	Decanoic acid	->	Methyldecanoate	+	Water
Decanol	+	Methanoic acid	->	Decylmethanoate	+	Water

During the formation of the ester, the “O” joining the alkanol and alkanolic acid comes from the alkanol.

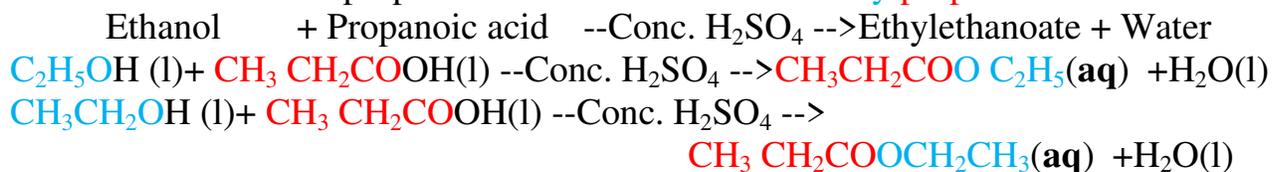


e.g.

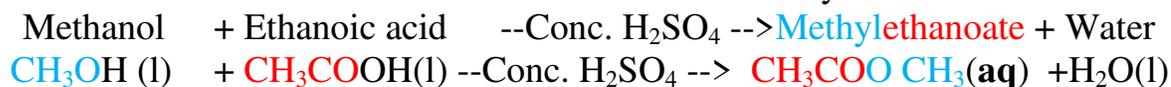
1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.



2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.



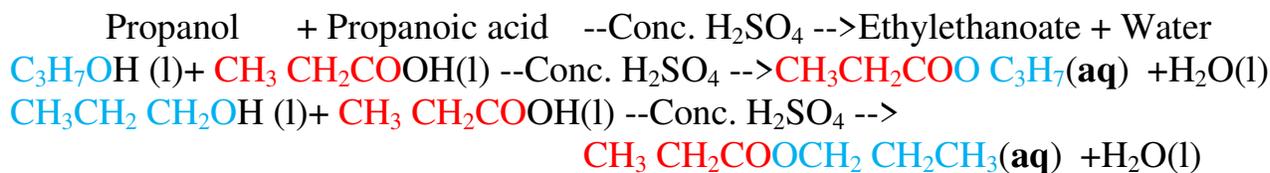
3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.



4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.



5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.



(j)Oxidation

Experiment

Place 5cm³ of absolute ethanol in a test tube. Add three drops of acidified potassium manganate(VII). Shake thoroughly for one minute/warm. Test the solution mixture using pH paper. Repeat by adding acidified potassium dichromate(VII).

Sample observation table

Substance/alkanol	Adding acidified KMnO ₄ /K ₂ Cr ₂ O ₇	pH of resulting solution/mixture	Nature of resulting solution/mixture
Pure ethanol	(i) Purple colour of KMnO ₄ decolorized	pH = 4/5/6	Weakly acidic
	(ii) Orange colour of K ₂ Cr ₂ O ₇ turns green.	pH = 4/5/6	Weakly acidic

Explanation

Both acidified KMnO₄ and K₂Cr₂O₇ are oxidizing agents (add oxygen to other compounds). They oxidize alkanols to a group of homologous series called alkanals then further oxidize them to alkanoic acids. The oxidizing agents are themselves reduced hence changing their colour:

(i) Purple KMnO₄ is reduced to colourless Mn²⁺

(ii) Orange K₂Cr₂O₇ is reduced to green Cr³⁺

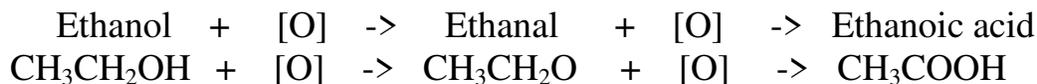
The pH of alkanolic acids show they have few H⁺ because they are weak acids i.e



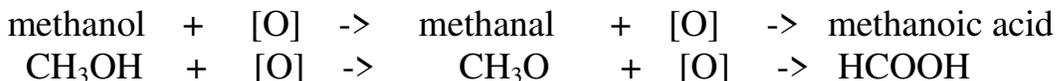
NB The [O] comes from the oxidizing agents acidified KMnO₄ or K₂Cr₂O₇

Examples

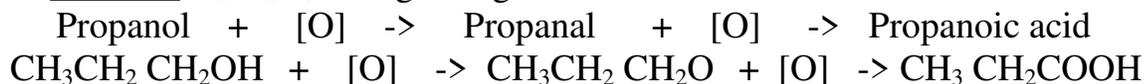
1. When ethanol is warmed with three drops of acidified KMnO₄ there is decolorization of KMnO₄



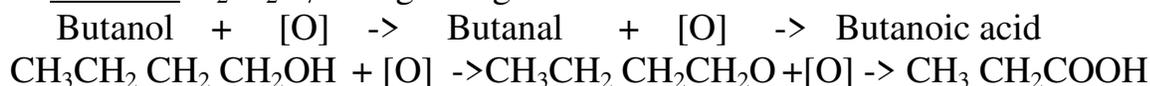
2. When methanol is warmed with three drops of acidified K₂Cr₂O₇, the orange colour of acidified K₂Cr₂O₇ changes to green.



3. When propanol is warmed with three drops of acidified $\text{K}_2\text{Cr}_2\text{O}_7$, the orange colour of acidified $\text{K}_2\text{Cr}_2\text{O}_7$ changes to green.



4. When butanol is warmed with three drops of acidified $\text{K}_2\text{Cr}_2\text{O}_7$, the orange colour of acidified $\text{K}_2\text{Cr}_2\text{O}_7$ changes to green.



Air slowly oxidizes ethanol to dilute ethanoic acid commonly called **vinegar**. If beer is not tightly corked, a lot of carbon(IV)oxide escapes and there is slow oxidation of the beer making it “flat”.

(k)Hydrolysis /Hydration and Dehydration

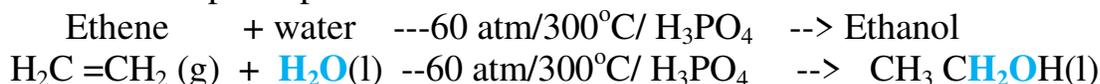
I. Hydrolysis/Hydration is the reaction of a compound/substance with water.

Alkenes react with water vapour/steam at high temperatures and high pressures in presence of phosphoric acid catalyst to form alkanols.i.e.

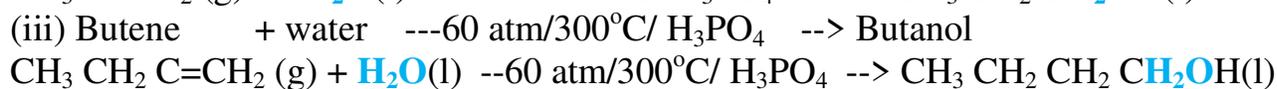


Examples

(i) Ethene is mixed with steam over a phosphoric acid catalyst at 300°C temperature and 60 atmosphere pressure to form ethanol

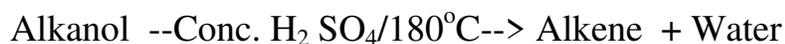


This is the main method of producing large quantities of ethanol instead of fermentation



II. Dehydration is the process which concentrated sulphuric(VI)acid (**dehydrating agent**) removes water from a compound/substances.

Concentrated sulphuric(VI)acid dehydrates alkanols to the corresponding alkenes at about 180°C . i.e

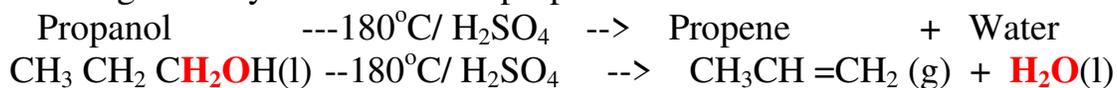


Examples

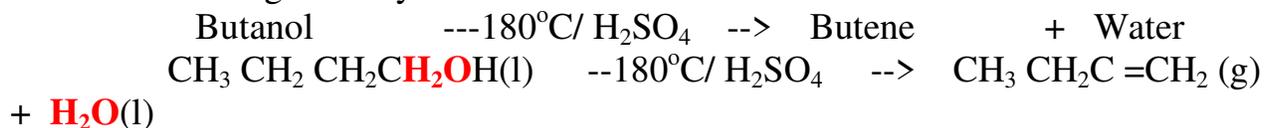
1. At 180°C and in presence of Concentrated sulphuric(VI)acid, ethanol undergoes dehydration to form ethene.



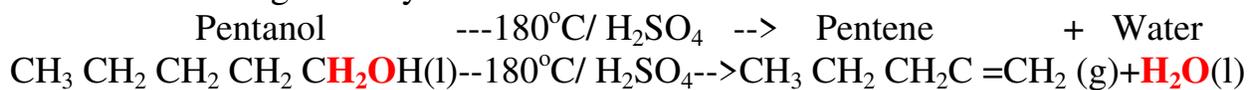
2. Propanol undergoes dehydration to form propene.



3. Butanol undergoes dehydration to form Butene.



3. Pentanol undergoes dehydration to form Pentene.



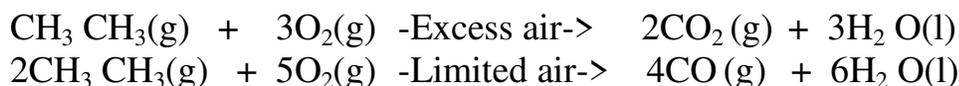
(I) Similarities of alkanols with Hydrocarbons

I. Similarity with alkanes

Both alkanols and alkanes burn with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water.

This shows they are saturated with high C:H ratio. e.g.

Both ethanol and ethane ignite and burns in air with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water.



II. Similarity with alkenes/alkynes

Both alkanols(R-OH) and alkenes/alkynes(with = C = C = double and - C ≡ C- triple) bond:

(i) decolorize acidified KMnO₄

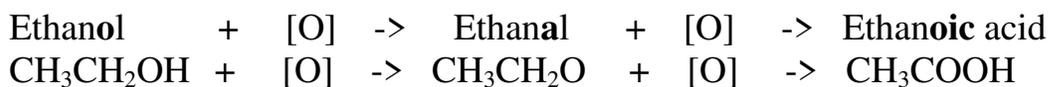
(ii) turns Orange acidified K₂Cr₂O₇ to green.

Alkanols(R-OH) are oxidized to alkanals(R-O) and then alkanonic acids(R-COOH).

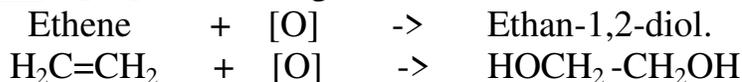
Alkenes are oxidized to alkanols with duo/double functional groups.

Examples

1. When ethanol is warmed with three drops of acidified K₂Cr₂O₇ the orange of acidified K₂Cr₂O₇ turns to green. Ethanol is oxidized to ethanol and then to ethanoic acid.



2. When ethene is bubbled in a test tube containing acidified $\text{K}_2\text{Cr}_2\text{O}_7$, the orange of acidified $\text{K}_2\text{Cr}_2\text{O}_7$ turns to green. Ethene is oxidized to ethan-1,2-diol.



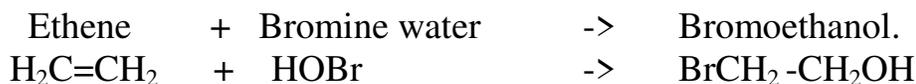
III. Differences with alkenes/alkynes

Alkanols do not decolorize bromine and chlorine water.

Alkenes decolorize bromine and chlorine water to form halogenoalkanes

Example

When ethene is bubbled in a test tube containing bromine water, the bromine water is decolorized. Ethene is oxidized to bromoethanol.



IV. Differences in melting and boiling point with Hydrocarbons

Alkanols have higher melting point than the corresponding hydrocarbon (alkane/alkene/alkyne)

This is because most alkanols exist as **dimer**. A dimer is a molecule made up of two other molecules joined usually by van-der-waals forces/hydrogen bond or dative bonding.

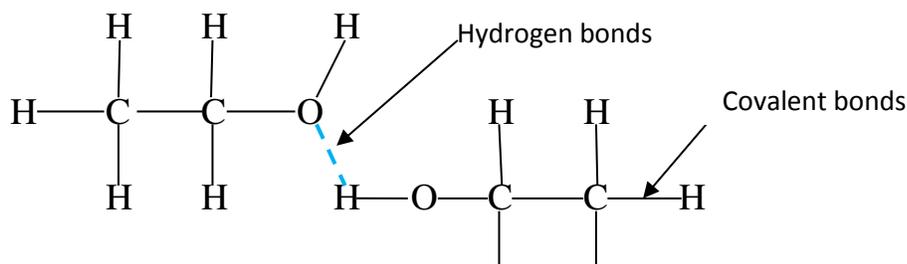
Two alkanol molecules form a dimer joined by hydrogen bonding.

Example

In Ethanol the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

This creates a partial negative charge (δ^-) on oxygen and partial positive charge (δ^+) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimer**.



H H

Dimerization of alkanols means more energy is needed to break/weaken the Hydrogen bonds before breaking/weakening the intermolecular forces joining the molecules of all organic compounds during boiling/melting.

E. USES OF SOME ALKANOLS

(a) Methanol is used as industrial alcohol and making methylated spirit

(b) Ethanol is used:

1. as alcohol in alcoholic drinks e.g Beer, wines and spirits.
2. as antiseptic to wash wounds
3. in manufacture of vanishes, ink, glue and paint because it is volatile and thus easily evaporate
4. as a fuel when blended with petrol to make gasohol.

B. ALKANOIC ACIDS (Carboxylic acids)

(A) INTRODUCTION.

Alkanoic acids belong to a homologous series of organic compounds with a general formula $C_nH_{2n+1}COOH$ and thus $-COOH$ as the functional group. The 1st ten alkanoic acids include:

Alkanoic acids like alkanols /alkanes/alkenes/alkynes form a homologous series where:

(i) the general name of an alkanoic acid is derived from the alkane name then ending with “-oic” acid as the table above shows.

(ii) the members have $R\text{-COOH}/R\text{-}\overset{\text{O}}{\parallel}{\text{C}}\text{-O-H}$ as the functional group.

n	General /molecular formular	Structural formula	IUPAC name
0	HCOOH	$\begin{array}{c} \text{H} - \text{C} - \text{O} - \text{H} \\ \parallel \\ \text{O} \end{array}$	Methanoic acid
1	CH ₃ COOH	$\begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \parallel \\ \text{H} \quad \text{O} \end{array}$	Ethanoic acid
2	CH ₃ CH ₂ COOH C ₂ H ₅ COOH	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{O} \end{array}$	Propanoic acid
3	CH ₃ CH ₂ CH ₂ COOH C ₃ H ₇ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \end{array}$	Butanoic acid
4	CH ₃ CH ₂ CH ₂ CH ₂ COOH C ₄ H ₉ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \end{array}$	Pentanoic acid
5	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ COOH C ₅ H ₁₁ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \end{array}$	Hexanoic acid
6	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ COOH C ₆ H ₁₃ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{O} - \text{H} \\ \quad \quad \quad \quad \quad \quad \parallel \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \end{array}$	Pentanoic acid

O

(iii) they have the same general formula represented by R-COOH where R is an alkyl group.

(iv) each member differs by $-\text{CH}_2-$ group from the next/previous.

(v) they show a similar and gradual change in their physical properties e.g. boiling and melting point.

(vi) they show similar and gradual change in their chemical properties.

(vii) since they are acids they show similar properties with mineral acids.

(B) ISOMERS OF ALKANOIC ACIDS.

Alkanoic acids exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines

(i) Like alkanes. identify the longest carbon chain to be the parent name.

(ii) Identify the position of the $\begin{array}{c} \text{-C-O-H} \\ || \\ \text{O} \end{array}$ functional group to give it the smallest

/lowest position.

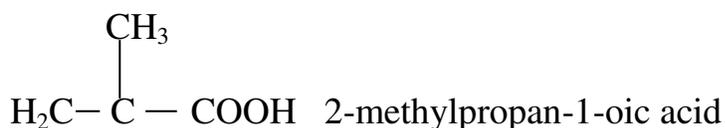
(iii) Identify the type and position of the side group branches.

Practice examples on isomers of alkanolic acids

1. Isomers of butanoic acid $\text{C}_3\text{H}_7\text{COOH}$

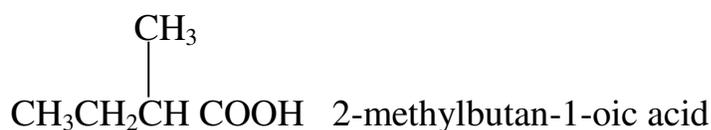


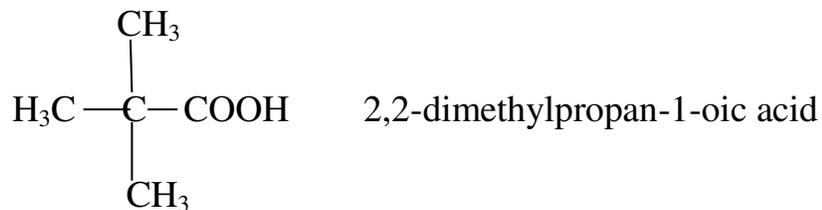
Butan-1-oic acid



2-methylpropan-1-oic acid and Butan-1-oic acid are structural isomers because the position of the functional group does not change but the arrangement of the atoms in the molecule does.

2. Isomers of pentanoic acid $\text{C}_4\text{H}_9\text{COOH}$

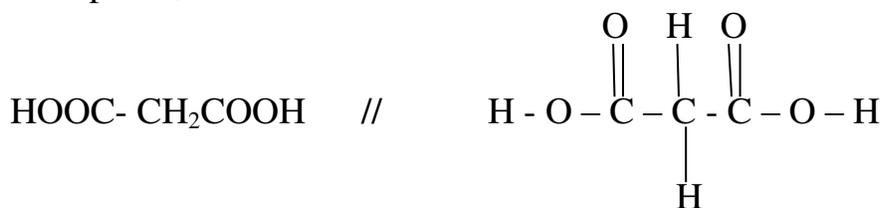




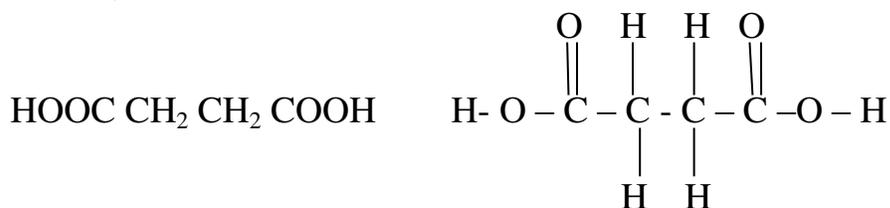
3. Ethan-1,2-dioic acid



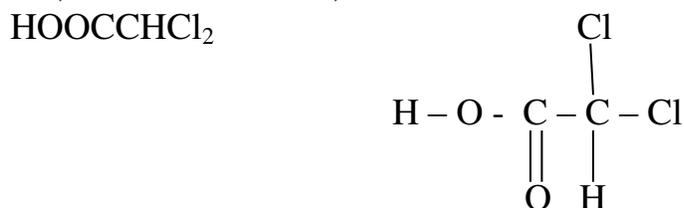
4. Propan-1,3-dioic acid



5. Butan-1,4-dioic acid



6. 2,2-dichloroethan-1,2-dioic acid



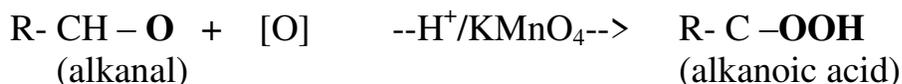
(C) LABORATORY AND INDUSTRIAL PREPARATION OF ALKANOIC ACIDS.

In a school laboratory, alkanolic acids can be prepared by adding an oxidizing agent (H^+/KMnO_4 or $\text{H}^+/\text{K}_2\text{Cr}_2\text{O}_7$) to the corresponding alkanol then warming.

The oxidation converts the alkanol first to an alkanal then the alkanonic acid.

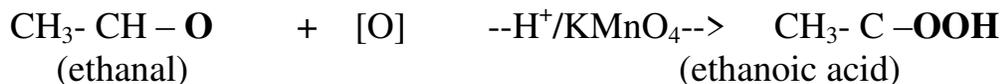
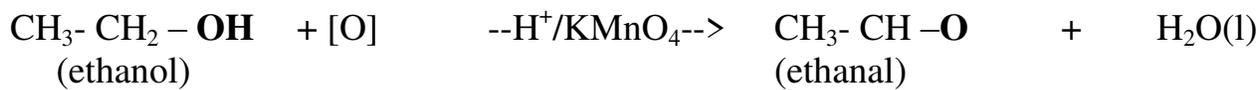
NB Acidified KMnO_4 is a stronger oxidizing agent than acidified $\text{K}_2\text{Cr}_2\text{O}_7$

General equation:

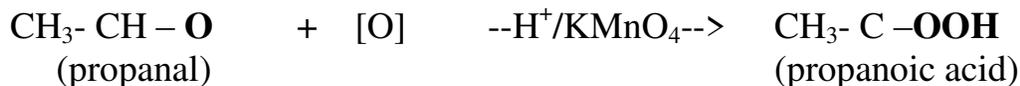
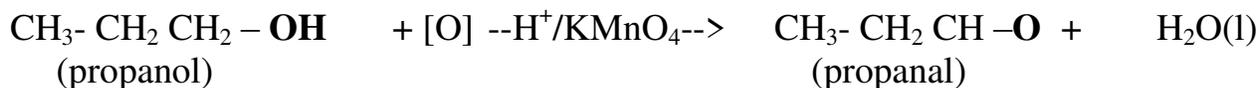


Examples

1. Ethanol on warming in acidified KMnO_4 is oxidized to ethanal then ethanoic acid.



2 Propanol on warming in acidified KMnO_4 is oxidized to propanal then propanoic acid



Industrially, large scale manufacture of alkanonic acid like ethanoic acid is obtained from:

(a) Alkenes reacting with steam at high temperatures and pressure in presence of phosphoric(V) acid catalyst and undergo hydrolysis to form alkanols. i.e.

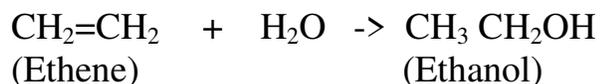
Alkenes + Steam/water $\xrightarrow{-\text{H}_2\text{PO}_4 \text{ Catalyst}}$ Alkanol

The alkanol is then oxidized by air at 5 atmosphere pressure with Manganese (II) sulphate(VI) catalyst to form the alkanonic acid.

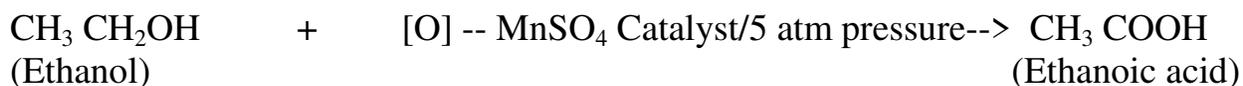
Alkanol + Air $\xrightarrow{-\text{MnSO}_4 \text{ Catalyst/5 atm pressure}}$ Alkanonic acid

Example

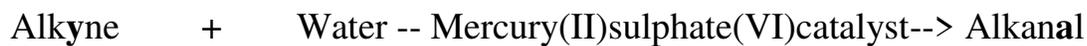
Ethene is mixed with steam over a phosphoric(V) acid catalyst, 300°C temperature and 60 atmosphere pressure to form ethanol.



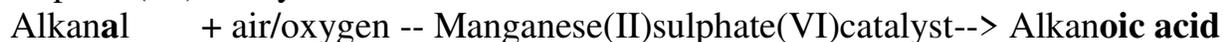
This is the industrial large scale method of manufacturing ethanol. Ethanol is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.



(b) Alkynes react with liquid water at high temperatures and pressure in presence of Mercury(II)sulphate(VI)catalyst and 30% concentrated sulphuric(VI)acid to form alkanals.

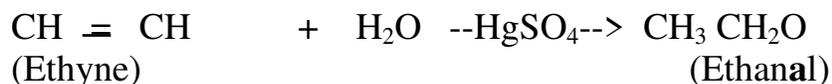


The alkanal is then oxidized by air at 5 atmosphere pressure with Manganese (II) sulphate(VI) catalyst to form the alkan \mathbf{oic} acid.



Example

Ethyne react with liquid water at high temperature and pressure with Mercury (II) sulphate (VI)catalyst and 30% concentrated sulphuric(VI)acid to form ethanal.



This is another industrial large scale method of manufacturing ethanol from large quantities of ethyne found in natural gas.

Ethanal is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.



(D) PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOIC ACIDS.

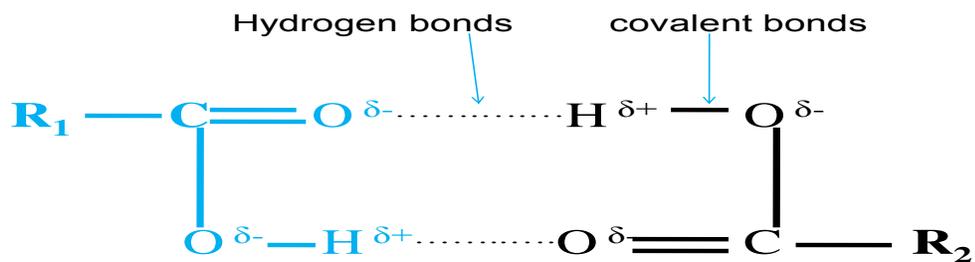
I.Physical properties of alkan \mathbf{oic} acids

The table below shows some physical properties of alkanolic acids

Alkanol	Melting point(°C)	Boiling point(°C)	Density(gcm ⁻³)	Solubility in water
Methanoic acid	18.4	101	1.22	soluble
Ethanoic acid	16.6	118	1.05	soluble
Propanoic acid	-2.8	141	0.992	soluble
Butanoic acid	-8.0	164	0.964	soluble
Pentanoic acid	-9.0	187	0.939	Slightly soluble
Hexanoic acid	-11	205	0.927	Slightly soluble
Heptanoic acid	-3	223	0.920	Slightly soluble
Octanoic acid	11	239	0.910	Slightly soluble
Nonanoic acid	16	253	0.907	Slightly soluble
Decanoic acid	31	269	0.905	Slightly soluble

From the table note the following:

- (i) Melting and boiling point decrease as the carbon chain increases due to increase in intermolecular forces of attraction between the molecules requiring more energy to separate the molecules.
- (ii) The density decreases as the carbon chain increases as the intermolecular forces of attraction increases between the molecules making the molecule very close reducing their volume in unit mass.
- (iii) Solubility decreases as the carbon chain increases as the soluble –COOH end is shielded by increasing insoluble alkyl/hydrocarbon chain.
- (iv) Like alkanols ,alkanoic acids exist as dimmers due to the hydrogen bonds within the molecule. i.e..

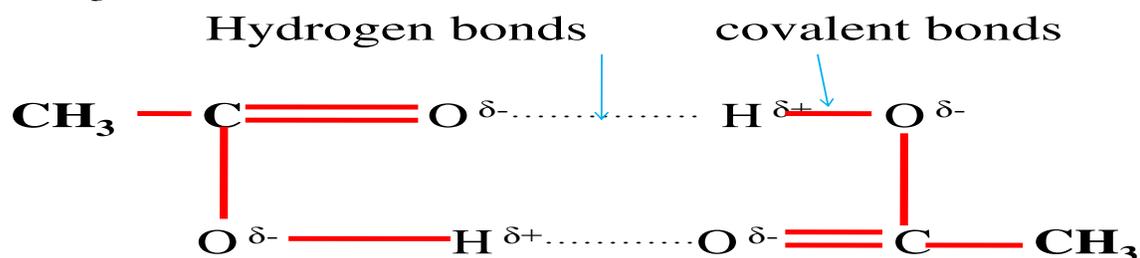


R_1 and R_2 are **extensions** of the molecule.
For ethanoic acid the extension is made up of **CH₃-** to make the structure;

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For ethanoic acid the extension is made up of **CH₃-** to make the structure;



Ethanoic acid has a **higher** melting/boiling point than ethanol. This is because ethanoic acid has **two/more** hydrogen bond than ethanol.

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II Chemical properties of alkanoic acids

The following experiments shows the main chemical properties of ethanoic (alkanoic) acid.

(a) Effect on litmus papers

Experiment

Dip both blue and red litmus papers in ethanoic acid. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute nitric(V) acid.

Sample observations

Solution/acid	Observations/effect on litmus papers	Inference
Ethanoic acid	Blue litmus paper turn red Red litmus paper remain red	H ₃ O ⁺ /H ⁺ (aq) ion

Succinic acid	Blue litmus paper turn red Red litmus paper remain red	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Citric acid	Blue litmus paper turn red Red litmus paper remain red	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Oxalic acid	Blue litmus paper turn red Red litmus paper remain red	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Tartaric acid	Blue litmus paper turn red Red litmus paper remain red	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Nitric(V)acid	Blue litmus paper turn red Red litmus paper remain red	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion

Explanation

All acidic solutions contains $\text{H}^+/\text{H}_3\text{O}^+(\text{aq})$ ions. The $\text{H}^+/\text{H}_3\text{O}^+(\text{aq})$ ions is responsible for turning blue litmus paper/solution to red

(b)pH

Experiment

Place 2cm³ of ethanoic acid in a test tube. Add 2 drops of universal indicator solution and determine its pH. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI)acid.

Sample observations

Solution/acid	pH	Inference
Ethanoic acid	4/5/6	Weakly acidic
Succinic acid	4/5/6	Weakly acidic
Citric acid	4/5/6	Weakly acidic
Oxalic acid	4/5/6	Weakly acidic
Tartaric acid	4/5/6	Weakly acidic
Sulphuric(VI)acid	1/2/3	Strongly acidic

Explanations

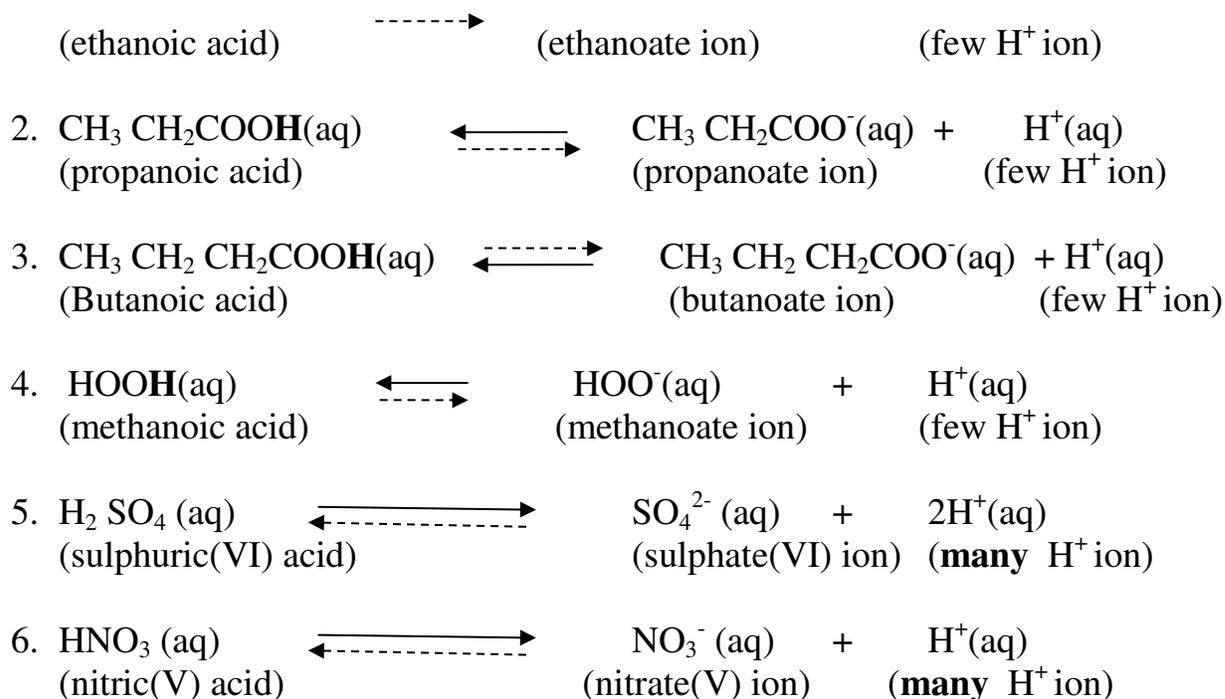
Alkanoic acids are weak acids that partially/partly dissociate to release few H^+ ions in solution. The pH of their solution is thus 4/5/6 showing they form weakly acidic solutions when dissolved in water.

All alkanoic acid dissociate to releases the “**H**” at the functional group in $-\text{COOH}$ to form the **alkanoate ion**; $-\text{COO}^-$

Mineral acids(Sulphuric(VI)acid, Nitric(V)acid and Hydrochloric acid) are strong acids that wholly/fully dissociate to release many H^+ ions in solution. The pH of their solution is thus 1/2/3 showing they form strongly acidic solutions when dissolved in water.i.e

Examples





(c) Reaction with metals

Experiment

Place about 4cm³ of ethanoic acid in a test tube. Put about 1cm length of polished magnesium ribbon. Test any gas produced using a burning splint. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

Solution/acid	Observations	Inference
Ethanoic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Succinic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Citric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Oxalic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
Tartaric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion

Nitric(V)acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with “pop” sound/explosion	$\text{H}_3\text{O}^+/\text{H}^+(\text{aq})$ ion
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Explanation

Metals higher in the reactivity series displace the hydrogen in all acids to evolve/produce hydrogen gas and form a salt. Alkanoic acids react with metals with metals to form alkanoates salt and produce/evolve hydrogen gas .Hydrogen extinguishes a burning splint with a pop sound/explosion. Only the “H” in the functional group $-\text{COOH}$ is /are displaced and not in the alkyl hydrocarbon chain.

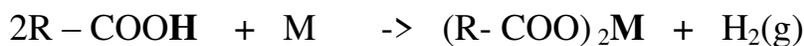
Alkanoic acid + Metal \rightarrow Alkanoate + Hydrogen gas. i.e.

Examples

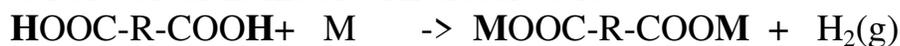
1. For a monovalent metal with monobasic acid



2.For a divalent metal with monobasic acid



3.For a divalent metal with dibasic acid

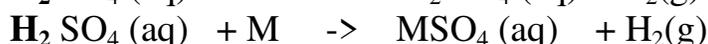
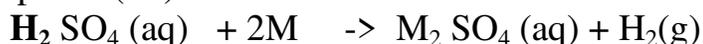


4.For a monovalent metal with dibasic acid



5 For mineral acids

(i)Sulphuric(VI)acid is a dibasic acid



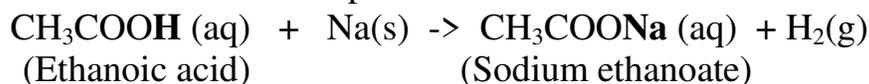
(ii)Nitric(V) and hydrochloric acid are monobasic acid



Examples

1.Sodium reacts with ethanoic acid to form sodium ethanoate and produce hydrogen gas.

Caution: This reaction is explosive.



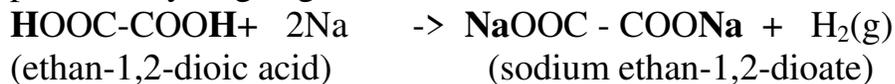
2.Calcium reacts with ethanoic acid to form calcium ethanoate and produce hydrogen gas.



(Ethanoic acid)

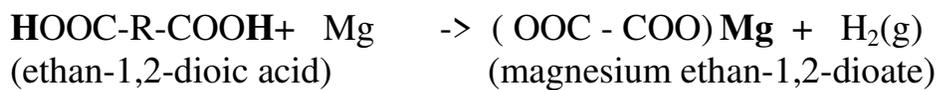
(Calcium ethanoate)

3. Sodium reacts with ethan-1,2-dioic acid to form sodium ethan-1,2-dioate and produce. hydrogen gas.



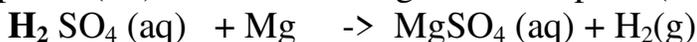
Commercial name of ethan-1,2-dioic acid is oxalic acid. The salt is sodium oxalate.

4. Magnesium reacts with ethan-1,2-dioic acid to form magnesium ethan-1,2-dioate and produce. hydrogen gas.

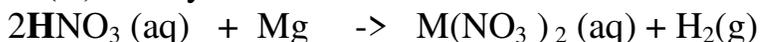


5. Magnesium reacts with

(i) Sulphuric(VI) acid to form Magnesium sulphate(VI)



(ii) Nitric(V) and hydrochloric acid are monobasic acid



(d) Reaction with hydrogen carbonates and carbonates

Experiment

Place about 3cm³ of ethanoic acid in a test tube. Add about 0.5g/ ½ spatula end full of sodium hydrogen carbonate/sodium carbonate. Test the gas produced using lime water. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

Solution/acid	Observations	Inference
Ethanoic acid	(i) effervescence, fizzing, bubbles (ii) colourless gas produced that forms a white precipitate with lime water	H ₃ O ⁺ /H ⁺ (aq) ion
Succinic acid	(i) effervescence, fizzing, bubbles (ii) colourless gas produced that forms a white precipitate with lime water	H ₃ O ⁺ /H ⁺ (aq) ion
Citric acid	(i) effervescence, fizzing, bubbles (ii) colourless gas produced that forms a white precipitate with lime water	H ₃ O ⁺ /H ⁺ (aq) ion
Oxalic acid	(i) effervescence, fizzing, bubbles	H ₃ O ⁺ /H ⁺ (aq) ion

(e) Esterification

Experiment

Place 4cm³ of ethanol acid in a boiling tube.

Add equal volume of ethanoic acid. To the mixture, add 2 drops of concentrated sulphuric(VI) acid **carefully**. Warm/heat gently on Bunsen flame.

Pour the mixture into a beaker containing 50cm³ of water. Smell the products.

Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

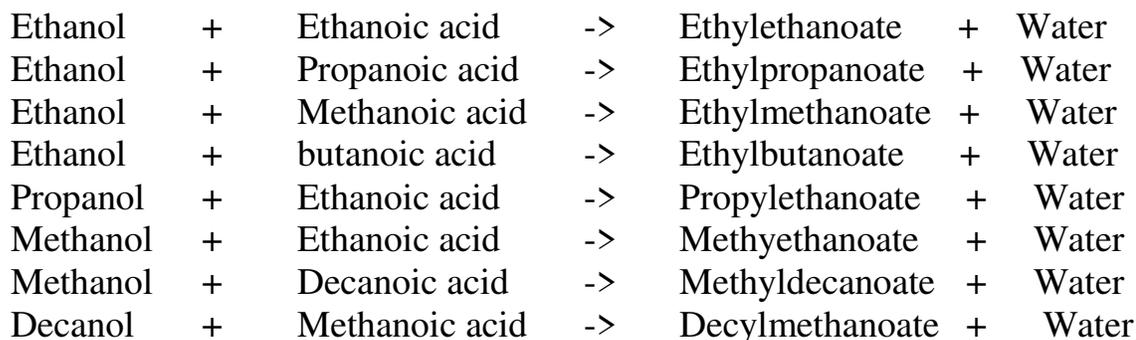
Solution/acid	Observations
Ethanoic acid	Sweet fruity smell
Succinic acid	Sweet fruity smell
Citric acid	Sweet fruity smell
Oxalic acid	Sweet fruity smell
Tartaric acid	Sweet fruity smell
Dilute sulphuric(VI) acid	No sweet fruity smell

Explanation

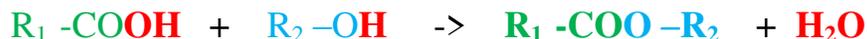
Alkanols react with alkanolic acid to form the sweet smelling homologous series of esters and water. The reaction is catalysed by concentrated sulphuric(VI) acid in the laboratory but naturally by sunlight /heat. Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanolic acids.



Esters derive their names from the alkanol first then alkanolic acids. The alkanol “becomes” an **alkyl** group and the alkanolic acid “becomes” **alkanoate** hence **alkylalkanoate**. e.g.

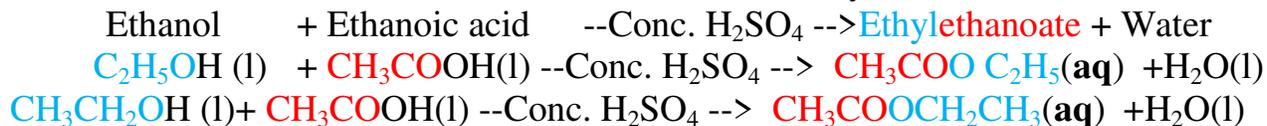


During the formation of the ester, the “O” joining the alkanol and alkanoic acid comes from the alkanol.

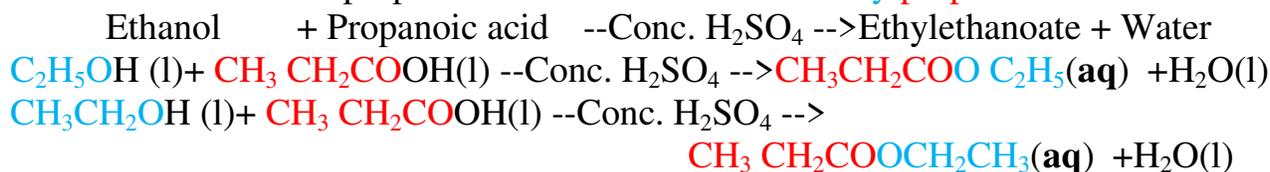


Examples

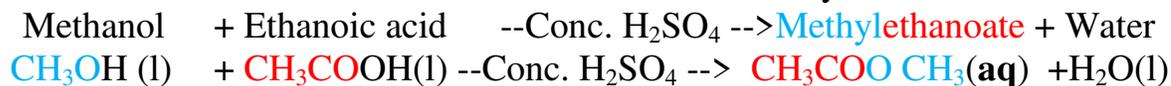
1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.



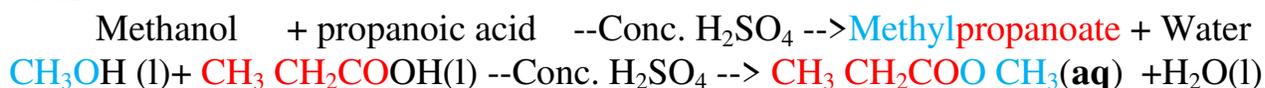
2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.



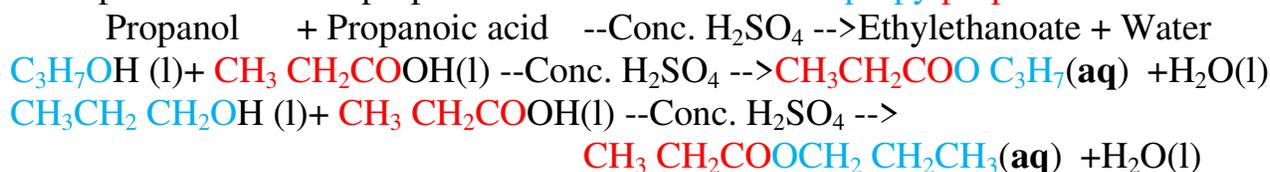
3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.



4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.



5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.



C. DETERGENTS

Detergents are cleaning agents that improve the cleaning power /properties of water. A detergent therefore should be able to:

- (i) dissolve substances which water can not e.g grease ,oil, fat
- (ii) be washed away after cleaning.

There are two types of detergents:

- (a) Soapy detergents
- (b) Soapless detergents

(a) SOAPY DETERGENTS

Soapy detergents usually called soap is long chain salt of organic alkanolic acids. Common soap is sodium octadecanoate .It is derived from reacting concentrated sodium hydroxide solution with octadecanoic acid(18 carbon alkanolic acid) i.e.

Sodium hydroxide + octadecanoic acid \rightarrow Sodium octadecanoate + water
 $\text{NaOH(aq)} + \text{CH}_3(\text{CH}_2)_{16}\text{COOH(aq)} \rightarrow \text{CH}_3(\text{CH}_2)_{16}\text{COO}^- \text{Na}^+ \text{(aq)} + \text{H}_2\text{O(l)}$
Commonly ,soap can thus be represented ;

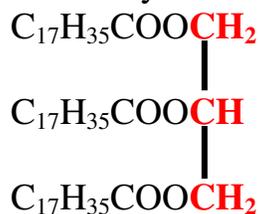


R is a long chain alkyl group and $-\text{COO}^- \text{Na}^+$ is the alkanolate ion.

In a school laboratory and at industrial and domestic level, soap is made by reacting concentrated sodium hydroxide solution with esters from (animal) **fat** and **oil**. The process of making soap is called **saponification**. During saponification ,the ester is **hydrolyzed** by the alkali to form sodium salt /soap and **glycerol/propan-1,2,3-triol** is produced.

Fat/oil(ester)+sodium/potassium hydroxide \rightarrow sodium/potassium salt(soap)+ glycerol

Fats/Oils are esters with fatty acids and glycerol parts in their structure;



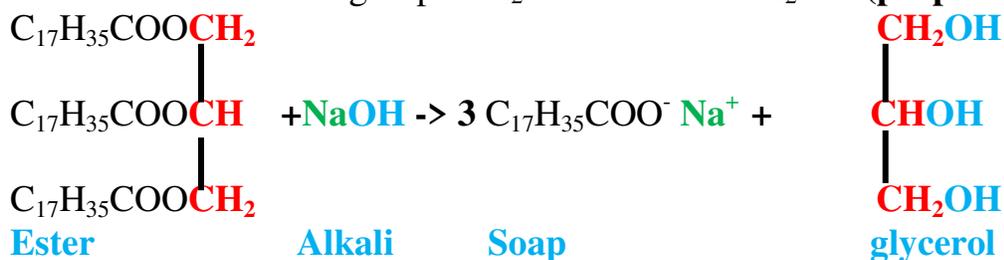
When boiled with concentrated sodium hydroxide solution NaOH;

(i) NaOH ionizes/dissociates into Na^+ and OH^- ions

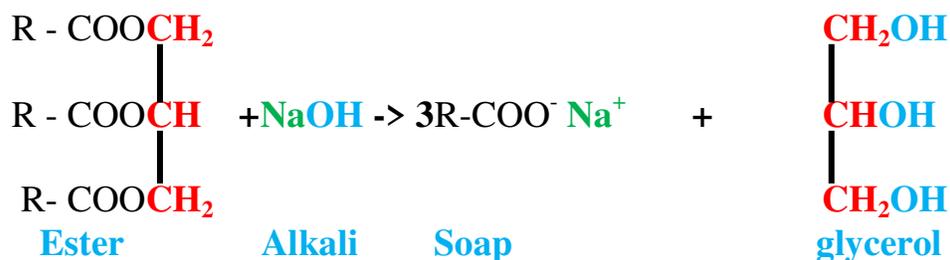
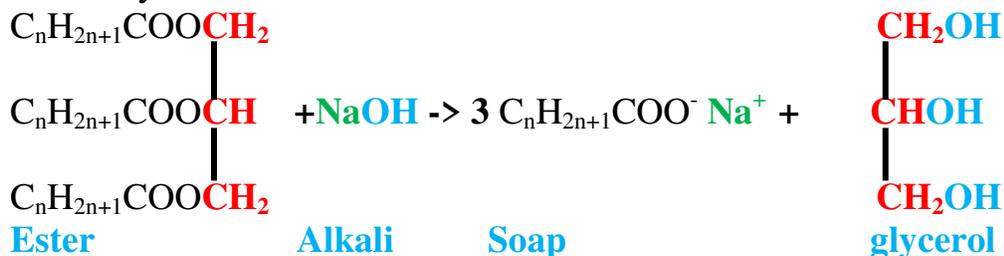
(ii) fat/oil split into **three** $\text{C}_{17}\text{H}_{35}\text{COO}^-$ and **one** CH_2CHCH_2

(iii) the three Na^+ combine with the three $\text{C}_{17}\text{H}_{35}\text{COO}^-$ to form the salt $\text{C}_{17}\text{H}_{35}\text{COO}^- \text{Na}^+$

(iv) the three OH^- ions combine with the CH_2CHCH_2 to form an alkanol with three functional groups $\text{CH}_2\text{OHCH(OH)CH}_2\text{OH}$ (propan-1,2,3-triol)



Generally:



During this process a little sodium chloride is added to **precipitate** the soap by reducing its solubility. This is called **salting out**.

The soap is then added colouring agents ,perfumes and herbs of choice.

School laboratory preparation of soap

Place about 40 g of fatty (animal fat)beef/meat in 100cm³ beaker .Add about 15cm³ of 4.0M sodium hydroxide solution. Boil the mixture for about 15minutes.Stir the mixture .Add about 5.0cm³ of distilled water as you boil to make up for evaporation. Boil for about another 15minutes.Add about four spatula end full of pure sodium chloride crystals. Continue stirring for another five minutes. Allow to cool. Filter of

/decant and wash off the residue with distilled water .Transfer the clean residue into a dry beaker. Preserve.

The action of soap

Soapy detergents:

(i)act by reducing the surface tension of water by forming a thin layer on top of the water.

(ii)is made of a **non-polar** alkyl /hydrocarbon tail and a **polar** $-\text{COO}^-\text{Na}^+$ head. The non-polar alkyl /hydrocarbon tail is **hydrophobic** (water hating) and thus does not dissolve in water .It dissolves in non-polar solvent like grease, oil and fat. The polar $-\text{COO}^-\text{Na}^+$ head is **hydrophilic** (water loving)and thus dissolve in water. When washing with soapy detergent, the non-polar tail of the soapy detergent surround/dissolve in the dirt on the garment /grease/oil while the polar head dissolve in water.

Through **mechanical agitation**/stirring/squeezing/rubbing/beating/kneading, some grease is dislodged/lifted of the surface of the garment. It is immediately surrounded by more soap molecules It float and spread in the water as tiny droplets that scatter light in form of emulsion making the water cloudy and shinny. It is removed from the garment by rinsing with fresh water.The repulsion of the soap head prevent /ensure the droplets do not mix.Once removed, the dirt molecules cannot be redeposited back because it is surrounded by soap molecules.

Advantages and disadvantages of using soapy detergents

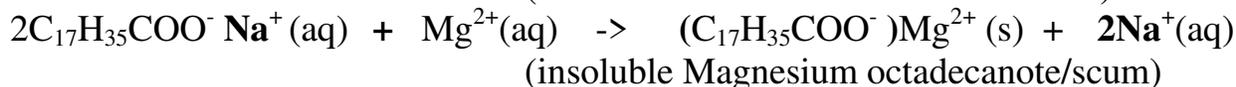
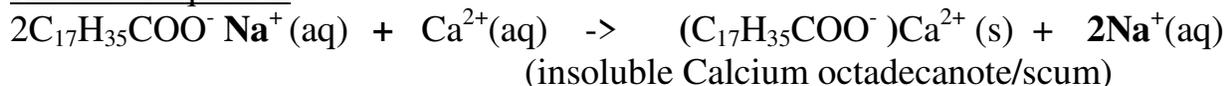
Soapy detergents are biodegradable. They are acted upon by bacteria and rot.They thus do not cause environmental pollution.

Soapy detergents have the diadvatage in that:

(i)they are made from fat and oils which are better eaten as food than make soap.

(ii)forms an insoluble precipitate with hard water called **scum**. Scum is insoluble calcium octadecanoate and Magnesium octadecanoate formed when soap reacts with Ca^{2+} and Mg^{2+} present in hard water.

Chemical equation



This causes wastage of soap.

Potassium soaps are better than Sodium soap. Potassium is more expensive than sodium and thus its soap is also more expensive.

(b)SOAPLESS DETERGENTS

Soapless detergent usually called detergent is a long chain salt formed from by-products of fractional distillation of crude oil. Commonly used soaps include:

- (i) washing agents
- (ii) toothpaste
- (iii) emulsifiers/wetting agents/shampoo

Soapless detergents are derived from reacting:

(i) concentrated sulphuric(VI) acid with a long chain alkanol e.g. Octadecanol (18 carbon alkanol) to form alkyl hydrogen sulphate(VI)

Alkanol + Conc sulphuric(VI) acid \rightarrow alkyl hydrogen sulphate(VI) + Water

$$R-OH + H_2SO_4 \rightarrow R-O-SO_3H + H_2O$$

(ii) the alkyl hydrogen sulphate(VI) is then neutralized with sodium/potassium hydroxide to form sodium/potassium alkyl hydrogen sulphate(VI). Sodium/potassium alkyl hydrogen sulphate(VI) is the soapless detergent.

alkyl hydrogen sulphate(VI) + Potassium/sodium hydroxide \rightarrow Sodium/potassium alkyl hydrogen sulphate(VI) + Water

$$R-O-SO_3H + NaOH \rightarrow R-O-SO_3^- Na^+ + H_2O$$

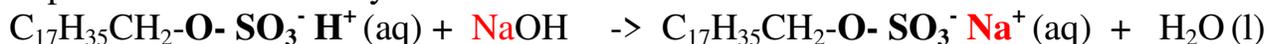
Example

Step I: Reaction of Octadecanol with Conc. H_2SO_4



octadecanol + sulphuric(VI) acid \rightarrow Octadecyl hydrogen sulphate(VI) + water

Step II: Neutralization by an alkali



Octadecyl hydrogen sulphate(VI) + sodium/potassium hydroxide \rightarrow sodium/potassium octadecyl hydrogen sulphate(VI) + Water

School laboratory preparation of soapless detergent

Place about 20g of olive oil in a 100cm³ beaker. Put it in a trough containing ice cold water.

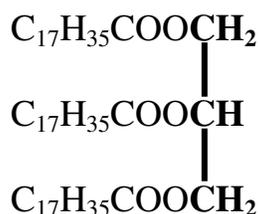
Add dropwise carefully 18M concentrated sulphuric(VI) acid stirring continuously into the olive oil until the oil turns brown. Add 30cm³ of 6M sodium hydroxide solution. Stir. This is a soapless detergent.

The action of soapless detergents

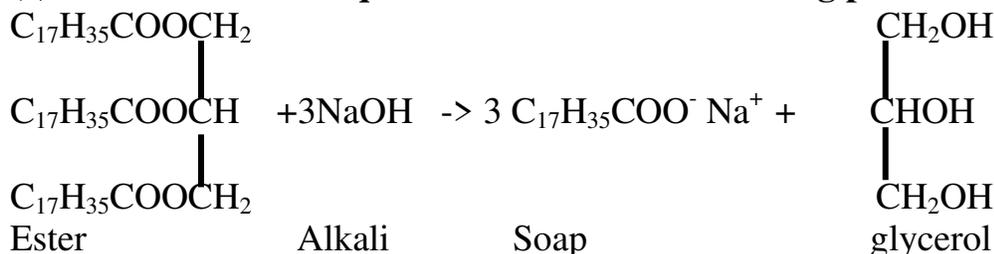
(a) Identify the process

Saponification

(b) Fats and oils are esters. Write the formula of the a common structure of ester



(c) Write a balanced equation for the reaction taking place during boiling



(d) Give the IUPAC name of:

(i) Residue X

Potassium octadecanoate

(ii) Filtrate Y

Propan-1,2,3-triol

(e) Give one use of filtrate Y

Making paint

(f) What is the function of sodium chloride

To reduce the solubility of the soap hence helping in precipitating it out

(g) Explain how residue X helps in washing.

Has a non-polar hydrophobic tail that dissolves in dirt/grease /oil/fat

Has a polar /ionic hydrophilic head that dissolves in water.

From mechanical agitation, the dirt is plucked out of the garment and surrounded by the tail end preventing it from being deposited back on the garment.

(h) State one:

(i) advantage of continued use of residue X on the environment

Is biodegradable and thus do not pollute the environment

(ii) disadvantage of using residue X

Uses fat/oil during preparation/manufacture which are better used for human consumption.

(i) Residue X was added dropwise to some water. The number of drops used before lather forms is as in the table below.

	Water sample		
	A	B	C
Drops of residue X	15	2	15
Drops of residue X in boiled water	2	2	15

(i) State and explain which sample of water is:

I. Soft

Sample B . Very little soap is used and no effect on amount of soap even on boiling/heating.

II. Permanent hard

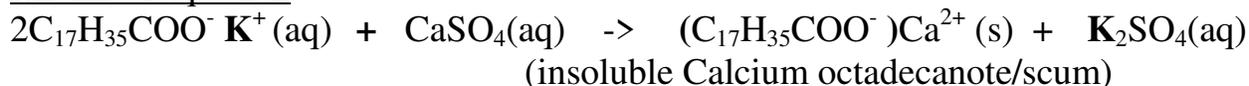
Sample C . A lot of soap is used and no effect on amount of soap even on boiling/heating. Boiling does not remove permanent hardness of water.

III. Temporary hard

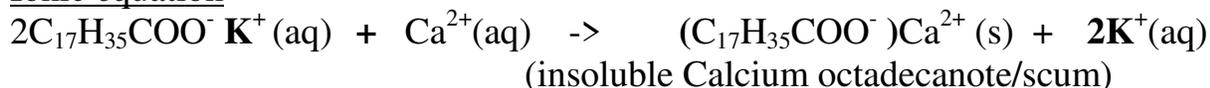
Sample A . A lot of soap is used before boiling. Very little soap is used on boiling/heating. Boiling remove temporary hardness of water.

(ii) Write the equation for the reaction at water sample C.

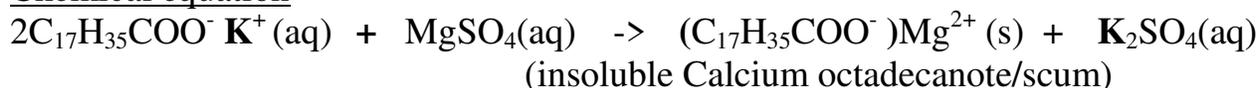
Chemical equation



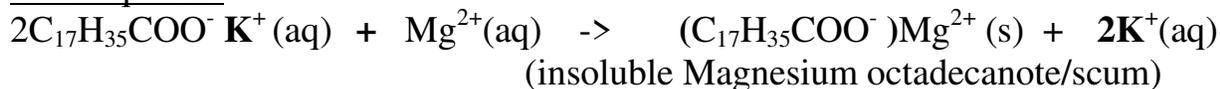
Ionic equation



Chemical equation

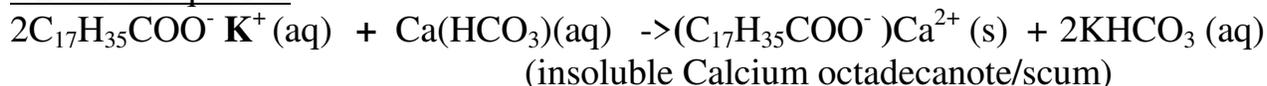


Ionic equation

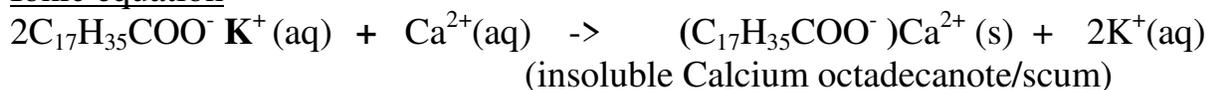


(iii) Write the equation for the reaction at water sample A before boiling.

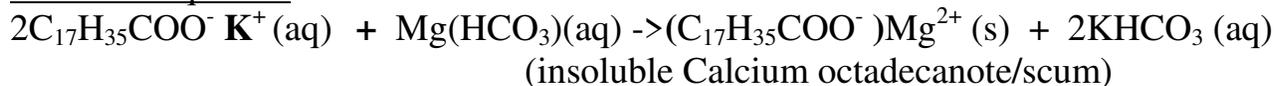
Chemical equation



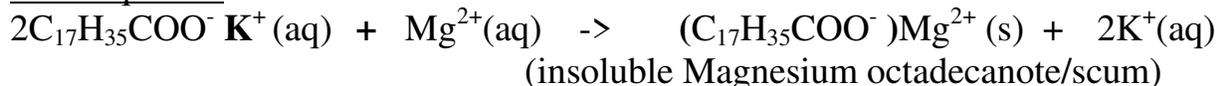
Ionic equation



Chemical equation



Ionic equation



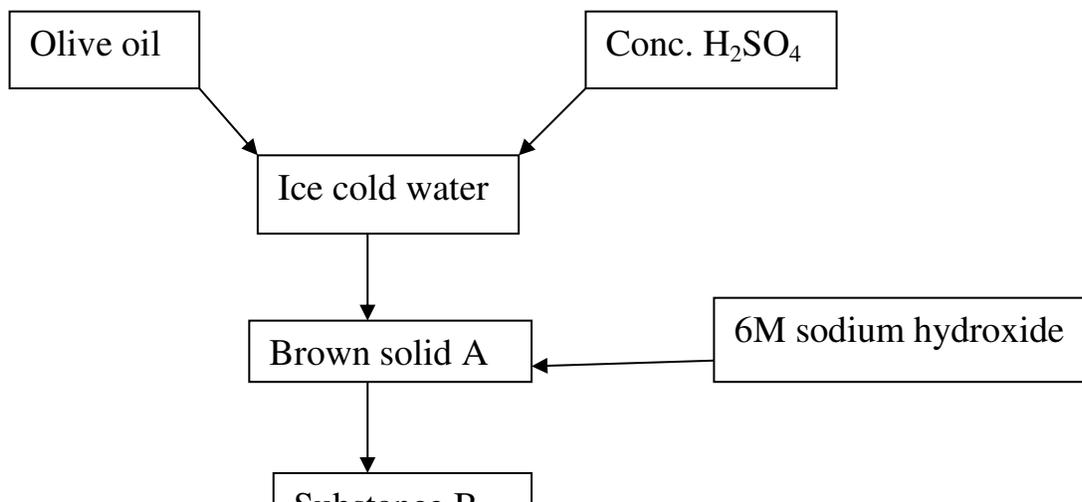
(iv) Explain how water becomes hard

Natural or rain water flowing /passing through rocks containing calcium (chalk, gypsum, limestone) and magnesium compounds (dolomite) dissolve them to form soluble Ca^{2+} and Mg^{2+} ions that causes water hardness.

(v) State two useful benefits of hard water

- Used in bone and teeth formation
- Coral polyps use hard water to form coral reefs
- Snails use hard water to make their shells

2. Study the scheme below and use it to answer the questions that follow.



(a) Identify :

(i) brown solid A

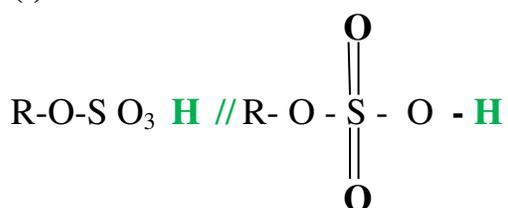
Alkyl hydrogen sulphate(VI)

(ii) substance B

Sodium alkyl hydrogen sulphate(VI)

(b) Write a general formula of:

(i) Substance A.



(ii) Substance B



(c) State one

(i) advantage of continued use of substance B

- Does not form scum with hard water
- Is cheap to make
- Does not use food for human as a raw material.

(ii) disadvantage of continued use of substance B.

Is non-biodegradable therefore do not pollute the environment

(d) Explain the action of B during washing.

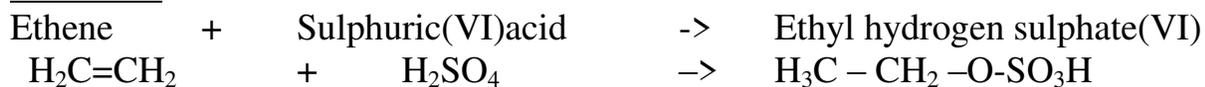
Has a non-polar hydrocarbon long tail that dissolves in dirt/grease/oil/fat.

Has a polar/ionic hydrophilic head that dissolves in water

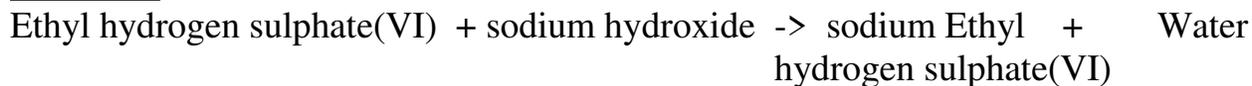
Through mechanical agitation the dirt is plucked /removed from the garment and surrounded by the tail end preventing it from being deposited back on the garment.

(e) Ethene was substituted for olive oil in the above process. Write the equation and name of the new products A and B.

Product A



Product B

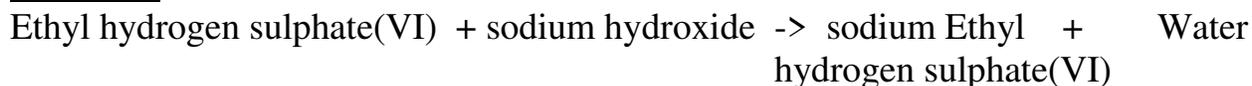


(f) Ethanol can also undergo similar reactions forming new products A and B. Show this using a chemical equation.

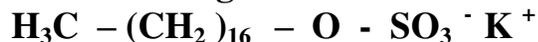
Product A



Product B



3. Below is part of a detergent



(a) Write the formula of the polar and non-polar end

Polar end



Non-polar end



(b) Is the molecule a soapy or soapless detergent?

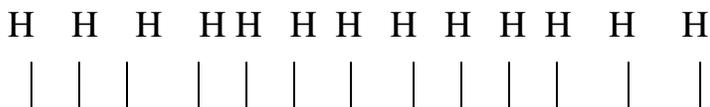
Soapless detergent

(c) State one advantage of using the above detergent

-does not form scum with hard water

-is cheap to manufacture

4. The structure of a detergent is



- polychloroethene
- polyphenylethene(polystyrene)
- Terylene(Dacron)
- Nylon-6,6
- Perspex(artificial glass)

Synthetic polymers and fibres have the following characteristic advantages over natural polymers

1. They are light and portable
2. They are easy to manufacture.
3. They can easily be molded into shape of choice.
4. They are resistant to corrosion, water, air , acids, bases and salts.
5. They are comparatively cheap, affordable, colourful and aesthetic

Synthetic polymers and fibres however have the following disadvantages over natural polymers

1. They are non-biodegradable and hence cause environmental pollution during disposal
2. They give out highly poisonous gases when burnt like chlorine/carbon(II)oxide
3. Some on burning produce Carbon(IV)oxide. Carbon(IV)oxide is a green house gas that cause global warming.
4. Compared to some metals, they are poor conductors of heat,electricity and have lower tensile strength.
- 5.

To reduce environmental pollution from synthetic polymers and fibres, the followitn methods of disposal should be used:

- 1.Recycling: Once produced all synthetic polymers and fibres should be recycled to a new product. This prevents accumulation of the synthetic polymers and fibres in the environment.
- 2.Production of biodegradable synthetic polymers and fibres that **rot** away.

There are two types of polymerization:

- (a)addition polymerization
- (b)condensation polymerization

(a)addition polymerization

Addition polymerization is the process where a small unsaturated monomer (alkene) molecule join together to form a large saturated molecule. Only alkenes undergo addition polymerization.

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix “**poly**” before the name of monomer to form a **polyalkene**

During addition polymerization

- (i) the double bond in alkenes break
- (ii) free radicals are formed
- (iii) the free radicals collide with each other and join to form a larger molecule.

The more collisions the larger the molecule.

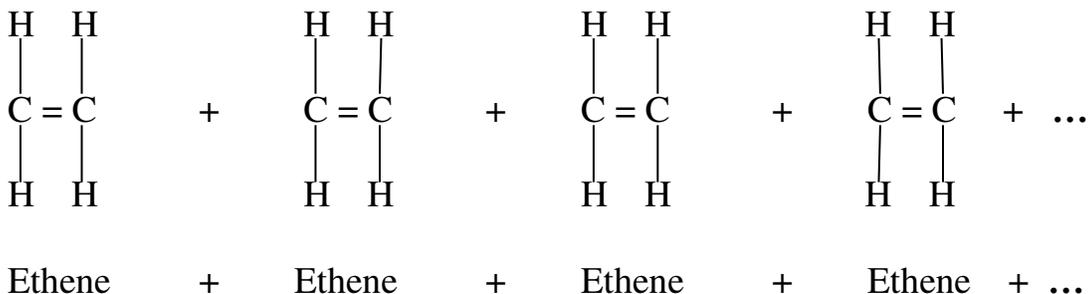
Examples of addition polymerization

1. Formation of Polyethene

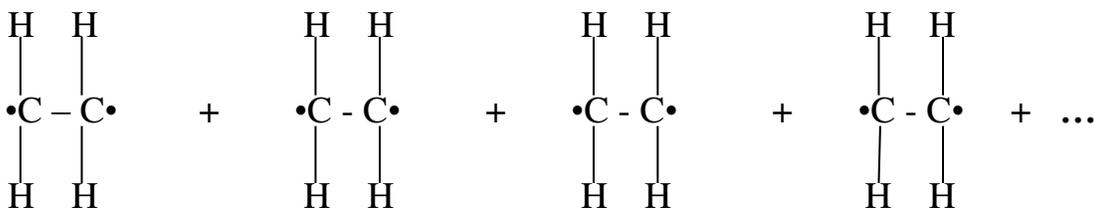
Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

- (i) many molecules are brought nearer to each other by the high pressure (which reduces the volume occupied by reacting particles)

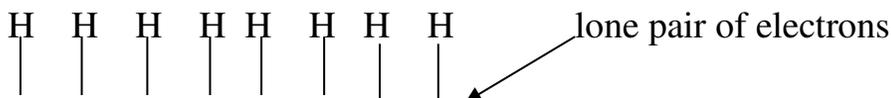


- (ii) the double bond joining the ethane molecule break to free radicals



Ethene radical + Ethene radical + Ethene radical + Ethene radical + ...

- (iii) the free radicals collide with each other and join to form a larger molecule

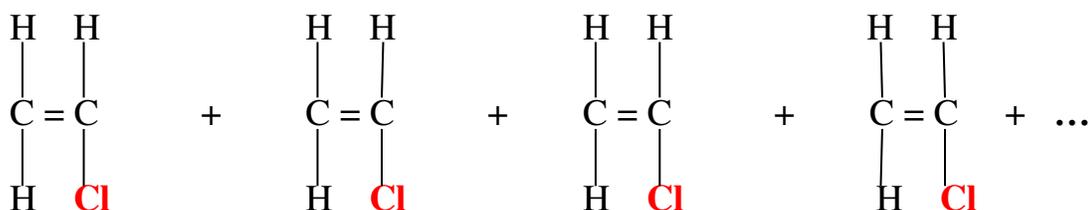


2. Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

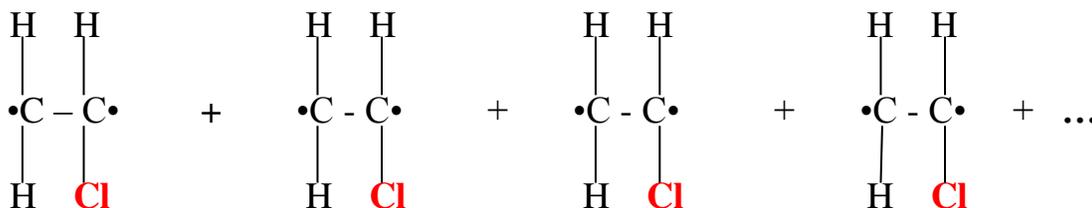
During polymerization:

(i) many molecules are brought nearer to each other by the high pressure (which reduces the volume occupied by reacting particles)

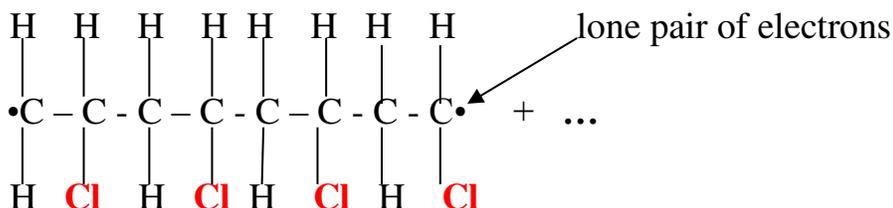


chloroethene + chloroethene + chloroethene + chloroethene + ...

(ii) the double bond joining the chloroethene molecule break to free radicals

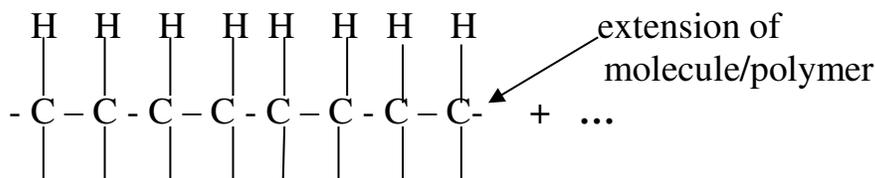


(iii) the free radicals collide with each other and join to form a larger molecule



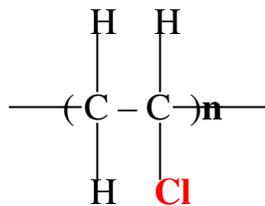
Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:





Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polychloroethene has a molar mass of 4760. Calculate the number of chlorethene molecules in the polymer(C=12.0, H=1.0,Cl=35.5)

Number of monomers/repeating units in monomer = $\frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$

=> Molar mass ethene (C₂H₃Cl) = 62.5 Molar mass polyethene = 4760

Substituting $\frac{4760}{62.5} = 77.16 \Rightarrow 77$ polychloroethene molecules(**whole number**)

The **commercial** name of polychloroethene is **polyvinylchloride(PVC)**. It is a tough, non-transparent and durable plastic. PVC is used:

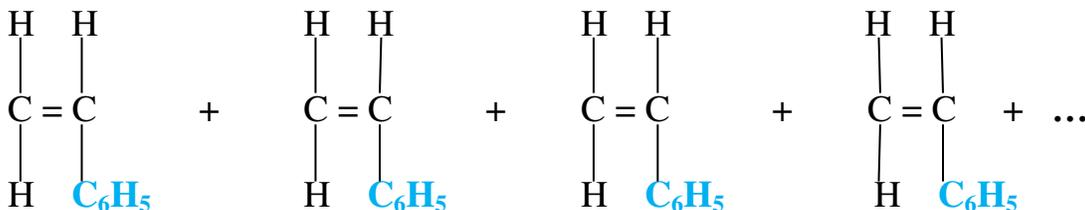
- (i)in making plastic rope
- (ii)water pipes
- (iii)crates and boxes

3.Formation of Polyphenylethene

Polyphenylethene is an addition polymer formed when phenylethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

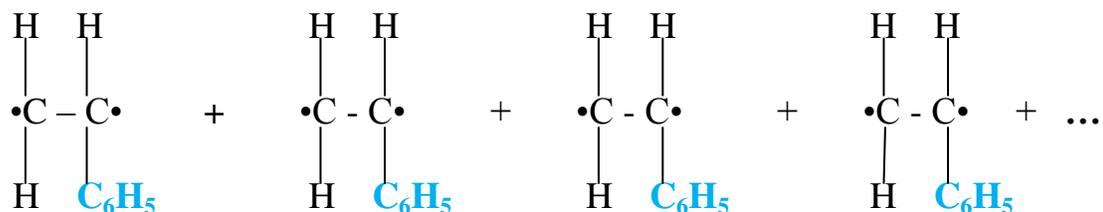
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

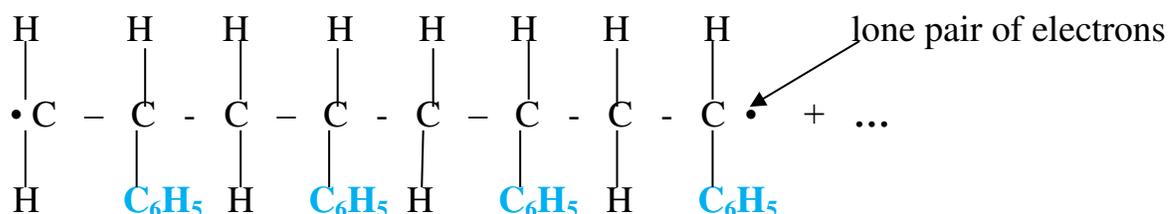


phenylethene + phenylethene + phenylethene + phenylethene + ...

(ii) the double bond joining the phenylethene molecule break to free radicals

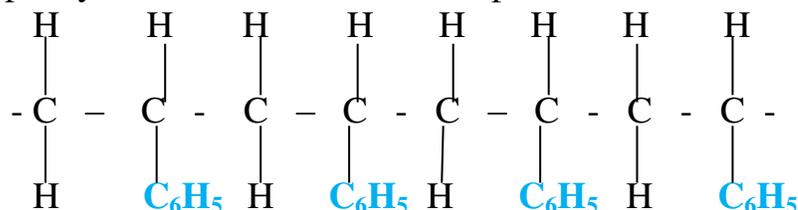


(iii) the free radicals collide with each other and join to form a larger molecule

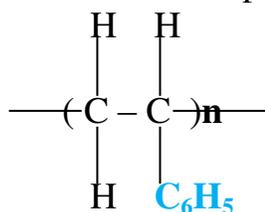


Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



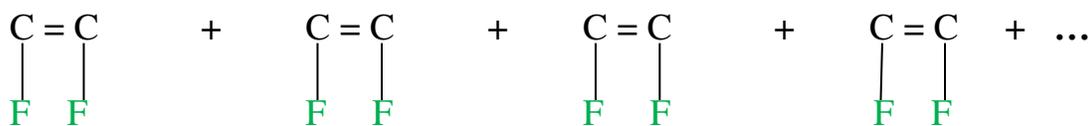
Examples

Polyphenylthene has a molar mass of 4760. Calculate the number of phenylethene molecules in the polymer (C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = $\frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$

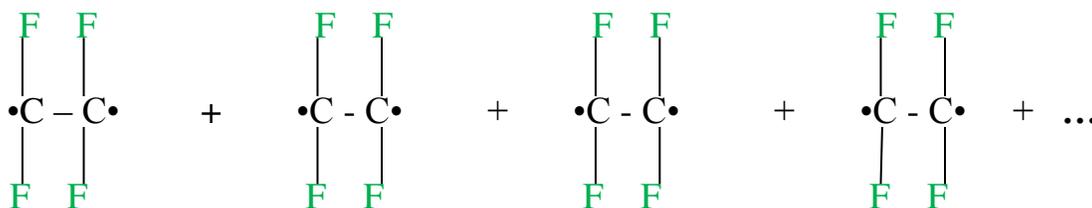
=> Molar mass ethene (C₈H₈) = 104 Molar mass polyethene = 4760

Substituting $\frac{4760}{104} = 45.7692 \Rightarrow 45$ polyphenylethene molecules (whole number)

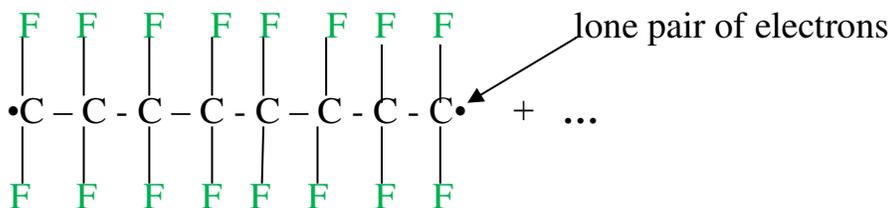


tetrafluoroethene + tetrafluoroethene+ tetrafluoroethene+ tetrafluoroethene + ...

(ii)the double bond joining the tetrafluoroethene molecule break to free radicals

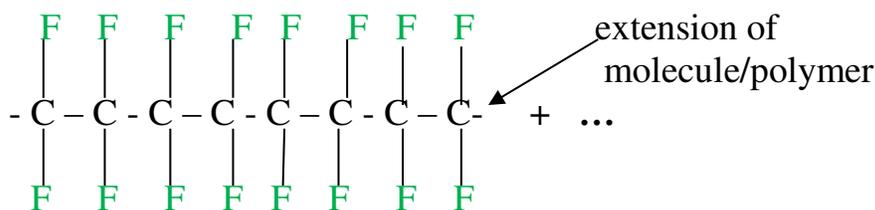


(iii)the free radicals collide with each other and join to form a larger molecule

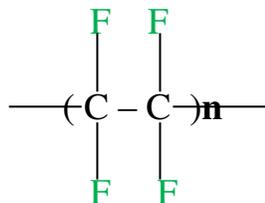


Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polytetrafluoroethene has a molar mass of 4760. Calculate the number of tetrafluoroethene molecules in the polymer (C=12.0, F=19)

Number of monomers/repeating units in monomer = $\frac{\text{Molar mass polymer}}{\text{Molar mass monomer}}$

=> Molar mass ethene (C₂F₄) = 62.5 Molar mass polyethene = 4760

Substituting $\frac{4760}{62.5} = 77.16 \Rightarrow 77$ polychloroethene molecules (whole number)

The **commercial** name of polytetrafluorethene (P.T.F.E) is **Teflon (P.T.F.E)**. It is a tough, non-transparent and durable plastic. PVC is used:

- (i) in making plastic rope
- (ii) water pipes
- (iii) crates and boxes

5. Formation of rubber from Latex

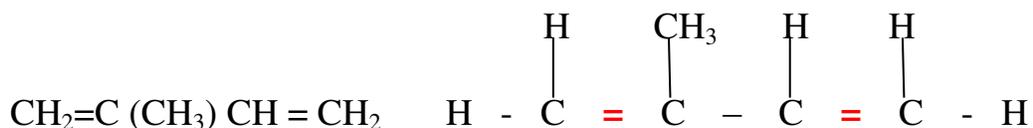
Natural rubber is obtained from rubber trees.

During harvesting an incision is made on the rubber tree to produce a milky white substance called **latex**.

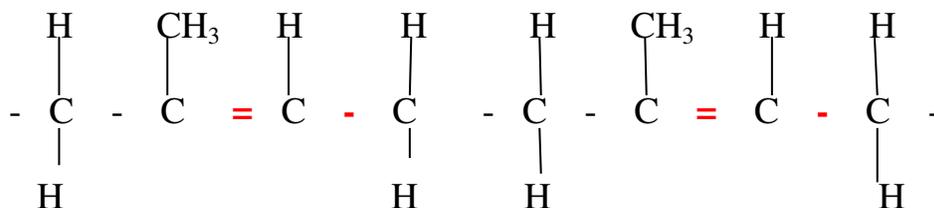
Latex is a mixture of rubber and lots of water.

The latex is then added an acid to coagulate the rubber.

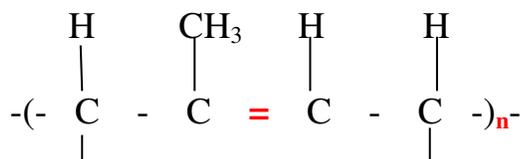
Natural rubber is a polymer of 2-methylbut-1,3-diene ;

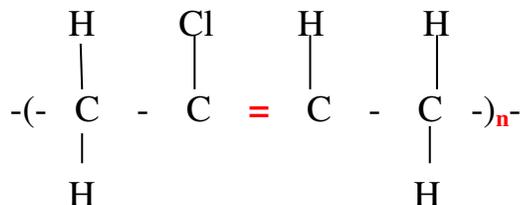


During natural polymerization to rubber, one double C=C bond break to self add to another molecule. The double bond remaining move to carbon "2" thus;



Generally the structure of rubber is thus;





Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

(b)Condensation polymerization

Condensation polymerization is the process where two or more small monomers join together to form a larger molecule by elimination/removal of a simple molecule. (usually water).

Condensation polymers acquire a different name from the monomers because the two monomers are two different compounds

During condensation polymerization:

- (i)the two monomers are brought together by high pressure to reduce distance between them.
- (ii)monomers realign themselves at the functional group.
- (iii)from each functional group an element is removed so as to form simple molecule (of usually H₂O/HCl)
- (iv)the two monomers join without the simple molecule of H₂O/HCl

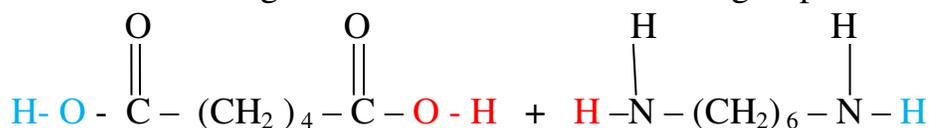
Examples of condensation polymerization

1.Formation of Nylon-6,6

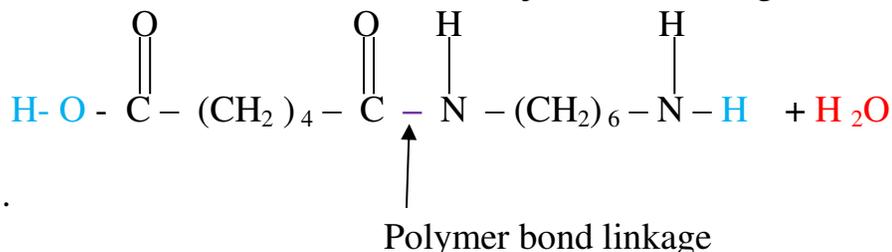
Method 1: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioic acid with hexan-1,6-diamine. Amines are a group of homologous series with a general formula R-NH₂ and thus -NH₂ as the functional group.

During the formation of Nylon-6,6:

- (i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii) from each functional group an element is removed so as to form a molecule of H₂O and the two monomers join at the linkage .



Nylon-6,6 derive its name from the two monomers each with **six** carbon chain

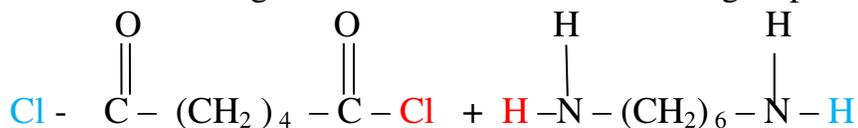
Method 2: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioyl dichloride with hexan-1,6-diamine.

Hexan-1,6-dioyl dichloride belong to a group of homologous series with a general formula R-OC_l and thus -OC_l as the functional group.

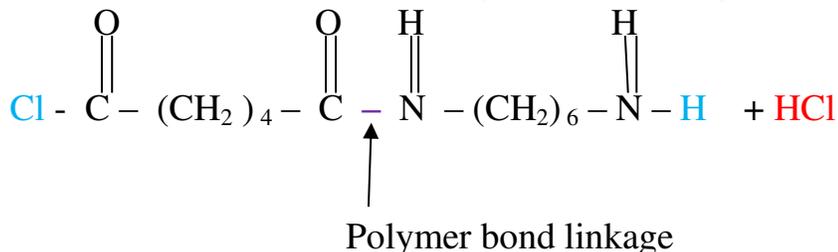
The R-OC_l is formed when the "OH" in R-OOH/alkanoic acid is replaced by Cl/chlorine/Halogen

During the formation of Nylon-6,6:

(i) the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii) from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .



The two monomers each has **six** carbon chain hence the name "nylon-6,6"

The commercial name of Nylon-6,6 is **Nylon** It is a a tough, elastic and durable plastic. It is used to make **clothes, plastic ropes and carpets.**

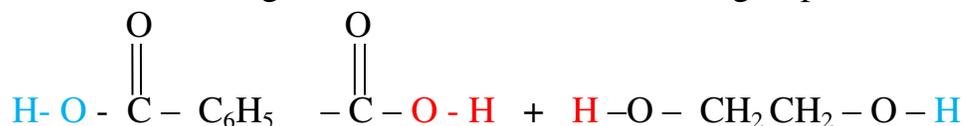
2. Formation of Terylene

Method 1: Terylene can be made from the condensation polymerization of ethan-1,2-diol with benzene-1,4-dicarboxylic acid.

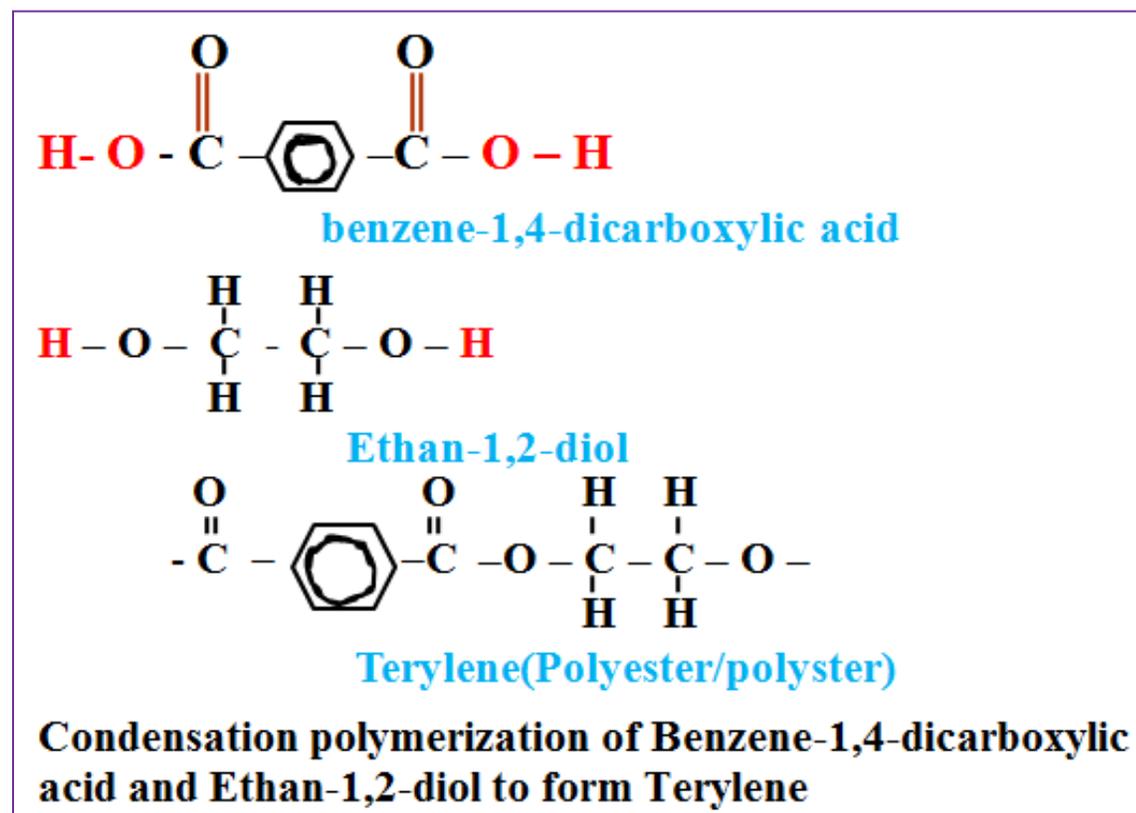
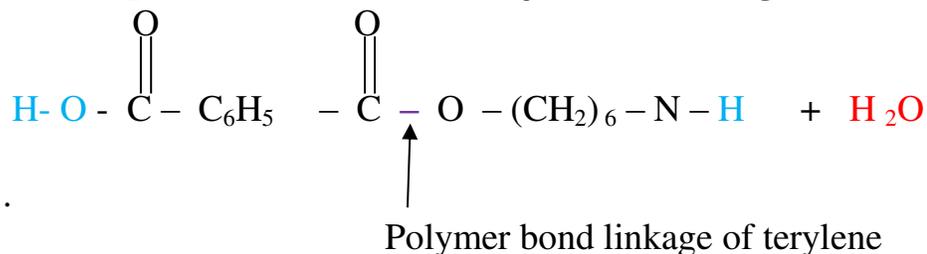
Benzene-1,4-dicarboxylic acid a group of homologous series with a general formula R-COOH where R is a ring of six carbon atom called Benzene ring .The functional group is -COOH.

During the formation of Terylene:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii)from each functional group an element is removed so as to form a molecule of H₂O and the two monomers join at the linkage .



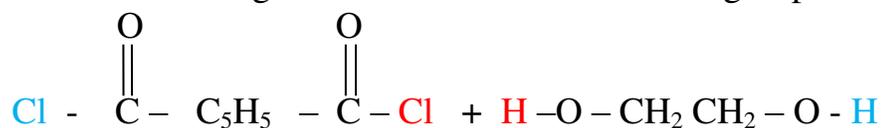
Method 2: Terylene can be made from the condensation polymerization of benzene-1,4-dioyl dichloride with ethan-1,2-diol.

Benzene-1,4-dioyl dichloride belong to a group of homologous series with a general formula R-OCl and thus -OCl as the functional group and R as a benzene ring.

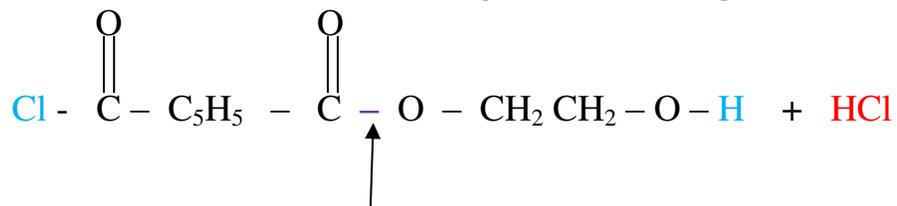
The R-OCl is formed when the "OH" in R-OOH is replaced by Cl/chlorine/Halogen

During the formation of Terylene

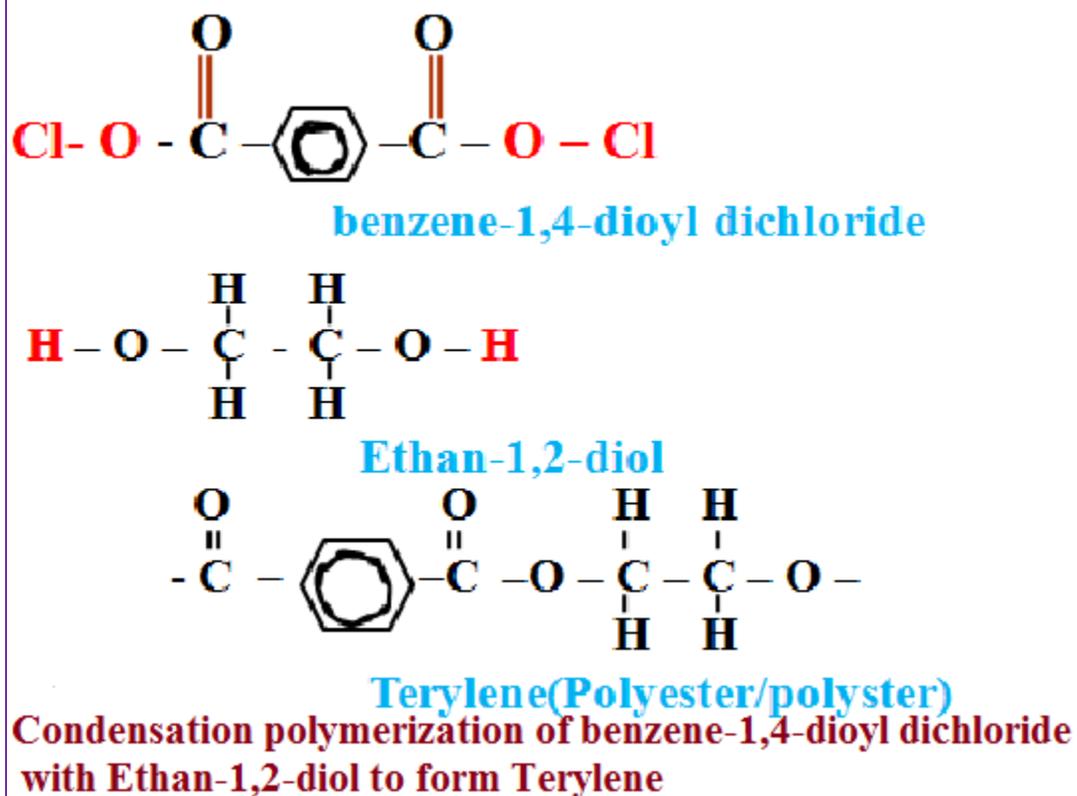
(i) the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii) from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .



↑
Polymer bond linkage of terylene



The commercial name of terylene is **Polyester /polyster** It is a a tough, elastic and durable plastic. It is used to make **clothes, plastic ropes and sails** and **plastic model kits**.

Practice questions Organic chemistry

1. A student mixed equal volumes of Ethanol and butanoic acid. He added a few drops of concentrated Sulphuric (VI) acid and warmed the mixture

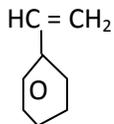
(i) Name and write the formula of the main products

Name.....

Formula.....

(ii) Which homologous series does the product named in (i) above belong?

2. The structure of the monomer phenyl ethene is given below:-



a) Give the structure of the polymer formed when four of the monomers are added together

b) Give the name of the polymer formed in (a) above

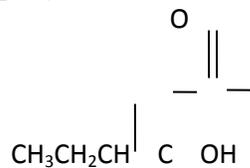
3. Explain the environmental effects of burning plastics in air as a disposal method

4. Write chemical equation to represent the effect of heat on ammonium carbonate

5. Sodium octadecanoate has a chemical formula $\text{CH}_3(\text{CH}_2)_6 \text{COO}^- \text{Na}^+$, which is used as soap.

Explain why a lot of soap is needed when washing with hard water

6. A natural polymer is made up of the monomer:



(a) Write the structural formula of the repeat unit of the polymer

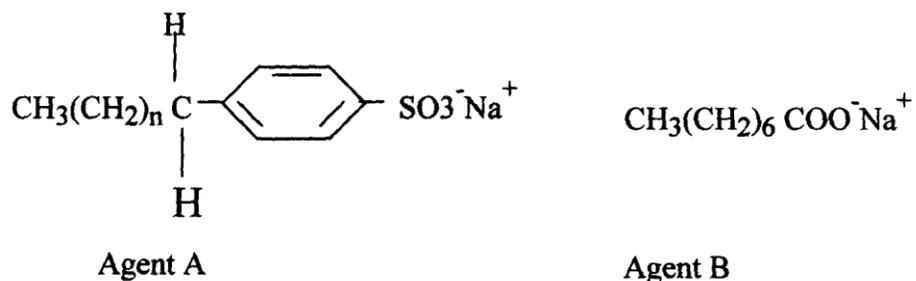
(b) When 5.0×10^{-5} moles of the polymer were hydrolysed, 0.515g of the monomer

were obtained.

Determine the number of the monomer molecules in this polymer.

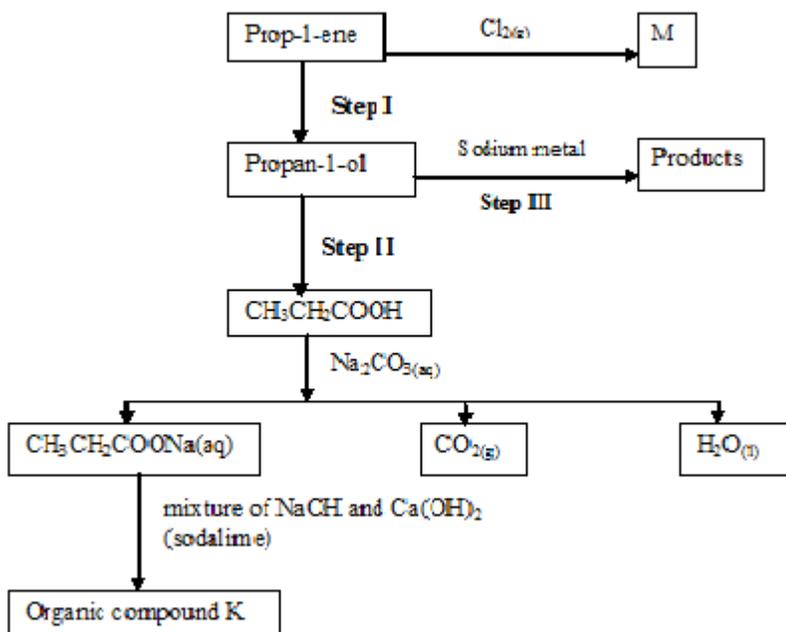
(C = 12; H = 1; N = 14; O = 16)

7. The formula below represents active ingredients of two cleansing agents **A** and **B**

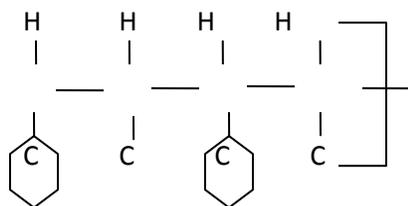


Which one of the cleansing agents would be suitable to be used in water containing magnesium hydrogen carbonate? Explain

(b) Study the flow chart below and use it to answer the questions that follow.



8. Study the polymer below and use it to answer the questions that follow:



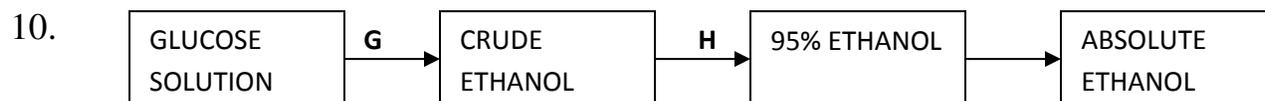
(a) Give the name of the monomer and draw its structures

(b) Identify the type of polymerization that takes place

(c) State **one** advantage of synthetic polymers

9. Ethanol and Pentane are miscible liquids. Explain how water can be used to separate a mixture

of ethanol and pentane



(a) What is absolute ethanol?

(b) State **two** conditions required for process **G** to take place efficiently

11. (a) (i) The table below shows the volume of oxygen obtained per unit time when hydrogen

peroxide was decomposed in the presence of manganese (IV) Oxide.

Use it to answer

the questions that follow:-

Time in seconds	Volume of Oxygen evolved (cm³)
0	0
30	10
60	19
90	27
120	34
150	38
180	43
210	45
240	45
270	45
300	45

(i) Plot a graph of volume of oxygen gas against time

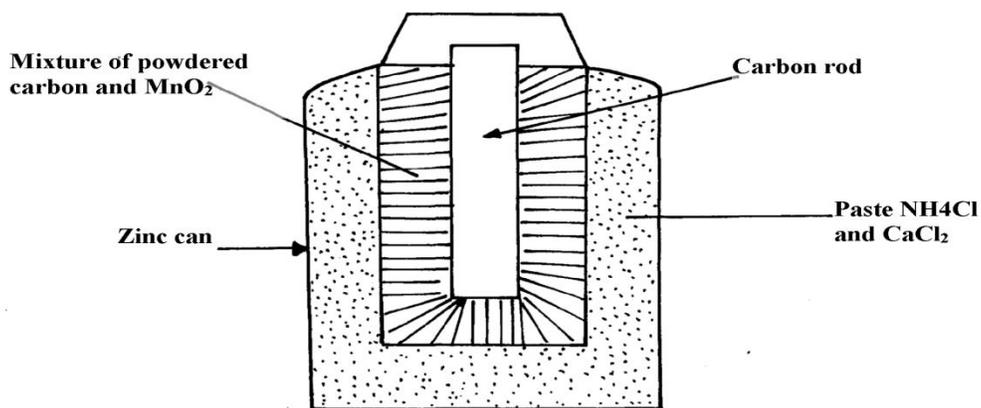
(ii) Determine the rate of reaction at time 156 seconds

(iii) From the graph, find the time taken for 18cm^3 of oxygen to be produced

(iv) Write a chemical equation to show how hydrogen peroxide decomposes in the presence

of manganese (IV) Oxide

(b) The diagram below shows how a Le'clanche (Dry cell) appears:-



(i) What is the function of MnO_2 in the cell above?

(ii) Write the equation of a reaction that occurs at the cathode

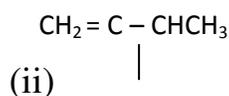
(iii) Calculate the mass of Zinc that is consumed when a current of 0.1 amperes flows

through the above cell for 30 minutes ($1F = 96500c$ $Zn = 65$)

12. (a) Give the IUPAC names of the following compounds:

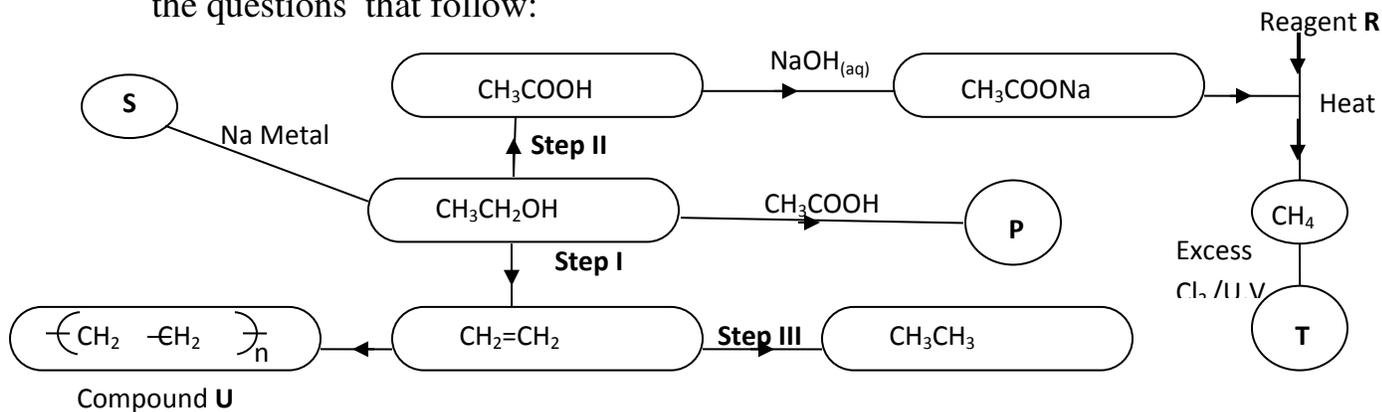


*



(b) The structure below shows some reactions starting with ethanol. Study it and answer

the questions that follow:



(i) Write the formula of the organic compounds **P** and **S**

*

(ii) Name the type of reaction, the reagent(s) and condition for the reactions in the following steps :-

(I) Step I

*

(II) Step II

*

(III) Step III

*

(iii) Name reagent **R**

..... *

(iv) Draw the structural formula of **T** and give its name

*

(v) (I) Name compound

U.....

(II) If the relative molecular mass of **U** is 42000, determine the value of **n** (**C**=12, **H**=1)

(c) State why C_2H_4 burns with a more smoky flame than C_2H_6

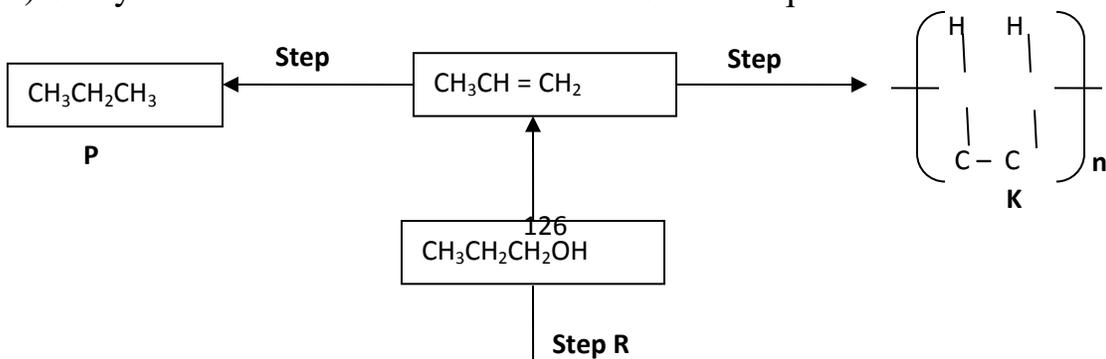
*

13. a) State **two** factors that affect the properties of a polymer

b) Name the compound with the formula below :



c) Study the scheme below and use it to answer the questions that follow:-



i) Name the following compounds:-

I. Product **T** II. **K**

ii) State **one** common physical property of substance **G**

iii) State the type of reaction that occurred in step **J**

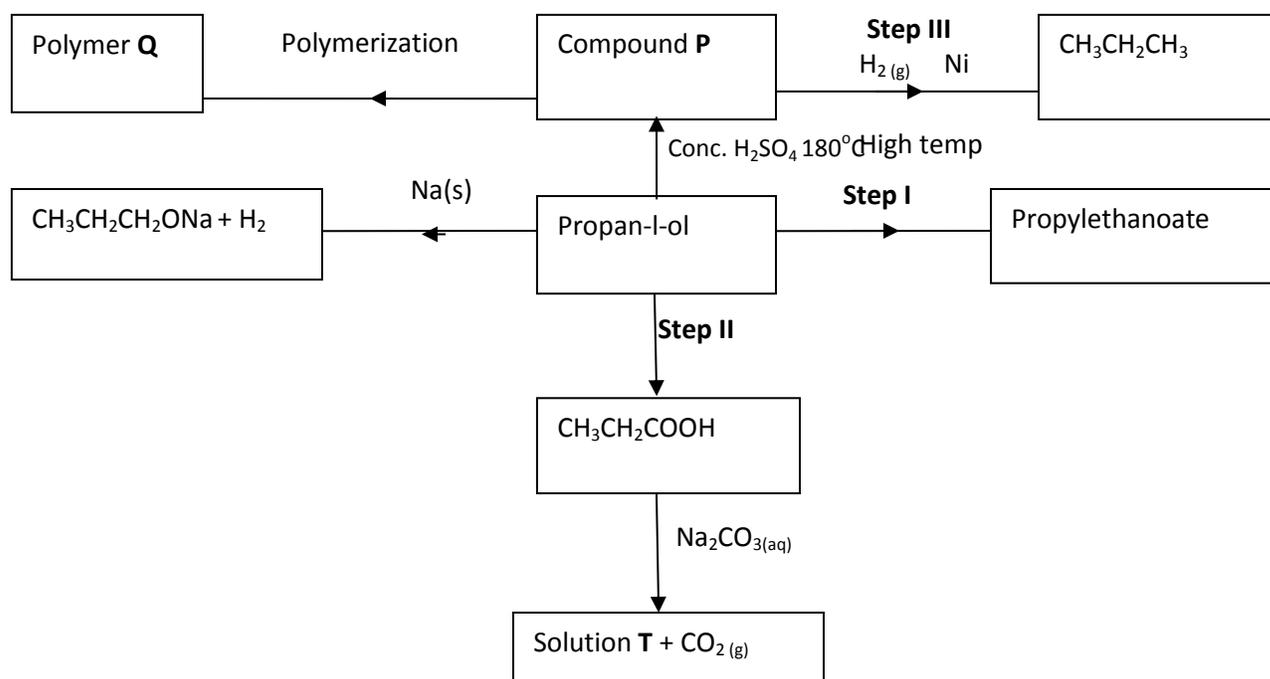
iv) Give **one** use of substance **K**

v) Write an equation for the combustion of compound **P**

vi) Explain how compounds $\text{CH}_3\text{CH}_2\text{COOH}$ and $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ can be distinguished chemically

vii) If a polymer **K** has relative molecular mass of 12,600, calculate the value of **n** (H=1 C =12)

14. Study the scheme given below and answer the questions that follow:-



(a) (i) Name compound **P**

.....

(ii) Write an equation for the reaction between $\text{CH}_3\text{CH}_2\text{COOH}$ and Na_2CO_3

(b) State **one** use of polymer **Q**

(c) Name **one** oxidising agent that can be used in **step II**

.....

(d) A sample of polymer **Q** is found to have a molecular mass of 4200. Determine the number of

monomers in the polymer (H = 1, C = 12)

(e) Name the type of reaction in **step I**

.....

(f) State **one** industrial application of **step III**

(g) State how burning can be used to distinguish between propane and propyne. Explain your

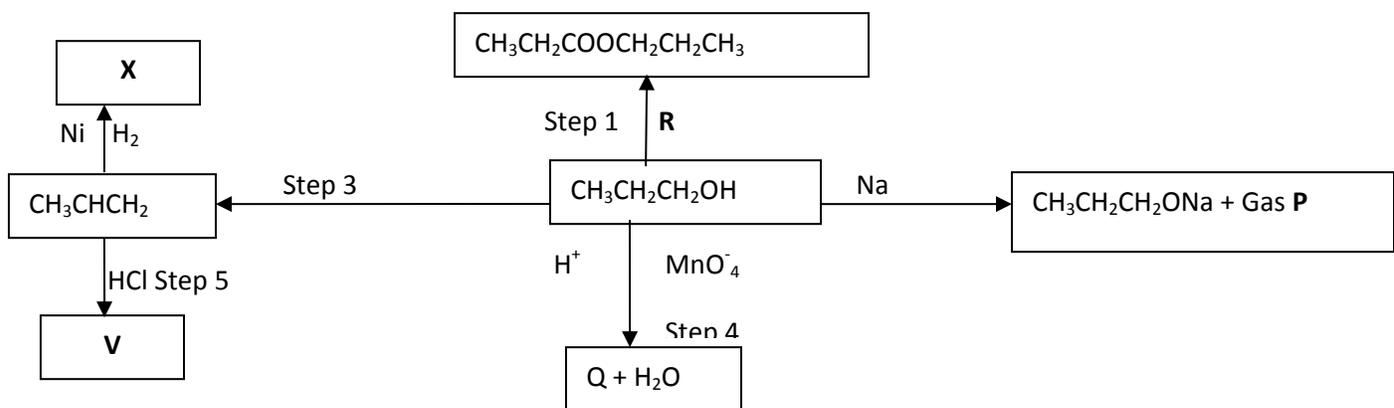
answer

(h) 1000cm^3 of ethene (C_2H_4) burnt in oxygen to produce Carbon (II) Oxide and water vapour.

Calculate the minimum volume of air needed for the complete combustion of ethene

(Air contains 20% by volume of oxygen)

15. (a) Study the schematic diagram below and answer the questions that follow:-



(i) Identify the following:

Substance **Q**

.....

Substance

R.....

Gas

P.....

(ii) Name:

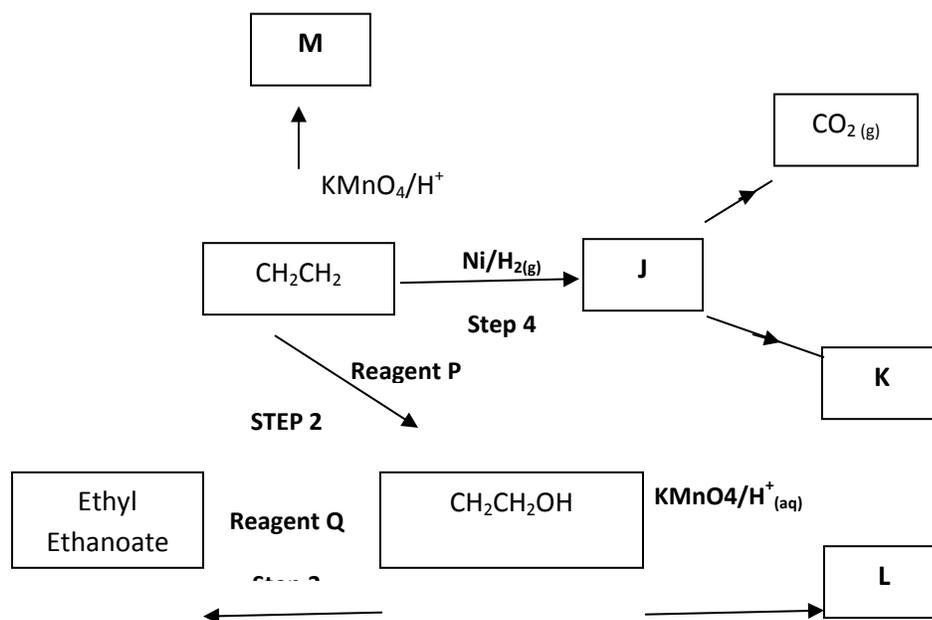
Step 1.....

Step 4.....

(iii) Draw the structural formula of the major product of step 5

(iv) State the condition and reagent in step 3

16. Study the flow chart below and answer the questions that follow



(a) (i) Name the following organic compounds:

M.....

L.....

(ii) Name the process in step:

Step 2

.....

Step 4

(iii) Identify the reagent **P** and **Q**

(iv) Write an equation for the reaction between $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ and sodium

17. a) Give the names of the following compounds:

i) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$

.....

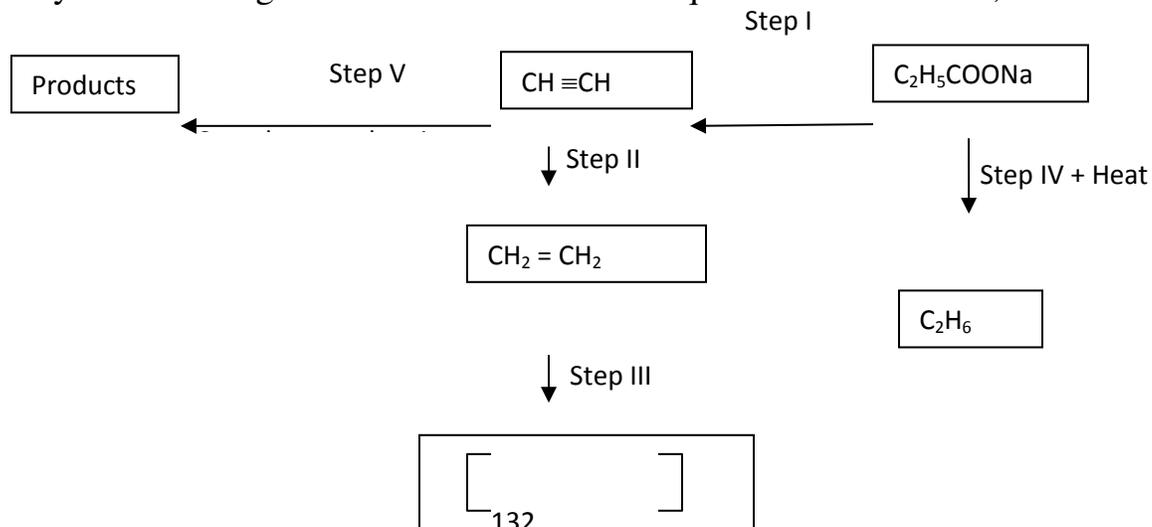
ii) $\text{CH}_3\text{CH}_2\text{COOH}$

.....

iii) $\text{CH}_3\text{C} - \text{O} - \text{CH}_2\text{CH}_3$

.....

18. Study the scheme given below and answer the questions that follow;



— $\frac{\quad}{n}$

i) Name the reagents used in:

Step I:

.....

Step II

.....

Step III

.....

ii) Write an equation to show products formed for the complete combustion of $\text{CH} = \text{CH}$

iii) Explain **one** disadvantage of continued use of items made from the compound formed

in step III

19. A hydrated salt has the following composition by mass. Iron 20.2 %, oxygen 23.0%,

sulphur 11.5%, water 45.3%

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

ii) 6.95g of the hydrated salt in **c(i)** above were dissolved in distilled water and the total

volume made to 250cm^3 of solution. Calculate the concentration of the resulting salt solution

in moles per litre. (Given that the molecular mass of the salt is 278)

20. Write an equation to show products formed for the complete combustion of $\text{CH} = \text{CH}$

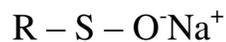
iii) Explain **one** disadvantage of continued use of items made from the compound formed

in step III

21. Give the IUPAC name for each of the following organic compounds;



22. The structure below represents a cleansing agent.



a) State the type of cleansing agent represented above

b) State **one** advantage and one disadvantage of using the above cleansing agent.

23. The structure below shows part of polymer .Use it to answer the questions that follow.

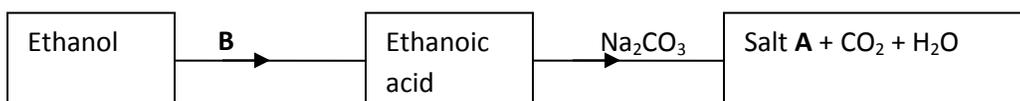




a) Derive the structure of the monomer

b) Name the type of polymerization represented above

24. The flow chart below represents a series of reactions starting with ethanoic acid:-



(a) Identify substances **A** and **B**

(b) Name the process **I**

25. a) Write an equation showing how ammonium nitrate may be prepared starting with

ammonia gas

(b) Calculate the maximum mass of ammonium nitrate that can be prepared using 5.3kg of

ammonia (H=1, N=14, O=16)

26. (a) What is meant by the term, esterification?

(b) Draw the structural formulae of **two** compounds that may be reacted to form ethylpropanoate

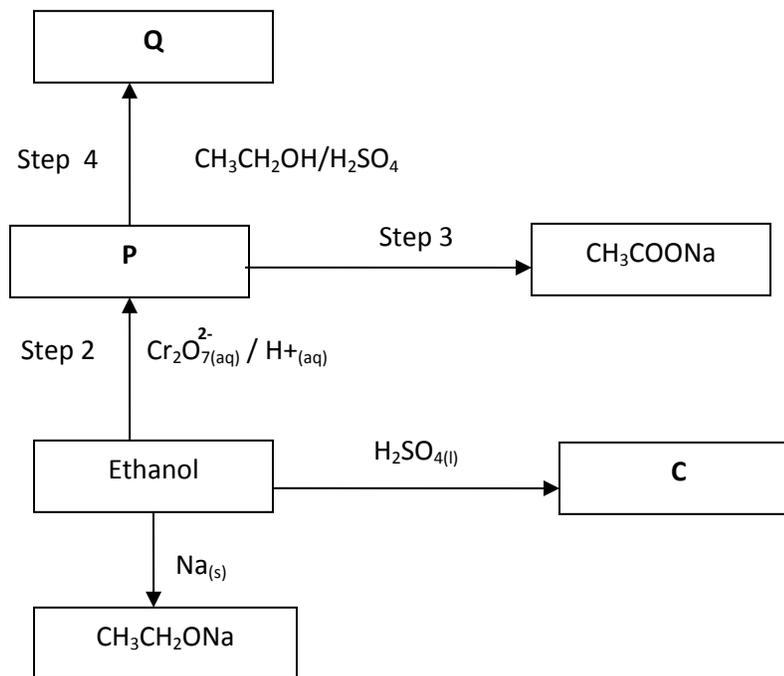
27. (a) Draw the structure of pentanoic acid

(b) Draw the structure and give the name of the organic compound formed when ethanol

reacts with pentanoic acid in presence of concentrated sulphuric acid

28. The scheme below shows some reactions starting with ethanol. Study it and answer the questions

that follow:-



(i) Name and draw the structure of substance **Q**

(ii) Give the names of the reactions that take place in **steps 2** and **4**

(iii) What reagent is necessary for reaction that takes place in step 3

29. Substances **A** and **B** are represented by the formulae **ROH** and **RCOOH** respectively.

They belong to two different homologous series of organic compounds. If both **A** and **B**

react with potassium metal:

(a) Name the common product produced by both

(b) State the observation made when each of the samples **A** and **B** are reacted with sodium

hydrogen carbonate

(i) **A**

(ii) **B**

30. Below are structures of particles. Use it to answer questions that follow. In each case only

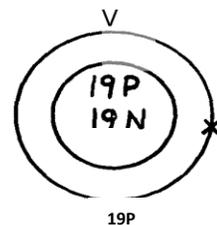
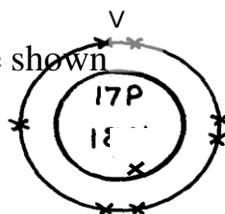
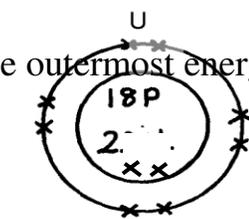
electrons in the outermost energy level are shown

key

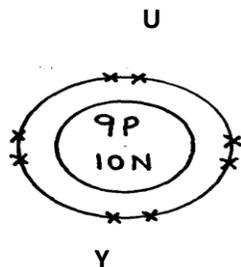
P = Proton

N = Neutron

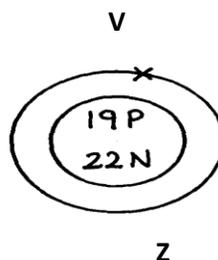
X = Electron



19P



Y



Z

W

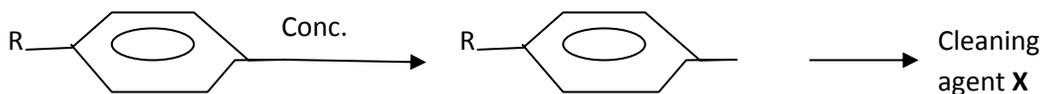
(a) Identify the particle which is an anion

31. Plastics and rubber are extensively used to cover electrical wires.

(a) What term is used to describe plastic and rubbers used in this way?

(b) Explain why plastics and rubbers are used this way

32. The scheme below represents the manufacture of a cleaning agent **X**



(a) Draw the structure of **X** and state the type of cleaning agent to which **X** belong

(b) State **one** disadvantage of using **X** as a cleaning agent

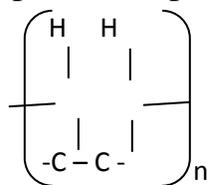
33. **Y** grams of a radioactive isotope take 120days to decay to 3.5grams. The half-life period

of the isotope is 20days

(a) Find the initial mass of the isotope

(b) Give **one** application of radioactivity in agriculture

34. The structure below represents a polymer. Study and answer the questions that follow:-



(i) Name the polymer
above.....

(ii) Determine the value of **n** if giant molecule had relative molecular mass of 4956

35. RCOO^-Na^+ and $\text{RCH}_2\text{OSO}_3^-\text{Na}^+$ are two types of cleansing agents;

i) Name the class of cleansing agents to which each belongs

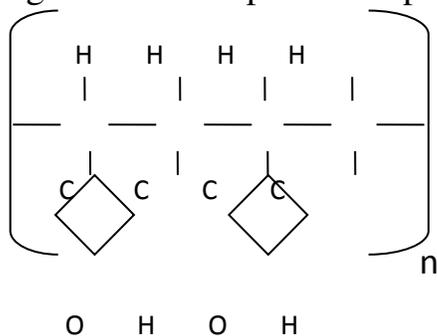
ii) Which one of these agents in (i) above would be more suitable when washing with water

from the Indian ocean. Explain

iii) Both sulphur (IV) oxide and chlorine are used bleaching agents. Explain the difference

in their bleaching properties

36. The formula given below represents a portion of a polymer



(a) Give the name of the polymer

(b) Draw the structure of the monomer used to manufacture the polymer

15.0.0 NITROGEN AND IT'S COMPOUNDS (30 LESSONS)

Introduction to oxides of nitrogen

Nitrogen has a position in second period of group V in the modern periodic table. It has molecular formula N_2 . It has atomic number 7 and atomic weight 14.08 and its electronic configuration of 2,5.

Besides combining with hydrogen and forming NH_3 , nitrogen combines with oxygen in different ratios and forms five different oxides.

The **oxides of nitrogen** and the details of the oxygen states of nitrogen and the N:O ratio can be presented in a tabular form as:

Name	Formula	Oxidation state of N	Ratio of N:O
1) Nitrogen Oxide or nitrous oxide	NO	+2	1:1
2) Nitrogen dioxide	NO ₂	+4	1:2
3) Nitrous oxide (also called laughing gas)	N ₂ O	+1	2:1

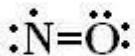
All of these oxides of nitrogen are gases, excepting NO and N₂O the other oxides are brownish gases. Except the oxides of NO and N₂O all the other oxides are acidic. NO and N₂O are neutral.

The other oxides are prepared in the laboratory using different methods characteristic to each oxide.

Example of Oxides of Nitrogen (nitric Oxide - NO)

Structure:

NO



Laboratory Preparation: The oxide is prepared in the laboratory by treating the metallic copper with a moderately concentrated nitric acid (1:1) at room temperature.

The reaction is given as :



The gas is collected by downward displacement of water. The apparatus used is Wolfe's apparatus. The purification is done by absorbing the NO gas in freshly prepared ferrous sulphate solution. Ferrous sulphate absorbs all the NO gas and forms

$\text{Fe}(\text{H}_2\text{O})_5 \text{NO}$ and the solution becomes brown. On heating this solution pure Nitric Oxide is obtained.

Physical Properties: NO is not a combustible gas. At high temperature around 1000°C it decomposes into N_2 and O_2 .



From the equation above we can see that once the decomposition starts 50% O_2 gets evolved and this O_2 supports combustion thus making the reaction more violent.

Chemical properties:

1) NO acts as an **oxidising agent**, oxidising SO_2 in presence of water to give H_2SO_4 :



2) NO acts as a **reducing agent**,

i) reducing an acidified solution of potassium permanganate (pink) to colorless manganous salt.



ii) It can also reduce aqueous solution of I₂ to HI



3) **With halogens** NO can form addition compounds as



It reacts in the same way with flourine and bromine.

4) With ferrous sulphate NO forms an addition compound as



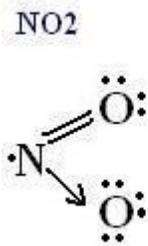
penta aqua nitrosyl iron (II) sulphate

This is the famous **brown ring test** used to identify the nitrate radical or the NO radical.

Uses: NO is used to prepare nitric acid.

Nitrogen Dioxide - NO₂

Structure:



Laboratory Preparation: In the laboratory NO₂ is prepared by thermal decomposition of Pb(NO₃)₂. Thus



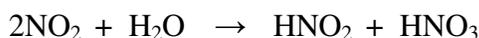
Care is taken to ensure the use of dried Pb(NO₃)₂ as hydrated nitrate salts on heating react violently and explode.

Physical Properties: NO₂ is a poisonous gas, main source being the exhaust of automobiles.

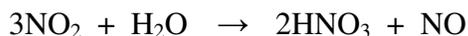
- 1) At room temperature it is a deep brown gas.
- 2) It does not support combustion.
- 3) It is not combustible.

Chemical Properties:

- 1) With cold water NO₂ reacts to give a mixture of HNO₂ and HNO₃ acid.



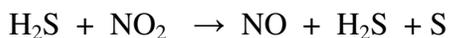
- 2) With hot water the reaction is



- 3) Being acidic it reacts with bases as



- 4) It is also a strong oxidising agent.



- 5) With excess oxygen and water NO₂ gives HNO₃.



- 6) It reacts with concentrated H₂SO₄ to give nitrosyl hydrogen sulphate

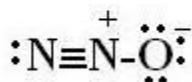


Uses: NO₂ is used as a fuel in rockets besides being used to prepare HNO₃.

Nitrous Oxide - N₂O (laughing Gas)

Structure:

N₂O



Laboratory Preparation: N₂O can be prepared in the laboratory by heating NH₄NO₃ below 200°C to avoid explosion. Sometimes as a safety measure instead of directly using NH₄NO₃, a mixture of (NH₄)₂SO₄ and NaNO₃ are heated to give NH₄NO₃ which decomposes further to give N₂O.

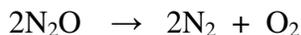


Physical Properties:

- 1) N₂O has a faint sweet smell and produces a tickling sensation on the neck when inhaled and makes people laugh hysterically. Excess of inhalation leads to unconsciousness.
- 2) Unlike other oxides of nitrogen, N₂O supports combustion though it does not burn itself.

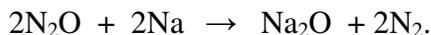
Chemical Properties:

- 1) At very high temperature N₂O decomposes to N₂ and O₂



If a glowing piece of Mg, Cu, or P is introduced in such an environment, these pieces burn brightly due to the O₂ produced from decomposition of N₂O.

- 2) With Sodium and potassium N₂O reacts to give the corresponding peroxides liberating N₂ in the process.



Na₂O is sodium peroxide

Uses:

- 1) It is used as propellant gas.
- 2) Used in combination with oxygen in the ratio N₂O : O₂ = 1:10 as a mild anaesthetic.

16.0.0 SULPHUR AND ITS COMPOUNDS

(25 LESSONS)

A.SULPHUR (S)

Sulphur is an element in Group VI Group 16)of the Periodic table . It has atomic number 16 and electronic configuration 16 and valency 2 /divalent and thus forms the ion S^{2-}

A. Occurrence.

Sulphur mainly occurs :

- (i) as **free** element in Texas and Louisiana in USA and Sicily in Italy.
- (ii)**Hydrogen sulphide** gas in active volcanic areas e.g. Olkaria near Naivasha in Kenya
- (iii)as **copper pyrites**($CuFeS_2$) ,**Galena** (PbS,**Zinc blende**(ZnS))and **iron pyrites**(FeS_2) in other parts of the world.

B. Extraction of Sulphur from Fraschs process

Sulphur occurs about 200 metres underground. The soil structure in these areas is usually **weak** and can easily **cave** in.

Digging of tunnels is thus discouraged in trying to extract the mineral. Sulphur is extracted by drilling three concentric /round pipes of diameter of ratios 2:8: 18 centimeters.

Superheated water at $170^{\circ}C$ and 10atmosphere pressure is forced through the outermost pipe.

The high pressures ensure the water remains as liquid at high temperatures instead of vapour of vapour /gas.

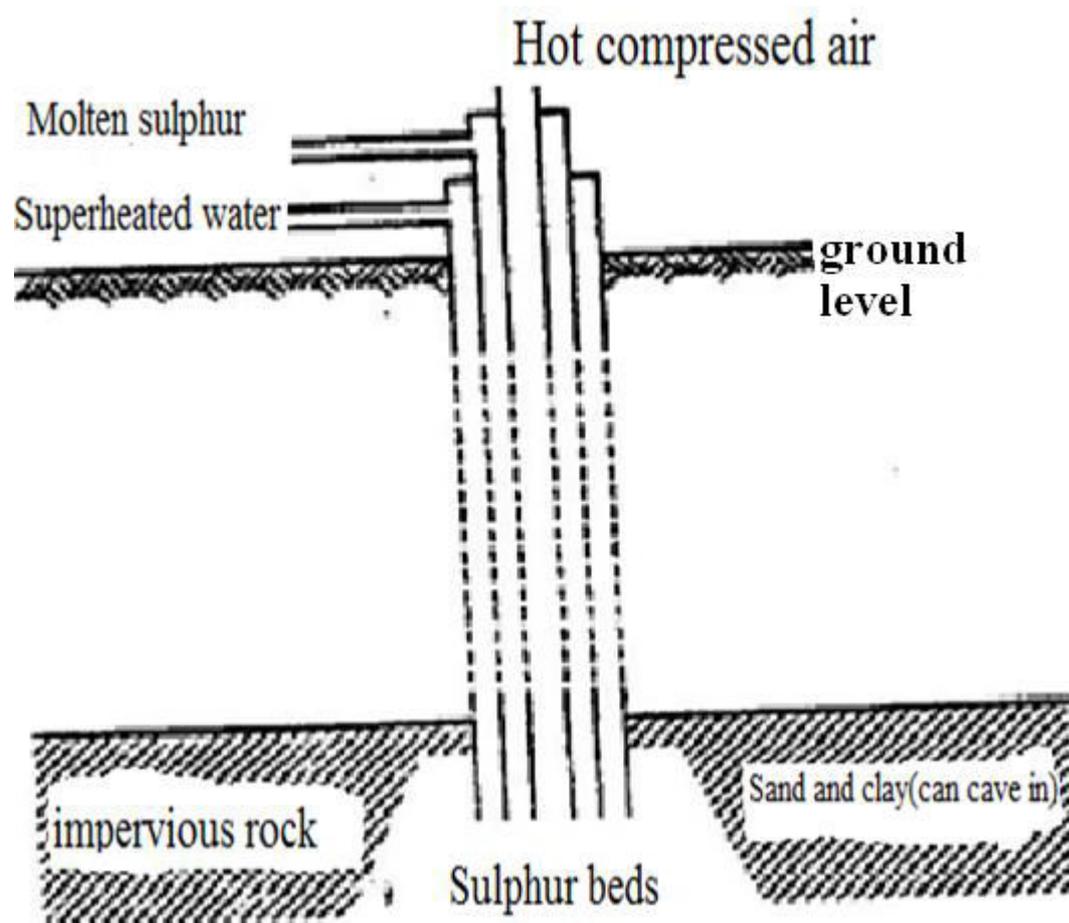
The superheated water melts the sulphur because the melting point of sulphur is lower at about at about $115^{\circ}C$.

A compressed air at 15 atmospheres is forced /pumped through the innermost pipe.

The hot air forces the molten sulphur up the middle pipe where it is collected and solidifies in a large tank.

It is about 99% pure.

Diagram showing extraction of Sulphur from Fraschs Process



C. Allotropes of Sulphur.

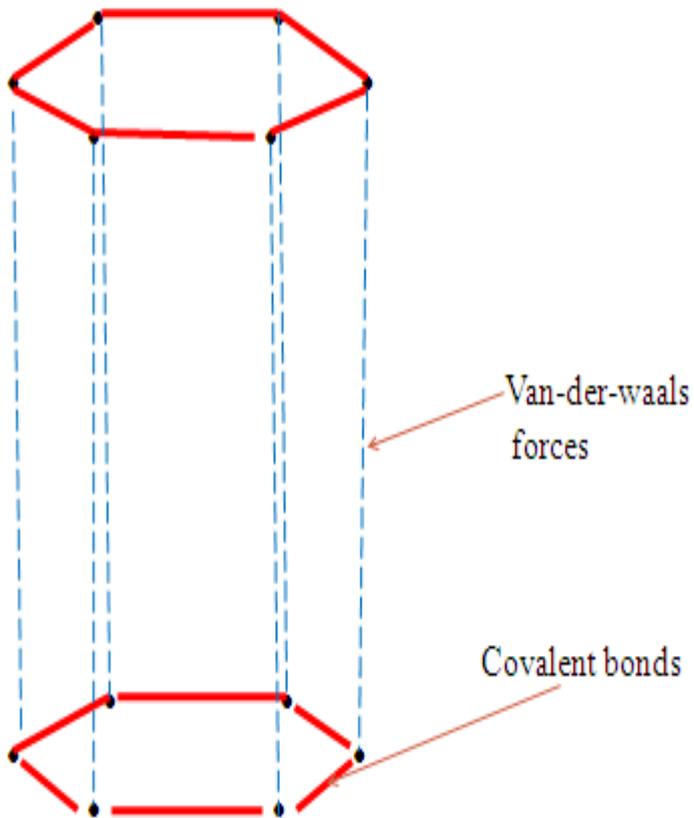
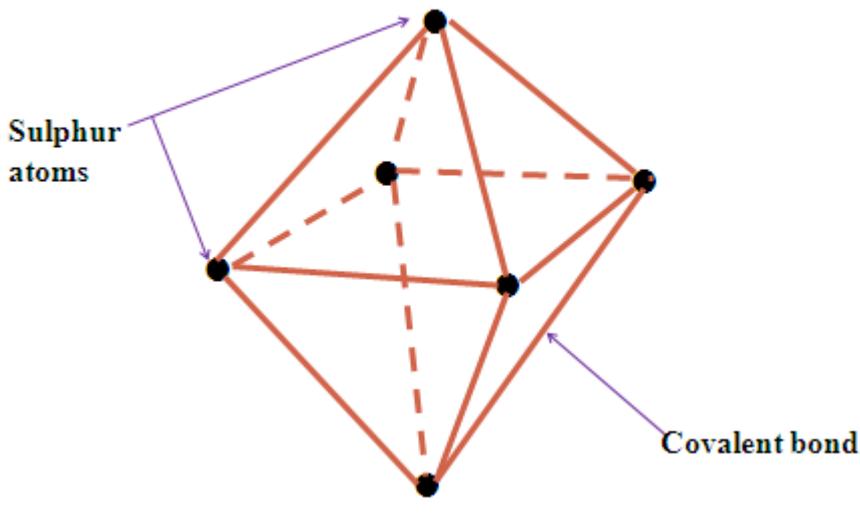
1. Sulphur exist as two crystalline allotropic forms:

- (i) Rhombic sulphur
- (ii) Monoclinic sulphur

Rhombic sulphur	Monoclinic sulphur
Bright yellow crystalline solid	Pale yellow crystalline solid
Has a melting point of 113°C	Has a melting point of 119°C
Has a density of 2.06gcm ⁻³	Has a density of 1.96gcm ⁻³
Stable below 96°C	Stable above 96°C
Has octahedral structure	Has a needle-like structure

Rhombic sulphur and Monoclinic sulphur have a **transition** temperature of 96°C. This is the temperature at which one allotrope changes to the other.

Sketch of Octahedral structure of Rhombic sulphur



Sketch of the needle-like structure of monoclinic sulphur

2. Sulphur exists in non-crystalline forms as:

(i) Plastic sulphur-

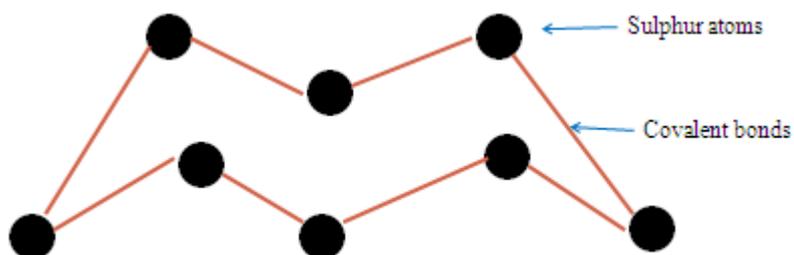
Plastic sulphur is prepared from heating powdered sulphur to boil then pouring a thin continuous stream in a beaker with cold water. A long thin elastic yellow thread of plastic sulphur is formed. If left for long it turns to bright yellow crystalline rhombic sulphur.

(ii) Colloidal sulphur-

Colloidal sulphur is formed when sodium thiosulphate ($\text{Na}_2\text{S}_2\text{O}_3$) is added to hydrochloric acid to form a yellow precipitate.

D. Heating Sulphur.

A molecule of sulphur exists as a puckered ring of eight atoms joined by covalent bonds as S_8 .



On heating the yellow sulphur powder melts at 113°C to a clear amber liquid with low viscosity and thus flows easily.

On further heating to 160°C the molten liquid darkens to a brown very viscous liquid that does not flow easily.

This is because the S_8 rings break into S_8 chains that join together to form very long chains made of over 100,000 atoms of sulphur.

The long chains **entangle** each other, reducing their mobility/flow and hence increasing their viscosity.

On continued further heating to above 160°C , the viscous liquid darkens but becomes more mobile/flows easily and thus less viscous.

This is because the long chains break to smaller/shorter chains.

At 444°C , the liquid boils and forms brown vapour of a mixture of S_8 , S_6 , S_2 molecules that solidifies to S_8 ring of "flowers of sulphur" on the cooler parts.

Summary of changes on heating sulphur

Observation on heating	Explanation/structure of Sulphur
Solid sulphur	Puckered S ₈ ring
Heat to 113°C Amber yellow liquid	Puckered S ₈ ring in liquid form (low viscosity/flow easily)
Heat to 160°C Liquid darkens	Puckered S ₈ ring break/opens then join to form long chains that entangle (very high viscosity/very low rate of flow)
Heat to 444°C Liquid boils to brown vapour	Mixture of S ₈ , S ₆ , S ₂ vapour
Cool to room temperature Yellow sublimate (Flowers of Sulphur)	Puckered S ₈ ring

E. Physical and Chemical properties of Sulphur.(Questions)

1. State three physical properties unique to Sulphur

Sulphur is a yellow solid, insoluble in water, soluble in carbon disulphide/tetrachloromethane/benzene, poor conductor of heat and electricity. It has a melting point of 115°C and a boiling point of 444°C.

2. Moist/damp/wet blue and red litmus papers were put in a gas jar containing air/oxygen. Burning sulphur was then lowered into the gas jar. State and explain the observation made.

Observations

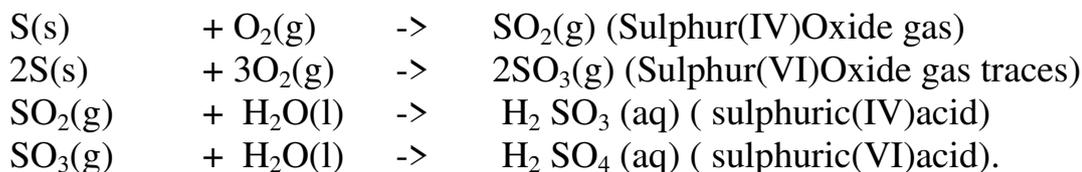
- Sulphur melts then burns with a blue flame
- Colourless gas produced that has a pungent smell
- Red litmus paper remains red. Blue litmus paper turns red.

Explanation

Sulphur burns in air and faster in Oxygen to form Sulphur(IV)Oxide gas and traces/small amount of Sulphur(VI)Oxide gas. Both oxides react with water to form the corresponding acidic solution i.e

- (i) Sulphur(IV)Oxide gas reacts with water to form sulphuric(IV)acid
- (ii) Sulphur(VI)Oxide gas reacts with water to form sulphuric(VI)acid

Chemical equation



3. Iron filings were put in a test tube containing powdered sulphur then heated on a Bunsen flame. Stop heating when reaction starts. State and explain the observations made. Test the effects of a magnet on the mixture before and after heating. Explain.

Observations

Before heating, the magnet attracts iron filings leaving sulphur

After heating, the magnet does not attract the mixture.

After heating, a red glow is observed that continues even when heating is stopped..

Black solid is formed.

Explanation

Iron is attracted to a magnet because it is ferromagnetic.

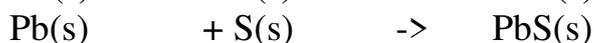
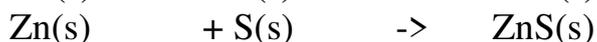
When a mixture of iron and sulphur is heated, the reaction is exothermic giving out heat energy that makes the mixture to continue glowing even after stopping heating.

Black Iron(II)sulphide is formed which is a compound and thus not ferromagnetic.

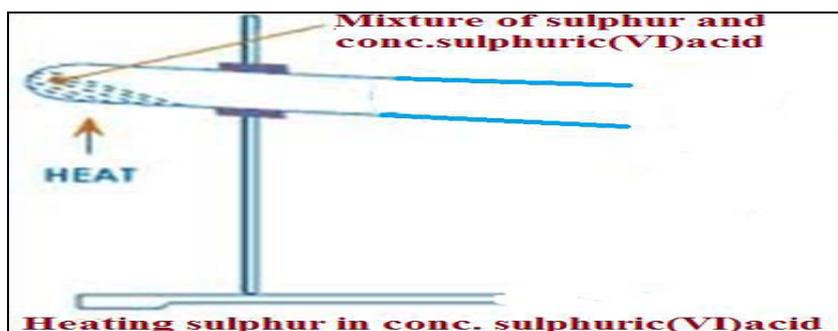
Chemical equation



Heated powdered heavy metals combine with sulphur to form **black** sulphides.



4.The set up below show the reaction of sulphur on heated concentrated sulphuric(VI)acid.



(i) State and explain the observation made.

Observation

Yellow colour of sulphur fades

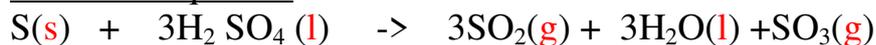
Orange colour of potassium dichromate(VI) paper turns to green.

Explanation

Hot concentrated sulphuric(VI) acid oxidizes sulphur to sulphur (IV) oxide gas.

The oxide is also reduced to water. Traces of sulphur (VI) oxide is formed.

Chemical equation



Sulphur (IV) oxide gas turns Orange potassium dichromate(VI) paper to green.

(ii) State and explain the observation made if concentrated sulphuric (VI) acid is replaced with concentrated Nitric (V) acid in the above set up.

Observation

Yellow colour of sulphur fades

Colourless solution formed

Brown fumes/gas produced.

Explanation

Hot concentrated Nitric(V) acid oxidizes sulphur to sulphuric (VI) acid. The

Nitric (V) acid is reduced to brown nitrogen(IV) oxide gas.

Chemical equation



NB:

Hydrochloric acid is a weaker oxidizing agent and thus cannot oxidize sulphur like the other mineral acids.

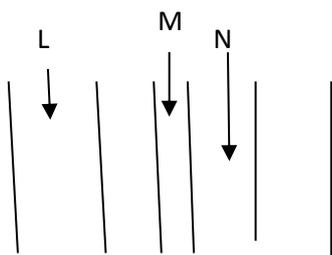
5.State three main uses of sulphur

Sulphur is mainly used in:

- (i) Contact process for the manufacture/industrial/large scale production of concentrated sulphuric(VI) acid.
- (ii) Vulcanization of rubber to make it harder, tougher, stronger, and more durable.
- (iii) Making gun powder and match stick heads
- (iv) As ointments to treat fungal infections

6. Revision Practice

The diagram below represents the extraction of sulphur by Frasch's process. Use it to answer the questions that follow.



(a) Name the substances that pass through:

- M Superheated water at 170°C and 10 atmosphere pressure
- L Hot compressed air
- N Molten sulphur

(b) What is the purpose of the substance that passes through L and M?

M- Superheated water at 170°C and 10 atmosphere pressure is used to melt the sulphur

L- Hot compressed air is used to force up the molten sulphur.

(c) The properties of the two main allotropes of sulphur represented by letters A and B are given in the table below. Use it to answer the questions that follow.

	A	B
Appearance	Bright yellow	Pale yellow
Density(gcm ⁻³)	1.93	2.08
Melting point(°C)	119	113
Stability	Above 96°C	Below 96°C

I. What are allotropes?

Different forms of the same element existing at the same temperature and pressure without change of state.

II. Identify allotrope:

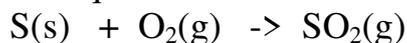
- A. Monoclinic sulphur
- B. Rhombic sulphur

III. State two main uses of sulphur.

- Manufacture of sulphuric(VI) acid
- as fungicide
- in vulcanization of rubber to make it harder/tougher/ stronger
- manufacture of dyes /fibres

(d) Calculate the volume of sulphur (IV)oxide produced when 0.4 g of sulphur is completely burnt in excess air. (S = 32.0 ,1 mole of a gas occupies 24 dm³ at room temperature)

Chemical equation



Mole ratio S: SO₂ = 1:1

Method 1

32.0 g of sulphur -> 24 dm³ of SO₂(g)

$$0.4 \text{ g of sulphur} \rightarrow \frac{0.4 \text{ g} \times 24 \text{ dm}^3}{32.0 \text{ g}} = \underline{\underline{0.3 \text{ dm}^3}}$$

Method 2

$$\text{Moles of sulphur used} = \frac{\text{Mass of sulphur}}{\text{Molar mass of sulphur}} \Rightarrow \frac{0.4}{32} = \underline{\underline{0.0125 \text{ moles}}}$$

Moles of sulphur used = Moles of sulphur(IV)oxide used => 0.0125 moles

Volume of sulphur(IV)oxide used = Moles of sulphur(IV)oxide x volume of one mole of gas => 0.0125 moles x 24 dm³ = 0.3 dm³

B.COMPOUNDS OF SULPHUR

The following are the main compounds of sulphur:

- (i) Sulphur(IV)oxide
- (ii) Sulphur(VI)oxide
- (iii) Sulphuric(VI)acid
- (iv) Hydrogen Sulphide
- (v) Sulphate(IV)/ SO_3^{2-} and Sulphate(VI)/ SO_4^{2-} salts

(i) Sulphur(IV)oxide(SO_2)

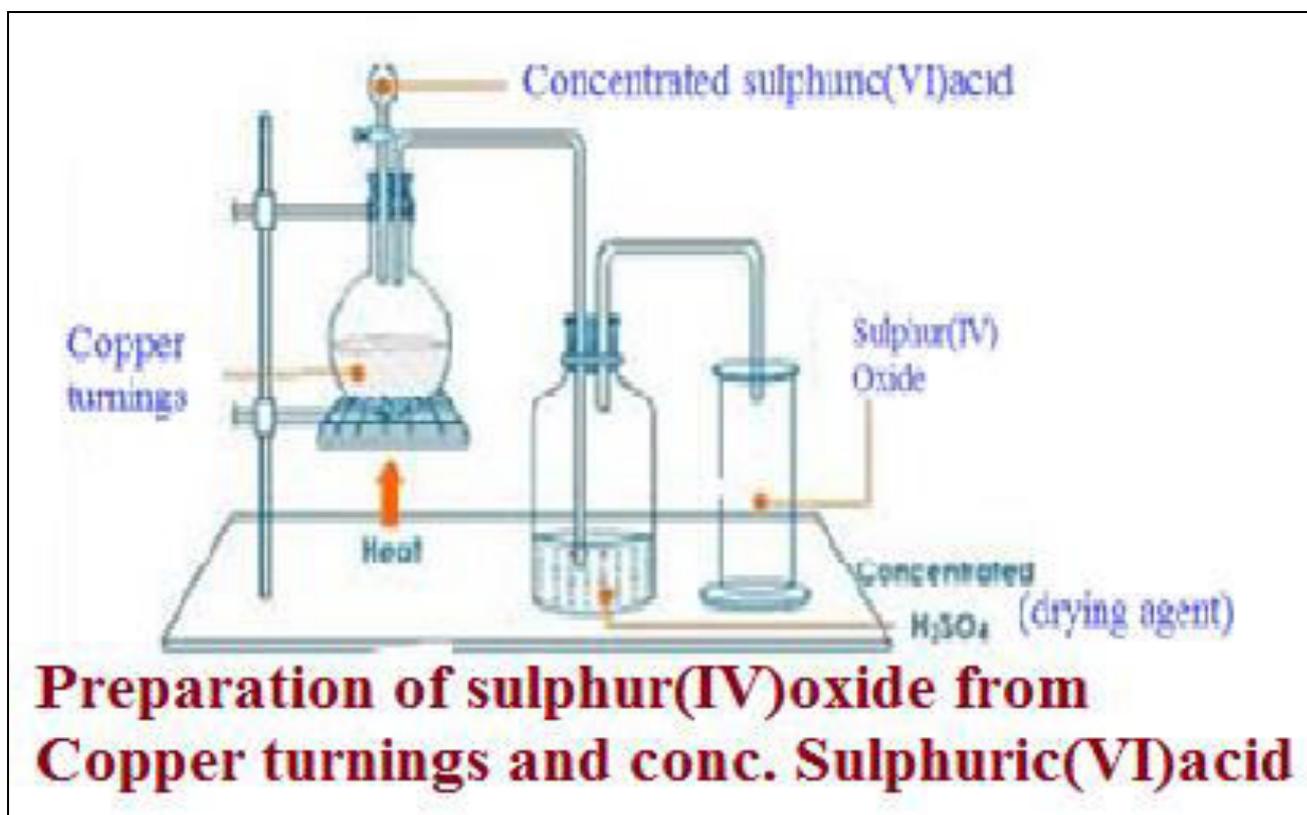
(a) Occurrence

Sulphur (IV)oxide is found in volcanic areas as a gas or dissolved in water from geysers and hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

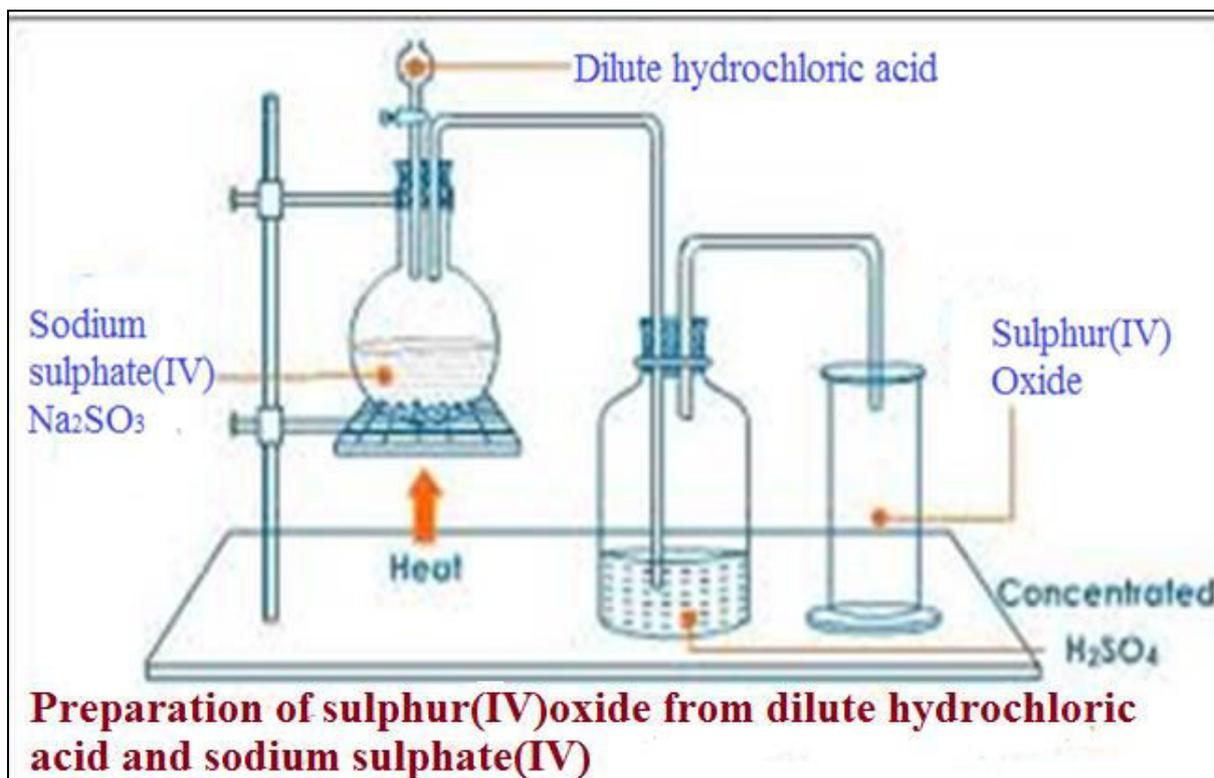
(b) School laboratory preparation

In a Chemistry school laboratory Sulphur (IV)oxide is prepared from the reaction of

Method 1:Using Copper and Sulphuric(VI)acid.



Method 2:Using Sodium Sulphate(IV) and hydrochloric acid.



(c) Properties of Sulphur(IV)oxide(Questions)

1. Write the equations for the reaction for the formation of sulphur (IV)oxide using:

(i) Method 1



Calcium, Lead and Barium will form insoluble sulphate(VI) salts that will cover unreacted metals stopping further reaction thus producing very small amount/quantity of sulphur (IV)oxide gas.

(ii) Method 2

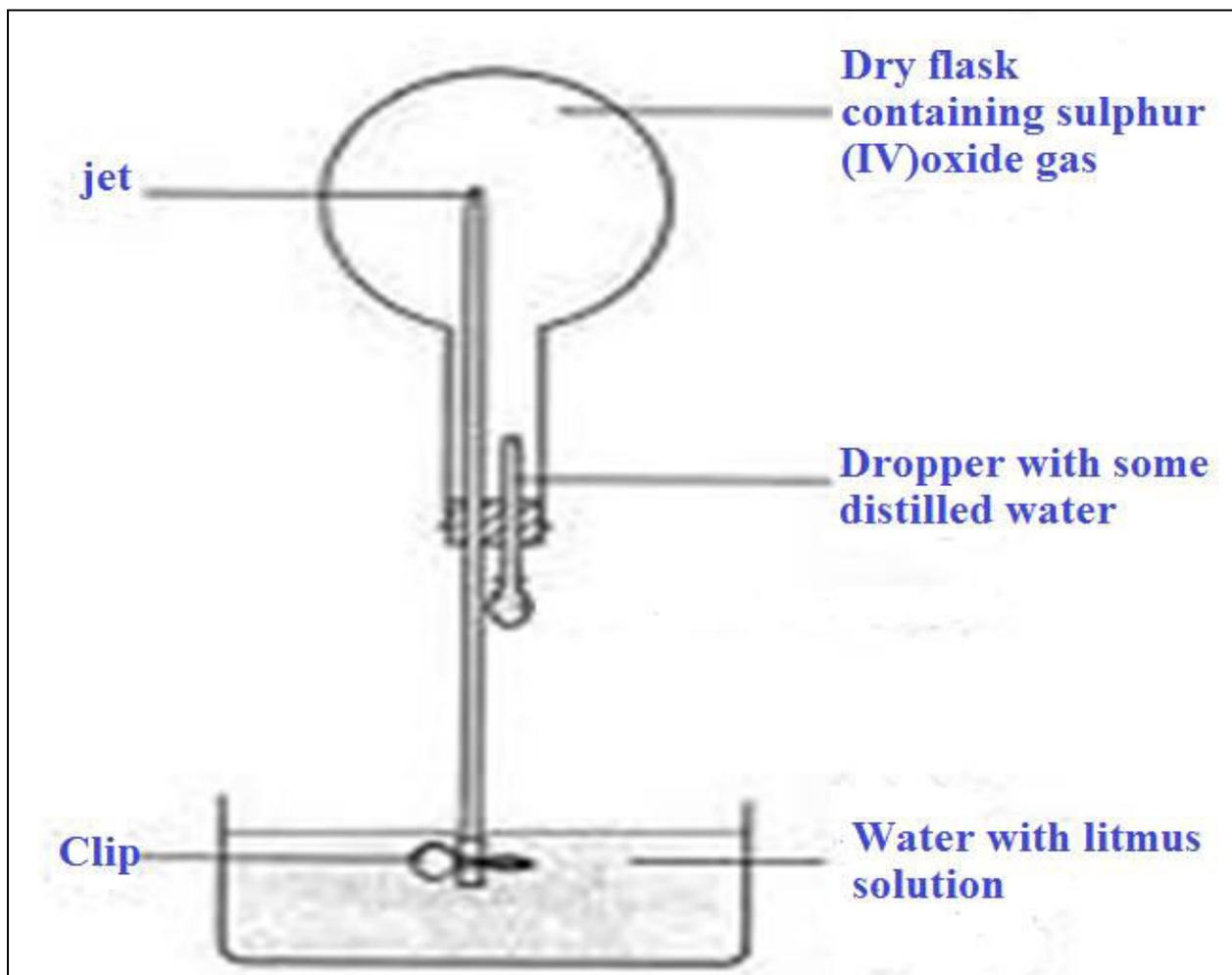


Lead(II)chloride is soluble on heating thus reactants should be heated to prevent it coating/covering unreacted $\text{PbSO}_3(\text{s})$

2.State the physical properties unique to sulphur (IV)oxide gas.

Sulphur (IV)oxide gas is a colourless gas with a pungent irritating and choking smell which **liquidifies** easily. It is about two times denser than air.

3. The diagram below show the solubility of sulphur (IV)oxide gas. Explain.



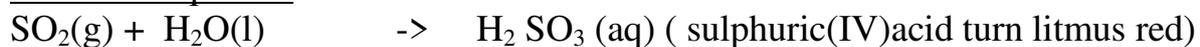
Sulphur(IV) oxide is very soluble in water.

One drop of water dissolves all the Sulphur (IV) oxide in the flask leaving a vacuum.

If the clip is removed, atmospheric pressure forces the water up through the narrow tube to form a fountain to occupy the vacuum.

An acidic solution of sulphuric (IV)acid is formed which turns litmus solution red.

Chemical equation



Sulphur(IV) oxide gas therefore bleaches by reduction /removing oxygen from a dye unlike chlorine that bleaches by oxidation /adding oxygen.

The bleaching by removing oxygen from Sulphur(IV) oxide gas is temporary.

This is because the bleached dye regains the atomic oxygen from the atmosphere/air in presence of sunlight as catalyst thus regaining/restoring its original colour. e.g.

Old newspapers turn brown on exposure to air on regaining the atomic oxygen. The bleaching through adding oxygen by chlorine gas is permanent.

(b)Turns Orange acidified potassium dichromate(VI) to green

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing acidified potassium dichromate(VI) solution. or;

(ii)Dip a filter paper soaked in acidified potassium dichromate(VI) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Orange acidified potassium dichromate(VI) turns to green.

Explanation:

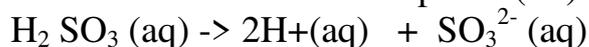
Sulphur(IV) oxide gas reduces acidified potassium dichromate(VI) from orange $\text{Cr}_2\text{O}_7^{2-}$ ions to green Cr^{3+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.

Chemical/ionic equation:

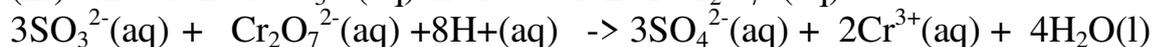
(i)Reaction of Sulphur(IV) oxide gas with water



(ii)Dissociation /ionization of Sulphuric(IV)acid.



(iii)Oxidation of $\text{SO}_3^{2-}(\text{aq})$ and reduction of $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$



This is a **confirmatory** test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also reduces acidified potassium dichromate(VI) from orange $\text{Cr}_2\text{O}_7^{2-}$ ions to green Cr^{3+} ions leaving a yellow residue.

(c)Decolorizes acidified potassium manganate(VII)

Experiment:

(i) Pass a stream of Sulphur(IV) oxide gas in a test tube containing acidified potassium manganate(VII) solution. or;

(ii) Dip a filter paper soaked in acidified potassium manganate(VII) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Purple acidified potassium manganate(VII) turns to colourless/ acidified potassium manganate(VII) is decolorized.

Explanation:

Sulphur(IV) oxide gas reduces acidified potassium manganate(VII) from purple MnO_4^- ions to green Mn^{2+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i) Reaction of Sulphur(IV) oxide gas with water
$$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$$

(ii) Dissociation /ionization of Sulphuric(IV) acid.
$$\text{H}_2\text{SO}_3(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$$

(iii) Oxidation of $\text{SO}_3^{2-}(\text{aq})$ and reduction of $\text{MnO}_4^-(\text{aq})$
$$5\text{SO}_3^{2-}(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 5\text{SO}_4^{2-}(\text{aq}) + 2\text{Mn}^{2+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$$

(purple) (colourless)

This is another test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also decolorizes acidified potassium manganate(VII) from purple MnO_4^- ions to colourless Mn^{2+} ions leaving a yellow residue.

(d) Decolorizes bromine water

Experiment:

(i) Pass a stream of Sulphur(IV) oxide gas in a test tube containing bromine water . or;

(ii) Put three drops of bromine water into a gas jar containing Sulphur(IV) oxide gas. Swirl.

Observation:

Yellow bromine water turns to colourless/ bromine water is decolorized.

Explanation:

Sulphur(IV) oxide gas reduces yellow bromine water to colourless hydrobromic acid (HBr) without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric (IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i) Reaction of Sulphur(IV) oxide gas with water
 $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$

(ii) Dissociation /ionization of Sulphuric(IV) acid.
 $\text{H}_2\text{SO}_3(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$

(iii) Oxidation of $\text{SO}_3^{2-}(\text{aq})$ and reduction of $\text{MnO}_4^-(\text{aq})$
 $\text{SO}_3^{2-}(\text{aq}) + \text{Br}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{HBr}(\text{aq})$
(yellow) (colourless)

This can also be used as another test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also decolorizes yellow bromine water to colourless leaving a yellow residue.

(e) Reduces Iron(III) Fe^{3+} salts to Iron(II) salts Fe^{2+}

Experiment:

(i) Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm³ of Iron (III) chloride solution. or;

(ii) Place about 3 cm³ of Iron (III) chloride solution into a gas jar containing Sulphur(IV) oxide gas. Swirl.

Observation:

Yellow/brown Iron (III) chloride solution turns to green

Explanation:

Sulphur(IV) oxide gas reduces Iron (III) chloride solution from yellow/brown Fe^{3+} ions to green Fe^{2+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i) Reaction of Sulphur(IV) oxide gas with water
 $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$

(ii) Dissociation /ionization of Sulphuric(IV) acid.
 $\text{H}_2\text{SO}_3(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_3^{2-}(\text{aq})$

(iii) Oxidation of $\text{SO}_3^{2-}(\text{aq})$ and reduction of $\text{Fe}^{3+}(\text{aq})$
 $\text{SO}_3^{2-}(\text{aq}) + 2\text{Fe}^{3+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{Fe}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq})$
(yellow) (green)

(f)Reduces Nitric(V)acid to Nitrogen(IV)oxide gas

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm³ of concentrated nitric(V)acid. or;

(ii)Place about 3cm³ of concentrated nitric(V)acid into a gas jar containing Sulphur(IV) oxide gas. Swirl.

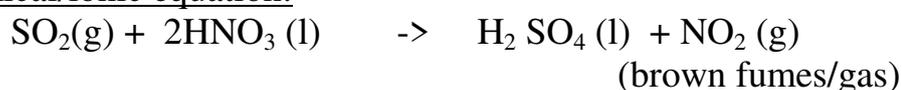
Observation:

Brown fumes of a gas evolved/produced.

Explanation:

Sulphur(IV) oxide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized from SO₃²⁻ ions in sulphuric(IV) acid to SO₄²⁻ ions in sulphuric(VI) acid.

Chemical/ionic equation:



(g)Reduces Hydrogen peroxide to water

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm³ of 20 volume hydrogen peroxide. Add four drops of Barium nitrate(V)or Barium chloride followed by five drops of 2M hydrochloric acid/ 2M nitric(V) acid.

Observation:

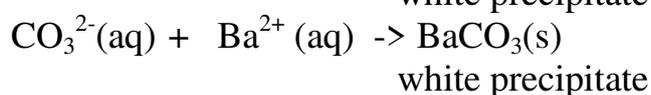
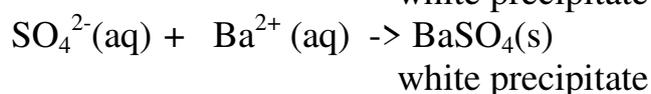
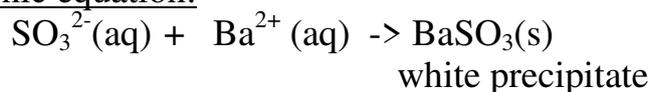
A white precipitate is formed that persist /remains on adding 2M hydrochloric acid/ 2M nitric(V) acid.

Explanation:

Sulphur(IV) oxide gas reduces 20 volume hydrogen peroxide and itself oxidized from SO₃²⁻ ions in sulphuric(IV) acid to SO₄²⁻ ions in sulphuric(VI) acid.

When Ba²⁺ ions in Barium Nitrate(V) or Barium chloride solution is added, a white precipitate of insoluble Barium salts is formed showing the presence of either SO₃²⁻, SO₄²⁻, CO₃²⁻ ions. i.e.

Chemical/ionic equation:

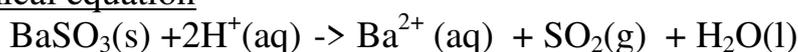


If nitric(V)/hydrochloric acid is added to the three suspected insoluble white precipitates above, the white precipitate:

- (i) persist/remains if $\text{SO}_4^{2-}(\text{aq})$ ions ($\text{BaSO}_4(\text{s})$) is present.
- (ii) dissolves if $\text{SO}_3^{2-}(\text{aq})$ ions ($\text{BaSO}_3(\text{s})$) and $\text{CO}_3^{2-}(\text{aq})$ ions ($\text{BaCO}_3(\text{s})$) is present. This is because:

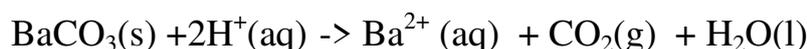
I. $\text{BaSO}_3(\text{s})$ reacts with Nitric(V)/hydrochloric acid to produce acidic SO_2 gas that turns Orange moist filter paper dipped in acidified Potassium dichromate to green.

Chemical equation



I. $\text{BaCO}_3(\text{s})$ reacts with Nitric(V)/hydrochloric acid to produce acidic CO_2 gas that forms a white precipitate when bubbled in lime water.

Chemical equation



5. Sulphur(IV)oxide also act as an oxidizing agent as in the following examples.

(a) Reduction by burning Magnesium

Experiment

Lower a burning Magnesium ribbon into a gas jar containing Sulphur(IV)oxide gas

Observation

Magnesium ribbon continues to burn with difficulty.
White ash and yellow powder/speck

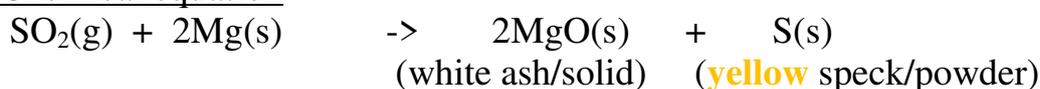
Explanation

Sulphur(IV)oxide does not support burning/combustion. Magnesium burns to produce enough heat energy to decompose Sulphur(IV)oxide to sulphur and oxygen.

The metal continues to burn on Oxygen forming white Magnesium oxide solid/ash.

Yellow specks of sulphur residue form on the sides of reaction flask/gas jar. During the reaction, Sulphur(IV)oxide is reduced(oxidizing agent)while the metal is oxidized (reducing agent)

Chemical equation



(b)Reduction by Hydrogen sulphide gas

Experiment

Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas
Bubble hydrogen sulphide gas into the gas jar containing Sulphur(IV)oxide gas.

Or

Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas
Invert a gas jar full of hydrogen sulphide gas over the gas jar containing Sulphur(IV)oxide gas. Swirl

Observation

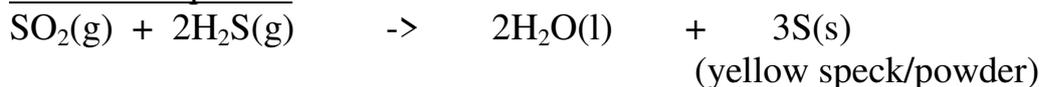
Yellow powder/speck

Explanation

Sulphur(IV)oxide oxidizes hydrogen sulphide to yellow specks of sulphur residue and itself reduced to also sulphur that form on the sides of reaction flask/gas jar.

A little moisture/water act as catalyst /speeds up the reaction.

Chemical equation



6.Sulphur(IV)oxide has many industrial uses. State three.

- (i)In the contact process for the manufacture of Sulphuric(VI)acid
- (ii)As a bleaching agent of pulp and paper.
- (iii)As a fungicide to kill microbes'
- (iv)As a preservative of jam, juices to prevent fermentation

(ii) Sulphur(VI)oxide(SO₃)

(a) Occurrence

Sulphur (VI)oxide is does not occur free in nature/atmosphere

(b) Preparation

In a Chemistry school laboratory Sulphur (VI)oxide may prepared from:

Method 1;Catalytic oxidation of sulphur(IV)oxide gas.

Sulphur(IV)oxide gas and oxygen mixture are first dried by being passed through Concentrated Sulphuric(VI)acid .

The dry mixture is then passed through platinised asbestos to catalyse/speed up the combination to form Sulphur (VI)oxide gas.

Sulphur (VI)oxide gas readily solidify as silky white needles if passed through a freezing mixture /ice cold water.

The solid fumes out on heating to a highly acidic poisonous gas.

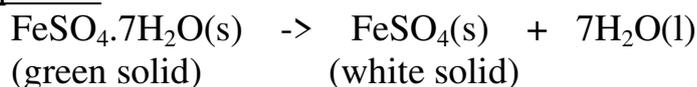
Chemical equation



Method 2; Heating Iron(II)sulphate(VI) heptahydrate

When green hydrated Iron(II)sulphate(VI) heptahydrate crystals are heated in a boiling tube ,it loses the water of crystallization and colour changes from green to white.

Chemical equation

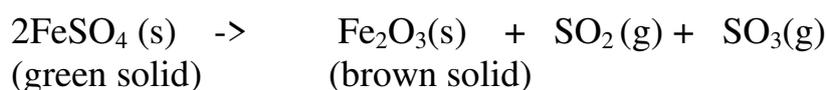


On further heating ,the white anhydrous Iron(II)sulphate(VI) solid decomposes to a mixture of Sulphur (VI)oxide and Sulphur (IV)oxide gas.

Sulphur (VI) oxide readily / easily solidify as white silky needles when the mixture is passed through a freezing mixture/ice cold water.

Iron(III)oxide is left as a brown residue/solid.

Chemical equation



Caution

On exposure to air Sulphur (VI)oxide gas produces highly corrosive poisonous fumes of concentrated sulphuric(VI)acid and thus its preparation in a school laboratory is very risky.

(c) Uses of sulphur(VI)oxide

One of the main uses of sulphur(VI)oxide gas is as an intermediate product in the contact process for industrial/manufacture/large scale/production of sulphuric(VI)acid.

(iii) Sulphuric(VI)acid(H₂SO₄)

(a) Occurrence

Sulphuric (VI)acid(H₂SO₄) is one of the three mineral acids. There are three mineral acids;

- Nitric(V)acid
- Sulphuric(VI)acid
- Hydrochloric acid.

Mineral acids do not occur naturally but are prepared in a school laboratory and manufactured at industrial level.

(b)The Contact process for industrial manufacture of H₂SO₄ .

I. Raw materials

The main raw materials for industrial preparation of Sulphur(VI)acid include:

- (i)**Sulphur** from Fraschs process or from heating metal sulphide ore like Galena(PbS),Zinc blende(ZnS)
- (ii)**Oxygen** from fractional distillation of air
- (iii)**Water** from rivers/lakes

II. Chemical processes

The contact process involves four main chemical processes:

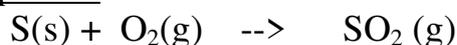
(i)Production of Sulphur (IV)oxide

As one of the raw materials, Sulphur (IV)oxide gas is got from the following sources;

I. Burning/roasting sulphur in air.

Sulphur from Fraschs process is roasted/burnt in air to form Sulphur (IV)oxide gas in the **burners**

Chemical equation



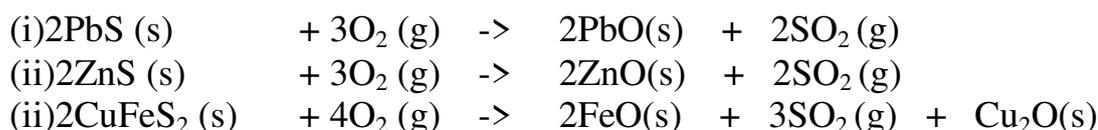
II. Burning/roasting sulphide ores in air.

Sulphur (IV)oxide gas is produced as a by product in extraction of some metals like:

- Lead from Lead(II)sulphide/Galena,(PbS)
- Zinc from zinc(II)sulphide/Zinc blende, (ZnS)
- Copper from Copper iron sulphide/Copper pyrites, (CuFeS₂)

On roasting/burning, large amount /quantity of sulphur(IV)oxide is generated/produced.

Chemical equation



Sulphur(IV)oxide easily/readily liquefies and thus can be transported to a far distance safely.

(ii) Purification of Sulphur(IV)oxide

Sulphur(IV)oxide gas contain dust particles and Arsenic(IV)oxide as impurities. These impurities “poison”/impair the catalyst by adhering on/covering its surface.

The impurities are removed by electrostatic precipitation method .

In the contact process Platinum or Vanadium(V)oxide may be used.

Vanadium(V)oxide is preferred because it is :

- (i) cheaper/less expensive
- (ii) less easily poisoned by impurities

(iii) Catalytic conversion of Sulphur(IV)oxide to Sulphur(VI)oxide

Pure and dry mixture of Sulphur (IV)oxide gas and Oxygen is heated to 450°C in a heat exchanger.

The heated mixture is passed through long pipes coated with pellets of Vanadium (V)oxide catalyst.

The close “contact” between the reacting gases and catalyst give the process its name.

Vanadium (V)oxide catalyse the conversion/oxidation of Sulphur(IV)oxide to Sulphur(VI)oxide gas.

Chemical equation



This reaction is exothermic (-ΔH) and the temperatures need to be maintained at around 450°C to ensure that:

(i) reaction rate/time taken for the formation of Sulphur(VI)oxide is not too **slow/long** at **lower** temperatures below 450°C

(ii) Sulphur(VI)oxide gas does not **decompose** back to Sulphur(IV)oxide gas and Oxygen gas at **higher** temperatures than 450°C.

(iv) Conversion of Sulphur(VI)oxide of Sulphuric(VI)acid

Sulphur(VI)oxide is the acid anhydride of concentrated Sulphuric(VI)acid.

Sulphur(VI)oxide reacts with water to form thick mist of fine droplets of very/highly corrosive concentrated Sulphuric(VI)acid because the reaction is highly exothermic.

To prevent this, Sulphur (VI)oxide is a passed up to meet downward flow of 98% Sulphuric(VI)acid in the absorption chamber/tower.

The reaction forms a very **viscous oily** liquid called **Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid**.

Chemical equation



Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid is diluted carefully with distilled water to give concentrated sulphuric (VI) acid .

Chemical equation



The acid is stored ready for market/sale.

III. Environmental effects of contact process

Sulphur(VI)oxide and Sulphur(IV)oxide gases are atmospheric pollutants that form acid rain if they escape to the atmosphere.

In the Contact process, about 2% of these gases do not form sulphuric (VI) acid.

The following precautions prevent/minimize pollution from Contact process:

(i)recycling back any unreacted Sulphur(IV)oxide gas back to the heat exchangers.

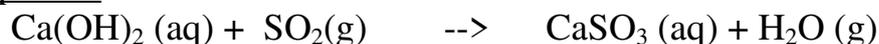
(ii)dissolving Sulphur(VI)oxide gas in concentrated sulphuric (VI) acid instead of water.

This prevents the formation of fine droplets of the corrosive/toxic/poisonous fumes of concentrated sulphuric (VI) acid.

(iii)**scrubbing**-This involves passing the exhaust gases through very tall chimneys lined with quicklime/calcium hydroxide solid.

This reacts with Sulphur (VI)oxide gas forming harmless calcium(II)sulphate (IV) /CaSO₃

Chemical equation



IV. Uses of Sulphur(VI)acid

Sulphuric (VI) acid is used:

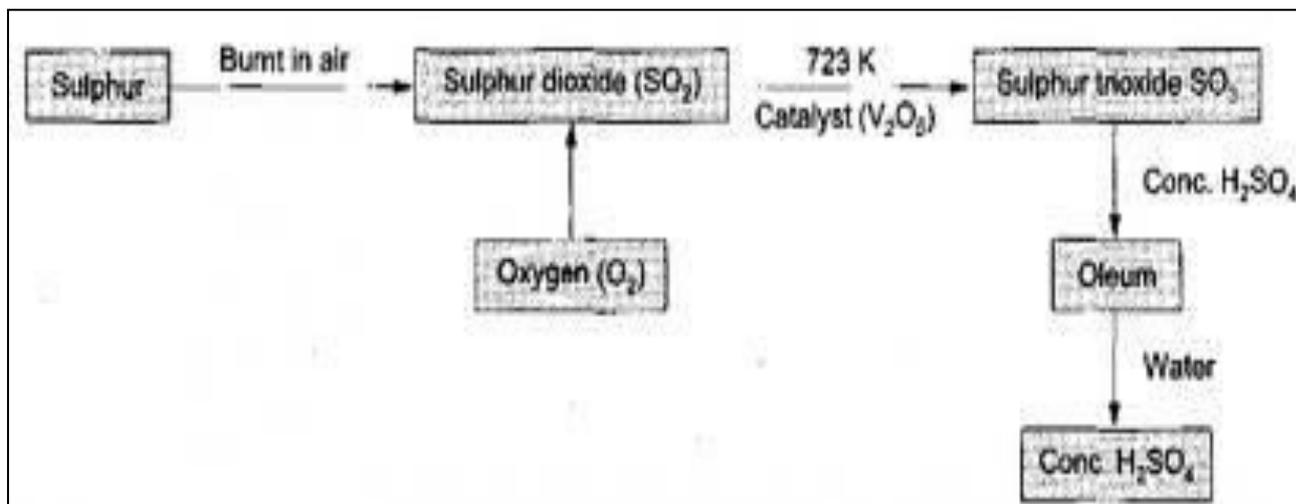
(i) in making dyes and paint

(ii)as acid in Lead-acid accumulator/battery

(iii) for making soapless detergents

(iv) for making sulphate agricultural fertilizers

VI. Sketch chart diagram showing the Contact process



(c) Properties of Concentrated sulphuric(VI)acid

(i) Concentrated sulphuric(VI)acid is a colourless oily liquid with a density of 1.84gcm^{-3} . It has a boiling point of 338°C .

(ii) Concentrated sulphuric(VI)acid is very soluble in water. The solubility /dissolution of the acid very highly exothermic. The concentrated acid should thus be diluted slowly in excess water. Water should never be added to the acid because the hot acid scatters highly corrosive fumes out of the container.

(iii) Concentrated sulphuric (VI)acid is a covalent compound. It has no free H^+ ions. Free H^+ ions are responsible for turning the blue litmus paper red. Concentrated sulphuric (VI) acid thus do not change the blue litmus paper red.

(iv) Concentrated sulphuric (VI)acid is hygroscopic. It absorbs water from the atmosphere and do not form a solution. This makes concentrated sulphuric (VI) acid very suitable as drying agent during preparation of gases.

(v) The following are some **chemical properties** of concentrated sulphuric (VI) acid:

I. As a dehydrating agent

Experiment I;

Put about four spatula end fulls of brown sugar and glucose in separate 10cm^3 beaker.

Carefully add about 10cm^3 of concentrated sulphuric (VI) acid. Allow to stand for about 10 minutes.

Observation;

Colour(in brown sugar)change from brown to **black**.

Colour (in glucose) change from white to **black**.

10cm^3 beaker becomes very hot.

Explanation

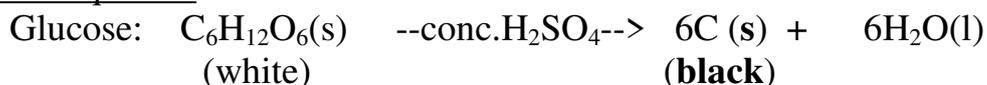
Concentrated sulphuric (VI) acid is strong dehydrating agent. It removes chemically and physically combined elements of water (Hydrogen and Oxygen in ratio 2:1) from compounds. When added to sugar /glucose a vigorous reaction that is highly exothermic takes place.

The sugar/glucose is **charred** to **black mass** of carbon because the acid dehydrates the sugar/glucose leaving carbon.

Caution

This reaction is highly exothermic that starts slowly but produces fine particles of carbon that if inhaled cause quick suffocation by blocking the lung villi.

Chemical equation



Experiment II:

Put about two spatula ends full of hydrated copper(II)sulphate(VI) crystals in a boiling tube. Carefully add about 10cm³ of concentrated sulphuric (VI) acid. Warm.

Observation:

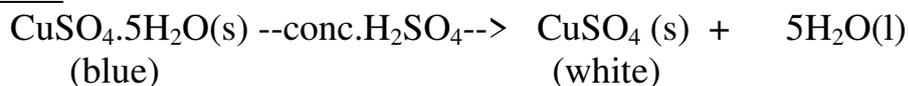
Colour change from blue to white.

Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent. It removes physically combined elements of water (Hydrogen and Oxygen in ratio 2:1) from hydrated compounds.

The acid dehydrates blue copper(II)sulphate to white anhydrous copper(II)sulphate.

Chemical equation



Experiment III:

Put about 4cm³ of absolute ethanol in a boiling tube. Carefully add about 10cm³ of concentrated sulphuric (VI) acid.

Place moist/damp/wet filter paper dipped in acidified potassium dichromate(VI) solution on the mouth of the boiling tube. Heat strongly.

Caution:

Absolute ethanol is highly flammable.

Observation:

Colourless gas produced.

Orange acidified potassium dichromate (VI) paper turns to green.

Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent.

It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1) from compounds.

The acid dehydrates ethanol to ethene gas at about 170°C.

Ethene with =C=C= double bond turns orange acidified potassium dichromate (VI) paper turns to green.

Chemical equation



NB: This reaction is used for the school laboratory preparation of ethene gas

Experiment IV;

Put about 4cm³ of methanoic acid in a boiling tube. Carefully add about 6 cm³ of concentrated sulphuric (VI) acid. Heat gently

Caution:

This should be done in a fume chamber/open

Observation;

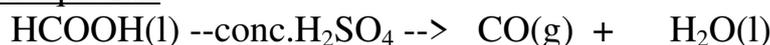
Colourless gas produced.

Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent. It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1) from compounds.

The acid dehydrates methanoic acid to poisonous/toxic carbon(II)oxide gas.

Chemical equation



NB: This reaction is used for the school laboratory preparation of small amount carbon (II)oxide gas

Experiment V;

Put about 4cm³ of ethan-1,2-dioic/oxalic acid in a boiling tube. Carefully add about 6 cm³ of concentrated sulphuric (VI) acid. Pass any gaseous product through lime water. Heat gently

Caution:

This should be done in a fume chamber/open

Observation;

Colourless gas produced.

Gas produced forms a white precipitate with lime water.

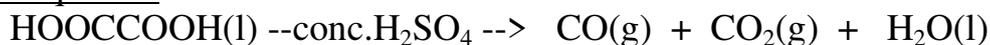
Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent.

It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1) from compounds.

The acid dehydrates ethan-1,2-dioic/oxalic acid to a mixture of poisonous/toxic carbon(II)oxide and carbon(IV)oxide gases.

Chemical equation



NB: This reaction is also used for the school laboratory preparation of small amount carbon (II) oxide gas.

Carbon (IV) oxide gas is removed by passing the mixture through concentrated sodium/potassium hydroxide solution.

II. As an Oxidizing agent

Experiment I

Put about 2cm³ of Concentrated sulphuric (VI) acid into three separate boiling tubes. Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube. Put about 0.5g of Copper turnings, Zinc granule and Iron filings to each boiling tube separately.

Observation:

Effervescence/fizzing/bubbles

Blue solution formed with copper,

Green solution formed with Iron

Colourless solution formed with Zinc

Colourless gas produced that has a pungent irritating choking smell.

Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Explanation

Concentrated sulphuric (VI) acid is strong oxidizing agent.

It oxidizes metals to metallic sulphate(VI) salts and itself reduced to sulphur(IV)oxide gas.

Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

CuSO₄(aq) is a blue solution. ZnSO₄(aq) is a colourless solution. FeSO₄(aq) is a green solution.

Chemical equation



Experiment II

Put about 2cm³ of Concentrated sulphuric (VI) acid into two separate boiling tubes. Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube.

Put about 0.5g of powdered charcoal and sulphur powder to each boiling tube separately.

Warm.

Observation:

Black solid charcoal dissolves/decrease

Yellow solid sulphur dissolves/decrease

Colourless gas produced that has a pungent irritating choking smell.

Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Explanation

Concentrated sulphuric (VI) acid is strong oxidizing agent. It oxidizes non-metals to non metallic oxides and itself reduced to sulphur(IV)oxide gas.

Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Charcoal is oxidized to carbon(IV)oxide. Sulphur is oxidized to Sulphur(IV)oxide .

Chemical equation



III. As the least volatile acid

Study the table below showing a comparison in boiling points of the three mineral acids

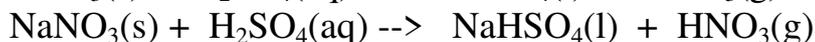
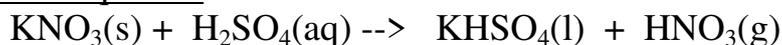
Mineral acid	Relative molecula mass	Boiling point(°C)
Hydrochloric acid(HCl)	36.5	35.0
Nitric(V)acid(HNO ₃)	63.0	83.0
Sulphuric(VI)acid(H ₂ SO ₄)	98.0	333

1. Which is the least volatile acid? Explain

Sulphuric(VI)acid(H₂SO₄) because it has the largest molecule and joined by Hydrogen bonds making it to have the highest boiling point/least volatile.

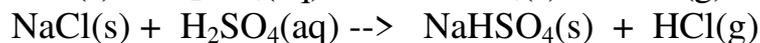
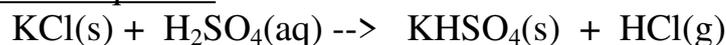
2. Using chemical equations, explain how sulphuric(VI)acid displaces the less volatile mineral acids.

(i)Chemical equation



This reaction is used in the school laboratory preparation of Nitric(V) acid (HNO₃).

(ii)Chemical equation



This reaction is used in the school laboratory **preparation** of Hydrochloric acid (HCl).

(d) Properties of dilute sulphuric(VI)acid.

Dilute sulphuric(VI)acid is made when about 10cm³ of concentrated sulphuric (VI) acid is carefully added to about 90cm³ of distilled water.

Diluting concentrated sulphuric (VI) acid should be done carefully because the reaction is highly exothermic.

Diluting concentrated sulphuric (VI) acid decreases the number of moles present in a given volume of solution which makes the acid less corrosive.

On diluting concentrated sulphuric(VI) acid, water ionizes /dissociates the acid fully/wholly into two(**dibasic**)free H⁺(aq) and SO₄²⁻(aq)ions:



The presence of free H⁺(aq)ions is responsible for ;

(i)turn litmus red because of the presence of free H⁺(aq)ions

(ii)have pH 1/2/3 because of the presence of many free H⁺(aq)ions hence a strongly acidic solution.

(iii)Reaction with metals

Experiment:

Place 5cm³ of 0.2M dilute sulphuric(VI)acid into four separate clean test tubes. Add about 0.1g of Magnesium ribbon to one test tube. Cover the mixture with a finger as stopper. Introduce a burning splint on top of the finger and release the finger “stopper”. Repeat by adding Zinc, Copper and Iron instead of the Magnesium ribbon.

Observation:

No effervescence/ bubbles/ fizzing with copper

Effervescence/ bubbles/ fizzing with Iron ,Zinc and Magnesium

Colourless gas produced that extinguishes burning splint with a “pop” sound.

Colourless solution formed with Zinc and Magnesium.

Green solution formed with Iron

Explanation:

When a metal higher than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid,

effervescence/ bubbling/ fizzing takes place with evolution of Hydrogen gas.

Impure hydrogen gas extinguishes burning splint with a “pop” sound.

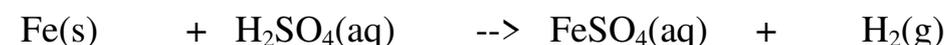
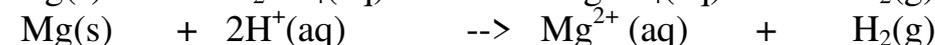
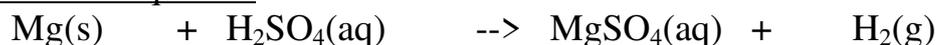
A sulphate (VI) salts is formed. Iron, Zinc and Magnesium are higher than hydrogen in the reactivity/electrochemical series.

They form Iron (II)sulphate(VI), Magnesium sulphate(VI) and Zinc sulphate(VI).

When a metal lower than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid, there is no effervescence/ bubbling/ fizzing that take place.

Copper thus do not react with dilute sulphuric(VI)acid.

Chemical/ionic equation



NB:(i) Calcium,Lead and Barium forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(ii)Sodium and Potassium react explosively with dilute sulphuric(VI)acid

(iv)Reaction with metal carbonates and hydrogen carbonates

Experiment:

Place 5cm³ of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of sodium carbonate to one boiling tube. Introduce a burning splint on top of the boiling tube. Repeat by adding Zinc carbonate, Copper (II)carbonate and Iron(II)Carbonate in place of the sodium hydrogen carbonate.

Observation:

Effervescence/ bubbles/ fizzing.

Colourless gas produced that extinguishes burning splint.

Colourless solution formed with Zinc carbonate, sodium hydrogen carbonate and sodium carbonate.

Green solution formed with Iron(II)Carbonate

Blue solution formed with Copper(II)Carbonate

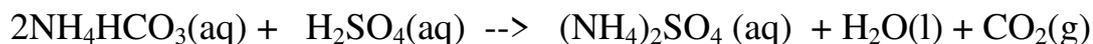
Explanation:

When a metal carbonate or a hydrogen carbonates is put in a test tube containing dilute sulphuric(VI)acid, effervescence/ bubbling/ fizzing takes place with evolution of carbon(IV)oxide gas. carbon(IV)oxide gas

extinguishes a burning splint and forms a white precipitate when bubbled in lime water.

A sulphate (VI) salts is formed.

Chemical/ionic equation



NB:

Calcium, Lead and Barium carbonates forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(v)Neutralization-reaction of metal oxides and alkalis/bases

Experiment I:

Place 5cm³ of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of copper(II)oxide to one boiling tube. Stir.

Repeat by adding Zinc oxide, calcium carbonate and Sodium (II)Oxide in place of the Copper(II)Oxide.

Observation:

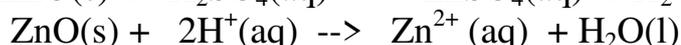
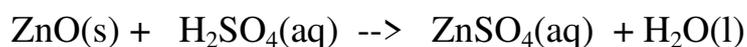
Blue solution formed with Copper(II)Oxide

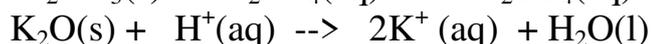
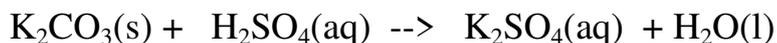
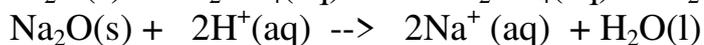
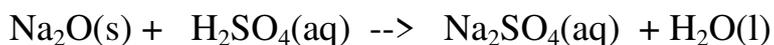
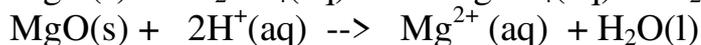
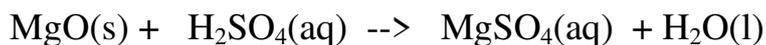
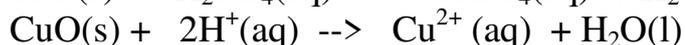
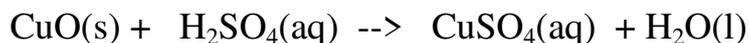
Colourless solution formed with other oxides

Explanation:

When a metal oxide is put in a test tube containing dilute sulphuric(VI)acid, the oxide dissolves forming a sulphate (VI) salt.

Chemical/ionic equation





NB:

Calcium, Lead and Barium oxides forms insoluble sulphate(VI)salts that cover/coat the unreacted metals oxides.

Experiment II:

Fill a burette with 0.1M dilute sulphuric(VI)acid. Pipette 20.0cm³ of 0.1Msodium hydroxide solution into a 250cm³ conical flask. Add three drops of phenolphthalein indicator. Titrate the acid to get a permanent colour change. Repeat with 0.1M potassium hydroxide solution in place of 0.1Msodium hydroxide solution

Observation:

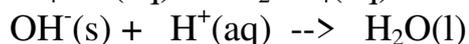
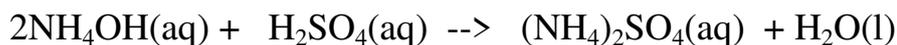
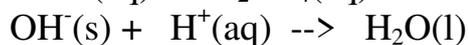
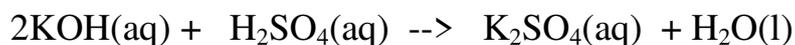
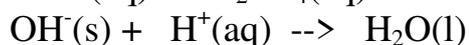
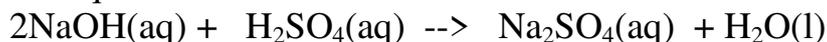
Colour of phenolphthalein changes from pink to colourless at the end point.

Explanation

Like other (mineral) acids dilute sulphuric(VI)acid neutralizes bases/alkalis to a sulphate salt and water only.

Colour of the indicator used changes when a slight excess of acid is added to the base at the end point

Chemical equation:



(iv) Hydrogen sulphide(H_2S)

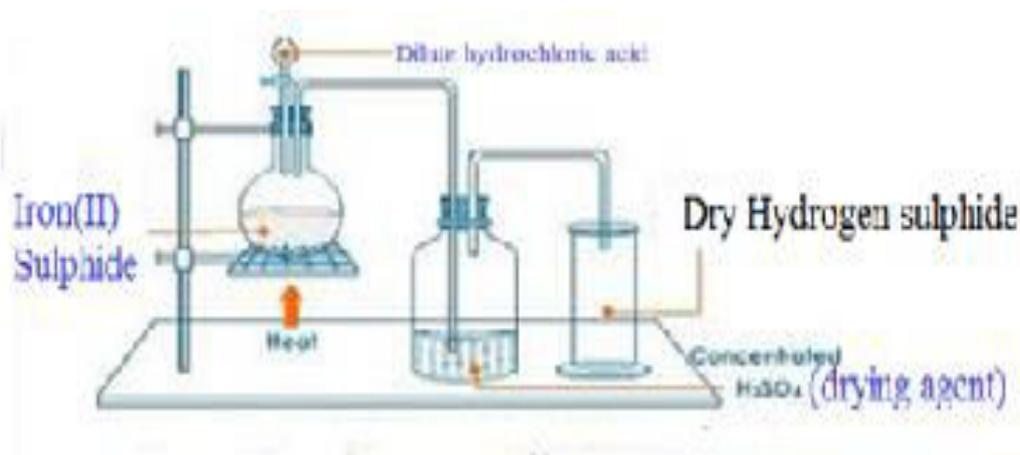
(a) Occurrence

Hydrogen sulphide is found in volcanic areas as a gas or dissolved in water from geysers and hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

It is present in rotten eggs and human excreta.

(b) Preparation

Hydrogen sulphide is prepared in a school laboratory by heating Iron (II) sulphide with dilute hydrochloric acid.



(c) Properties of Hydrogen sulphide(Questions)

1. Write the equation for the reaction for the school laboratory preparation of Hydrogen sulphide.

Chemical equation: $\text{FeS (s)} + 2\text{HCl (aq)} \rightarrow \text{H}_2\text{S (g)} + \text{FeCl}_2 \text{ (aq)}$

2. State three physical properties unique to Hydrogen sulphide.

Hydrogen sulphide is a colourless gas with characteristic pungent poisonous smell of rotten eggs. It is soluble in cold water but insoluble in warm water. It is denser than water and turns blue litmus paper red.

3. Hydrogen sulphide exist as a dibasic acid when dissolved in water. Using a chemical equation show how it ionizes in aqueous state.



Hydrogen sulphide therefore can form both normal and acid salts e.g Sodium hydrogen sulphide and sodium sulphide both exist

4. State and explain one gaseous impurity likely to be present in the gas jar containing hydrogen sulphide above.

Hydrogen/ H_2

Iron(II)sulphide contains Iron as impurity .The iron will react with dilute hydrochloric acid to form iron(II)chloride and produce hydrogen gas that mixes with hydrogen sulphide gas.

5. State and explain the observations made when a filter paper dipped in Lead(II) ethanoate /Lead (II) nitrate(V) solution is put in a gas jar containing hydrogen sulphide gas.

Observations

Moist Lead(II) ethanoate /Lead (II) nitrate(V) paper turns black.

Explanation

When hydrogen sulphide is bubbled in a metallic salt solution, a metallic sulphide is formed.

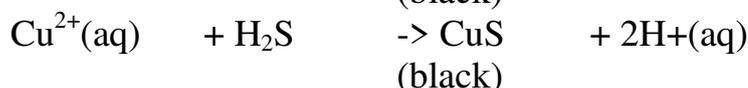
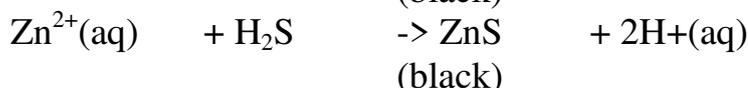
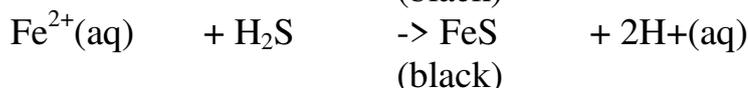
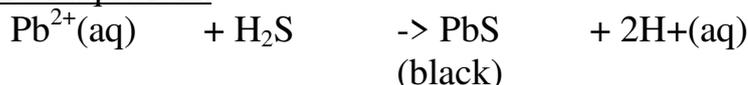
All sulphides are insoluble black salts except sodium sulphide, potassium sulphide and ammonium sulphides.

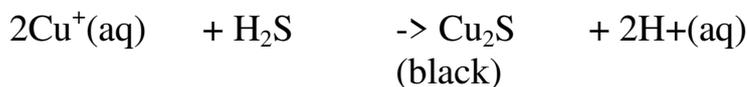
Hydrogen sulphide gas blackens moist Lead (II) ethanoate /Lead (II) nitrate(V) paper .

The gas reacts with Pb^{2+} in the paper to form black Lead(II)sulphide.

This is the chemical test for the presence of H_2S other than the physical smell of rotten eggs.

Chemical equations





6. Dry hydrogen sulphide was ignited as below.



(i) State the observations made in flame A

Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.



Hydrogen sulphide burns in limited air with a blue flame to form sulphur solid and water.



7. Hydrogen sulphide is a strong reducing agent that is oxidized to yellow solid sulphur as precipitate. The following experiments illustrate the reducing properties of Hydrogen sulphide.

(a) Turns Orange acidified potassium dichromate(VI) to green

Experiment:

(i) Pass a stream of Hydrogen sulphide gas in a test tube containing acidified potassium dichromate (VI) solution. or;

(ii) Dip a filter paper soaked in acidified potassium dichromate (VI) into a gas jar containing Hydrogen sulphide gas.

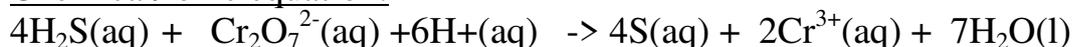
Observation:

Orange acidified potassium dichromate (VI) turns to green.
Yellow solid residue.

Explanation:

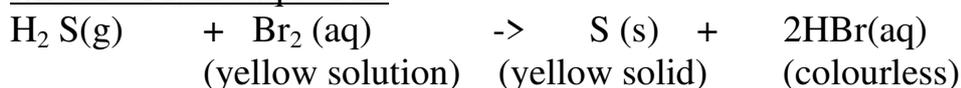
Hydrogen sulphide gas reduces acidified potassium dichromate(VI) from orange $\text{Cr}_2\text{O}_7^{2-}$ ions to green Cr^{3+} ions leaving a yellow solid residue as itself is oxidized to sulphur.

Chemical/ionic equation:



This test is used for differentiating Hydrogen sulphide and sulphur (IV)oxide gas.

Chemical/ionic equation:



This is another test for differentiating Hydrogen sulphide and Sulphur(IV) oxide gas.

Sulphur(IV) oxide also decolorizes acidified potassium manganate(VII) from purple MnO_4^- ions to colourless Mn^{2+} ions leaving no yellow residue.

(d) Reduces Iron(III) Fe^{3+} salts to Iron(II) salts Fe^{2+}

Experiment:

(i) Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm³ of Iron (III)chloride solution. or;

(ii) Place about 3cm³ of Iron (III)chloride solution into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

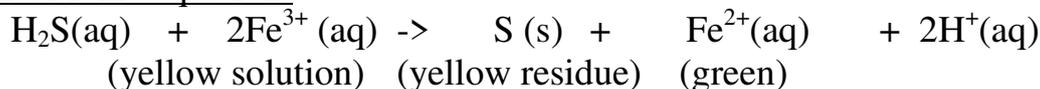
Yellow/brown Iron (III)chloride solution turns to green.

Yellow solid solid

Explanation:

Hydrogen sulphide gas reduces Iron (III)chloride solution from yellow/brown Fe^{3+} ions to green Fe^{2+} ions leaving a yellow residue. The gas is itself oxidized to sulphur.

Chemical/ionic equation:



(e) Reduces Nitric(V)acid to Nitrogen(IV)oxide gas

Experiment:

(i) Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm³ of concentrated nitric(V)acid. or;

(ii) Place about 3cm³ of concentrated nitric(V)acid into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Brown fumes of a gas evolved/produced.

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized to yellow sulphur.

Chemical/ionic equation:



(yellow residue) (brown fumes)

(f) Reduces sulphuric(VI) acid to Sulphur

Experiment:

(i) Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm³ of concentrated sulphuric(VI) acid. or;

(ii) Place about 3 cm³ of concentrated sulphuric (VI) acid into a gas jar containing Hydrogen sulphide gas. Swirl.

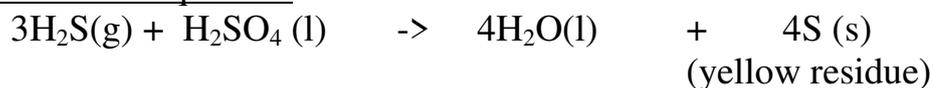
Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces concentrated sulphuric(VI) acid to yellow sulphur.

Chemical/ionic equation:



(g) Reduces Hydrogen peroxide to water

Experiment:

(i) Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm³ of 20 volume hydrogen peroxide.

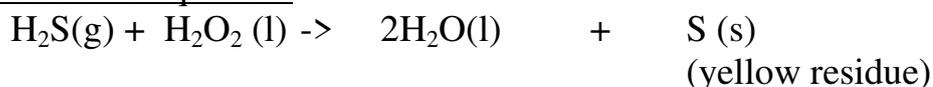
Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces 20 volume hydrogen peroxide to water and itself oxidized to yellow sulphur

Chemical/ionic equation:

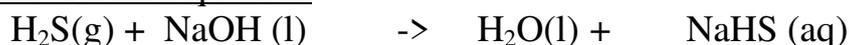


8. Name the salt formed when:

(i) equal volumes of equimolar hydrogen sulphide neutralizes sodium hydroxide solution:

Sodium hydrogen sulphide

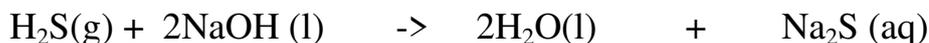
Chemical/ionic equation:



(ii) hydrogen sulphide neutralizes excess concentrated sodium hydroxide solution:

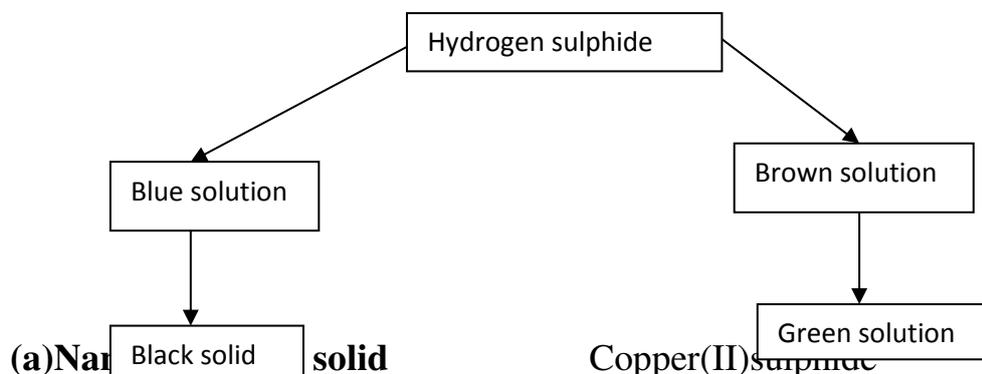
Sodium sulphide

Chemical/ionic equation:



Practice

Hydrogen sulphide gas was bubbled into a solution of metallic nitrate(V)salts as in the flow chart below



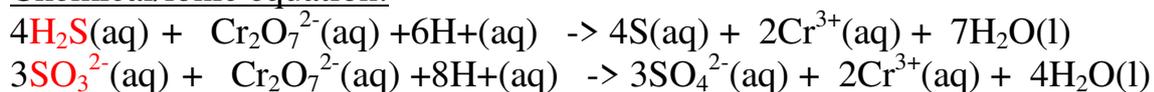
(b)Identify the cation responsible for the formation of:

- | | |
|---------------------|-----------------------------|
| I. Blue solution | $\text{Cu}^{2+}(\text{aq})$ |
| II. Green solution | $\text{Fe}^{2+}(\text{aq})$ |
| III. Brown solution | $\text{Fe}^{3+}(\text{aq})$ |

(c)Using acidified potassium dichromate(VI) describe how you would differentiate between sulphur(IV)Oxide and hydrogen sulphide

- Bubble the gases in separate test tubes containing acidified Potassium dichromate(VI) solution.
- Both changes the Orange colour of acidified Potassium dichromate(VI) solution to green.
- Yellow solid residue/deposit is formed with Hydrogen sulphide

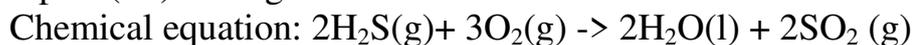
Chemical/ionic equation:



(d)State and explain the observations made if a burning splint is introduced at the mouth of a hydrogen sulphide generator.

Observation Gas continues burning with a blue flame

Explanation: Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.



(v) Sulphate (VI) (SO_4^{2-}) and Sulphate (IV) (SO_3^{2-}) salts

1. Sulphate (VI) (SO_4^{2-}) salts are normal and acid salts derived from Sulphuric (VI) acid H_2SO_4 .
2. Sulphate (IV) (SO_3^{2-}) salts are normal and acid salts derived from Sulphuric (IV) acid H_2SO_3 .
3. Sulphuric (VI) acid H_2SO_4 is formed when sulphur(VI) oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (VI) (SO_4^{2-}) and hydrogen sulphate (VI) (HSO_4^-) salts.

i.e.



All Sulphate (VI) (SO_4^{2-}) salts **dissolve** in water/are soluble except Calcium (II) sulphate (VI) (CaSO_4), Barium (II) sulphate (VI) (BaSO_4) and Lead (II) sulphate (VI) (PbSO_4)

All Hydrogen sulphate (VI) (HSO_3^-) salts **exist** in solution/dissolved in water.

Sodium (I) hydrogen sulphate (VI) (NaHSO_4), Potassium (I) hydrogen sulphate (VI) (KHSO_4) and Ammonium hydrogen sulphate (VI) (NH_4HSO_4) exist also as solids.

Other Hydrogen sulphate (VI) (HSO_4^-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (VI) ($\text{Ca}(\text{HSO}_4)_2$) and Magnesium (II) hydrogen sulphate (VI) ($\text{Mg}(\text{HSO}_4)_2$).

4. Sulphuric (IV) acid H_2SO_3 is formed when sulphur(IV) oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (IV) (SO_3^{2-}) and hydrogen sulphate (VI) (HSO_4^-) salts.

i.e.





All Sulphate (IV) (SO_3^{2-}) salts **dissolve** in water/are soluble except Calcium (II) sulphate (IV) (CaSO_3), Barium (II) sulphate (IV) (BaSO_3) and Lead (II) sulphate (IV) (PbSO_3)

All Hydrogen sulphate (IV) (HSO_3^-) salts **exist** in solution/dissolved in water. Sodium (I) hydrogen sulphate (IV) (NaHSO_3), Potassium (I) hydrogen sulphate (IV) (KHSO_3) and Ammonium hydrogen sulphate (IV) (NH_4HSO_3) exist also as solids.

Other Hydrogen sulphate (IV) (HSO_3^-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (IV) ($\text{Ca}(\text{HSO}_3)_2$) and Magnesium (II) hydrogen sulphate (IV) ($\text{Mg}(\text{HSO}_3)_2$).

5. The following experiments show the effect of heat on sulphate(VI) (SO_4^{2-}) and sulphate(IV) (SO_3^{2-}) salts:

Experiment:

In a clean dry test tube place separately about 1.0g of :

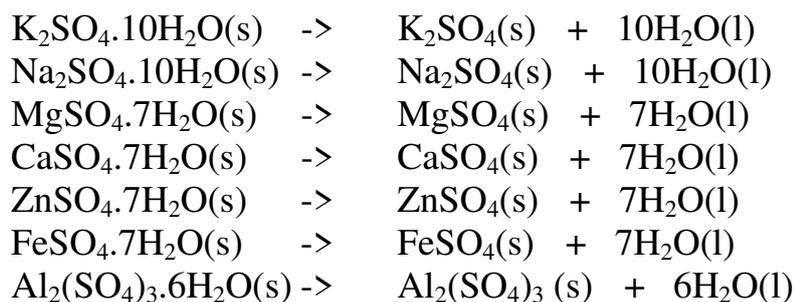
Zinc(II)sulphate (VI), Iron(II)sulphate(VI), Copper(II)sulphate(VI), Sodium (I) sulphate (VI), Sodium (I) sulphate (IV). Heat gently then strongly. Test any gases produced using litmus papers.

Observations:

- Colourless droplets of liquid forms on the cooler parts of the test tube in all cases.
- White solid residue is left in case of Zinc (II)sulphate(VI), Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV).
- Colour changes from green to brown /yellow in case of Iron (II)sulphate(VI)
- Colour changes from blue to white then black in case of Copper (II) sulphate (VI)
- Blue litmus paper remain blue and red litmus paper remain red in case of Zinc(II)sulphate(VI), Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV)
- Blue litmus paper turns red and red litmus paper remain red in case of Iron (II)sulphate(VI) and Copper (II) sulphate (VI).

Explanation

(i) All Sulphate (VI) (SO_4^{2-}) salts exist as **hydrated** salts with water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid on gentle heating. e.g.

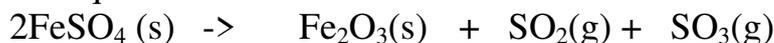




All Sulphate (VI) (SO_4^{2-}) salts do not decompose on heating **except** Iron (II) sulphate (VI) and Copper (II) sulphate (VI).

(i) Iron (II) sulphate (VI) decomposes on strong heating to produce acidic sulphur (IV) oxide and sulphur(VI) oxide gases. Iron(III) oxide is formed as a brown /yellow residue.

Chemical equation



This reaction is used for the school laboratory preparation of small amount of sulphur(VI) oxide gas.

Sulphur (VI) oxide readily /easily solidifies as white silky needles when the mixture is passed through freezing mixture/ice cold water.

Sulphur (IV) oxide does not.

(ii) Copper(II) sulphate(VI) decomposes on strong heating to black copper (II) oxide and Sulphur (VI) oxide gas.

Chemical equation



This reaction is used for the school laboratory preparation of small amount of sulphur(VI) oxide gas.

6. The following experiments show the test for the presence of sulphate (VI) (SO_4^{2-}) and sulphate(IV) (SO_3^{2-}) ions in a sample of a salt/compound:

Experiments/Observations:

(a) Using Lead(II) nitrate(V)

I. To about 5cm³ of a salt solution in a test tube add four drops of Lead(II) nitrate(V) solution. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} , Cl^- ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , Cl^- ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.(a) To the preserved sample observation 1 in (II) above, Heat to boil.

Observation 1

Observation	Inference
White precipitate/ppt persists on boiling	SO_4^{2-} ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves on boiling	Cl^- ions

(b) To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i) acidified potassium manganate(VII) decolorized (ii) Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

Observation	Inference
(i) acidified potassium manganate(VII) not decolorized (ii) Orange colour of acidified potassium dichromate(VI) does not turn to green	CO_3^{2-} ions

Experiments/Observations:

(b) Using Barium(II)nitrate(V) / Barium(II)chloride

I. To about 5cm³ of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III. To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i) acidified potassium manganate(VII) decolorized (ii) Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

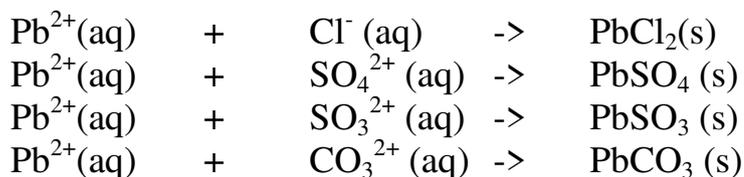
Observation	Inference
(i) acidified potassium manganate(VII) not decolorized (ii) Orange colour of acidified potassium dichromate(VI) does not turn to green	CO_3^{2-} ions

Explanations

Using Lead(II)nitrate(V)

(i) Lead(II)nitrate(V) solution reacts with chlorides(Cl^-), Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV) salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

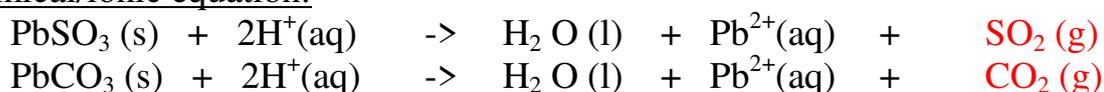


(ii) When the insoluble precipitates are acidified with nitric(V) acid,

- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.

- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

Chemical/ionic equation:



(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;

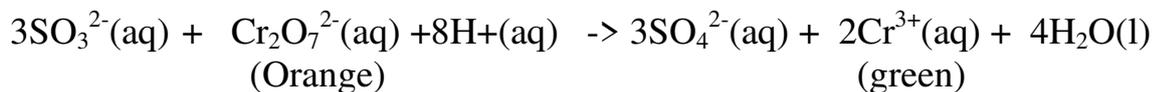
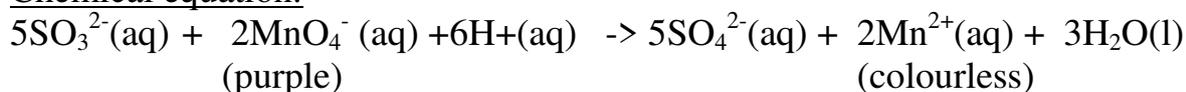
- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)

- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

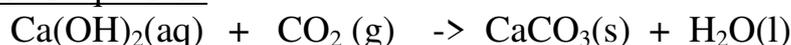
- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:



- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

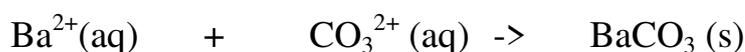
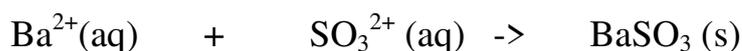
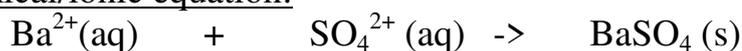


These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Using Barium(II)nitrate(V)/ Barium(II)Chloride

(i)Barium(II)nitrate(V) and/ or Barium(II)chloride solution reacts with Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

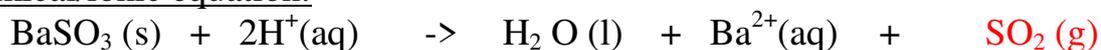


(ii) When the insoluble precipitates are acidified with nitric(V) acid,

- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.

- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

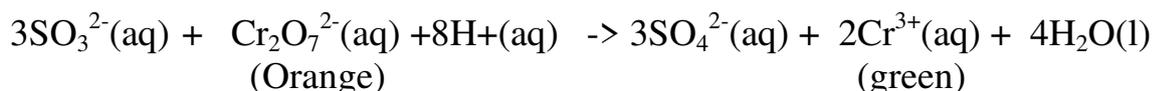
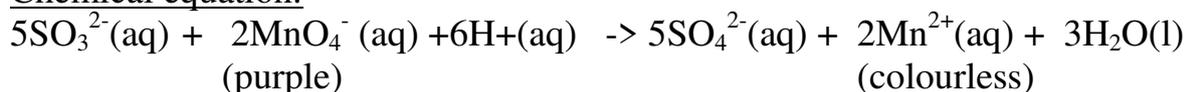
. Chemical/ionic equation:



(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

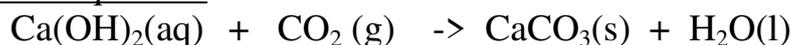
- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green. **Carbon(IV)oxide will not.**

Chemical equation:



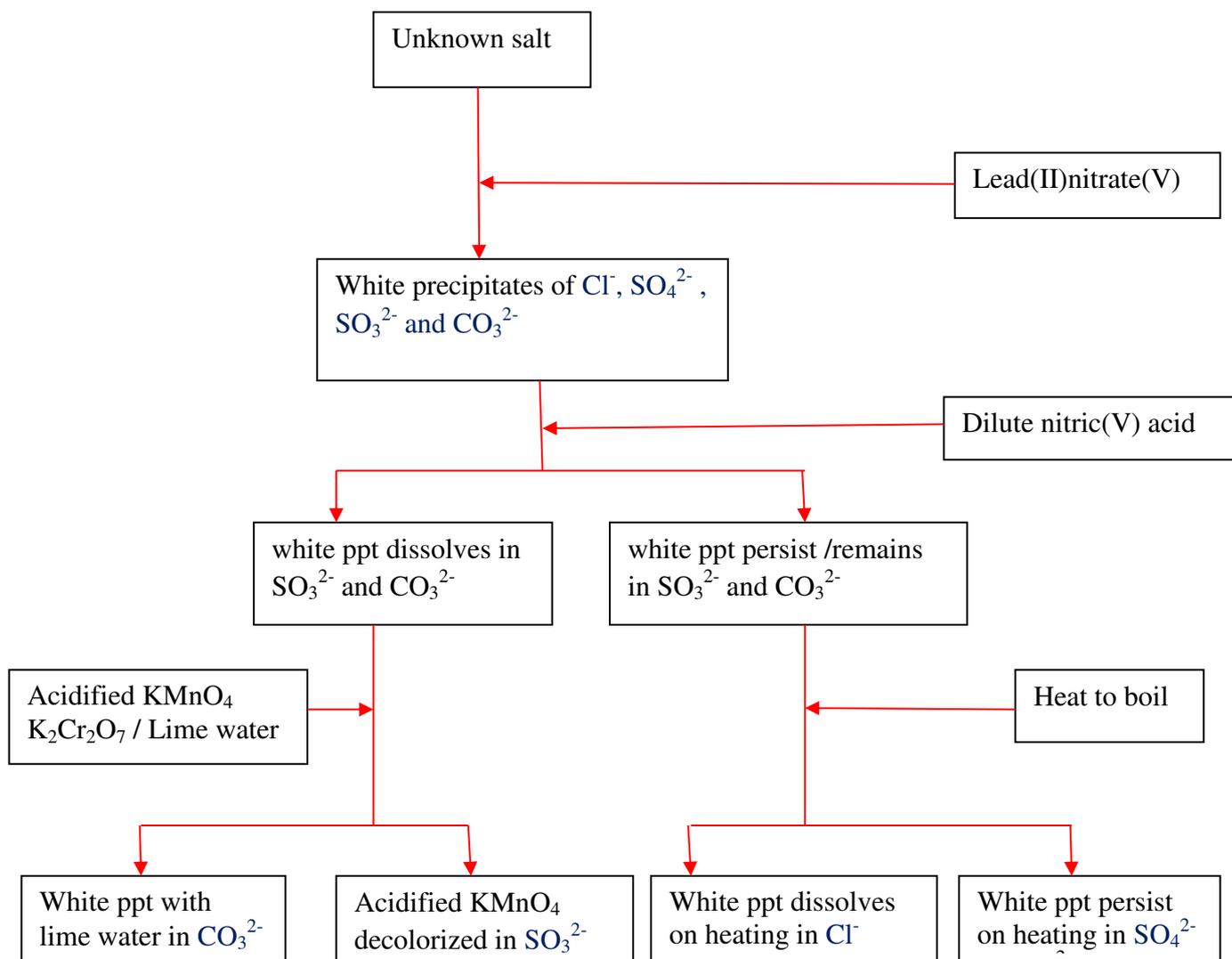
- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:



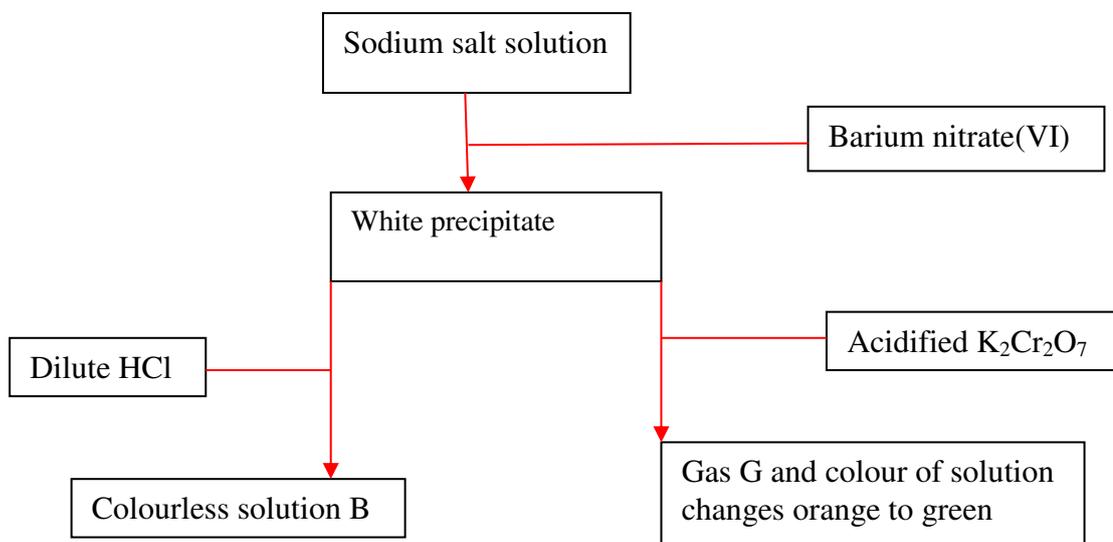
These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Summary test for Sulphate (VI) (SO_4^{2-}) and Sulphate(IV) (SO_3^{2-}) salts



Practice revision question

1. Study the flow chart below and use it to answer the questions that follow

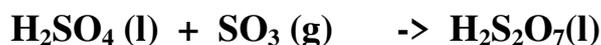


(i) Write equation for the reaction taking place at:

I. The roasting furnace (1mk)



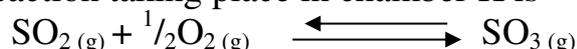
II. The absorption tower (1mk)



III. The diluter (1mk)



(ii) The reaction taking place in chamber K is



I. Explain why it is necessary to use excess air in chamber K

To ensure all the SO₂ reacts

II. Name another substance used in chamber K

Vanadium(V)oxide

3.(a) Describe a chemical test that can be used to differentiate between sodium sulphate (IV) and sodium sulphate (VI).

Add acidified Barium nitrate(V)/chloride.

White precipitate formed with sodium sulphate (VI)

No white precipitate formed with sodium sulphate (IV)

(b) Calculate the volume of sulphur (IV) oxide formed when 120 kg of copper is reacted with excess concentrated sulphuric(VI) acid. (Cu = 63.5, 1 mole of a gas at s.t.p = 22.4 dm³)

Chemical equation



Mole ratio Cu(s) : SO₂(g) = 1:1

Method 1

$$\begin{aligned} 1 \text{ Mole Cu} &= 63.5 \text{ g} & \rightarrow & & 1 \text{ mole SO}_2 &= 22.4 \text{ dm}^3 \\ (120 \times 1000) \text{ g} & & \rightarrow & & \frac{(120 \times 1000) \text{ g} \times 22.4 \text{ dm}^3}{63.5 \text{ g}} \end{aligned}$$

$$= \underline{\underline{42330.7087}}$$

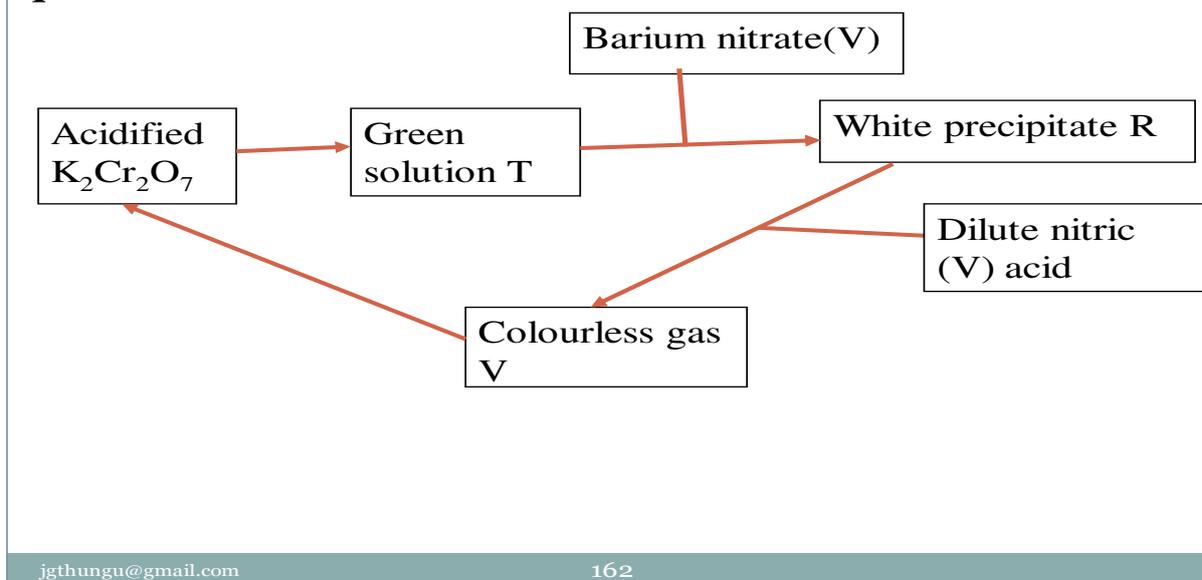
Method 2

$$\text{Moles of Cu} = \frac{(120 \times 1000) \text{ g}}{63.5} = \mathbf{1889.7639 \text{ moles}}$$

$$\text{Moles SO}_2 = \text{Moles of Cu} = \mathbf{1889.7639 \text{ moles}}$$

$$\text{Volume of SO}_2 = \text{Mole} \times \text{molar gas volume} = (1889.7639 \text{ moles} \times 22.4) \\ = \mathbf{42330.7114}$$

4. Use the reaction scheme below to answer the questions that follow.



(a) Identify the:

(i) cation responsible for the green solution T



(ii) possible anions present in white precipitate R



(b) Name gas V

Sulphur (IV) oxide

(c) Write a possible ionic equation for the formation of white precipitate R.

