

(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

- K -273 = °C substituting:** $10 - 273 = -263\text{ °C}$
- (ii) (i) 1 K
K -273 = °C substituting: $1 - 273 = -272\text{ °C}$
- (iii) 110 K
K -273 = °C substituting: $110 - 273 = -163\text{ °C}$
- (iv) -24 K
K -273 = °C substituting: $-24 - 273 = -297\text{ °C}$

The **standard** temperature is **273K = 0°C**.

The **room** temperature is assumed to be **298K = 25°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**”

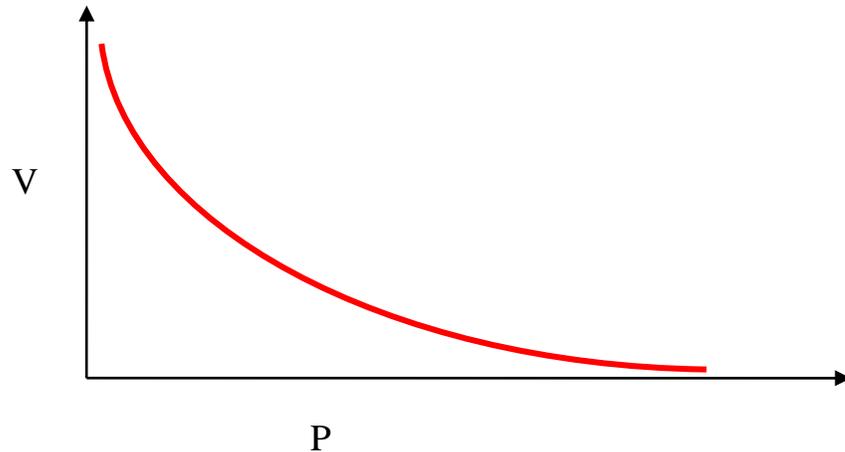
Mathematically:

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

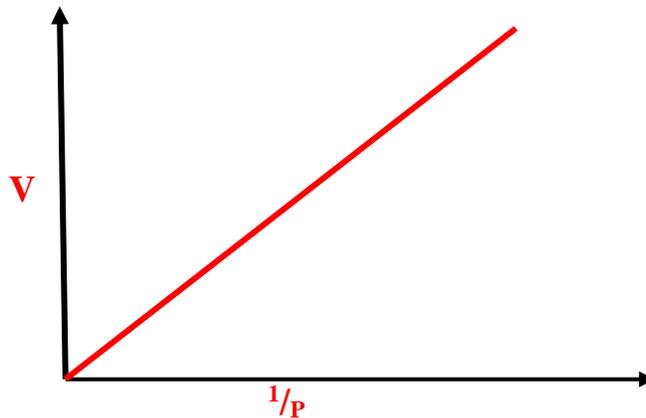
$$V \propto \frac{1}{P} \text{ (Fixed /constant T) i.e. } \mathbf{PV = Constant(k)}$$

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 balloon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the balloon 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** balloon 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**.

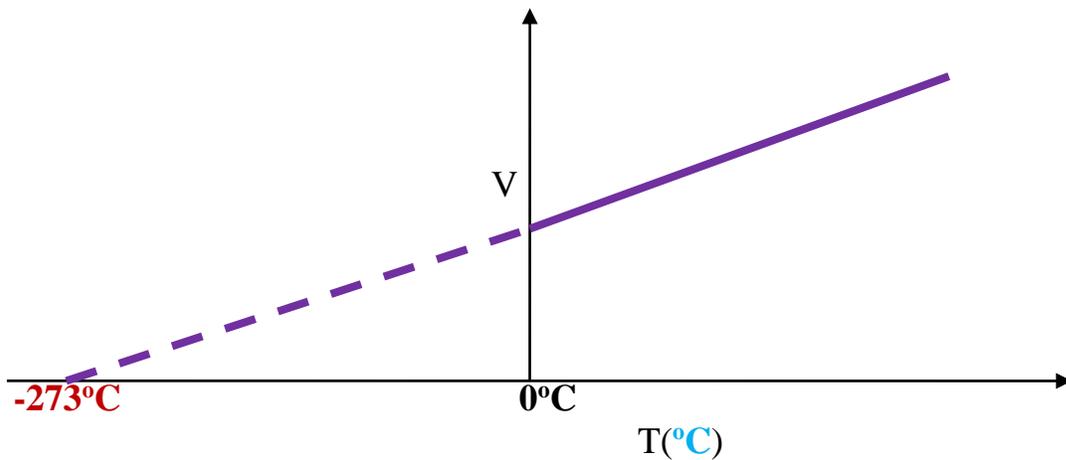
Gases expand/increase by $\frac{1}{273}$ by volume on heating. Gases contract/decrease by $\frac{1}{273}$ by volume on cooling at constant/fixed pressure.

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

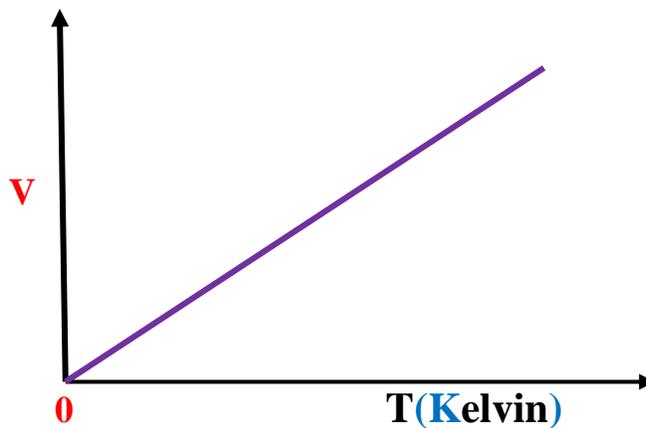
This temperature is called **absolute zero**. It is the **lowest** temperature at which a gas **can** exist.

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

(i)°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.

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

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.

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

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.

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

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.**

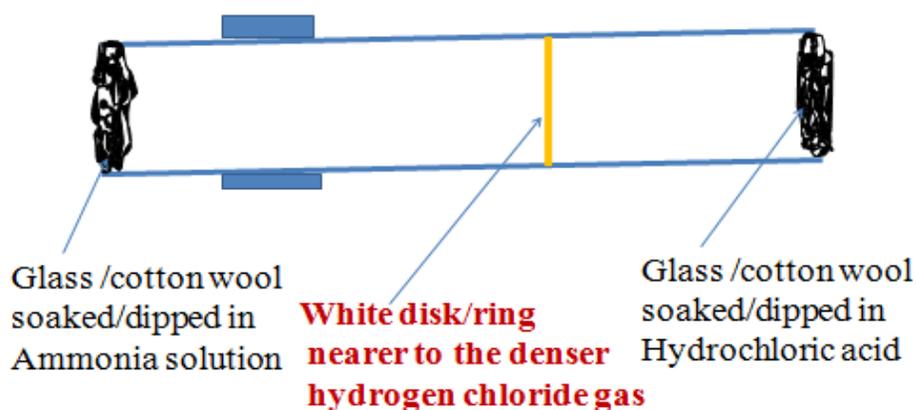
Examples

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

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

For two gases then:

$$\frac{\mathbf{R}_1}{\sqrt{\mathbf{M}_2}} = \frac{\mathbf{R}_2}{\sqrt{\mathbf{M}_1}} \quad \text{where: } \mathbf{R}_1 \text{ and } \mathbf{R}_2 \text{ is the } \underline{\text{rate}} \text{ of diffusion of } \mathbf{1}^{\text{st}} \text{ and } \mathbf{2}^{\text{nd}} \text{ gas.}$$

$$\mathbf{M}_1 \text{ and } \mathbf{M}_2 \text{ is the } \underline{\text{molar mass}} \text{ of } \mathbf{1}^{\text{st}} \text{ and } \mathbf{2}^{\text{nd}} \text{ gas.}$$

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

For two gases then:

$$\frac{\mathbf{T}_1}{\sqrt{\mathbf{M}_1}} = \frac{\mathbf{T}_2}{\sqrt{\mathbf{M}_2}} \quad \text{where: } \mathbf{T}_1 \text{ and } \mathbf{T}_2 \text{ is the } \underline{\text{time taken}} \text{ for } \mathbf{1}^{\text{st}} \text{ and } \mathbf{2}^{\text{nd}} \text{ gas to diffuse.}$$

$$\mathbf{M}_1 \text{ and } \mathbf{M}_2 \text{ is the } \underline{\text{molar mass}} \text{ of } \mathbf{1}^{\text{st}} \text{ and } \mathbf{2}^{\text{nd}} \text{ gas.}$$

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

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

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

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

Method 2

$$100\text{cm}^3 \text{ CO}_2 \text{ takes} \quad 30\text{seconds}$$

$$1\text{cm}^3 \quad \text{takes} \quad \frac{100 \times 1}{30} = \underline{\underline{3.3333\text{cm}^3\text{sec}^{-1}}}$$

$$\frac{R \text{ CO}_2}{R \text{ NO}_2} = \frac{\sqrt{\text{molar mass NO}_2}}{\sqrt{\text{molar mass CO}_2}} \Rightarrow \frac{3.3333\text{cm}^3\text{sec}^{-1}}{R \text{ NO}_2} = \frac{\sqrt{46.0}}{\sqrt{44.0}}$$

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

$$\frac{3.2601\text{cm}^3}{150\text{cm}^3} \quad \text{takes} \quad \frac{1\text{seconds}}{\text{take}} = \underline{\underline{46.0109\text{seconds}}}$$

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

Molar mass CO₂ = 44g Molar mass HCl = 36.5g

$$\frac{T \text{ CO}_2}{T \text{ HCl}} = \frac{\sqrt{\text{molar mass CO}_2}}{\sqrt{\text{molar mass HCl}}} \Rightarrow \frac{200 \text{ seconds}}{T \text{ HCl}} = \frac{\sqrt{44.0}}{\sqrt{36.5}}$$

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

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.

Molar mass O₂ = 32g Molar mass Z = x g

$$\frac{T \text{ O}_2}{T \text{ Z}} = \frac{\sqrt{\text{molar mass O}_2}}{\sqrt{\text{molar mass Z}}} \Rightarrow \frac{250 \text{ seconds}}{227\text{seconds}} = \frac{\sqrt{32.0}}{\sqrt{x}}$$

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

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)

Molar mass CO₂ = 44.0 Molar mass CO = 28.0

Method 1

$$\frac{25\text{cm}^3 \text{ CO}}{75\text{cm}^3} \quad \text{takes} \quad \frac{25\text{seconds}}{\text{take}} = \underline{\underline{75\text{seconds}}}$$

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

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

Method 2

25cm³ CO₂ takes 25seconds

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

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

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

$$\begin{array}{l} 0.7977\text{cm}^3 \quad \text{takes} \quad 1 \text{ seconds} \\ 75\text{cm}^3 \quad \quad \text{takes} \quad \frac{75\text{cm}^3}{0.7977\text{cm}^3} = \underline{\underline{94.0203\text{seconds}}} \end{array}$$

(b) 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 **Avogadro's Constant**. It is denoted "**L**".

The Avogadro's Constant contains **6.023×10^{23}** 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.

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

$$\begin{array}{l} \text{Molar mass chlorine molecule (Cl}_2\text{)} = \text{relative molecular mass} = (35.5 \times 2) \text{ g} = 71.0 \text{ g} \\ 6.023 \times 10^{23} \text{ particles of chlorine molecule} = 1 \text{ mole} = 71.0 \text{ g} \end{array}$$

$$\text{Molar mass Nitrogen molecule (N}_2\text{)} = \text{relative molecular mass} = (14.0 \times 2) \text{ g} = 28.0 \text{ g}$$

6.023×10^{23} 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 = $[(1.0 \times 2) + 16.0]$ g = 18.0g

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

6.023×10^{23} particles of Water molecule has:

- **2** x 6.023×10^{23} particles of Hydrogen atoms

- **1** x 6.023×10^{23} particles of Oxygen atoms

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

6.023×10^{23} particles of sulphuric(VI)acid(H₂SO₄) = 1 mole = 98.0g

6.023×10^{23} particles of sulphuric(VI)acid(H₂SO₄) has:

- **2** x 6.023×10^{23} particles of **H**ydrogen atoms

- **1** x 6.023×10^{23} particles of **S**ulphur atoms

- **4** x 6.023×10^{23} particles of **O**xxygen atoms

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

6.023×10^{23} particles of sodium carbonate(IV)(Na₂CO₃) = 1 mole = 106.0g

6.023×10^{23} particles of sodium carbonate(IV)(Na₂CO₃) has:

- **2** x 6.023×10^{23} particles of **S**odium atoms

- **1** x 6.023×10^{23} particles of **C**arbon atoms

- **3** x 6.023×10^{23} particles of **O**xxygen atoms

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

6.023×10^{23} particles of Calcium carbonate(IV)(CaCO₃) = 1 mole = 100.0g

6.023×10^{23} particles of Calcium carbonate(IV)(CaCO₃) has:

- **1** x 6.023×10^{23} particles of **C**alcium atoms

- **1** x 6.023×10^{23} particles of **C**arbon atoms

- **3** x 6.023×10^{23} particles of **O**xxygen atoms

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

6.023 x 10²³ particles of Water(H₂O) = 1 mole = 18.0g

6.023 x 10²³ particles of Water(H₂O) has:

- 2 x 6.023 x 10²³ particles of Hydrogen atoms

- 2 x 6.023 x 10²³ 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 x 2) = **71.0 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 H₂SO₄ = [(2 x 1) + 32 + (4 x 14)] = **98.0g**

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

2. Calculate the number of atoms present in: (Avogadro's constant L = 6.0 x 10²³)

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

Method I

Molar mass of H₂SO₄ = [(2 x 1) + 32 + (4 x 14)] = **98.0g**

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

1 mole has 6.0 x 10²³ atoms

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

1

Method II

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

98.0g = 1 mole has 6.0×10^{23} atoms

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

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

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

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

Method I

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

1 mole has 6.0×10^{23} atoms

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

Method II

276.0g = 1 mole has 6.0×10^{23} atoms

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

(iii) 0.23 g of Oxygen gas

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

Method I

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

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

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

Method II

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

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

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

Molar mass of CO₂ = [12 + (2 x 16)] = **44.0 g**

Method I

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

1 mole has **3 x 6.0 x 10²³ atoms** in CO₂

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

Method II

44.0g = 1 mole has **3 x 6.0 x 10²³ atoms** in CO₂

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

(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

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

Method II

$$\begin{array}{rcl}
 \text{Mass of crucible + copper after heating } (M_3) & = & 19.1g \\
 \text{Mass of crucible} & = & \underline{- 15.6g}
 \end{array}$$

Mass of copper(II)Oxide	=	3.5 g
Mass of copper(II)Oxide	=	3.5 g
Mass of copper	=	- <u>2.8 g</u>
Mass of Oxygen	=	0.7 g

3. Calculate the number of moles of:

(i) copper used (Cu = 63.5)

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

(ii) Oxygen used (O = 16.0)

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

4. Determine the mole ratio of the reactants

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

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

CuO (copper(II)oxide)

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

Observation

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

Explanation

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

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

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

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

Procedure.

Weigh a clean dry porcelain boat (**M₁**). Put two spatula full of copper(II)oxide **powder** into the crucible. Reweigh the porcelain boat (**M₂**). Put the porcelain boat in a glass tube and set up the apparatus as below;

(i) Oxygen.

$$\begin{array}{rcl}
 \text{Mass of boat before heating (M}_2\text{)} & = & 19.1 \\
 \text{Mass of boat after heating (M}_3\text{)} & = & \underline{- 18.4\text{g}} \\
 \text{Mass of oxygen} & = & \mathbf{0.7\text{ g}}
 \end{array}$$

(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**

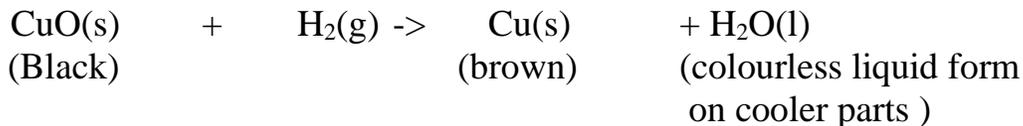
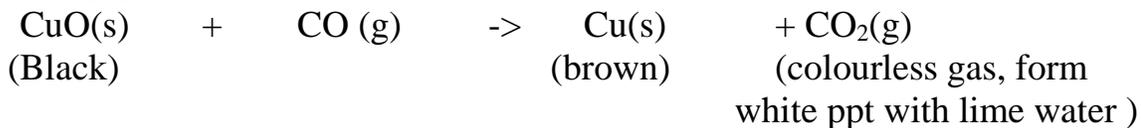
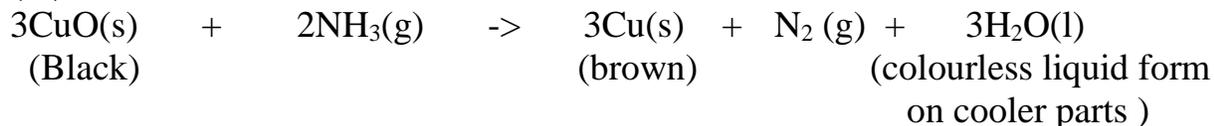
Explanation

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 = $1.60 - 0.84 \Rightarrow 0.56$ 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 = $100 - 47 \Rightarrow 53\%$ 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 contain 70% by mass of Iron and 30% Oxygen. What is its empirical formula (Fe = 56.0, 16.0)

Mass of Oxygen = $100 - 47 \Rightarrow 53\%$ 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 = $305.0\text{g} - 300.0\text{g} = \underline{5.0\text{g}}$

Mass of anhydrous salt = $303.2\text{g} - 300.0\text{g} = \underline{3.2\text{g}}$

Mass of water in hydrated salt = $5.0\text{g} - 3.2\text{g} = \underline{1.8\text{g}}$

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 = $\text{CuSO}_4 \cdot \underline{5}\text{H}_2\text{O}$

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 :

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

The molecular formula is (C H) x 6 = **C₆H₆**

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 **C₂H₄O**

The molecular formula is thus determined :

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

The molecular formula is (C₂H₄O) x 2 = **C₄H₈O₂**.

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:

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

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

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	$\frac{0.12}{0.12}$	$\frac{0.24}{0.12}$

	0.12	0.12
Mole ratios	1	2✓

Empirical formula is **CH₂**✓

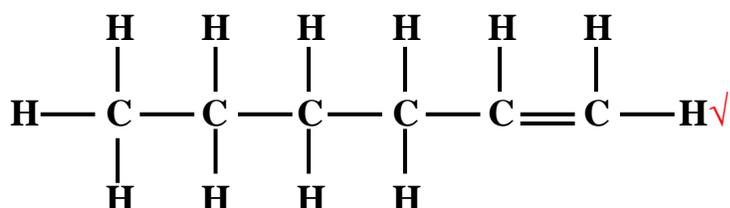
The molecular formula is thus determined :

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

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	$\frac{0.00032}{0.00032}$	$\frac{0.00509}{0.00032}$	$\frac{0.0041\text{g}}{0.00032}$	$\frac{0.00096}{0.00032}$
	1	16	13	3
Mole ratios	1	16	13	3

Empirical formula = C₁₆H₁₃NO₃

(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 x 10²³ particles = 24dm³/24litres/24000cm³ at r.t.p

1mole of H₂ = 2g = 6.0 x 10²³ particles = 24dm³/24litres/24000cm³ at r.t.p

1mole of CO₂ = 44g = 6.0 x 10²³ particles = 24dm³/24litres/24000cm³ at r.t.p

1mole of NH₃ = 17g = 6.0 x 10²³ particles = 24dm³/24litres/24000cm³ at r.t.p

1mole of CH₄ = 16g = 6.0 x 10²³ 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 x 10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of H₂ = 2g = 6.0 x 10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of CO₂ = 44g = 6.0 x 10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of NH₃ = 17g = 6.0 x 10²³ particles = 22.4dm³/22.4litres/22400cm³ at s.t.p

1mole of CH₄ = 16g = 6.0 x 10²³ 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 x 10²³ mole⁻¹)

(i) 2.24dm³ of Oxygen.

22.4dm³ -> 6.0 x 10²³

2.24dm³ -> 2.24 x 6.0 x 10²³

22.4

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

(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)

$$\text{Molar mass CO}_2 = 44 \text{ gmole}^{-1} \checkmark \quad \text{Molar mass H}_2\text{O} = 18 \text{ gmole}^{-1} \checkmark$$

$$\text{Molar mass X} = \frac{0.29 \times (24 \times 1000)\text{cm}^3}{120\text{cm}^3} = 58 \text{ gmole}^{-1} \checkmark$$

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

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

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

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 \checkmark

Empirical formula is **C₂H₅** \checkmark

The molecular formula is thus determined :

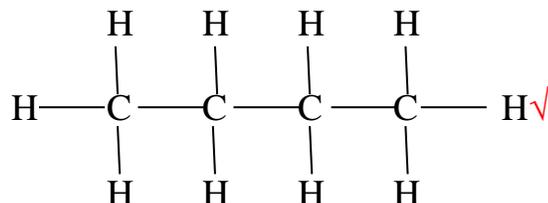
$$n = \frac{\text{Relative formula mass}}{\text{Empirical formula mass}} = \frac{58}{29} = \mathbf{2} \checkmark$$

Relative empirical formula 29

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

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

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

Method 2

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

Mole ratio 1:1

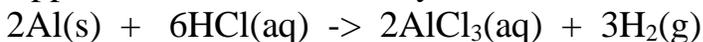
Moles of $\text{CO}_2(\text{g}) = \mathbf{0.05}$ moles

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

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

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

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

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

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

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

Method 2

Mole ratio 2Al: 3H₂ = 2:3

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

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

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

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

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

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

(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.

Chemical equation: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

Volume ratios 2 : 1 : 0

Reacting volumes 50cm³ : **25cm³**

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)

Chemical equation: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

Volume ratios 2 : 1 : 0

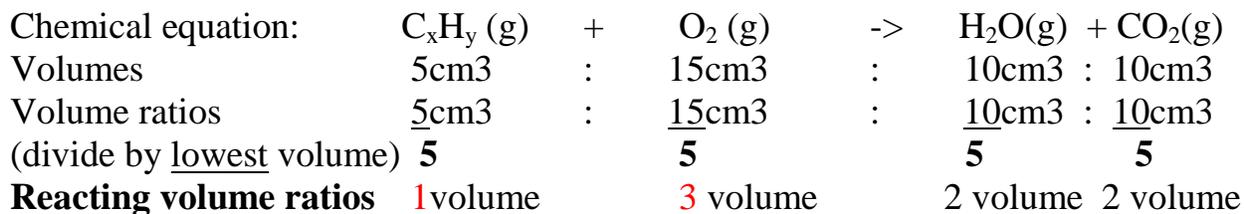
Reacting volumes 50cm³ : **25cm³**

50cm³ of Oxygen is used

21% = 25cm³

$$100\% = \frac{100 \times 25}{21} =$$

3.If 5cm³ of a hydrocarbon C_xH_y burn in 15cm³ of Oxygen to form 10cm³ of Carbon(IV)oxide and 10cm³ of water vapour/steam, obtain the equation for the reaction and hence find the value of x and y in C_xH_y.



Balanced chemical equation: **C₂H₄ (g) + 3O₂ (g) -> 2H₂O(g) + 2CO₂(g)**

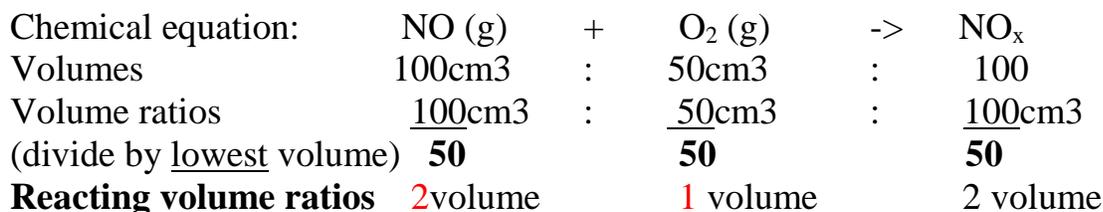
If "4H" are in **2H₂O(g)** the y=4

If "2C" are in **2CO₂ (g)** the x=2

Thus(i) chemical formula of hydrocarbon = **C₂H₄**

(ii) chemical name of hydrocarbon = **Ethene**

4.100cm³ of nitrogen (II)oxide NO combine with 50cm³ of Oxygen to form 100cm³ of a single gaseous compound of nitrogen. All volumes measured at the same temperature and pressure. Obtain the equation for the reaction and name the gaseous product.



Balanced chemical equation: **2 NO (g) + O₂ (g) -> 2NO₂(g)**

Thus(i) chemical formula of the nitrogen compound = **2 NO₂**

(ii) chemical name of compound = **Nitrogen(IV)oxide**

5.When 15cm³ of a gaseous hydrocarbon was burnt in 100cm³ of Oxygen ,the resulting gaseous mixture occupied70cm³ at room temperature and pressure. 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)

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

(P was completely burnt)

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



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

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

$$1 \quad : \quad 3\sqrt{\quad}$$

=> **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.

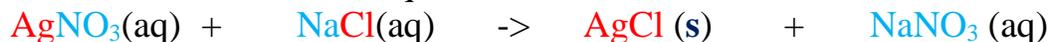
Practice

(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

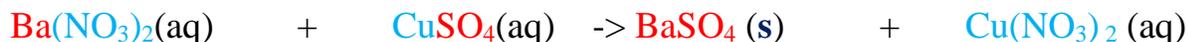


Rewrite the equation

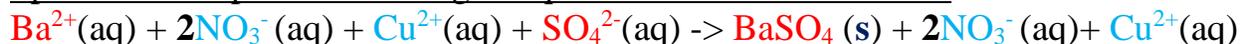


2. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

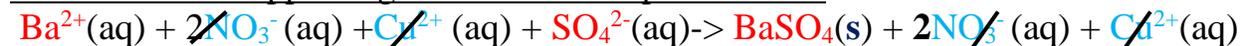
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion



Cancel out ions appearing on reactant and product side



Rewrite the equation

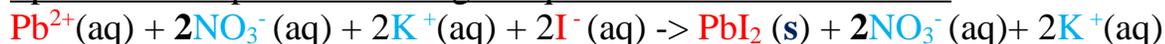


3. A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.

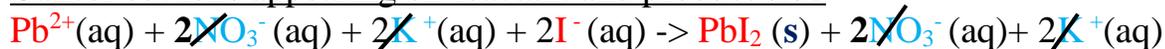
Balanced stoichiometric equation



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.

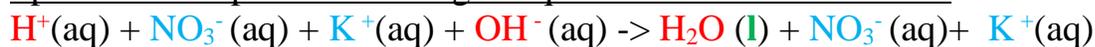
(i)Reaction of alkalis with acids

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

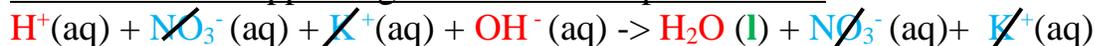
Balanced stoichiometric equation



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

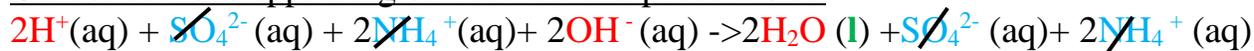
Balanced stoichiometric equation



Split reactants product existing in aqueous state as cation/anion

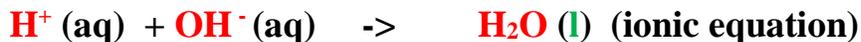


Cancel out ions appearing on reactant and product side



Rewrite the equation



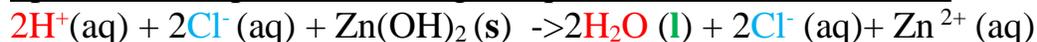


3. Reaction of hydrochloric acid with Zinc hydroxide

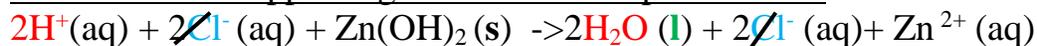
Balanced stoichiometric equation



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.

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

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

(ii) 250 cm³ solution in the volumetric flask.

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

(iii) one decimeter of solution

Method 1

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

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

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

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

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

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

Method 1

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

Method 2

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

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

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

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

Method 1

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

Method 2

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

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

Molar mass of $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ = $(23 \times 2 + 12 + 16 \times 3 + 20 \times 1 + 10 \times 16)$ = 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 1.85×10^{-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 Na_2SO_4 = 142 g

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

Method 2

Molar mass Na_2SO_4 = 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

$$1.0\text{cm}^3 \text{ solution} \rightarrow 1.84 \text{ g}$$

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

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

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

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

Method 2

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

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

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)

$$20 \text{ cm}^3 \text{ solution contain } 1.88 \times 10^{-2} \text{ moles}$$

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

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

2. Calculate the mass of :

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

$$\text{Molar mass NaOH} = \underline{\underline{40\text{g}}}$$

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

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

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

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

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

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

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

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

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

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

3. Calculate the volume required to dissolve :

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

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

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

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

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

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

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

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

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

$$C_2 = \frac{0.8\text{M} \times 312.5\text{cm}^3}{412.5} = \underline{\underline{0.6061\text{M}}}$$

(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:

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

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

$$\underline{\underline{M_1 V_1}} = \underline{\underline{M_2 V_2}} \quad \rightarrow \quad \underline{\underline{M_1 \times 5.0\text{cm}^3}} = \underline{\underline{0.5\text{M} \times 25.0\text{cm}^3}}$$

$$n_1 \quad n_2 \quad \Rightarrow M_1 = \frac{1}{0.5 \times 25.0 \times 1} = \frac{2}{1.25M / 1.25 \text{ moledm}^{-3}}$$

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 $MCO_3 : HCl = 1:2$

$$\text{Moles } MCO_3 = \frac{0.0125 / 1.25 \times 10^{-2} \text{ moles}}{2} = \underline{\underline{0.00625 / 6.25 \times 10^{-3} \text{ moles}}}$$

$$\text{Molarity } MCO_3 = \frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$$

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

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

Stoichiometric equation: $M_2CO_3(aq) + 2HCl(aq) \rightarrow 2MCl(aq) + CO_2(g) + H_2O(l)$

Method 1

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

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

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 $M_2CO_3 : HCl = 1:2$

$$\text{Moles } M_2CO_3 = \frac{0.0125 / 1.25 \times 10^{-2} \text{ moles}}{2} = \underline{\underline{0.00625 / 6.25 \times 10^{-3} \text{ moles}}}$$

$$\text{Molarity } M_2CO_3 = \frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$$

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

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

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2} \quad \rightarrow \quad \frac{5 \times 0.1 \text{cm}^3}{2} = \frac{0.1 \text{M} \times V_2 \text{cm}^3}{1}$$

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

Method 2

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

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

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

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

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

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

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 = $\frac{\text{molarity} \times \text{volume}}{1000}$

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

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

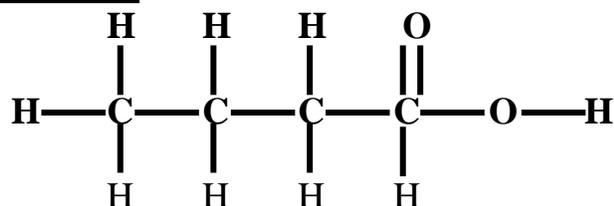
$$X = 87.8324 - 42 = 45.8324 \approx 46$$

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

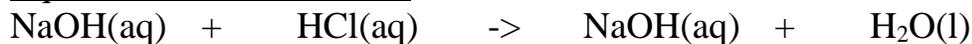
Molecule name : Butan-1-oic acid

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

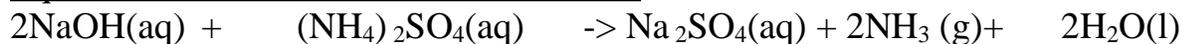


Mole ratio NaOH(aq):HCl(aq) = 1:1

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

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

Equation for reaction with ammonium salt



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

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

Moles of NaOH that reacted with (NH₄)₂SO₄ = 0.2 - 0.0425 = **0.1575 moles**

Moles (NH₄)₂SO₄ = $\frac{1}{2} \times 0.1575 \text{ moles} = \underline{\underline{0.07875 \text{ moles}}}$

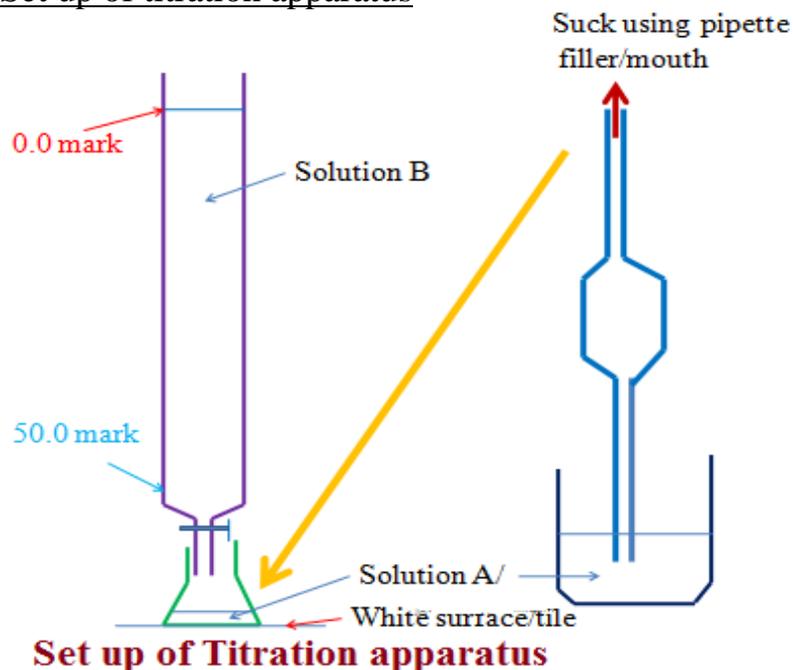
Molar mass (NH₄)₂SO₄ = **132 gmole⁻¹**

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

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 involve 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

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

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**

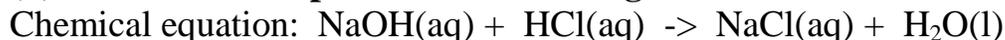
$$\text{Average titre} = \frac{\text{Titre 1} + \text{Titre 2} + \text{Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3} \right) = \underline{\underline{20.0\text{cm}^3}}$$

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.



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

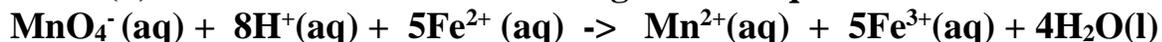
$$\text{Average titre} = \frac{\text{Titre 1} + \text{Titre 2} + \text{Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3} \right) = \mathbf{20.0\text{cm}^3}$$

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:



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

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

(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 (NH₄)₂SO₄FeSO₄.xH₂O.

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

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

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

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

(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

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

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

(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.

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

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)

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

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

$$\begin{aligned} \text{Moles} &= \frac{\text{molarity} \times \text{average burette volume}}{1000} \Rightarrow \frac{0.5 \times 24.5}{1000} \\ &= \underline{\underline{0.01225 / 1.225 \times 10^{-2} \text{ moles}}} \end{aligned}$$

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

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

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

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

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

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

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

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

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

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

$$\begin{aligned} \Rightarrow \text{Moles CaCO}_3(\text{s}) &= \frac{1}{2} \text{ moles HCl} \\ &= \frac{1}{2} \times 0.03582 / 3.582 \times 10^{-2} \text{ moles} \\ &= \underline{0.01791 / 1.791 \times 10^{-2} \text{ moles}} \end{aligned}$$

(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 cm^3 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)

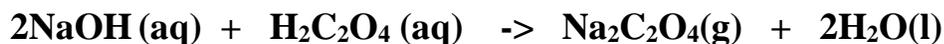
$$\frac{24.9 + 24.9 + 24.9}{3} = 24.9 \text{ cm}^3$$

(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?

25cm³ pipette volume → 0.001245/1.245 × 10⁻³ moles

250cm³ →

$$\frac{0.001245/1.245 \times 10^{-3} \text{ moles} \times 250}{25} = \underline{\underline{0.01245/1.245 \times 10^{-2} \text{ moles}}}$$

(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)

Molar mass H₂C₂O₄ = **90.0g**

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

% by mass of sodium ethanedioate

$$= \frac{(\text{Mass of mixture} - \text{mass of H}_2\text{C}_2\text{O}_4)}{\text{Mass of mixture}} \times 100\%$$

$$\Rightarrow \frac{2.0 - 1.1205 \text{ g}}{2.0} = \underline{\underline{43.975\%}}$$

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

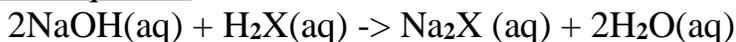
Sample questions

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 = **0.036M/Mole l⁻¹**

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

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

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

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

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

Method 2

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

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

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

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

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

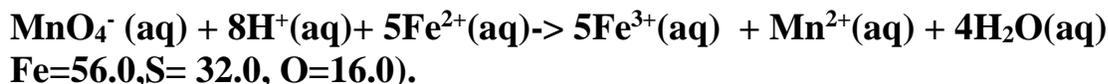
$$0.036 \text{ M/ Mole l}^{-1} \rightarrow 4.0 \text{ g/l}$$

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

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

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

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).



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

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

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

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

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

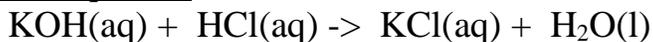
$$\text{Mass of in the mixture} = \text{Moles} \times \text{molar mass} \Rightarrow 0.2 \times 152 = \underline{\underline{30.4 \text{ g}}}$$

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

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

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

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

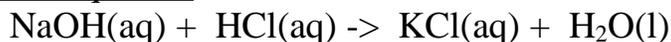
$$\text{Molar mass of GCO}_3 = \frac{\text{mass}}{\text{moles}} \Rightarrow \frac{1}{0.01 \text{ moles}} = \underline{\underline{100 \text{ g}}}$$

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

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:

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

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

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

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