Gas laws and THE MOLE

(a)Gas laws

1. Matter is made up of small particle 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(**^oC**) are also used.
The two units can be interconverted from the relationship:

$$^oC + 273 = K$$
$$K - 273 = {}^oC$$

Practice examples

1. Convert the following into Kelvin.

(i) O **^oC**
^oC + 273 = K substituting : O ^oC + 273 = **273 K**

(ii) -273 **^oC**
^oC + 273 = K substituting : -273^oC + 273 = **0 K**

(iii) 25 **^oC**
^oC + 273 = K substituting : 25 ^oC + 273 = **298 K**

(iv) 100 **^oC**
^oC + 273 = K substituting : 100 ^oC + 273 = **373 K**

2. Convert the following into degrees Celsius/Centigrade(^oC).

(i) 10 K

K -273 = ^{o}C substituting: 10 – 273 = **-263 ^{o}C**

(ii) (i) 1 K

K -273 = ^{o}C substituting: 1 – 273 = **-272 ^{o}C**

(iii) 110 K

K -273 = ^{o}C substituting: 110 – 273 = -163 ^{o}C

(iv) -24 K

K -273 = ^{o}C substituting: -24 – 273 = -297 ^{o}C

The **standard** temperature is **273K** = **0 ^{o}C**.
The **room** temperature is assumed to be **298K** = **25^{o}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:

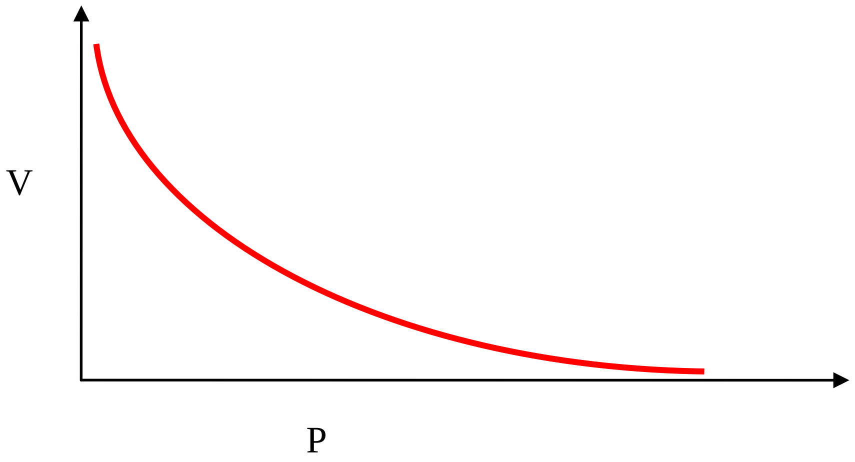
$$\text{Volume} \quad \alpha \quad \underline{\qquad 1 \qquad}$$ **(Fixed /constant Temperature)**

Pressure

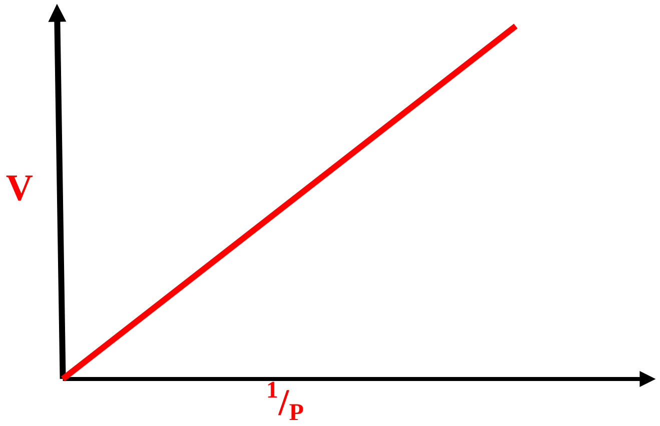
$V \;\; \alpha \;\; \frac{1}{P}$ **(Fixed /constant T)** ie **PV = Constant**(k)

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 $P_1 V_1 = P_2 V_2$

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

$P_1 V_1 = P_2 V_2$ Substituting :$102300 \times 25 = (102300 \times 2) \times V_2$

$$V_2 = \frac{102300 \times 25}{(102300 \times 2)} = 12.5\text{cm3}$$

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

$P_1 V_1 = P_2 V_2$ Substituting :$100000 \times 100 = P_2 \times (100 \times 3)$

$$V_2 = \frac{100000 \times 100}{(100 \times 3)} = 33333.3333 \text{ Nm}^{-2}$$

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

$P_1 V_1 = P_2 V_2$ Substituting :$101325 \times 60 = 90000 \times V_2$

$$V_2 = \frac{101325 \times 60}{90000} = 67.55 \text{ cm3}$$

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

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

Mathematically:

Volume α Pressure **(Fixed /constant pressure)**

$V \propto T$ **(Fixed /constant P)** ie $\frac{V}{T}$ **= 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**.

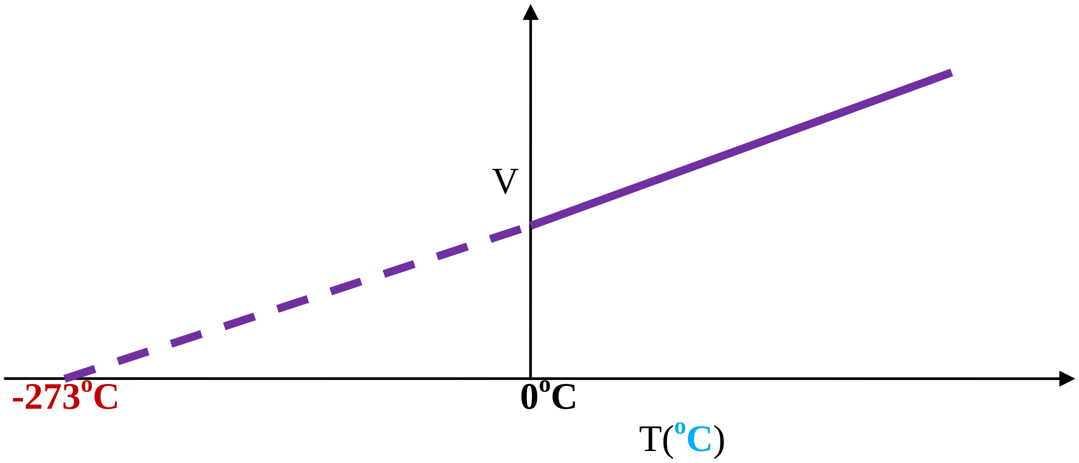
Gases expand/increase by $^{1}/_{273}$ by volume on heating.Gases contact/decrease by $^{1}/_{273}$ by volume on cooling at constant/fixed pressure.

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.

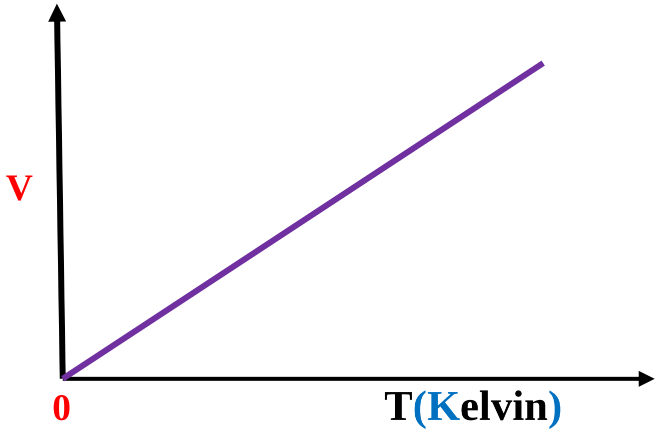
This temperature is called **absolute zero.** It is the **lowest** temperature at which a gas **can** exist.

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 **O 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^{o}C was transfered into a cylinder at -4^{o}C. If the capacity of the cylinder is 450 cm3,explain what happened.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{substituting} \quad \frac{500}{(0+273)} = \frac{V_2}{(-4+273)}$$

$$= \frac{500 \text{ x } (-4 \text{ x } 273)}{(0+273)} = \mathbf{492.674cm3}$$

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^{o}C.Explain what happens.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{substituting} \quad \frac{40000}{(\mathbf{25}+273)} = \frac{V_2}{(\mathbf{30}+273)}$$

$$= \frac{40000 \text{ x } (30 \text{ x } 273)}{(25+273)} = \mathbf{40671.1409cm3}$$

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^{o}C , explain what happened.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{substituting} \quad \frac{80}{(\mathbf{25}+273)} = \frac{V_2}{(-\mathbf{30}+273)}$$

$$= \frac{80 \times (-30 \times 273)}{(25 + 273)} \quad = \quad \mathbf{65.2349cm3}$$

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** <u>higher</u> **the density the** <u>higher</u> **the molar mass/relative atomic mass and thus the** <u>lower</u> **the rate of diffusion.**

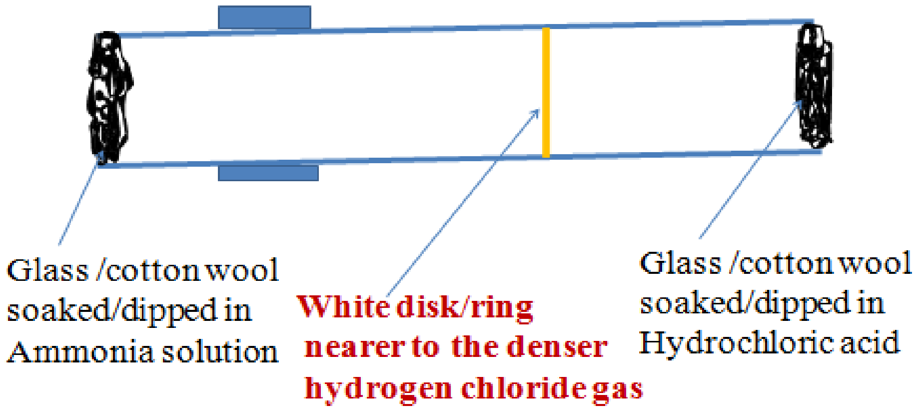
<u>Examples</u>

1.Carbon (IV)oxide(CO_2) has a molar mass of 44g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus lighter/less dense than Carbon (IV)oxide(CO_2). N_2 diffuses faster than $CO_{2.}$

2.Ammonia(NH_3) has a molar mass of 17g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus about <u>**twice**</u> lighter/less dense than Ammonia(NH_3). Ammonia(NH_3) diffuses twice faster than $N_{2.}$

3. Ammonia(NH_3) has a molar mass of 17g.Hydrogen chloride gas has a molar mass of 36.5g.Both gases on contact react to form **white fumes** of ammonium chloride .When a 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



Chemical equation: $HCl(g) + NH_3(s) \rightarrow NH_3Cl(s)$

The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**.
Grahams law states that:
"the rate of diffusion of a gas is inversely proportional to the square root of its density, at the same/constant/fixed temperature and pressure"
Mathematically

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

For two gases then:

$$\frac{R_1}{\sqrt{M_2}} = \frac{R_2}{\sqrt{M_1}}$$

where: R_1 and R_2 is the **rate** of diffusion of 1st and 2nd gas.
M_1 and M_2 is the **molar mass** of 1st and 2nd gas.

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

For two gases then:

$$\frac{T_1}{\sqrt{M_1}} = \frac{T_2}{\sqrt{M_2}}$$

where: T_1 and T_2 is the **time taken** for 1st and 2nd gas to diffuse.
M_1 and M_2 is the **molar mass** of 1st and 2nd 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)

Molar mass CO_2=44.0 Molar mass NO_2=46.0

Method 1

100cm3 CO_2 takes 30seconds

150cm3 takes $\frac{150 \text{ x}30}{100}$ = 45seconds

$$\frac{T\ CO_2}{T\ NO_2} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } NO_2}} \Rightarrow \frac{45\text{seconds}}{T\ NO_2} = \frac{\sqrt{44.0}}{\sqrt{46.0}}$$

$$T\ NO_2 = \frac{45\text{seconds x } \sqrt{46.0}}{\sqrt{44.0}} = \mathbf{46.0114}\text{ seconds}$$

Method 2

100cm3 CO_2 takes 30seconds

1cm3 takes $\frac{100 \text{ x}1}{30}$ = $\mathbf{3.3333cm3sec^{-1}}$

$$\frac{R\ CO_2}{R\ NO_2} = \frac{\sqrt{\text{molar mass } NO_2}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{3.3333\text{cm3sec}^{-1}}{R\ NO_2} = \frac{\sqrt{46.0}}{\sqrt{44.0}}$$

$$R\ NO_2 = \frac{3.3333\text{cm3sec}^{-1} \text{ x } \sqrt{44.0}}{\sqrt{46.0}} = \mathbf{3.2601}\text{cm3sec}^{-1}$$

3.2601cm3 takes 1seconds

150cm3 take $\frac{150\text{cm3}}{3.2601\text{cm3}}$ = **46.0109seconds**

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 CO_2 = 44g Molar mass HCl = 36.5g

$$\frac{T\ CO_2}{T\ HCl} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } HCl}} \Rightarrow \frac{200 \text{ seconds}}{T\ HCl} = \frac{\sqrt{44.0}}{\sqrt{36.5}}$$

$$T\ HCl = \frac{200\text{seconds x } \sqrt{36.5}}{\sqrt{44.0}} = \mathbf{182.1588}\text{ seconds}$$

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 O_2 = 32g Molar mass Z = x g

$$\frac{T\ O_2}{T\ Z} = \frac{\sqrt{\text{molar mass } O_2}}{\sqrt{\text{molar mass } Z}} \Rightarrow \frac{250\text{ seconds}}{227\text{seconds}} = \frac{\sqrt{32.0}}{\sqrt{x}}$$

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

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=12.0,0=16.0)

Molar mass CO_2 = 44.0 Molar mass CO = 28.0

Method 1

25cm3 CO takes 25seconds

75cm3 takes $\frac{75\text{ x}25}{25} = \underline{75\text{seconds}}$

$$\frac{T\ CO_2}{T\ CO} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } CO}} \Rightarrow \frac{T\ CO_2\text{seconds}}{75} = \frac{\sqrt{44.0}}{\sqrt{28.0}}$$

$$T\ CO_2 = \frac{75\text{seconds x } \sqrt{44.0}}{\sqrt{28.0}} = \underline{\mathbf{94.0175}}\text{ seconds}$$

Method 2

25cm3 CO_2 takes 25seconds

1cm3 takes $\frac{25\text{ x}1}{25} = \underline{\mathbf{1.0cm3sec^{-1}}}$

$$\frac{R\ CO_2}{R\ CO} = \frac{\sqrt{\text{molar mass } CO}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{x\text{ cm3sec}^{-1}}{1.0\text{cm3sec}^{-1}} = \frac{\sqrt{28.0}}{\sqrt{44.0}}$$

$$R\ CO_2 = \frac{1.0\text{cm3sec}^{-1}\text{ x } \sqrt{28.0}}{\sqrt{44.0}} = \underline{\mathbf{0.7977}\text{cm3sec}^{-1}}$$

0.7977cm3 takes 1 seconds

75cm3 takes $\frac{75\text{cm3}}{0.7977\text{cm3}} = \underline{\mathbf{94.0203seconds}}$

(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 **Avogadros Constant**. It is denoted "**L**".

The Avogadros Constant contain **6.023×10^{23}** particles. i.e.

1mole = 6.023×10^{23} particles = 6.023×10^{23}
2 moles = $2 \times 6.023 \times 10^{23}$ particles = 1.205×10^{24}
0.2 moles = 0.2 x 6.023 x10 23 particles = 1.205×10^{22}
0.0065 moles = $0.0065 \times 6.023 \times 10^{23}$ particles = 3.914×10^{21}

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

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

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

Molar mass of Iron (Fe) = relative atomic mass = 56.0g
6.023 x10 ^{23} particles of iron = 1 mole =56.0 g

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

Molar mass **Oxygen molecule(O_2)** =relative molecular mass =(16.0x 2)g =32.0g
6.023 x10 ^{23} particles of Oxygen molecule = 1 mole = 32.0 g

Molar mass **chlorine molecule(Cl_2)** =relative molecular mass =(35.5x 2)g =71.0g
6.023 x10 ^{23} particles of chlorine molecule = 1 mole = 71.0 g

Molar mass **Nitrogen molecule(N_2)** =relative molecular mass =(14.0x 2)g =28.0g
6.023 x10 ^{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(H_2O)** = relative formular mass =[(1.0 x 2) + 16.0]g =18.0g

6.023 x10 ^{23} particles of Water molecule = 1 mole = 18.0 g
6.023 x10 ^{23} particles of Water molecule has:
- **2** x 6.023 x10 ^{23} particles of Hydrogen atoms
- **1** x 6.023 x10 ^{23} particles of Oxygen atoms

(ii)Molar mass **sulphuric(VI)acid(H_2SO_4)** = relative formular mass
=[(1.0 x 2) + 32.0 + (16.0 x 4)]g =98.0g

6.023 x10 ^{23} particles of sulphuric(VI)acid(H_2SO_4) = 1 mole = 98.0g
6.023 x10 ^{23} particles of sulphuric(VI)acid(H_2SO_4) has:
- **2** x 6.023 x10 ^{23} particles of **H**ydrogen atoms
- **1** x 6.023 x10 ^{23} particles of **S**ulphur atoms
- **4** x 6.023 x10 ^{23} particles of **O**xygen atoms

(iii)Molar mass **sodium carbonate(IV)(Na_2CO_3)** = relative formular mass

=[(23.0 x 2) + 12.0 + (16.0 x 3)]g =106.0g

6.023 x10 ^{23} particles of sodium carbonate(IV)(Na_2CO_3) = 1 mole = 106.0g
6.023 x10 ^{23} particles of sodium carbonate(IV)(Na_2CO_3) has:

- **2** x 6.023 x10 ^{23} particles of **S**odium atoms
- **1** x 6.023 x10 ^{23} particles of **C**arbon atoms
- **3** x 6.023 x10 ^{23} particles of **O**xygen atoms

(iv)Molar mass **Calcium carbonate(IV)($CaCO_3$)** = relative formular mass
=[(40.0+ 12.0 + (16.0 x 3)]g =100.0g.

6.023 x10 ^{23} particles of Calcium carbonate(IV)($CaCO_3$) = 1 mole = 100.0g
6.023 x10 ^{23} particles of Calcium carbonate(IV)($CaCO_3$) has:

- **1** x 6.023 x10 ^{23} particles of **C**alcium atoms
- **1** x 6.023 x10 ^{23} particles of **C**arbon atoms
- **3** x 6.023 x10 ^{23} particles of **O**xygen atoms

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

6.023 x10 ^{23} particles of Water(H_2O) = 1 mole = 18.0g
6.023 x10 ^{23} particles of Water(H_2O) has:

- **2** x 6.023 x10 ^{23} particles of Hydrogen atoms
- **2** x 6.023 x10 ^{23} particles of **O**xygen atoms

Practice

1. Calculate the number of moles present in:

(i)0.23 g of Sodium atoms

Molar mass of Sodium atoms = 23g

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

(ii) 0.23 g of Chlorine atoms

Molar mass of Chlorine atoms = 35.5 g

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

(iii) 0.23 g of Chlorine molecules

Molar mass of Chlorine molecules =(35.5 x 2) = **71.0** g

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

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

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

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

2. Calculate the number of atoms present in:(Avogadros constant L = 6.0×10^{23})

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

Method I

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

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

1 mole has 6.0×10^{23} atoms

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

Method II

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

98.0g = 1 mole has 6.0×10^{23} atoms

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

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

Molar mass of $Na_2CO_3.10H_2O$=
[(2 x 23) + 12 + (3 x16) + (10 x 1.0) + (10 x 16)] = **276.0**g

Method I

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

1 mole has 6.0×10^{23} atoms

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

Method II

276.0g = 1 mole has 6.0×10^{23} atoms

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

(iii)0.23 g of Oxygen gas

Molar mass of O_2 = (2 x16) = **32.0** g

Method I

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

1 mole has **2** x 6.0 x 10 ^{23} **atoms** in O_2

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

Method II

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

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

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

Molar mass of CO_2 = [12 + (2 x16)] = **44.0** g

Method I

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

1 mole has **3** x 6.0 x 10 ^{23} **atoms** in CO_2

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

Method II

44.0g = 1 mole has **3** x 6.0 x 10 ^{23} **atoms** in CO_2

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

(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.

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

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

Mass of copper **2.8 g**

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

Method I

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

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

Mass of Oxygen = **0.7 g**

Method II

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

Mass of crucible = - 15.6g

Mass of copper(II)Oxide = 3.5 g

Mass of copper(II)Oxide = 3.5 g

Mass of copper = - 2.8 g

Mass of Oxygen = **0.7 g**

3. Calculate the number of moles of:

(i) copper used (Cu = 63.5)

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

(ii) Oxygen used (O = 16.0)

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

4. Determine the mole ratio of the reactants

$\frac{\text{Moles of copper}}{\text{Moles of oxygen}} = \frac{0.0441\text{moles}}{0.0441\text{moles}} = \frac{1}{1}$ => Mole ratio Cu: O = **1:1**

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

CuO (copper(II)oxide

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

Observation

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

Explanation

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

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

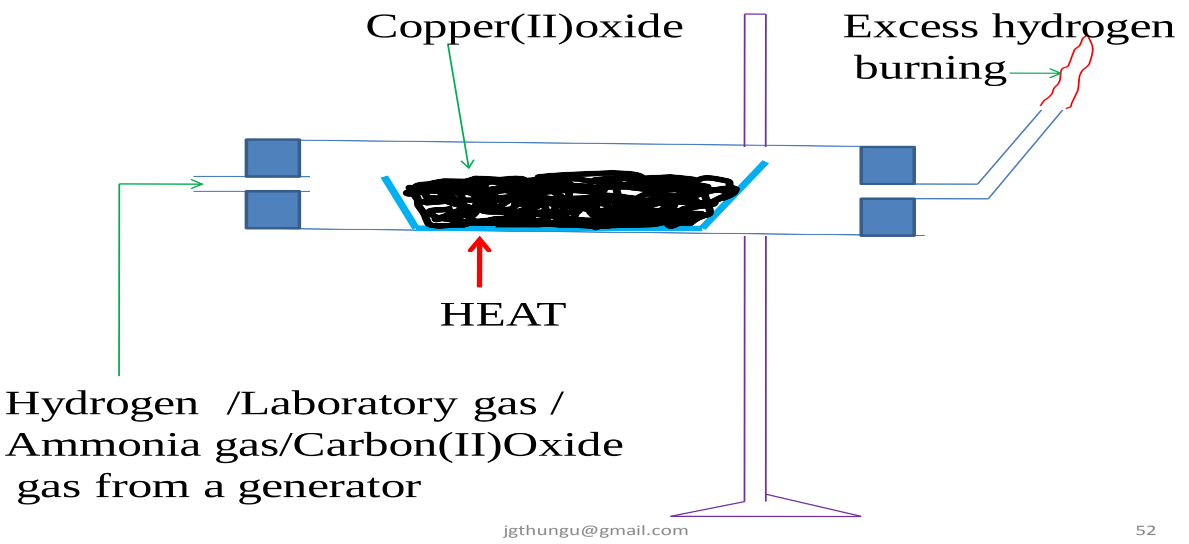
Hot magnesium generates enough heat energy to react with **both** Oxygen and Nitrogen in the air forming a white solid mixture of Magnesium **oxide** and magnesium **nitride**. This causes experimental mass errors.

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

Procedure.

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

Determining empirical formula from copper(II)oxide **to copper**



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

When all the in the apparatus set up is driven out ,heat the copper(II)oxide strongly for about five minutes until there is no further change. Stop heating.
Continue passing the gases until the glass tube is cool.
Turn off the gas generator.
Carefully remove the porcelain boat form the combustion tube.
Reweigh (**M_3**).

Sample results

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

Sample questions

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

Mass of boat before heating(**M_2**) = 19.1
Mass of empty boat(**M_1**) = - 15.6g
Mass of copper(II)Oxide **3.5 g**

2. Calculate the mass of

(i) Oxygen.

Mass of boat before heating(**M_2**) = 19.1
Mass of boat after heating (**M_3**) = - 18.4g
Mass of oxygen = **0.7 g**

(ii)Copper

Mass of copper(II)Oxide = 3.5 g
Mass of oxygen = 0.7 g
Mass of oxygen = **2.8 g**

3. Calculate the number of moles of:

(i) Copper used (Cu = 63.5)

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

(ii) Oxygen used (O = 16.0)

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

4. Determine the mole ratio of the reactants

$\frac{\text{Moles of copper}}{\text{Moles of oxygen}} = \frac{0.0441\text{moles}}{0.0441\text{moles}} = \frac{1}{1}$ => Mole ratio Cu: O = **1:1**

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

CuO (copper(II)oxide

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

Observation

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

Explanation

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

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

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

(Black) (brown) (colourless liquid form on cooler parts)

(ii)Carbon(II)oxide

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$

(Black) (brown) (colourless gas, form white ppt with lime water)

(iii)Ammonia

$3CuO(s) + 2NH_3(g) \rightarrow 3Cu(s) + N_2(g) + 3H_2O(l)$

(Black) (brown) (colourless liquid form on cooler parts)

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

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

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

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

Empirical formula is **CuO**

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

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

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

Mole ratios	1	1

Empirical formula is **MgO**

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

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

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

Empirical formula is **SiO_2**

(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 => **53%** of Oxygen

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

Empirical formula is **SiO_2**

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

Mass of evaporating dish =300.0g

Mass of evaporating dish + hydrated salt = 305.0g

Mass of evaporating dish + anhydrous salt = 303.2g

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

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

Working

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

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

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

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

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

Element/compound	anhydrous copper (II) sulphate(VI)	Oxygen
Symbol	$CuSO_4$	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 = $CuSO_4.\mathbf{5}H_2O$

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) $$\mathbf{n} = \frac{\text{Relative formular mass}}{\text{Relative empirical formula}}$$

where **n** is a whole number.

(ii) Relative empirical formula x **n** = Relative formular mass

where **n** is a whole number.

Practice sample examples

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 = **$\underline{C_6H_6}$**

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

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

Empirical formula is **C_2H_4O**
The molecular formula is thus determined :

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

The molecular formula is $(C_2H_4O) \times 2 = \mathbf{C_4H_8O_2}$.

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

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

Mass of carbon in CO_2 = $\frac{\text{Mass of C in } CO_2}{\text{Molar mass of } CO_2} \times \text{mass of } CO_2$ =>

$\frac{12 \times 5.28}{44} = \mathbf{1.44g}$√

Mass of Hydrogen in H_2O = $\frac{\text{Mass of C in } H_2O}{\text{Molar mass of } H_2O} \times \text{mass of } H_2O$ =>

$\frac{2 \times 2.16}{18} = \mathbf{0.24g}$√

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

Empirical formula is $\mathbf{CH_2}$√

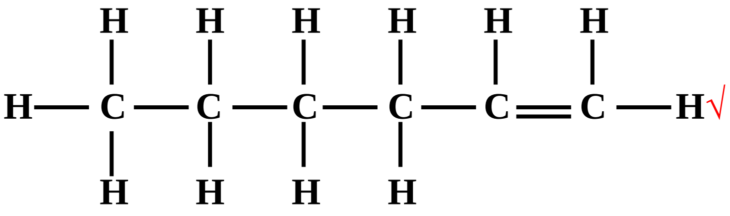
The molecular formular is thus determined :

n = $\frac{\text{Relative formular mass}}{\text{Relative empirical formula}} = \frac{84}{14} = \mathbf{6}$√

The molecular formula is $(CH_2) \times 6 = \mathbf{C_6H_{12}}$. √

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)

Mass of **N** in A = 5.2% x 0.085 **= 0.00442 g**

Mass of C in A = $\frac{12}{44}$ x 0.224 = **0.0611g**

Mass of H in A = $\frac{2}{18}$ x 0.0372 = **0.0041g**

Mass of O in A = 0.085g – 0.004442g = **0.0806g** (Mass of C,H,O)
=> 0.0611g + 0.0041g = **0.0652g (Mass of C,H)**
0.0806g (Mass of C,H,O)- 0.0652g (Mass of C,H) = **0.0154 g**

Element	Nitrogen	Carbon	Hydrogen	Oxygen
Symbol	N	C	H	O
Moles present = $\frac{\text{mass}}{\text{Molar mass}}$	$\frac{0.00442\ g}{14}$	$\frac{0.0611g}{12}$	$\frac{0.0041g}{1}$	$\frac{0.0154\ g}{16}$
Divide by the smallest value	$\frac{0.00032}{0.00032}$	$\frac{0.00509}{0.00032}$	$\frac{0.0041g}{0.00032}$	$\frac{0.00096}{0.00032}$
Mole ratios	1	16	13	3

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

(d)Molar gas volume

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

(i) 24dm3/24litres/24000cm3 at **r**oom **t**emperature(25°C/298K)and **p**ressure(**r.t.p**).
i.e. 1mole of all gases =24dm3/24litres/24000cm3 at r.t.p
Examples

1mole of O_2 = 32g =6.0 x10^{23} particles**= 24dm3/24litres/24000cm3 at r.t.p**
1mole of H_2 = 2g =6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**
1mole of CO_2 = 44g = 6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**
1mole of NH_3 = 17g =6.0 x10^{23} particles **= 24dm3/24litres/24000cm3 at r.t.p**
1mole of CH_4 = 16g =6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**

(**ii**)22.4dm3/22.4litres/22400cm3 at **s**tandard **t**emperature(0°C/273K) and **p**ressure(**s.t.p**)

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

Examples

1mole of O_2 = 32g =6.0 x10^{23} particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**
1mole of H_2 = 2g =6.0 x10^{23} particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**
1mole of CO_2 = 44g = 6.0 x10^{23} particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**
1mole of NH_3 = 17g =6.0 x10^{23} particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**
1mole of CH_4 = 16g =6.0 x10^{23} 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:

(Avogadros constant =6.0 x10^{23}mole^{-1} **)**

(i) 2.24dm3 of Oxygen.

22.4dm3 -> 6.0 x10^{23}
2.24dm3 -> $\frac{2.24 \times 6.0 \times 10^{23}}{22.4}$

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

(ii) 2.24dm3 of Carbon(IV)oxide.

22.4dm3 -> 6.0 x10^{23}
2.24dm3 -> $\frac{2.24 \times 6.0 \times 10^{23}}{22.4}$

=$\underline{6.0 \times 10^{22}}$ **molecules** = (CO_2) = 3 x 6.0 x$10^{22.}$ = $\underline{1.8 \times 10^{23}}$**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.O,H=1.0,1 mole of gas occupies 24dm3 at r.t.p)

Molar mass CO_2= 44 gmole^{-1}√ Molar mass H_2O = 18 gmole^{-1}√

Molar mass X = $\frac{0.29 \text{ x } (24 \text{ x } 1000)\text{cm3}}{120\text{cm3}}$ = 58 gmole^{-1}√

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

Mass of carbon in CO_2 = $\frac{\text{Mass of C in } CO_2}{\text{Molar mass of } CO_2}$ x mass of CO_2 =>

$\frac{12 \text{ x } 0.41}{44}$ = **0.1118g**√

Mass of Hydrogen in H_2O = $\frac{\text{Mass of C in } H_2O}{\text{Molar mass of } H_2O}$ x mass of H_2O =>

$\frac{2 \text{ x } 0.209}{18}$ = **0.0232g**√

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = $\frac{\text{\% composition}}{\text{Molar mass}}$	$\frac{0.g118}{12}$	$\frac{0.0232g}{1}$√
Divide by the smallest value	$\frac{0.0093}{0.0093}$	$\frac{0.0232}{0.0093}$√
Mole ratios	1 x2	2.5x2
	2	5√

Empirical formula is $\mathbf{C_2H_5}$√

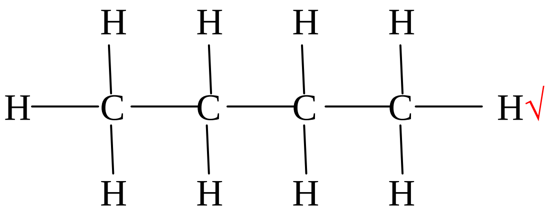
The molecular formular is thus determined :

n = $\frac{\text{Relative formular mass}}{\text{Relative empirical formula}}$ = $\frac{58}{29}$ = **2**√

The molecular formula is $(\mathbf{C_2H_5})$ x 2 = $\mathbf{C_4H_{10}}$.√

Molecule name **Butane**

Molecula 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 ration 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

	$CaCO_3(s)$	->	$CaO(s)$	+	$CO_2(g)$
Mole ratios	1:		1:		1

Molar Mass $CaCO_3$ =100g

Method 1

100g $CaCO_3(s)$ -> 24dm3 $CO_2(g)$ at r.t.p

5.0 g $CaCO_3(s)$ -> $\frac{5.0\text{ g x }24\text{dm3}}{100\text{g}}$ = **1.2dm3/1200cm3**

Method 2

Moles of 5.0 g $CaCO_3(s)$ = $\frac{5.0\text{ g}}{100\text{ g}}$ = **0.05** moles

Mole ratio 1:1

Moles of $CO_2(g)$ = **0.05**moles

Volume of $CO_2(g)$ = 0.05 x 24000cm3 =**1200cm3 /1.2dm3**

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

$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

Method 1

$3H_2(g)$ = 3 moles x (22.4 x 1000)cm3 => 2 x 27 g Al

840cm3 => $\underline{840\text{cm3 x 2 x 27}}$ = **0.675**g of Aluminium

3 x 22.4 x 1000

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g =**0.325**g of copper

$$\% \text{ copper} = \frac{\text{mass of copper x100\%}}{\text{Mass of alloy}} = \mathbf{32.5\%}$$

Method 2

Mole ratio $2Al : 3H_2$ = 2:3

$$\text{Moles of Hydrogen gas} = \frac{\text{volume of gas}}{\text{Molar gas volume}} \Rightarrow \frac{840\text{cm3}}{22400\text{cm3}} = \mathbf{0.0375moles}$$

Moles of Al = $^2/_3$ moles of H_2 => $^2/_3$x 0.0375moles = **0.025moles**

Mass of Al = moles x molar mass =>0.025moles x 27 = **0.675g**

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g = **0.325** g of copper

$$\% \text{ copper} = \frac{\text{mass of copper x100\%}}{\text{Mass of alloy}} = \mathbf{32.5\%}$$

(f)Gay Lussac's law

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

Gay Lussacs law thus only apply to <u>gases</u>

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

<u>Examples</u>

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

Chemical equation:	$2H_2$ (g)	+	O_2 (g)	->	$2H_2O$(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:	$2H_2$ (g)	+	O_2 (g)	->	$2H_2O$(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.

Chemical equation:	C_xH_y (g)	+	O_2 (g)	->	H_2O(g)	+	CO_2(g)
Volumes	5cm3	:	15cm3	:	10cm3	:	10cm3
Volume ratios (divide by <u>lowest</u> volume)	$\frac{5\text{cm3}}{\mathbf{5}}$	:	$\frac{15\text{cm3}}{\mathbf{5}}$	:	$\frac{10\text{cm3}}{\mathbf{5}}$	:	$\frac{10\text{cm3}}{\mathbf{5}}$

Reacting volume ratios	1volume	3 volume	2 volume	2 volume

Balanced chemical equation: **C_xH_y (g) + 3O_2 (g) -> 2H_2O(g) + 2CO_2(g)**

If "4H" are in 2H_2O(g) the y=4

If "2C" are in 2CO_2 (g) the x=2

Thus(i) chemical formula of hydrocarbon = **C_2H_4**

(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.

Chemical equation:	NO (g)	+	O_2 (g)	->	NO_x
Volumes	100cm3	:	50cm3	:	100
Volume ratios	100cm3	:	50cm3	:	100cm3
(divide by lowest volume)	**50**		**50**		**50**
Reacting volume ratios	2volume		1 volume		2 volume

Balanced chemical equation: 2 **NO (g) + O_2 (g) -> 2NO_x(g)**

Thus(i) chemical formula of the nitrogen compound = **2 NO_2**

(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)

Volume of Oxygen used =100-25 =75cm3√

(P was completely burnt)

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

C_xH_y + O_2 -> xCO_2 + yH_2O

15cm3 : 75cm3

15 15

1 : 3√

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

Thus x = 1 and y = 2 => P has molecula formula CH_4√

(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

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

Split reactants product existing in aqueous state as cation/anion

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

Cancel out ions appearing on reactant and product side

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

Rewrite the equation

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

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

Balanced stoichiometric equation

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

Split reactants product existing in aqueous state as cation/anion

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

Cancel out ions appearing on reactant and product side

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

Rewrite the equation

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

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

Balanced stoichiometric equation

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

Split reactants product existing in aqueous state as cation/anion

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

Cancel out ions appearing on reactant and product side

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

Rewrite the equation

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

HNO_3(aq) + KOH (aq) -> H_2O (l) + KNO_3 (aq)

Split reactants product existing in aqueous state as cation/anion

H^+(aq) + NO_3^- (aq) + K^+(aq) + OH^- (aq) -> H_2O (l) + NO_3^- (aq)+ K^+(aq)

Cancel out ions appearing on reactant and product side

H^+(aq) + $\cancel{NO_3^-}$ (aq) + $\cancel{K^+}$(aq) + OH^- (aq) -> H_2O (l) + $\cancel{NO_3^-}$ (aq)+ $\cancel{K^+}$(aq)

Rewrite the equation

H^+ (aq) + OH^- (aq) -> H_2O (l) (ionic equation)

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

Balanced stoichiometric equation

H_2SO_4(aq) + $2NH_4OH$ (aq) -> H_2O (l) + $(NH_4)_2SO_4$ (aq)

Split reactants product existing in aqueous state as cation/anion

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

Cancel out ions appearing on reactant and product side

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

Rewrite the equation

$\underline{2}H^+$ (aq) + $\underline{2}OH^-$ (aq) -> $\underline{2}H_2O$ (l)
H^+ (aq) + OH^- (aq) -> H_2O (l) (ionic equation)

3.Reaction of hydrochloric acid with Zinc hydroxide

Balanced stoichiometric equation

2HCl(aq) + $Zn(OH)_2$ (**s**) -> $2H_2O$ (l) + $ZnCl_2$ (aq)

Split reactants product existing in aqueous state as cation/anion

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

Cancel out ions appearing on reactant and product side

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

Rewrite the equation

$2H^+(aq) + Zn(OH)_2(\mathbf{s}) -> 2H_2O(\mathbf{l}) + Zn^{2+}(aq)$ **(ionic 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.

Molar mass of NaOH = (23 + 16 + 1) = 40g

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

(ii) 250 cm3 solution in the volumetric flask.

Moles in 250 cm3 = **0.1 / 1.0 x 10^{-1}** moles

(iii) one decimeter of solution

<u>Method 1</u>

$$\text{Moles in decimeters} = \mathbf{Molarity} = \frac{\text{Moles x 1000cm3/1dm3}}{\text{Volume of solution}}$$

$$\Rightarrow \frac{1.0 \times 10^{-1}\ \text{moles x 1000cm3}}{\text{250cm3}} =$$

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

<u>Method 2</u>

250cm3 solution contain 1.0 x 10^{-1} moles

$$\text{1000cm3 solution = Molarity contain } \frac{1000 \times 1.0 \times 10^{-1}\ \text{moles}}{\text{250 cm3}}$$

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

Theoretical sample practice

1. Calculate the molarity of a solution containing:

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

Molar mass of NaOH = (23 + 16 + 1) = 40g

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

<u>Method 1</u>

$$\text{Moles in decimeters} = \mathbf{Molarity} = \frac{\text{Moles x 1000cm3/1dm3}}{\text{Volume of solution}}$$

$$\Rightarrow \frac{1.0 \times 10^{-1}\ \text{moles x 1000cm3}}{\text{500cm3}}$$

= **0.2 M / 0.2 molesdm^{-3}**

Method 2

500 cm3 solution contain 1.0×10^{-1} moles

1000cm3 solution = Molarity contain $\frac{1000 \times 1.0 \times 10^{-1} \text{ moles}}{500 \text{ cm3}}$

= **0.2 M / 0.2 molesdm^{-3}**

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

Molar mass of Na_2CO_3 = (23 x 2 + 12 + 16 x 3) = 106 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

$$\text{Moles in decimeters} = \textbf{Molarity} = \frac{\text{Moles x 1000cm3/1dm3}}{\text{Volume of solution}}$$

$$\Rightarrow \frac{\text{1.0 moles x 1000cm3}}{\text{50cm3}} =$$

=**1.0 M**

Method 2

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

1000cm3 solution = Molarity contain $\frac{1000 \times 5.0 \times 10^{-2} \text{ moles}}{50 \text{ cm3}}$

= **1.0M / 1.0 molesdm^{-3}**

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

Molar mass of $Na_2CO_3.10H_2O$ = (23 x 2 + 12 + 16 x 3 + 20 x 1 + 10 x 16) =286g

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

Method 1

$$\text{Moles in decimeters} = \textbf{Molarity} = \frac{\text{Moles x 1000cm3/1dm3}}{\text{Volume of solution}}$$

$$\Rightarrow \frac{1.85 \times 10^{-2} \text{ moles x 1000cm3}}{\text{50cm3}} =$$

= **0.37 M/0.37 molesdm^{-3}**

Method 2

50 cm3 solution contain 1.85 x 10 $^{-2}$ moles

1000cm3 solution = Molarity contain $\frac{1000 \times 1.85 \times 10^{-2} \text{ moles}}{50 \text{ cm3}}$

= **3.7 x 10^{-1} M / 3.7 x 10^{-1} molesdm^{-3}**

(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

1000cm3 solution -> $\frac{1000 \times 71}{20}$ = 3550 g **dm^{-3}**

Molar mass Na_2SO_4 = 142 g

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

Method 2

Molar mass Na_2SO_4 = 142 g

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

Method 2(a)

Moles in decimeters = **Molarity** = $\frac{\text{Moles} \times 1000\text{cm3/1dm3}}{\text{Volume of solution}}$

=> $\frac{5.0 \times 10^{-2} \text{ moles} \times 1000\text{cm3}}{20\text{cm3}}$

= **2.5 M/2.5 molesdm^{-3}**

Method 2(b)

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

1000cm3 solution = Molarity contain $\frac{1000 \times 5.0 \times 10^{-2} \text{ moles}}{20 \text{ cm3}}$

= **2.5 M/2.5 molesdm^{-3}**

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

Method 1

1.0cm3 solution ->1.84 g

1000cm3 solution -> $\frac{1000 \times 1.84}{1}$ = 1840 g **dm^{-3}**

Molar mass H_2SO_4 = 98 g

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

= **18.7755 M**/ moles**dm**$^{-3}$

Method 2

Molar mass H_2SO_4 = 98 g

Moles = $\frac{\text{Mass}}{\text{Molar mass}}$ => $\frac{1.84}{98}$ = **0.0188 / 1.88 x 10** $^{-2}$ moles

Method 2(**a**)

Moles in decimeters = **Molarity** = $\frac{\text{Moles x 1000cm3/1dm3}}{\text{Volume of solution}}$

=> $\frac{1.88 \times 10^{-2} \text{ moles x 1000cm3}}{\text{1.0cm3}}$

= **18.8M/18.8 molesdm**$^{-3}$

Method 2(**b**)

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

1000cm3 solution = Molarity contain $\frac{1000 \times 1.88 \times 10^{-2} \text{ moles}}{\text{1.0 cm3}}$

= **18.8M/18.8 molesdm**$^{-3}$

2. Calculate the mass of :

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

Molar mass NaOH = **40g**

Moles in 25 cm3 = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.2 \times 25}{1000}$ = **0.005/5.0 x 10**$^{-3}$moles

Mass of NaOH =Moles x molar mass = 5.0 x 10^{-3} x 40 = **0.2 g**

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

Molar mass H_2SO_4 = **98g**

Moles in 20 cm3 = $\frac{\text{Molarity x volume}}{1000}$=> $\frac{0.625 \times 20}{1000}$ = **0.0125/1.25.0 x 10**$^{-3}$moles

Mass of H_2SO_4 =Moles x molar mass => 5.0×10^{-3} x 40 = **<u>0.2 g</u>**

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

Molar mass HNO_3 = **<u>63 g</u>**

Moles in 1 cm3 = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{2.5 \times 1}{1000}$ = **<u>0.0025 / 2.5. x 10^{-3}</u>**moles

Mass of HNO_3 =Moles x molar mass => 2.5×10^{-3} x 40 = **<u>0.1 g</u>**

3. Calculate the volume required to dissolve :
(a)(i) 0.25moles of sodium hydroxide solution to form a 0.8M solution

Volume (in cm3) = $\frac{\text{moles x 1000}}{\text{Molarity}}$ => $\frac{0.25 \times 1000}{0.8}$ = **<u>312.5cm3</u>**

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

$C_1 \times V_1 = C_2 \times V_2$ where:
C_1 = molarity/concentration <u>before</u> diluting/adding water
C_2 = molarity/concentration <u>after</u> diluting/adding water
V_1 = volume <u>before</u> diluting/adding water
V_2 = volume <u>after</u> diluting/adding water

=> 0.8M x 312.5cm3 = C_2 x (312.5 + 100**)**

C_2 = $\frac{\text{0.8M x 312.5cm3}}{412.5}$ = **0.6061M**

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

Volume (in cm3) = $\frac{\text{moles x 1000}}{\text{Molarity}}$ => $\frac{0.01 \times 1000}{0.01}$ = **<u>1000 cm3</u>**

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

$C_1 \times V_1 = C_2 \times V_2$ where:
C_1 = molarity/concentration <u>before</u> diluting/adding water
C_2 = molarity/concentration <u>after</u> diluting/adding water
V_1 = volume <u>before</u> diluting/adding water

V_2 = volume after diluting/adding water

$$\Rightarrow 0.01\text{M} \times 1000 \text{ cm3} = 0.008 \times V_2$$

$$V_2 = \frac{0.01\text{M} \times 1000\text{cm3}}{0.008} = \mathbf{1250cm3}$$

Volume added = 1250 - 1000 = **250cm3**

(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:

$$\frac{\mathbf{M_1V_1}}{\mathbf{n_1}} = \frac{\mathbf{M_2V_2}}{\mathbf{n_2}}$$ where**:**

M_1 = Molarity of 1st reactant
M_2 = Molarity of 2nd reactant
V_1 = Volume of 1st reactant
V_1 = 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 5.0cm3 of MCO_3 react with 25.0cm3 of 0.5M hydrochloric acid.(C=12.0 ,O =16.0)

Stoichiometric equation: $MCO_3(s) + 2HCl(aq) \rightarrow MCl_2(aq) + CO_2(g) + H_2O(l)$

Method 1

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2} \quad \rightarrow \quad \frac{M_1 \times 5.0\text{cm3}}{1} = \frac{0.5\text{M} \times 25.0\text{cm3}}{2}$$

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

Method 2

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

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

Mole ratio MCO_3 : HCl = 1:2

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

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

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

2. 2.0cm3 of 0.5M hydrochloric acid react with 0.1M of M_2CO_3. Calculate the volume of 0.1M M_2CO_3 used.

Stoichiometric equation: M_2CO_3 (aq) + 2HCl(aq) -> 2MCl (aq) + CO_2(g) + H_2O(l)

Method 1

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2} \quad \rightarrow \quad \frac{0.5 \text{ x } 2.0\text{cm3}}{2} = \frac{0.1\text{M x } V_2 \text{ cm3}}{1}$$

$$\Rightarrow V_2 = \frac{0.5 \text{ x } 2.0 \text{ x1}}{0.1 \text{ x } 2} = \underline{\mathbf{1.25M}} \ \mathbf{/} \ \underline{\mathbf{1.25 \ moledm^{-3}}}$$

Method 2

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

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

Mole ratio M_2CO_3 : HCl = 1:2

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

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

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

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

(ii)the mass of Lead(II)Iodide formed

(Pb=207.0, I =127.0)

Stoichiometric equation: 2NaI(aq) + $Pb(NO_3)_2$(aq) -> $2NaNO_3$(aq) + PbI_2(**s**)

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

Method 1

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2} \quad -> \quad \frac{5 \times 0.1cm3}{2} = \frac{0.1M \times V_2\,cm3}{1}$$

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

Method 2

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

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

Mole ratio M_2CO_3 : HCl = 1:2

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

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

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

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

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

Mole ratio B : NaOH = 1:1

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

Molar mass B = $\frac{\text{mass}}{\text{moles}}$ => $\frac{0.388}{0.0044175 /4.4175 \times 10^{-3}moles}$

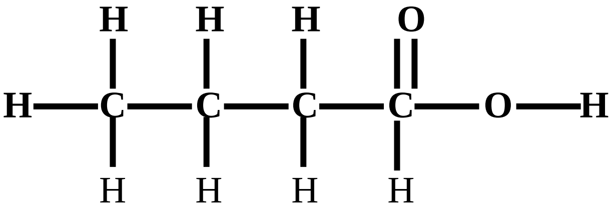
$$= \underline{\mathbf{87.8324}}\ \mathbf{gmole^{-1}}$$

X-COOH = 87.8324 where X is an alkyl group

X =87.8324- 42 = 42.8324=43

By elimination: CH_3 = 15 CH_3CH_2 = 29 $\underline{\mathbf{CH_3CH_2\ CH_2 = 43}}$

Molecula formula : CH_3CH_2 CH_2COOH
Molecule name : Butan-1-oic acid
Molecular structure



$$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{O}{\overset{\|}{C}}}-O-H$$

5. 10.5 g of an impure sample containing ammonium sulphate (VI) fertilizer was warmed with 250cm3 of o.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

$NaOH(aq) + HCl(aq) \rightarrow NaOH(aq) + H_2O(l)$

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

Moles of HCl = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.5 \times 85}{1000}$ = **0.0425 moles**

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

Equation for reaction with ammonium salt

$2NaOH(aq) + (NH_4)_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2NH_3(g) + 2H_2O(l)$

Mole ratio NaOH(aq): $(NH_4)_2SO_4(aq)$= 2:1

Total moles of NaOH = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.8 \times 250}{1000}$ = **0.2 moles**

Moles of NaOH that reacted with$(NH_4)_2SO_4$ = 0.2 - 0.0425 = **0.1575moles**

Moles $(NH_4)_2SO_4$ = ½ x 0.1575moles = **0. 07875moles**

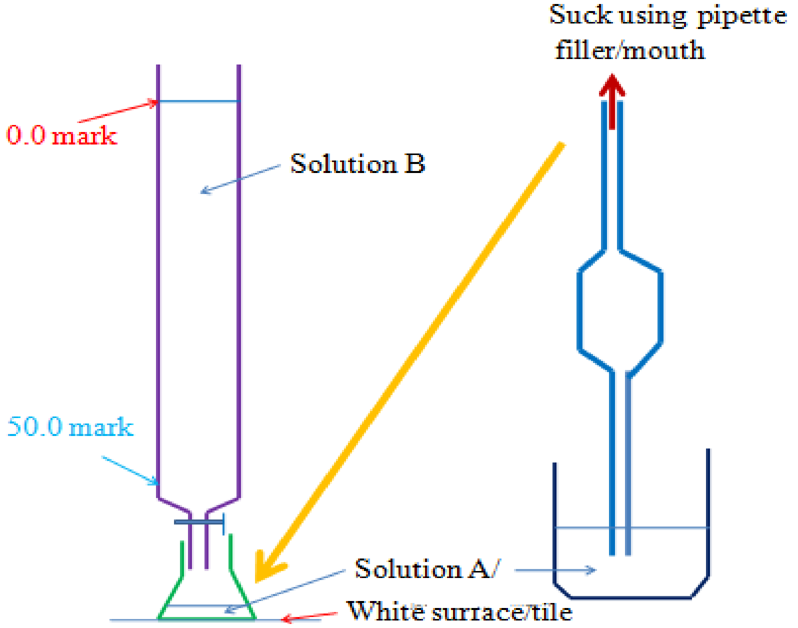
Molar mass $(NH_4)_2SO_4$= **132 gmole^{-1}**

Mass of in impure sample = moles x molar mass =>0. 07875 x 132 = **10.395 g**

Mass of impurities = 10.5 -10.395 = **0.105 g**

% impurities = $\frac{0.105 \times 100}{10.5}$ = **1.0 %**

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



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 (cm3)	20.0	20.0	20.0
Initial burette reading (cm3)	0.0	0.0	0.0
Volume of solution used(cm3)	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

$$\frac{24.0 + 24.0 + 24.0}{3} = \mathbf{24.0\ cm3}$$

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 (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution B used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution B used

$$\text{Average titre} = \frac{\text{Titre 1 + Titre 2 +Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3}\right) = \underline{\mathbf{20.0cm3}}$$

2. How many moles of:

(i)solution A were present in 25cm3 solution.

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

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

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

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

(iii) solution B in moles per litre.

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

(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$(NH_4)_2\ SO_4FeSO_4.xH_2O$
solid C

You are required to

(i)standardize acidified potassium manganate(VII)

(ii)determine the value of x in the formula $(NH_4)_2\ SO_4FeSO_4.xH_2O$.

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 (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution A used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution A used

$$\text{Average titre} = \frac{\text{Titre 1 + Titre 2 +Titre 3}}{3} \Rightarrow \left(\frac{20.0 + 20.0 + 20.0}{3}\right) = \underline{\mathbf{20.0cm3}}$$

2. How many moles of:

(i)solution B were present in 25cm3 solution.

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

(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

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

(iii) solution B in moles per litre.

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

$$= \mathbf{0.025}\ \text{M /moles per litre /moles l}^{-1}$$

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 (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution A used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution A used

$$\text{Average titre} = \frac{\text{Titre 1 + Titre 2 +Titre 3}}{3} \Rightarrow \left(\frac{20.0 +20.0 +20.0}{3} \right) = \underline{\mathbf{20.0cm3}}$$

2. How many moles of:

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

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

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

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

Mole ratio $MnO_4^-(aq): 5Fe^{2+}(aq) = 1:5$ =>

$$\text{Moles of } 5Fe^{2+}(aq) = \frac{\text{Moles of } MnO_4^-(aq)}{5} = \frac{5.0 \times 10^{-4}}{5} \text{ moles} = \underline{\mathbf{1.0 \times 10^{-4}}}\ \mathbf{moles}$$

(iii) solution B in 250cm3.

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

3. Calculate the molar mass of solid C and hence the value of x in the chemical formula $(NH_4)_2SO_4FeSO_4.xH_2O$.
(N=14.0, S=32.0, Fe=56.0, H=1.0 O=16.0)

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

$NH_4)_2SO_4FeSO_4.xH_2O = 8500$

$$284 + 18x = 8500$$

$$8500 - 284 = \frac{8216}{18} = \frac{18x}{18} = \underline{\mathbf{454.4444}}$$

x = **454** (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 (cm3)	**6.5**	**6.5**	**6.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**6.5**	**6.5**	**6.5**

Sample questions

(a) Calculate the average volume of solution N used

$$\frac{6.5 + 6.5 + 6.5}{3} = \mathbf{6.5}\ \text{cm3}$$

(b) How many moles of sodium hydroxide are contained in 25cm3of solution L

Molar mass NaOH =**40**g

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

$$\text{Moles NaOH in 25cm3} = \frac{\text{molarity x volume}}{1000} \Rightarrow \frac{0.5M \text{ x } 25cm3}{1000} = \underline{\mathbf{0.0125}\ \text{moles}}$$

(c)Calculate:

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

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

Moles HCl =Moles NaOH => **0.0125** moles

(ii)the molarity of hydrochloric acid solution N.

$$\text{Molarity} = \frac{\text{moles x 1000}}{6.5} \Rightarrow \frac{\text{0.0125 moles x 1000}}{6.5} = \mathbf{1.9231}\text{M/moledm}^{-3}$$

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 waterto make up the solution up to 200cm3.Label this solution K.Using a clean pipettetransfer 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 (cm3)	**24.5**	**24.5**	**24.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.5**	**24.5**	**24.5**

Sample calculations

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

$$\frac{24.5 + 24.5 + 24.5}{3} = \mathbf{24.5}\text{cm3}$$

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

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

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

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

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

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

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

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

Original moles = $\frac{\text{Original molarity x pipetted volume}}{1000\text{cm3}}$ =>

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

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

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

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

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

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

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

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

From the equation $CaCO_3(s):2HCl(aq)$ = 1:2

=> Moles $CaCO_3(s)$ = $^1/_2$moles HCl
= $^1/_2 \times 0.03582/3.582 \times 10^{-2}$ moles
= **$0.01791 / 1.791 \times 10^{-2}$**moles

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

Molar mass $CaCO_3$ = 100g

Mass $CaCO_3$ = moles x molar mass => 0.01791 /1.791×10^{-2}moles x 100g
= **1.791**g

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

$$\% \, CaCO_3 = \frac{\text{mass of pure} \times 100\%}{\text{Mass of impure}} \Rightarrow \frac{1.791\text{g} \times 100\%}{4.0} = \underline{44.775\%}$$

(d)Sample titration practice 4 (**Multiple titration**)

You are provided with:

(i)sodium L containing 5.0g per litre of a dibasic organic acid $H_2X.2H_2O$.
(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 (cm3)	**24.0**	**24.0**	**24.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.0**	**24.0**	**24.0**

Sample calculations

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

$$\underline{24.0 + 24.0 + 24.0} = \underline{\mathbf{24.0}}\text{cm3}$$

3

(b)Given that the concentration of the dibasic acid is 0.05molesdm^{-3}.determine the value of x in the formula $H_2X.2H_2O$ (H=1.0,O=16.0)

$$\text{Molar mass } H_2X.2H_2O = \frac{\text{mass per litre}}{\text{Moles/litre}} \Rightarrow \frac{5.0\text{g/litre}}{0.05\text{molesdm}^{-3}} = \mathbf{100}\text{g}$$

$H_2X.2H_2O$ =100

X = 100 – ((2 x1) + 2 x (2 x1) + (2 x 16) => 100 – 34 = **66**

(c) Calculate the number of moles of the dibasic acid $H_2X.2H_2O$.

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

(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.

Moles H_2X = $^2/_5$ moles of MnO_4^-

=> $^2/_5$ x 0.0125/1.25 x10^{-2} moles

= **0.005/5.0 x 10^{-3}**moles

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

$$\text{Moles per litre/molarity} = \frac{\text{moles x 1000}}{\text{average burette volume}}$$

$$\Rightarrow \frac{0.005/5.0 \text{ x } 10^{-3}\text{moles x } 1000}{24.0} = \mathbf{0.2083} \text{ molesl}^{-1}\text{/M}$$

Procedure 2

With solution M still in the burette ,pipette 25.0cm3 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 (cm3)	**12.5**	**12.5**	**12.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**12.5**	**12.5**	**12.5**

Sample calculations

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

$$\frac{12.5\ +\ 12.5\ +\ 12.5}{3} = \mathbf{12.5}\text{cm3}$$

(b)Calculations:

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

$$\text{Moles} = \frac{\text{molarity of solution M x average burette volume}}{1000}$$

$$=> \frac{0.2083\text{ molesl}^{-1}/\text{ M x } 12.5}{1000} = \mathbf{0.0026\ /\ 2.5\ x\ 10^{-3}}\text{ moles}$$

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

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$

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 $MnO_4^-(aq)$: $C_2O_4^{2-}(aq)$ = 2:5

=> moles $C_2O_4^{2-}$ = $^5/_2$ moles MnO_4^- => $^5/_2$ x 0.0026 / 2.5 x 10^{-3} moles

= **0.0065 /6.5 x10^{-3}** moles

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

25cm3 pipette volume -> 0.0065 /6.5 x10^{-3} moles
250cm3 ->

$$\frac{0.0065\ /6.5\ x10^{-3}\text{ moles x } 250}{25} = \mathbf{0.065\ /\ 6.5\ x10^{-2}}\text{ moles}$$

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 (cm3)	**24.9**	**24.9**	**24.9**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.9**	**24.9**	**24.9**

Sample calculations

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

$$\frac{24.9 + 24.9 + 24.9}{3} = \mathbf{24.9\ cm3}$$

(b)Calculations:

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

$$\text{Moles} = \frac{\text{molarity of solution P x average burette volume}}{1000}$$

$$\Rightarrow \frac{0.1\ \text{molesl}^{-1} \ \text{x}\ 24.9}{1000} = \underline{\mathbf{0.00249\ /\ 2.49\ x\ 10^{-3}}}\ \text{moles}$$

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

$2NaOH(aq) + H_2C_2O_4(aq) \rightarrow Na_2C_2O_4(g) + 2H_2O(l)$

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

From the stoichiometric equation,mole ratio NaOH(aq): $H_2C_2O_4$ (aq) = 2:1

=> moles $H_2C_2O_4$ = $^1/_2$ moles NaOH => $^1/_2$ x 0.00249 / 2.49 x 10^{-3} moles

= $\underline{\mathbf{0.001245/1.245\ x10^{-3}}}$ moles.

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

25cm3 pipette volume -> 0.001245/1.245 x10^{-3} moles

250cm3 ->

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

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

Molar mass $H_2C_2O_4$ = **90.0**g

Mass of $H_2C_2O_4$ in 250cm3 = moles in 250cm3 x molar mass $H_2C_2O_4$

=>0.01245/1.245 x10^{-2} moles x 90.0

= **1.1205**g

% by mass of sodium ethanedioate

$$=\frac{(\text{Mass of mixture} - \text{mass of } H_2C_2O_4)}{\text{Mass of mixture}} \times 100\%$$

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

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 thereaction 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(cm3)	**12.5**	**25.0**	**37.5**
Initial volume(cm3)	**0.0**	**12.5**	**25.0**

Volume of solution A used(cm3)	**12.5**	**12.5**	**12.5**

Sample questions

Calculate the average volume of solution A used.

$$\frac{12.5+12.5+12.5}{3} = \mathbf{12.5cm3}$$

Theoretical Practice examples

1. 1.0g of dibasic acid $HOOC(CH_2)_xCOOH$ 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 $HOOC(CH_2)_xCOOH$. (C=12.0,H=1.0,O=16.)

Chemical equation

$2NaOH(aq) + H_2X(aq) \rightarrow Na_2X (aq) + 2H_2O(aq)$

Mole ratio $NaOH(aq) : H_2X(aq) = 2:1$

Method 1

$$\frac{Ma\ Va = na}{Mb\ Vb = nb} \Rightarrow \frac{Ma \times 25.0}{0.06 \times 30.0} = \frac{1}{2} \Rightarrow Ma = \frac{0.06 \times 30.0 \times 1}{25.0 \times 2}$$

Molarity of acid = **0.036M/Mole l^{-1}**

Mass of acid per lite = $\frac{1.0 \times 1000}{250}$ = **4.0 g/l**

0.036M/ Mole l^{-1} -> 4.0 g /l

1 mole= molar mass of $HOOC(CH_2)_xCOOH$ = $\frac{4.0 \times 1}{0.036}$ = **111.1111** g

Molar mass $(CH_2)_x$ = 111.1111 – (HOOCCOOH = 90.0) = **21.1111**

$(CH_2)_x = 14x = \frac{21.1111}{14} =$ **1.5 = 1 (whole number)**

Method 2

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

Moles of Hydrochloric acid = $^1/_2$ x 1.8 x 10^{-3}moles = 9.0 x10^{-4}moles

Molarity of Hydrochloric acid = $\frac{\text{moles x 1000}}{\text{Volume}}$ = $\frac{9.0\text{ x}10^{-4}\text{moles x1000}}{25}$

Molarity of acid = **0.036M/Mole l^{-1}**

Mass of acid per lite = $\frac{1.0\text{ x1000}}{250}$ = **4.0 g/l**

0.036M/ Mole l^{-1} -> 4.0 g /l

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$(CH_2)_x$ = 14x = $\frac{21.1111}{14}$ = **1.5 = 1 (whole number)**

2. 20.0cm3 of 0.05 M acidified potassium manganate(VII)solution oxidized 25.0cm3 of Fe^{2+}(aq) ions in 40.0g/l of impure Iron (II)sulphate(VI) to Fe^{3+}(aq) ions. Calculate the percentage impurities in the Iron (II)sulphate(VI).

MnO_4^- (aq) + $8H^+$(aq)+ $5Fe^{2+}$(aq)-> $5Fe^{3+}$(aq) + Mn^{2+}(aq) + $4H_2O$(aq)
Fe=56.0,S= 32.0, O=16.0).

Moles of MnO_4^- (aq) = $\frac{\text{Molarity x volume}}{1000}$ = $\frac{0.05\text{ x }20.0}{1000}$ = **0.001 Moles**

Mole ratio MnO_4^- (aq): $5Fe^{2+}$(aq)= 1:5

Moles $5Fe^{2+}$(aq) **=** 5 x0.001_ **=** **0.005 Moles**

Moles of $5Fe^{2+}$(aq) per litre/molarity = $\frac{\text{Moles x 1000}}{\text{Volume}}$ = $\frac{0005\text{ x }1000}{25.0}$

= 0.2 M/ Moles/litre

Molar mass =$FeSO_4$=**152 g**

Mass of in the mixture = Moles x molar mass => 0.2 x 152 = **30.4** g

Mass of impurity = 40.0 – 30.4 =**9.6** **g**

% impurity = $\frac{9.6\text{ g x}100}{40.0}$ = **24.0 %** **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

KOH(aq) + HCl(aq) -> KCl(aq) + H_2O(l)

Moles of HCl = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.12\text{ x }25.0}{1000}$ = **0.003**/**3.0 x 10 $^{-3}$** moles

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

Moles KOH =**0.003**/**3.0 x 10 $^{-3}$** moles

Method 1

Molar mass KOH =**56.0**g

Mass KOH in 25cm3 =0.003/3.0 x 10 $^{-3}$ moles x56.0 = 0.168g

Mass KOH in 1000cm3/1 litre = $\frac{0.168\text{ x}1000}{20}$= **8.4 g/l**

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = $\frac{1.3\text{ x }100}{9.7}$ = **13.4021%**

Method 2

Moles KOH in 1000cm3 /1 litre = $\frac{\text{Moles in 20cm3 x 1000}}{20}$ =>$\frac{0.003\text{ x }1000}{20}$

=**0.15M/Moles /litre**

Molar mass KOH =**56.0**g

Mass KOH in 1000/1 litre = 0.15M/Moles /litre x 56.0 = **8.4g/l**

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = $\frac{1.3\text{ x }100}{9.7}$ = **13.4021%**

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

$GCO_{3(s)}$) + 2$HCl_{(aq)}$ -> $GCl_{2\,(aq)}$ + CO_2 (g) + $H_2O_{(l)}$

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)

Moles of HCl = $\frac{\text{Molarity x volume}}{1000}$=> $\frac{1 \text{ x}20}{1000}$ = **0.02 moles**

Mole ratio HCl; GCO_3 = 2:1

Moles of GCO_3= $\frac{0.02 \text{ moles}}{2}$ = **0.01moles**

Molar mass of GCO_3 = $\frac{\text{mass}}{\text{moles}}$ => $\frac{1}{0.01\text{moles}}$ = **100** g

G= GCO_3 - CO_3 =>100g – (12+ 16 x3 = 60) = **40**(**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

$NaOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

Moles of NaOH = $\frac{\text{Molarity x volume}}{1000}$=> $\frac{0.1 \text{ x}20}{1000}$ = **0.002 moles**

Mole ratio HCl; NaOH = 1:1

Excess moles of HCl = **0.002 moles**

25cm3 -> 0.002 moles

1000cm3 -> $\frac{1000 \text{ x } 0.002}{25\text{cm}3}$ = **0.08moles**

Original moles of HCl = Molarity x volume => 1M x 1litre = **1.0 moles**

Moles of HCl reacted with MCO_3 = **1.0 - 0.08 moles** = **0.92moles**

Chemical equation

$MCO_{3(s)}) + 2HCl_{(aq)} \rightarrow MCl_{2\ (aq)} + CO_2\ (g) + H_2O_{(l)}$

Mole ratio $MCO_{3(s)}$) : $HCl_{(aq)}$ =1:2

Moles of MCO_3 = 0.92moles => **0.46moles**

2

Molar mass of MCO_3= $\frac{\text{mass}}{\text{moles}}$ => $\frac{46\text{g}}{0.46\text{moles}}$ = **100 g**

M= $MCO_3 - CO_3$ =>100g – (12+ 16 x3 = 60) = **40**

6. 25.0cm3 of a mixture of Fe^{2+} and Fe^{3+} ions in an aqueous salt was acidified with sulphuric(VI)acid then titrated against potassium manganate(VI).The salt required 15cm3 ofe0.02M potassium manganate(VI) for complete reaction.

A second 25cm3 portion of the Fe^{2+} and Fe^{3+} 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 Fe^{2+} and Fe^{3+} ion in the solution on moles per litre.

Mole ratio Fe^{2+} :$Mn0_4^-$ = 5:1

Moles $Mn0_4^-$ used = $\frac{0.02 \times 15}{1000}$ = **3.0×10^{-4}** moles

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

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

Since Zinc reduces **Fe^{3+}** to **Fe^{2+}** in the mixture:

Moles $Mn0_4^-$ that reacted with all **Fe^{2+}**= $\frac{0.02 \times 19}{1000}$ = **3.8×10^{-4}** moles

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

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

Molarity of **Fe^{3+}** **=** $\frac{1.6 \times 10^{-5} \text{ moles} \times 1000}{25}$ = **4.0×10^{-4} moles l^{-1}**

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