CHAPTER #4 : STOICHOMETRY

'Branch of chemistry which deals with the study of quantitative relationship between reactant and products in a balanced chemical equation.'

↳ Derived from greek word stoicheion meaning element, metron means measurement.

I.E.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

This reaction can be explained in 4 ways

1. Through moles
2. Through molecules
3. Through mass
4. Through standard temperature and pressure (STP)

MOLES : It is a measuring unit of amount of substance.

So, 1 mol = 6.022×10^{23} atoms / particles / molecules / ions

↳ Atomic mass : It is the mass of a single atom of an element. i.e. H^1, N^7.

↳ Formula mass : It is the mass of an ionic compound / molecule and requires calculation. i.e. H_3O.

↳ Molecular mass : It is the mass of a covalent compound / molecule and requires calculation. i.e. H_2O.

MOLAR VOLUME : 'It is the one mole of any gas at STP [standard temperature and pressure] occupying a volume of 22.414 dm^3/mol.'

At STP, T = 0°C / 273 k, P = 1 atm

At RTP, T = 25°C / 298 k, P = 1 atm

RTP is regular temperature and pressure.

$n = \frac{V}{V_m}$ where n is number of moles, V is volume and V_m is molar volume.

Example 4.1.

Q What is the volume of 2.5 moles of Cl_2 at STP?

Sol

$n = 2.5$

$V = ?$

$V_m = 22.414\ dm^3$

So,

$n = \frac{V}{V_m} \Rightarrow n \times V_m = V$

$2.5 \times 22.414 = V \Rightarrow \boxed{56.035\ dm^3 = V}$

$22.414\ dm^3 \quad 22.414 \times 1000 \quad 22414\ cm^3$

- Multiply when going from big to small
- Divide when going from small to big

MOLAR MASS AND DENSITY OF GAS:

Molar mass is basically formula mass and molecular mass. 'Molar mass is the mass of one mole of a substance'.

As we know from, $\rho = m/v$ that $\boxed{\text{density} \propto \text{mass}}$ if density is higher than that means the mass of the gas must be high. As we know that the molar mass does not change since its basically sum of atomic mass of all the elements involved in a molecule and atomic mass stays the same since each element's atomic mass is already experimentally determined and fixed.

$\boxed{\rho = m/v}$ .....(i)

and $n = \frac{\text{mass in g}}{\text{molar mass}}$ so $n \times \text{molar mass} = \text{mass}$

n x molar mass = mass can be written as

$[n \times M = m] \cdots (ii)$

Ideal gas law

$$\boxed{V = \frac{nRT}{P}} \cdots (iii)$$

n is no. of moles.
R is gas constant.
T is temperature.
P is pressure.

From eq (i) we know that $V = \frac{m}{\rho}$

$$\frac{m}{\rho} = \frac{nRT}{P}$$

we also know that $n = \frac{m}{M}$ from eq (ii)

$$\frac{m}{\rho} = \frac{m}{M}\frac{RT}{P}$$

$$\frac{\rho}{m} = \frac{MP}{mRT} \quad \therefore \text{Since we want } \rho \text{ isolated.}$$

$$\rho = \frac{\cancel{m}MP}{\cancel{m}RT}$$

$$\boxed{\rho = \frac{MP}{RT}} \cdots (iv)$$

So, $\rho \propto M$

However molar mass cannot increase or decrease but a higher molar mass indicates higher density whereas lower molar mass indicates lower density.

Example 4.2 :

Q. Calculate gram molecular mass of a gas which has density of $1.43\ g/dm^3$ at STP.

Sol :

$\rho = 1.43\ g/dm^3$

$V = 22.414\ dm^3$

$m = ?$

$$\rho = \frac{m}{V} \Rightarrow \rho \times V = m$$

$$1.43 \times 22.414 = m \Rightarrow \boxed{32.05g = m}$$

* formula mass, molecular mass expressed in grams is molar mass

Self Assessment : 4.1

Q. How many moles of O_2 molecules are there in $20.0dm^3$ of O_2 gas at STP?

$V = 20.0dm^3$

$V_m = 22.414\ dm^3$

$n = ?$

Sol:

$$n = \frac{V}{V_m} \Rightarrow \frac{20}{22.414} = n$$

$$\boxed{0.892 = n}$$

Q. What volume does 0.6 mole of H_2 gas occupy at STP?

Sol:

$n = 0.6$

$V_m = 22.414\ dm^3$

$V = ?$

$$n = \frac{V}{V_m} \Rightarrow n \times V_m = V$$

$$0.6 \times 22.414 = V \Rightarrow \boxed{13.45\ dm^3 = V}$$

Additional topic for understanding :

Listed below are the laws that lay out the foundation for stoichiometry.

STOICHIOMETRY AND LAW OF CONSERVATION OF MASS :

Law of conservation of mass states that :
"Mass can neither be created nor be destroyed" in a chemical reaction.
Stoichiometry is based on this law as it is the study of quantitative relation between reactant and products. Hence we can say that

Amount of reactants = Amount of products

So stoichiometry ensures that amount of reactants and amount of products are equal in terms of moles and mass.

↳ Chemists accurately calculate the amount of reactant needed and the amount of product formed by stoichiometry.

I.E.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Through stoichiometric calculations

This balanced chemical equation tells us about : ↑

(i) 2 mol of Hydrogen react with 1 mol of oxygen to produce 2 mol of water.
(ii) $2 \times 6.022 \times 10^{23}$ molecules of Hydrogen react with 6.022×10^{23} molecules of oxygen to produce $2 \times 6.022 \times 10^{23}$ molecules of water.
(iii) 4g of hydrogen react with 32g of oxygen to produce 36g of water.
(iv) At STP $2 \times 22.414\ dm^3$ of Hydrogen gas react with $22.414\ dm^3$ of oxygen to produce $2 \times 22.414\ dm^3$ of water vapour.

GAY LUSSAC'S LAW : 1st And 2nd

↳ This law is formulated by a french scientist Joseph gay Lussac in 1808.
↳ This is law of pressure- temperature relationship.

This law states that :
'The pressure exerted by a gas is proportional to the temperature of the gas when the mass is fixed and the volume is constant.'

Mathematically,

↳ $P \propto T$

↳ $P = kT$ ∴ k is proportionality constant

For initial pressure and temperature,

↳ $P_i = kT_i$

↳ $\frac{P_i}{T_i} = k$

For final pressure and temperature,

↳ $P_f = kT_f$

↳ $\frac{P_f}{T_f} = k$

Since both R.H.S are same/equal then,

↳ $\boxed{\frac{P_i}{T_i} = \frac{P_f}{T_f}}$

The second gay Lussac's law formulated around the same year in 1808 - 1809.

↳ This is law of combining volumes. (only for gases)

* This law states that ;
'When gases react together at constant temperature and pressure, the volumes of the reacting gases and the volumes of the products (if gaseous) can be expressed in simple whole-number ratios.'

By simple whole ratio, the numbers will not be in decimal or fractional form rather they will be in normal form i.e. 1, 2, 3, 4......

In this way the ratio can be easily further simplified if possible i.e. 4:2 can be 2:1.

AVOGADRO'S LAW :

This law is formulated by Amedeo Avogadro in 1811.

↳ It is a hypothesis on gas volumes and molecules.
↳ Gray lussac's law provided a framework for this law.

*. This law states that ;
' At constant temperature and pressure, equal volumes of gases contain an equal number of molecules.'
This law is applicable to both same and different gases.

Mathematically,

↳ $V \propto n$

↳ $V = kn$ ∴ k is proportionality constant

For initial Volume and molecules,

↳ $V_i = kn_i$

↳ $\frac{V_i}{n_i} = k$

For final Volume and molecules,

↳ $V_f = kn_f$

↳ $\frac{V_f}{n_f} = k$

Since both R.H.S are same/equal then,

↳ $\boxed{\frac{V_i}{n_i} = \frac{V_f}{n_f}}$

So we can say that moles increase with the increase in volume and moles decrease with the decrease in volume hence in pressure-temperature relationship law, the moles will be constant since volume is constant.

BOYLE'S LAW ;

This law was formulated in 1662 by Robert boyle.

↳ This law states that ;

'The volume of a gas is inversely proportional to the pressure at constant temperature.'

Mathematically,

↳ $V \propto \frac{1}{P}$

↳ $V = k/P$ ∴ k is proportionality constant

For initial volume and pressure,

↳ $V_i = k/P_i$

↳ $V_i P_i = k$

For final volume and pressure,

↳ $V_f = \frac{k}{P_f}$

↳ $V_f P_f = k$

Since both R.H.S are same / equal then,

↳ $\boxed{V_i P_i = V_f P_f}$

CHARLE'S LAW ;

This law was formulated in 1787 by Jacques Charles.

↳ This law states that ;

'The volume of a gas is directly proportional to the temperature at constant pressure.'

Mathematically,

↳ $V \propto T$

↳ $V = kT$ ∴ k is proportionality constant

For Initial volume and temperature,

↳ $V_i = kT_i$

$$\hookrightarrow \frac{V_i}{T_i} = k$$

For final volume and temperature,

$$\hookrightarrow V_f = kT_f$$

$$\hookrightarrow \frac{V_f}{T_f} = k$$

Since both R.H.S are same / equal then,

$$\hookrightarrow \boxed{\frac{V_i}{T_i} = \frac{V_f}{T_f}}$$

IDEAL GAS LAW :

This law was formulated in 1857 and by Rudolf Clausius.

↳ This law combines Avogadro's law, Boyle's law and Charles law.

'This law describes the relationship between pressure (P), volume (V), number of moles (n), gas constant (R), temperature (T) for an ideal gas.'

Ideal gas means : A hypothetical gas

* Has no volume
* Does not have attractive / IMF.
* Moves randomly with no specific direction
* Behaves predictably ; Follows Boyle's law, Charle's law and Avogadro's law with no exceptions.

Mathematically,

$V \propto n$ ∴ Avogadro's law

$V \propto T$ ∴ Charle's law

$V \propto \frac{1}{P}$ ∴ Boyles law

If we combine these

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

By introducing the gas constant to remove proportionality and change it into equality :

↳ $V = \frac{RnT}{P}$

where $R = 8.32\ J/mol \cdot k$

Rearrange ,

$\boxed{PV = nRT}$ ... Ideal gas law

Dalton's law :

↳ This law was formulated by John Dalton in 1803.

↳ This law is named as the law of multiple proportions.

It states that ;

'When two elements combine to form more than one compound , the masses of one element that combine with a fixed mass of the other element are in simple whole number ratios'.

I.E.

Carbon combines with oxygen to form two compounds ; carbon monoxide (CO) and carbon dioxide (CO_2).

CO ; 12 grams of carbon combine with 16 grams of oxygen.

CO_2 ; 12 grams of carbon combine with 32 grams of oxygen.

↳ The mass of oxygen can be written in simple whole number ratio ,

$\frac{32}{16}$ or $32 : 16$

This can be further simplified

$\frac{2}{1}$ or 2 : 1

↳ This law is for same elements involved in two different reactions where mass of one element stays the same and mass of the second element varies.

PROUST'S LAW :

↳ This law was formulated by Joseph Proust in 1799.

↳ It is also known as law of definite proportions. It states that;

'A chemical compound always contains its component elements in fixed ratio by mass, regardless of its source or method of preparation.'

I.E.

$H_2O = 2(1) + 16 = 2 + 16 = 18$ = Molar mass

$H_2 = \frac{2}{18}$ mass ratio

2 : 1

$O = \frac{16}{18}$ (2 hydrogen atoms) : (1 oxygen atom)

⇒ Percentage composition of $H_2 = \frac{2}{18} \times 100$

$= 11.1\%$

Percentage composition of $O = \frac{16}{18} \times 100$

$= 88.9\%$

This whole composition of water will remain the same regardless of ;

Source = rainwater, seawater, synthetic (laboratory)

Method = Chemical reaction, Electrolysis (both methods are for purification of water).

* Listed above are laws directly or indirectly related to stoichiometry. These laws apply completely to stoichiometric calculations but can have exceptions for non-stoichiometric calculations.

↳ These laws hold true with no exceptions only for stoichiometric calculations.

STOICHIOMETRIC CALCULATIONS AND MOLE RATIO:

Mole ratio refers to the ratio between the amount of moles of either reactants with each other or of reactant with product or of products with each other.

I.E.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

(i) Reactant to reactant mole ratio:

mole ratio between CH_4 and O_2

1 : 2 or 1/2

(ii) Reactant to product mole ratio:

mole ratio between CH_4 and CO_2

1 : 1 or 1/1

mole ratio between CH_4 and H_2O

1 : 2 or 1/2

mole ratio between O_2 and CO_2

2 : 1 or 2/1

mole ratio between O_2 and H_2O

2 : 2 or 2/2

1 : 1 or 1/1

(iii) Product to product mole ratio:

mole ratio between CO_2 and H_2O

1 : 2 or 1/2

↳ We can also take mole ratio between O_2 and CH_4, we donot need to fallow a specific order for taking mole ratio. i.e. 1 : 2 can be 2 : 1
Stoichiometric Calculations basically have the principle that amount of reactants = amount of products. So basically for stoichiometric calculations and for mole ratios a balanced chemical equation is a must.

• **CONVERSION FACTOR**: It is the mole ratio of the quantity of the required substance to the quantity of the given substance in a balanced chemical equation.
Mathematically,
Conversion factor = coefficient of substance whose quantity is to be determined ÷ coefficient of substance whose quantity is given.

$$\text{Conversion factor} = \frac{\text{coefficient of required substance}}{\text{coefficient of given substance}}$$

$$C.F = \frac{\text{find}}{\text{given}}$$

• **DESIRED QUANTITY**: The amount of substance that is required / desired can be found by multiplying the quantity of the given substance (given value) with the conversion factor of a balanced chemical equation.
Mathematically,
Desired quantity = given quantity × conversion factor

$$D.q = g.q \times C.F$$

*. C.F and D.q or C.F and g.q must always be in same units.

Hence mole ratio / conversion factor and desired quantity are components of stoichiometric calculations which ensure accurate stoichiometric predictions in chemical reactions. (refers to using the balanced chemical equation to predict the quantities of reactants or products.)

Example 4.3 :

Q. When 100g of magnesium is treated with dilute hydrochloric acid. What volume of hydrogen can be collected at STP?

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

Sol :

$$n = \frac{V}{V_m}, \quad n = \frac{m}{M}$$

we need V of H_2 = ? , mg has $m = 100g$, $V_m = 22.414\ dm^3$

$$\hookrightarrow n = \frac{\text{mass in g}}{\text{Molar mass}} = \frac{100g}{24 g/mol} = 4.17$$

$$n = 4.2\ mol \quad \because g \div g/mol \Rightarrow \cancel{g} \times mol/\cancel{g} \Rightarrow mol$$

$$n = \frac{V}{V_m} \Rightarrow n \times V_m = V \Rightarrow 4.2 \times 22.414 = V$$

$$\boxed{94.14\ dm^3 = V}$$

Self Assessment : 4.2

Q. Quantitative information of the following reaction

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

We can explain this reaction in 4 ways :

(i) 1 mole of C_3H_8 react with 5 moles of O_2 to produce 3 moles of CO_2 and 4 moles of H_2O.

(ii) 6.022×10^{23} particles / molecules of C_3H_8 reacts with $5 \times 6.022 \times 10^{23}$ particles / molecules of O_2 to produce $3 \times 6.022 \times 10^{23}$ particles / molecules of CO_2 and $4 \times 6.022 \times 10^{23}$ particles / molecules of H_2O.

i. At STP, 22.414 dm^3 of C_3H_8 reacts with 5×22.414 dm^3 of O_2 to produce 3×22.414 dm^3 of CO_2 and 4×22.414 dm^3 of H_2O.

ii. 44g of C_3H_8 reacts with $5\times 32g$ (160g) of O_2 to produce $3\times 44g$ (132g) of CO_2 and $4\times 18g$ (72g) of H_2O.

2. Differentiate between molecular mass and molar mass

Molecular Mass	Molar Mass
1. The molecular mass or molecular weight is the mass of a single molecule of a substance.	1. The molar mass is the mass of a mole of a substance. i.e it can be mass of a molecule or an atom.
2. It remains constant for each particular substance.	2. It also remains constant for each particular substance.
3. It is typically expressed in atomic mass units (amu).	3. It is expressed in grams per mole (g/mol).
4. It is calculated only for molecules.	4. It is calculated for both molecules and atoms.

3. What mass of Zn is needed to produce 100cm^3 of H_2 at STP?

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

$$n = \frac{V}{V_m} \quad , \quad n = \frac{m}{M}$$

V_m of H_2 = 100 cm^3 , m of Zn = ?

$$n = \frac{100cm^3}{22414\ cm^3} \Rightarrow 0.00446\ mol$$

$$n\times M = m \Rightarrow 0.00446\times 65.4 = \text{mass in of Zn}$$

$$\boxed{0.292g = m}$$

Mole - Mole Ratio :

Mole - Mole ratio or conversion factor basically serves as a way to correctly calculate the required amount in a balanced chemical equation.

Example 4.5 :

Q. How much MgO is formed when 4 moles of magnesium react with excess of O_2. Also calculate moles of O_2 required.

$$2Mg + O_2 \longrightarrow 2MgO$$

① ↳ $C.F = \frac{\text{find}}{\text{given}}$, desired quantity = g.q × C.F

$C.F = \frac{2}{2} = 1$ mole

desired quantity = given quantity × conversion factor

D.q = 4 × 1

D.q = 4 = MgO formed

② ↳ $C.F = \frac{\text{find}}{\text{given}}$, desired quantity = g.q × C.F

$C.F = \frac{1}{2} = 0.5$ mole

D.q = g.q × C.F = 4 × 0.5 = 2

= 2 moles of O_2 required

Conclusion :

Hence 4 moles of Mg react with 2 moles of O_2 to produce 4 moles of MgO.

Mole - Mass Calculations :

Example 4.6

Q. Potassium Chlorate is used in making matches and dyes, On decomposition it produces KCl and oxygen gas.

If 25g of $KClO_3$ is decomposed then calculate:

(i). O_2 moles produced.

(ii). mass in gram of O_2 produced.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

$C.F = \frac{\text{find}}{\text{given}}$, desired quantity $= g.q \times C.F$

Since we know that given quantity and conversion factor must be in same units.

$$n = \frac{m}{M} \Rightarrow n = \frac{25}{39+35.5+48} = \frac{25}{122.5}$$

$n = 0.204$ moles $=$ given quantity

↳ $C.F = \frac{\text{find}}{\text{given}} = \frac{3}{2} = 1.5$ moles

$D.q = g.q \times C.F = 0.204 \times 1.5$

$\boxed{D.q = 0.306 \text{ moles of } O_2}$

↳ Now to convert moles into mass or derive mass from moles,

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

$0.306 \times 32 =$ mass in grams

$\boxed{9.8g = m \text{ of } O_2}$

Self Assessment : 4.3

1. CO can reduce Iron oxide into Iron metal. How many moles of CO are needed to reduce 5 moles of Fe_2O_3.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

$C.F = \frac{\text{find}}{\text{given}}$, desired quantity $= g.q \times C.F$

$C.F = \frac{3}{1} = 3$ moles

$D.q = 5 \times 3 = 15$ moles of CO are required

Conclusion.
15 moles of CO are required to reduce Fe_2O_3 5 mol into Fe.

Q. Calculate mass amount of Phosphine (PH_3) that can be prepare when 1 mole of calcium phosphide (Ca_3P_2) reacts with excess of H_2O.

$\hookrightarrow \; Ca_3P_2 + 6H_2O \longrightarrow 3Ca(OH)_2 + 2PH_3$

$C.F = \frac{\text{find}}{\text{given}}$, Desired quantity $= g.q \times C.F$

$C.F = \frac{2}{1} = 2$ moles.

$D.q = g.q \times C.F = 1 \times 2 = 2$

$D.q = 2$ moles of PH_3.

Now calculate mass in grams from moles

$n = \frac{m}{M} \Rightarrow 2 = \frac{m}{31 + 3.024}$

$2 = \frac{m}{34} \Rightarrow 2 \times 34 = m$

$\boxed{68g} = m$ of PH_3

Q. Methanol burns according to the equation

$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$

If 3.50 moles of methanol are burnt in oxygen, calculate
(i) moles of oxygen used
(ii) moles of water produced

$C.F = \frac{\text{find}}{\text{given}}$, $D.Q = g.q \times C.F$

(i) $C.F = \frac{3}{2} = 1.5$ moles.

$D.Q = 3.50 \times 1.5 = \boxed{5.25 \text{ moles of } O_2}$

2. $C \cdot F = \frac{4}{2} = 2$ moles.

$D \cdot Q = 3.5 \times 2 = \boxed{7 \text{ moles of } H_2O}$

2. Calculate the number of molecules and volume of O_2 at STP produced by thermal decomposition of 490 grams of $KClO_3$.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

$n = \frac{V}{V_m}$, $C \cdot F = \frac{\text{find}}{\text{given}}$, $D \cdot q = g \cdot q \times C \cdot F$

$n = \frac{m}{M}$, $n = \frac{N}{N_A}$

As we know that C.F and g.q must be in same units.

$n = \frac{490}{39 + 35.5 + 48} = \frac{490}{122.5} = 4$

$n = 4$ moles of $KClO_3$

Now to find number of molecules and volume of O_2 at STP.

1. $C \cdot F = \frac{\text{find}}{\text{given}} = \frac{3}{2} = 1.5$ moles

$D \cdot q = g \cdot q \times C \cdot F = 4 \times 1.5 = 6$

$\boxed{D \cdot q = 6 \text{ moles}}$

So for number of particles,

$n \times N_A = N \Rightarrow 6 \times 6.022 \times 10^{23} = N$

$3.61 \times 10^{24} = N$ of O_2

So for volume,

$n \times V_m = V$, $6 \times 22.414 \, dm^3 = V$

$\boxed{134.5 \, dm^3 = V}$

Example 4.7

Q. 20g of H_2SO_4 on dissolving in water ionizes completely. Calculate

i). No. of H_2SO_4 molecules

ii). No. of H^+ and SO_4^{-2}

iii). Mass of invidual ions

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{-2}$$

$$n = \frac{m}{M}, \quad n = \frac{N}{N_A}, \quad C.F = \frac{\text{find}}{\text{given}}$$

$$D.q = g.q \times C.F$$

As we need C.F and g.q in same units so,

$$n = \frac{m}{M} \Rightarrow n = \frac{20}{2.02 + 32.07 + 64}$$

$$n = \frac{20}{98.09} \Rightarrow 0.204 \text{ moles}$$

$$n = 0.204 \text{ moles of } H_2SO_4$$

①. $n \times N_A = N \Rightarrow 0.204 \times 6.022 \times 10^{23} = N$

$$\Rightarrow \boxed{N = 1.23 \times 10^{23} \text{ malecules of } H_2SO_4}$$

②. $C.F = \frac{\text{find}}{\text{given}} = \frac{2}{1} = 2 \text{ moles}$

$D.q = g.q \times C.F = 0.204 \times 2 = 0.408 \text{ moles}$

$n \times N_A = N \Rightarrow 0.408 \times 6.022 \times 10^{23} = N$

$$\boxed{N = 2.46 \times 10^{23} \text{ molecules of } H^+}$$

$C.F = \frac{\text{find}}{\text{given}} = \frac{1}{1} = 1 \text{ mole}$

$D.q = g.q \times C.F = 1 \times 0.204 = 0.204 \text{ moles}$

$n \times N_A = N \Rightarrow 0.204 \times 6.022 \times 10^{23} = N$

$$\boxed{N = 1.23 \times 10^{23} \text{ malecules of } SO_4^{-2}}$$

i) $n = \frac{m}{M} \Rightarrow n \times M = m$

$0.408 \times 1.008 = m \Rightarrow 0.4113\,g = m$

$\boxed{m = 0.4113\,g \text{ of } H^+}$

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

$0.204 \times 96.07 = m \Rightarrow 19.598\,g = m$

$\boxed{m = 19.6\,g \text{ of } SO_4^{-2}}$

MOLARITY: It is defined as the number of moles of solute per dm^3 of solution.

Mathematically,

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in } dm^3}$$

$$M = \frac{n}{V} = \frac{\text{moles}}{dm^3} = M \qquad \because cm^3 \div 1000 \text{ gives } dm^3$$

↳ It is basically a measure of the concentration of a solute in a solution. Hence used to express the concentration of solutions.

↳ It tells us how many moles of a substance are dissolved in $1\,dm^3$ of solution.

SOLUTION STOICHIOMETRY:

Solution stoichiometry involves calculating the amounts of reactants and products in chemical reactions that take place in solutions. It is based on the principle of stoichiometry but specifically deals with reactions where the substances are dissolved in a solvent i.e. solute-solvent relationship.

SOLUTION is a homogenous mixture where a solute (the substance being dissolved) is dissolved in a solvent (the liquid in which the solute dissolves).

PERCENTAGE COMPOSITION expresses the concentration of the solute in a mixture / solution as a percentage of the total amount.

↳ Four types of % composition are discussed below:

1. Mass Over Mass percentage (m/m %):

⇒ Compares the mass of the solute to the total mass of the solution.

↳ Useful for solid-solid or solid-liquid mixtures.

Mathematically,

$$\text{Mass } \% = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

2. Mass Over Volume Percentage (m/v %):

⇒ Compares the mass of the solute to the volume of the solution.

↳ Useful for solid-solid mixture.

$$\text{Mass / Volume } \% = \frac{\text{mass of solute}}{\text{volume of solution}} \times 100$$

3. Volume Over Volume percentage (v/v %):

⇒ Compares the volume of the solute to the total volume of the solution.

↳ Useful for liquid-liquid mixtures.

$$\text{Volume } \% = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

4. Volume Over Mass Percentage (v/m %):

⇒ Compares the volume of the solute to the total mass of the solution.

↳ Useful in specific industrial contexts.

$$\text{Volume / Mass } \% = \frac{\text{volume of solute}}{\text{mass of solution}} \times 100$$

Example 4.8 :

Q. Urea (NH_2CONH_2) is a white solid used as fertilizer and starting material for synthetic plastic. A solution contains 40g urea dissolved in $500\ cm^3$ of Solution. Calculate the molarity of this solution.

$$n = \frac{m}{M} \quad , \quad M = \frac{n}{V}$$

$$n = \frac{40g}{14+2.02+12+16+14+2.02} = \frac{40g}{60.04\ g/mol}$$

$$n = 0.67 \text{ moles}$$

$$M = \frac{n}{V} = \frac{0.67\ mol}{0.5\ dm^3} \qquad \because \frac{500\ cm^3}{1000} = 0.5\ dm^3$$

$$\boxed{M = 1.34M}$$

Example 4.9 :

Q. When $250cm^3$ of 0.1M $AgNO_3$ solution is added to an excess of NaCl solution. What mass of AgCl will be formed?

$$\hookrightarrow M = \frac{n}{V} \quad , \quad n = \frac{m}{M} \quad , \quad C.F = \frac{\text{find}}{\text{given}} \quad , \quad D.q = g.q \times C.F$$

$$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$$

$$M \times V = n \Rightarrow 0.1M \times 0.25\ dm^3 = n$$

$$0.025 \text{ moles} = n$$

$$C.F = \frac{\text{find}}{\text{given}} = \frac{1}{1} = 1 \text{ mole}$$

$$D.q = g.q \times C.F = 0.025 \times 1 = 0.025$$

$\because 250 \div 1000 \rightarrow 0.025$ $\quad D.q = 0.025$ moles of AgCl

$$n = \frac{m}{M} \Rightarrow n \times M = m \Rightarrow 0.025 \times (107.9 + 35.5)$$

$$\boxed{m = 3.6\ g \text{ of AgCl}}$$

MOLARITY AND DILUTION

Molarity is basically concentration of a solution while dilution (reducing the concentration of a solution or mixture by adding more solvent), basically decreases with the increase in concentration (increasing the concentration of a solution by adding more solute or by removing the solvent) By the formula of molarity,

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

the 'n' number of moles basically stay the same throughout the solution because of the principle of stoichiometry. However 'M' and 'v' molarity and volume can vary, if we increase molarity it means the volume is less. Basically more volume decreases molarity and less volume increases molarity. Thus,

$$M \propto \frac{1}{V}$$

=> Because of the principle of stoichiometry the total number of moles stay the same throughout the solution but if volume increases / more solvent is added, the number of moles of solute decreases per dm^3 of solution since the solute is now spread over a large volume of solution and the total number of moles are constant so now the same amount of solute has to arrange itself on increased amount of solution. The 'n' can be derived from the formula of molarity,

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

$$MV = n$$

Within the same reaction 'n' of reactants can be found by,

$$M_1V_1 = n_1$$

Now within the same reaction 'n' of products can be found by,

$$M_2V_2 = n_2$$

Since we know that Amount of reactants = Amount of products so we can say that,

$$n_1 = n_2$$

$$M_1V_1 = M_2V_2$$

⇒ If n_1 and n_2 both are reactants then by mole ratio,

$$\frac{n_1}{n_2} \quad \therefore \quad n_1 = M_1V_1, \quad n_2 = M_2V_2$$

⇒ Thus it can be written as,

$$\frac{n_1}{n_2} = \frac{M_1V_1}{M_2V_2}$$

$$M_2V_2\,n_1 = M_1V_1\,n_2$$

$$\frac{M_2V_2}{n_2} = \frac{M_1V_1}{n_1}$$

⇒ Rearranging,

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

Example 4.10

Q. What volume of 0.5M sodium sulphate will react with 275cm^3 of 0.25M barium chloride solution to completely precipitate Ba^{2+} in solution.

$$Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2NaCl$$

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

$$0.5M \times V_1 = 0.25M \times 275cm^3$$

$$V_1 = \frac{0.25M \times 275cm^3}{0.5M}$$

$$V_1 = 137.5cm^3$$

Self Assessment : 4.3

Q. Calculate the molarity of a solution that contains

(i). 1.5 moles of $KClO_3$ dissolved in $250cm^3$ of solution.

(ii). 75g of $KClO_3$ dissolved to produce $1.25\ dm^3$ of solution.

(iii). $50cm^3$ of $KClO_3$ solution that yields 0.25g residue after evaporation of the water.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

$$M = \frac{n}{V}, \quad n = \frac{m}{M}$$

① $M = \frac{1.5\ moles}{0.25\ dm^3} = 6\ M \quad \because \frac{250\ cm^3}{1000} = 0.25\ dm^3$

$M = 6M$

② $n = \frac{75g}{39 + 35.5 + 48} = \frac{75g}{122.5\ g/mol} = 0.612$

$n = 0.612$ moles

$M = \frac{0.612\ moles}{1.25\ dm^3} = 0.5\ M$

$M = 0.5M$

③ $n = \frac{0.25g}{39 + 35.5 + 48} = \frac{0.25\ g}{122.5\ g/mol} = 0.00201$

$n = 0.002$ moles

$M = \frac{0.002\ moles}{0.05dm^3} = 0.04M \quad \because \frac{50cm^3}{1000} = 0.05dm^3$

$M = 0.04M$

Q. What volume of 0.25M LiOH will completely react with $0.500dm^3$ of 0.25M H_2SO_4.

$$2LiOH + H_2SO_4 \longrightarrow Li_2SO_4 + 2H_2O$$

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

$$\Rightarrow \frac{0.25M \times V_1}{2} = \frac{0.25M \times 0.500dm^3}{1}$$

$$V_1 = \frac{0.25 \times 0.500 \times 2}{0.25M}$$

$$V_1 = \frac{0.25}{0.25}$$

$$V_1 = 1\ dm^3$$

Q. How many grams of NaOH are required to neutralize $40cm^3$ of 0.5M HCl solution.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

$$\Rightarrow M = \frac{n}{V}, \quad D.q = g.q \times C.F, \quad C.F = \frac{find}{given}, \quad n = \frac{m}{M}$$

$$M \times V = n \Rightarrow 0.5M \times 0.04\ dm^3 = n$$

$$\Rightarrow 0.02\ moles = n \text{ of } HCl$$

$$C.F = \frac{find}{given} = \frac{1}{1} = 1\ moles$$

$$D.q = g.q \times C.F = 0.02 \times 1 = 0.02\ moles$$

$$D.q = 0.02 \text{ moles of } NaOH$$

$$n \times M = m \Rightarrow 0.02 \times (23 + 16 + 1.008) = m$$

$$0.02 \times 40 = m \Rightarrow 0.8$$

$$m = 0.8g$$

LIMITING AND NON-LIMITING REAGENT : (4.2)

Limiting reagent is the reactant which is consumed earlier during a chemical reaction.

OR

Limiting reagent is the reactant which produces least number of moles of product in a chemical reaction.

Non-limiting reagent is the reactant which is left unreacted after the completion of reaction.

OR

Non-Limiting reagent is the reactant which produces greater number of moles of product in a chemical reaction.

↳ Limiting reagent determines the maximum amount of product that can be formed.

↳ Limiting reagent basically controls the reaction.

↳ Limiting reagent affects the reaction rate and yield.

⇒ Non-limiting reagent can be recovered or reused.

⇒ Non-limiting reagent may affect the reaction rate but not the yield.

⇒ Non-limiting reagent does not controls the reaction because it is in excess.

So, the reaction can only continue if the limiting reagent is present since the non-limiting reagent is in excess. Hence limiting reagent controls the reaction stoichiometry, non-limiting reagent does'nt.

⇒ The classification of a limiting or non-limiting reagent is not inherent to the substance itself. It depends upon the specific reaction and the amount of reactants present in the reaction / in that specific reaction.

↳ One substance can act as a limiting reagent in one reaction and also can act as a non-limiting reagent in another reaction.

Hence there is no fixed nature for a substance to be limiting or non-limiting reagent. This classification is purely reaction dependent and amount of each reactant present in that reaction dependent.

I.E.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

$$H_2 + \frac{1}{2} O_2 \longrightarrow H_2O$$

In real life application, oxygen is always in excess however hydrogen needs to be provided. In the above chemical reactions for complete reaction of O_2 we need 2 moles of H_2 thus O_2 is the non-limiting / excess reactant while H_2 is the limiting reactant.

Example: 4.11

Q. Calculate the mass of Calcium Nitrate (Ca_3N_2) prepared from 54.9 grams of Ca and 43.2 grams of N_2.

$$3Ca + N_2 \longrightarrow Ca_3N_2$$

$n = \frac{m}{M}$, C.F = $\frac{\text{find}}{\text{given}}$, D.q = g.q × C.F

① $n = \frac{54.9g}{40\,g/mol} = 1.4$ moles

$n = 1.4$ moles of Ca

② $n = \frac{43.2g}{28\,g/mol} = 1.5$ moles

$n = 1.5$ moles of N_2

Since there are less moles of Ca it is the limiting reactant while N_2 is the excess reactant. As we know that limiting reactant controls the reaction

$C.F = \frac{1}{3} = 0.33$ moles

$D.q = 1.4 \times 0.33 = 0.462$ moles of Ca_3N_2

③ $n \times M = m \Rightarrow 0.462 \times 148 = m$

$m = 68.4g$ of Ca_3N_2

Self Assessment : 4.5

Q. How much volume is occupied by CH_4 at STP containing 4.8×10^{24} molecules?

$$n = \frac{V}{V_m} \quad , \quad n = \frac{N}{N_A}$$

① Since we need n $\Rightarrow$ $n = \frac{4.8\times10^{24}}{6.022\times10^{23}}$

$$n = 0.797\times10^{1} = 7.97$$

$$n = 7.97 \text{ moles of } CH_4$$

$$n = \frac{V}{V_m} \Rightarrow n\times V_m = V$$

$$7.97\times 22.414\,dm^3 = V$$

$$\Rightarrow V = 178.6\,dm^3 \text{ of } CH_4$$

Q. Calculate the molecular weight of a gas with the density of $1.25g/dm^3$.

$$PV = nRT \quad , \quad \rho = \frac{m}{V}$$

In terms of STP we can assume that $V = 22.414\,dm^3$

$$\Rightarrow \rho\times V = m \Rightarrow 1.25g/dm^3 \times 22.414dm^3/mol$$

$$m = 28g/mol$$

In terms of RTP we can assume that $T = 298K$, $P = 1atm$, $R = 0.0821\,dm^3.atm/mol.K$

$$\Rightarrow PV = nRT$$

$$\frac{Pm}{\rho} = \frac{m}{M}RT \qquad \because V = \frac{m}{\rho} \;,\; n = \frac{m}{M}$$

$$M = \frac{\cancel{m}RT\rho}{\cancel{m}P}$$

$$M = \frac{RT\rho}{P}$$

Putting in the values,

$$M = \frac{0.0821 \times 298 \times 1.25}{1}$$

$$M = 30.6 g/mol$$

★ At STP $V_m = 22.4 dm^3/mol$

At RTP $V_m = 24.5 dm^3/mol$

Q. Calculate molecules of H_2O produced by 500 molecules each of H_2 and O_2. Which one is the limiting reactant. How much excess reactant is left unreacted. What would happen if molecules of O_2 were doubled?

$$2H_2 + O_2 \longrightarrow 2H_2O$$

$$n = \frac{N}{N_A}, \quad C.F = \frac{\text{find}}{\text{given}}, \quad D.q = g.q \times C.F$$

① $$n = \frac{5 \times 10^2}{6.022 \times 10^{23}} = 8.3 \times 10^{-22}$$

$$n = 8.3 \times 10^{-22} \text{ moles of } O_2$$

Since 500 molecules of H_2 are used hence there are 8.3×10^{-22} moles of H_2.

② To find the limiting reactant we will find moles of O_2 required in the reaction.

$$C.F = \frac{\text{find}}{\text{given}} = \frac{1}{2} = 0.5 \text{ moles}$$

$$D.q = 8.3 \times 10^{-22} \times 0.5 = 4.15 \times 10^{-22} \text{ moles of } O_2$$

So the moles of O_2 required to completely react with H_2 are 4.15×10^{-22}, O_2 is present in excess while H_2 is the limiting reactant. Hence,

③ Excess reactant is left unreacted,

$$8.3 \times 10^{-22} \times 4.15 \times 10^{-22} = 4.15 \times 10^{-22}$$

So 4.15×10^{-22} moles of O_2 are left unreacted.

Limiting reactant is considered when calculating the amount of product so,

$$C.F = \frac{2}{2} = 1 \text{ mol}$$

$$D.q = 8.3\times10^{-22} \times 1 = 8.3\times10^{-22} \text{ moles of } H_2O$$

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

$$8.3\times10^{-22} \times 6.022\times10^{23} = N \Rightarrow 500 = N$$

$\hookrightarrow N = 5\times10^2$ molecules of H_2O

④. If O_2 molecules were doubled

$2 \times 500 = 1000$

1000 molecules of O_2

$$n = \frac{1\times10^3}{6.022\times10^{23}} = 1.66\times10^{-21} \text{ moles of } O_2$$

Since the amount of H_2 stays the same so only 4.15×10^{-22} moles of O_2 are needed to completely react with 8.3×10^{-22} moles of H_2. Hence,

$$1.66\times10^{-21} - 4.15\times10^{-22} = 1.245\times10^{-21}$$

1.245×10^{-21} moles of O_2 will be left unreacted.

And product's amount is calculated by considering the limiting reactant then,

8.3×10^{-22} moles of H_2 will produce 8.3×10^{-22} moles of H_2O, i.e. 5×10^2 molecules of H_2O

OR

Since we have 500 molecules of H_2 and O_2 from the following balanced chemical equation we can say that:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

If no. of moles are equal then no. of molecules will also be equal.

$\hookrightarrow$ Hence, 2 moles of H_2 produces 2 moles of H_2O

$\hookrightarrow$ And 500 molecules of H_2 produces 500 molecules of H_2O

In terms of O_2

Since 2 moles of H_2O are produced by 1 mole of O_2.

1 mole of O_2 completely reacts with 2 moles of H_2.

So for 500 molecules of H_2 we need half the amount of O_2.
500 molecules of H_2 will react completely with 250 molecules of O_2.
So we can conclude that O_2 is the excess reactant and H_2 is the limiting reactant.
250 molecules of O_2 are required for the reaction thus,

$$500 - 250 = 250$$

250 molecules of O_2 will be left over
If molecules of O_2 are doubled,

$$2 \times 500 = 1000$$

H_2 remains the same in amount hence only 250 molecules of O_2 are required to react with 500 molecules of H_2 and produce 500 molecules of H_2O.

$$1000 - 250 = 750$$

750 molecules of O_2 will be left over.

OR

$$2H_2 + O_2 \longrightarrow 2H_2O$$

By unity method : H_2O molecules
2 moles of H_2 produce H_2O = 2 moles
1 mole of H_2 produce H_2O = 1 mole

2 : 2 => 1 : 1

500:500 molecules of H_2O will be 500
1 mole of O_2 produce H_2O = 2 moles
1/2 mole of O_2 produce H_2O = 1 mole

1 : 2 => 500 : 1000 molecules of H_2O will be 1000

O_2 and H_2 Relation :
2 moles of H_2 react with O_2 = 1 mole
1 mole of H_2 react with O_2 = 1/2 mole

2 : 1 => 1 : 1/2

500 : 250 molecules of O_2 required will be 250

Excess reactant and molecules left over:

For complete reaction of O_2 1000 molecules of H_2 are required and 1000 molecules of H_2O will be produced. Since only 250 molecules of O_2 reacted with 500 molecules of H_2 then H_2 is non limiting reactant while O_2 is non-limiting / excess reactant.

$500 - 250 = 250$ molecules

250 molecules of O_2 are left over.

O_2 molecules doubled:

If O_2 molecules were doubled then,

$2 \times 500 = 1000$ molecules

Since only 250 molecules are required because molecules of H_2 did not change so,

$1000 - 250 = 750$ molecules

750 molecules of O_2 are left unreacted.

Q. Excess of $AgNO_3$ was added to $FeCl_3$ and 2.02g of AgCl was produced. Calculate amount of $FeCl_3$ produced in the reaction.

$$3AgNO_3 + FeCl_3 \quad 3AgCl + Fe(NO_3)_3$$

$n = \frac{m}{M}$, C.F = find , D.q = g.q x C.F given

$$C.F = \frac{1}{3} = 0.33 \text{ moles}$$

D.q = g.q x C.F {Since g.q and C.F should be in same units so,}

$$n = \frac{2.02g}{107.9 + 35.45} = \frac{2.02g}{143.4 g/mol} = 0.014 \text{ mol}$$

$n = 0.014$ mol of AgCl

$$D.q = 0.014 \times 0.33 = 4.62 \times 10^{-3}$$

$D.q = 4.62 \times 10^{-3}$ moles of $FeCl_3$

$n \times M = m \Rightarrow 4.62 \times 10^{-3} \times 55.9 + 106.5 = m$

$4.62 \times 10^{-3} + 162.5 = m$

$m = 0.75g$ of $FeCl_3$

4.3. Theoretical, Actual And Percent Yield:

Theoretical yield is the maximum amount of product that can be obtained from a given amount of reactants.

↳ It is determined based on the stoichiometry with the help of balanced chemical equation.

↳ It is a calculated value.

↳ Limiting reactant controls the maximum amount of products that can be formed. So theoretical yield is calculated from the limiting reactant. So,

Theoretical yield ∝ Amount of limiting reactant

Since non-limiting reactant is present in sufficient quantity / it is more than enough or required to react with the limiting reactant hence it does not effect the amount of product formed.

↳ Through stoichiometric balanced chemical equations it lays out a reference point for estimating amount of products ideally formed just by looking at the amount of reactants.

↳ It can be calculated in terms of mass, volume and moles, and litres.

* For solids it is usually calculated in grams / mass.
* For liquids it is usually calculated in litres $(L) = (dm^3)$ or moles (mol) ∴ $1dm^3 = 1L$ ∴ liq generally calculated in Litres.
* For gases it is usually calculated in volume (dm^3 / L) or moles (mol)

↳ For reactions with more than one product formed, theoretical yield of each product must be calculated seperately based on the mole-mole stoichiometric ratios for each product in the balanced chemical equation.

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Assumptions while calculating theoretical yield

1. Pure Reactants : Theoretical yield is calculated assuming that the reactants are pure and uncontaminated because impurities can reduce the effective amount of reactant available for the reaction, in the following way :

* Impurities can block reactants from interacting.
* Impurities can react with reactants.
* Impurities can neutralize reactant charges.
* Impurities can bind to reactants molecules altering their shape.
* Impurities can attract or repel the reactants. preventing the reactant molecules from coming into contact
* Impurities deactivate catalysts.

So the reactant that has the impurity can basically perform the above listed functions / same functions as the impurity present in that reactant which can either enhance or inhibit the reaction speed. Enhancing of reaction rate can be done by :

* Impurities can act as catalysts.
* Impurities can enhance catalyst activity.
* Impurities can modify reactant surfaces by adding or removing functional groups, increasing oxygen concentration, increasing surface area.

2. Perfect efficiency : Theoretical yield assumes that the reaction goes to completion such that all of the limiting reactant is converted to product.

3. No losses : Theoretical yield assumes that no reactant or product is evaporated, spilled, leaked, lost due to equipment issues. And no unwanted reactions like side reactions occur thus no reactant is wasted.

4. Ideal Conditions : Theoretical yield assumes that temperature, pressure and volume is constant throughout the reaction. Also that the system is closed thus no product or reactant escapes.

In conclusion, theoretical yield is based on 100% efficiency which is impossible in real world.

Actual yield is the amount of product obtained from reactants in a chemical reaction.

- It is determined experimentally.
- It is often usually less than the theoretical yield.
- It is a measured value.
- The more pure the reactants will be the higher will be the efficiency of the reaction thus higher will be the actual yield. Because pure reactants result in more reactants fully converting into the desired product. Thus,

Pure reactants $\propto$ Efficiency of the reaction $\propto$ Actual yield

- Actual yield can be measured in grams (mass), milliliters (ml) or litres (L) and moles.
- For solids it is measured in grams or moles.
- For liquids it is measured in grams, moles or ml or L
- For gases it is measured in Litres (L), dm^3 or moles.

Actual yield not showing 100% Efficiency

- Incomplete Reactions: If the reaction does not reach completion meaning that not all of the reactant has converted to the product due to unideal conditions, then the actual yield will be low. i.e. decreasing the efficiency of the reaction.
- Side Reactions: If the reaction is producing byproducts through side reactions, which will consume the reactants in a wasteful way reducing the amount of desired product formed then the actual yield will be less i.e. the efficiency of the reaction will be decreased.
- Impurities: If a reactant contains impurities, the impurities can enhance or inhibit the reaction rate and they will always reduce the actual yield. Impurities might assist the reaction or increase its rate but they will always disadvantage the actual yield by the following ways: (decreased Efficiency)

1. Interfere with the reaction either by slowing it down or causing side reactions.
2. Form by-products by reacting with reactants.
3. Poison or deactivate the catalysts by chemically reacting with the catalysts or by binding to the active sites (where reactants usually bind) of the catalyst.
4. Impurities mixed with the product can make purification difficult leading to loss of product during seperation process.

4. Equilibrium limitations : If the reaction reaches equilibrium without fully completing meaning that not all of the reactant has converted into the product then the actual yield will be less i.e. Efficiency of the reaction will be decreased.

5. Vaporization : In some reactions, some products mostly volatile compounds (substances with low boiling points that can easily vaporize at room temperature or during reactions) can evaporate either during the reaction or during the purification process or during the collection (of product) process causing the actual yield to be less i.e. Efficiency will decrease.

6. Seperation : After the reaction, during the seperation or purification process some product might remain mixed with impurities or byproduct making it difficult to isolate the product fully, some products mainly sticky or reactive ones (might bond with the surface of the material it is present in through chemical interaction) can stick to the surface of the container or filter etc, some products can be lost / spilled during transfer of product from one container to another, leading to decrease in actual yield i.e. Efficiency will decrease.

7. Practically inexperienced person : Human error or inexperience resulting in a mistake by using improper techniques such as incorrect mixing, timing, temperature control can lead to reduced actual yield i.e. Efficiency will decrease.

Reaction time : If the reaction requires more time to reach completion and the reaction is stopped prematurely then the actual yield will be less because not all reactant will be fully converted to product. i.e. Efficiency will decrease.

Catalyst deactivation / poisoning : In a reaction, if the catalyst used deactivates or becomes poisoned due to prolonged use or impurities the reaction will proceed slowly thus producing reduced amount of product in the same amount of time / given time hence the actual yield will be less because the reaction will take more time and produce less product in a given period of time and if a reaction proceeds long enough other factors like vaporization, side reactions etc can lessen the actual yield i.e. Efficiency will decrease.

Degradation of product or reactant can also take place decreasing the actual yield because of catalyst deactivation / poisoning. (degradation is basically decomposition)

If there is measurement error during measurement of reactant or products it may appear that the actual yield of the reaction or the efficiency of the reaction is less but this will be a misconception. i.e. Efficiency will not decrease.

In conclusion, theoretical yield is based on 100% efficiency which is not possible in real world thus actual yield is always less than theoretical yield.

Percent yield is a percentage that represents the ratio of the actual yield to the theoretical yield.

→ It is a calculated value.

→ It determines the reaction efficiency and indicates how much product was actually obtained compared to the maximum amount of product that could be theoretically obtained.

→ Its value depends upon actual yield and theoretical yield.

→ Its value can never exceed 100%.

↳ It is dimensionless and always expressed as a percentage.
↳ It is often lower than 100% percent because of all the factors discussed previously.
↳ Percent yield close to 100% suggests a highly efficient reaction while lower than 50% suggests inefficiency in the reaction process. (below 50% is lower yield, above 50% is considered a higher yield).

Example: 4.12

Q. Calculate the percent yield of Ozone that could be produced by 10g of O_2. Actual yield of O_3 is 1.05g.

$$3O_2 \longrightarrow 2O_3$$

$n = \frac{m}{M}$, $C.F = \frac{find}{given}$, $D.q = g.q \times C.F$

As we know that C.F and g.q should be in same units so,

① $n = \frac{10g}{32g/mol} \Rightarrow 0.3125$ moles

$n = 0.3125$ moles of O_2

$C.F = \frac{find}{given} = \frac{2}{3} = 0.67$ moles

$D.q = 0.67 \times 0.3125 = 0.209$

$D.q = 0.209$ moles of O_3

Since actual yield and theoretical yield must be in same units while calculating percent yield so,

$n \times M = m \Rightarrow 0.209 \times 48 = m$ ∵ $3 \times 16 = 48$

$m = 10.032g$ of O_3

② $\text{Percent yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100\%$

$\%\ \text{yield} = \frac{1.05g}{10.03g} \times 100\%$

$= 0.105 \times 100\%$

$= 10.5\%$

Percent yield = 10.5 % (of Ozone)

Self Assessment : 4.6

Q. Determine the ferric chloride produced by the reaction of $KMnO_4$, 10 moles of $FeCl_2$ and 22 moles of HCl.

$KMnO_4 + 5FeCl_2 + 8HCl \longrightarrow MnCl_2 + KCl + 5FeCl_3 + 4H_2O$

$C.F = \frac{\text{find}}{\text{given}}$, $D.q = g.q \times C.F$

To find moles of $KMnO_4$

$C.F = \frac{1}{5} \Rightarrow 0.2$ moles

$D.q = 0.2 \times 10 \Rightarrow 2$ moles of $KMnO_4$

Amount of HCl required

$C.F = \frac{8}{5} \Rightarrow 1.6$ moles

$D.q = 1.6 \times 10 \Rightarrow 16$ moles of HCl

So to oxidize 10 moles of $FeCl_2$, we need 2 moles of $KMnO_4$ and 16 moles of HCl.

As 16 moles of HCl are required and 22 moles of HCl are present, HCl is the excess reactant and $KMnO_4$ is the limiting reactant.

Since limiting reactant is considered when calculating the amount of a product so,

$C.F = \frac{\text{find}}{\text{given}} = \frac{5}{1} = 5$ moles

$D.q = 2 \times 5 = 10$ moles

$D.q = 10$ moles of $FeCl_3$ produced

Q. If 20g NH_3 and 30g CO_2 produced 40g baking soda then calculate percentage yield of baking soda taking NaCl and H_2O in excess.

$NaCl + H_2O + CO_2 + NH_3 \longrightarrow NaHCO_3 + NH_4Cl$

$n = \frac{m}{M}$, $C.F = \frac{find}{given}$, $D.q = g.q \times C.F$

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100\%$$

To find theoretical yield

$$n = \frac{20g}{14+3(1.008)} = \frac{20g}{17.024 g/mol} = 1.18 \text{ moles}$$

$$n = 1.18 \text{ moles of } NH_3$$

$$n = \frac{30g}{12+2(16)} = \frac{30g}{44 g/mol} = 0.68 \text{ moles}$$

$$n = 0.68 \text{ moles of } CO_2$$

Since limiting reactant should be considered when calculating the amount of the product.

And CO_2 is the limiting reactant

$$C.F = \frac{find}{given} = \frac{1}{1} = 1 \text{ moles}$$

$$D.q = 0.68 \times 1 = 0.68 \text{ moles}$$

$$D.q = 0.68 \text{ moles of } NaHCO_3$$

As we know that for percent yield actual yield and theoretical yield should be in same units.

$$n \times M = m \Rightarrow 0.68 \times 23 + 1.008 + 12 + 3(16) = m$$

$$0.68 \text{ moles} \times 84 g/mol = m$$

$$m = 57.13g \text{ of } NaHCO_3$$

Calculating Percent yield :

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{40g}{57.13g} \times 100\%$$

$$= 0.7002 \times 100\%$$

$$= 70.02\%$$

Percent yield = 70.02% (of $NaHCO_3$)

..OR UNDERSTANDING [LAWS RELEVANCY TO STOICHIOMETRY]

Gay - Lussac's 1st Law : (Constant Volume)

It helps you understand the relation of temperature and pressure so that you can adjust them to get the desired amount of product.

Gay - Lussac's 2nd Law : (Constant temperature, Pressure)

It helps you understand how gases react in simple whole ratios so that you can calculate how much gas reactants are needed to produce gaseous product.

Avogadro's Law : (Constant temperature and Pressure)

It helps you compare multiple gases with same volume or equal moles to calculate moles or volume respectively.

Boyle's Law : (Constant temperature)

It helps you understand the relation of pressure and volume so that you can adjust them to get desired amount of product.

.. Charle's Law : (Constant Pressure)

It helps you understand the relation of volume and temperature so that you can adjust them to get the desired amount of product.

6. Ideal gas law :

It helps you understand the relation of pressure, volume, temperature and moles so that you can simply calculate thermodynamic variables of a gas and you can also adjust these properties to get desired amount of product.

7. Dalton's Law :

It helps you understand how varying the amount of reactants can possibly result in different products thus allowing you to calculate exact proportions needed for a specific product.

2. Proust's Law :

It ensures that for any compound / product, the ratio of the reactants / elements is fixed hence if you know the amount of reactant you can get to know the specific product produced.

Importance Of Stoichiometry In Medicine :

1. Antibiotics : Stoichiometry makes sure that each dose has correct amount of active ingredients to effectively target bacteria.
2. Insulin : Stoichiometry ensures correct amount of insulin dose to properly control blood sugar levels.
3. Vaccines : Stoichiometry ensures correct concentration of viral antigens in the preparation of vaccines to get desired results.

Basically stoichiometry ensures correct input for desired output.