CHAPTER # 1
STOICHIOMETRY

Q1. Define stoichiometry.

Ans: Stoichiometry:

The study of relative amounts of substances involved in a chemical reaction is called Stoichiometry. Such phenomenon is studied through the knowledge of Stoichiometry (Greek word ***Stoicheion*** means element and ***metry*** means measurement).

Importance of Stoichiometry:

Stoichiometry is essential when quantitative information about a chemical reaction is required. Moreover it is important to predict yields of chemical products.

Q2. Explain the significance of stoichiometry with the help of example.

OR

How will you explain law of conservation of mass in the case of combustion of hydrogen fuel in rockets?

Ans: Combustion of hydrogen fuel in rockets:

1. Consider a rocket manufacturer uses liquid hydrogen as a fuel. He may have to determine how much fuel is necessary for a particular flight. Hydrogen burns in oxygen (of air) to produce water.

$$2H_{2\,(g)} + O_{2\,(g)} \longrightarrow 2H_2O_{(g)}$$

It states that

(i) Two moles of Hydrogen react with one mole of oxygen to form two moles of steam.

(ii) Two molecules of Hydrogen react with one molecule of oxygen to produce two molecules of steam.

(iii) Four grams of hydrogen react with thirty-two grams of oxygen to produce thirty-six grams of water. Here the total mass of reactants is equal to the total mass of products. Thus it confirms the Law of conservation of mass.

2. Another example is the reaction taking place in a gas barbecue. This is the example of combustion to form carbon dioxide and water. The balanced chemical reaction is

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

Q3. What do you understand by the term Mole?

Ans: Mole:

The atomic mass, formula mass and molecular mass of a substance expressed in grams is called Mole.

$$Number\ of\ moles = \frac{mass\ in\ grams}{molecular\ mass}$$

Example: One mole of O = 16 g

One mole of O_2 = 32 g
One mole of H_2O = 18 g

Explanation of one mole of NaCl:

The explanation of one mole of NaCl i.e. 58.5 g is quite different as it is ionic in nature and will be called as formula mass which produce ions on dissolving in water. Therefore,

one mole of $NaCl_{(s)} \xrightarrow{H_2O}$ 1 mole of $Na^+_{(aq)}$ ions + 1 mole of $Cl^-_{(aq)}$ ions.

Q4. How will you explain representative particles by using Avogadro's number?

OR

How is mole related to Avogadro's number?

Ans: Representative Particles (Avogadro's Number):

"The number of atoms, ions or molecules present in one mole of a substance is called Avogadro's Number". Its numerical value is 6.023×10^{23}. One mole of any gas at S.T.P occupies 22.414 dm^3 and contains 6.02×10^{23} particles.

e.g. $$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

This reaction can also be expressed in terms of Avogadro's number. The equation states that $2\times6.02\times10^{23}$ molecules of hydrogen react with 6.02×10^{23} molecules of oxygen to produce $2\times6.02\times10^{23}$ molecules of water.

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substance}}{N_A}$$

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substance}}{6.02\times10^{23}}$$

Relation between moles and Avogadro's number:

The relationship between moles and Avogadro's number in an atom and covalent molecules is as follows:

e.g. 1 mole of O $= 6.02\times10^{23}$ atoms

1 mole of O_2(gas) $= 6.02\times10^{23}$ molecules
= 22.414 dm^3 (volume occupied by 1 mole of gas at S.T.P)

1 mole of $H_2O_{(1)}$ $= 6.02\times10^{23}$ molecules

In the case of ionic compounds, the explanation is somewhat different.

For example, **(i)** $$NaCl \xrightarrow{H_2O} Na^{+1}_{(aq)} + Cl^{-1}_{(aq)}$$

It shows that 1 mole of NaCl when dissolves in water gives 1 mole of Na^{1+} ions and 1 mole of Cl^{-1} ions. So According to Avogadro's number we can say that when 6.02×10^{23} formula units of NaCl are dissolved in water there are produced 6.02×10^{23} Na^{1+} and 6.023×10^{23} Cl^{-1} ions.

(ii) $$C + O_2 \longrightarrow CO_2$$

This equation shows that 1 mole of C and according to Avogadro's 6.02×10^{23} atoms reacts with 1 mole of O_2 i.e 6.02×10^{23} molecules of O_2 to produce 2 moles or 6.02×10^{23} molecules of CO_2

Q5. Describe construction of mole ratios as conversion factors in stoichiometric calculations.

Ans: Construction of Mole ratios as Conversion Factors in Stoichiometric Calculations:

Mole Ratios:

Mole ratios mean the ratios of number of moles of reactants taking part and the number of moles of products formed.

Example Combustion of propane:

For example, combustion of propane

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

The mole ratios between the reactants and products can be shown as, one mole of C_3H_8 reacts with five moles of oxygen to give three moles of CO_2 and four moles of water. The amount of propane used will not affect these ratios.

Sample Problem No. 1.1

Methanol burns according to the following equation.

$$2CH_3OH + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 4H_2O_{(g)}$$

If 3.50 moles of methanol are burned in oxygen, calculate

(a) How many moles of oxygen are used

(b) How many moles of water are produced

Solution: Mole ratios = conversion factor

The problem can be solved by using correct conversion factors which are obtained from the balanced chemical reaction.

(a) Moles of methanol (Given quantity) = 3.50 moles

Moles of oxygen (Desired) =?

Conversion Factor = Mole ratios = $\dfrac{3\text{moles } O_2}{2\text{moles } CH_3OH}$

Desired quantity (of O_2) = Given quantity × conversion factor

$$= 3.50 \text{ moles of } \cancel{CH_3OH} \times \frac{3 \text{ moles } O_2}{2 \text{ moles } \cancel{CH_3OH}}$$

$$= \frac{3.50 \times 3}{2} \text{ moles of } O_2$$

Desired quantity = 5.25 moles of O_2

So the number of moles of O_2 consumed (Desired) = 5.25 moles

(b) Given quantity of CH_3OH = 3.5 moles

Desired quantity of H_2O =?

Conversion factor (or mole ratio) = $\dfrac{4 \text{ moles } H_2O}{2 \text{ moles } CH_3OH}$

Desired quantity (i.e. No of moles of H_2O formed)

$$= \text{Given quantity of } CH_3OH \times \frac{4 \text{ moles } H_2O}{2 \text{ moles } CH_3OH}$$

$$= 3.50 \text{ moles of } \cancel{CH_3OH} \times \frac{4 \text{ moles } H_2O}{2 \text{ moles } \cancel{CH_3OH}}$$

$$\text{Required} = \frac{3.50 \times 4}{2} \text{ moles of } H_2O$$

Quantity of H_2O = 7.00 moles of H_2O

Self Check Exercise 1.1

NH_3 is an important raw material in the manufacture of fertilizers. It is obtained by the combination of N_2 and H_2 as shown by the following balanced equation.

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

How many moles of the following are required to manufacture 5.0 moles of NH_3. (a) Nitrogen (b) Hydrogen

(Ans: (a) N_2 = 2.5 Moles (b) H_2 = 7.5 Moles)

Solution: Stoichiometric Calculation:

(a) From given balanced equation it is clear that:

$$2 \text{ moles of } NH_3 = 1 \text{ mole of } N_2$$

$$1 \text{ moles of } NH_3 = \frac{1}{2} \text{ moles of } N_2$$

$$5 \text{ moles of } NH_3 = 5 \times \frac{1}{2} \text{ moles of } N_2 = 2.5 \text{ moles of } N_2$$

(b) From given balanced equation it is clear that:

$$2 \text{ moles of } NH_3 = 3 \text{ mole of } H_2$$

$$1 \text{ moles of } NH_3 = \frac{3}{2} \text{ moles of } H_2$$

$$5 \text{ moles of } NH_3 = 5 \times \frac{3}{2} \text{ moles of } H_2 = 7.5 \text{ moles of } H_2$$

Q6. Define molar volume.

Ans: Molar volume:

Molar quantities of gases can be expressed in terms of volumes. It has been experimentally proved that one mole of any gas at STP occupies a volume of 22.4 dm^3. This volume is called molar volume.

Sample Problem No. 1.2

Iron can be produced from iron ore Fe_2O_3 by reacting the ore with carbon monoxide (CO).Carbon dioxide (CO_2) is produced in this reaction as by product. What mass of iron can be formed from 425 g of iron ore?

Solution:

The balanced equation can be written as

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

Mass of iron ore = 425 g (given mass)

$$\text{No of moles of iron ore} = \frac{\text{mass}}{\text{molecular mass}} = \frac{425\cancel{g}}{159.6\cancel{g}\text{ moles}^{-1}}$$

$= 2.66$ moles of Fe_2O_3

Number of moles of iron produced:

i.e Desired quantity = Given quantity x conversion factor

$$\text{No of moles of Fe} = \text{No of moles of } Fe_2O_3 \times \frac{\text{No of moles of Fe}}{\text{No of moles of } Fe_2O_3}$$

$$= 2.66\ \cancel{\text{Moles } Fe_2O_3} \times \frac{2\text{ Moles Fe}}{1\ \cancel{\text{Mole } Fe_2O_3}}$$

$$= 2.66 \times 2\text{ Moles of Fe} = 5.32\text{ Moles of Fe}$$

How to convert number of moles of iron to mass of Fe in grams:

Desired quantity = Given quantity × conversion factor

$$(\text{Mass of Fe produced}) = 5.32\ \cancel{\text{Mole Fe}} \times \frac{55.9\text{g}}{1\ \cancel{\text{Mole Fe}}}$$

$= 5.32 \times 55.9$ g

Mass of iron produced = 297.388g

Self Check Exercise 1.2

The main engines of the U.S. space shuttle are powered by liquid hydrogen and liquid oxygen. If 1.02×10^5 kg of liquid hydrogen is carried on a particular launch, what mass of liquid oxygen is necessary for all the hydrogen to burn. The equation for the reaction is,

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

(Ans: 8.16×10^5 kg oxygen)

Solution:

$$\textit{Mass of liquid hydrogen} = 1.02 \times 10^5\ kg = 1.02 \times 10^5 \times 10^3\ g$$
$$= 1.02 \times 10^8\ g$$

$$\textit{Number of moles of hydrogen} = \frac{\textit{mass in gram}}{\textit{molar mass}}$$

$$\textit{Number of moles of hydrogen} = \frac{1.02 \times 10^8}{2} = 0.51 \times 10^8\ moles$$
$$= 5.1 \times 10^7\ moles$$

From given balanced equation it is clear that:

$$2\ moles\ of\ H_2 = 1\ mole\ of\ O_2$$

$$1\ mole\ of\ H_2 = \frac{1}{2}\ mole\ of\ O_2$$

$$5.1 \times 10^7\ moles\ of\ H_2 = \frac{1}{2} \times 5.1 \times 10^7\ moles\ of\ O_2$$

$$= 2.55 \times 10^7\ moles\ of\ O_2$$

Now, $Mass\ of\ Oxygen = Number\ of\ moles \times Molar\ mass$

$$= 2.55 \times 10^7 \times 32\ g = 8.16 \times 10^8\ g$$

$$Mass\ of\ liquid\ oxygen\ in\ Kg = \frac{8.16 \times 10^8}{10^3} = 8.16 \times 10^5\ Kg$$

Sample Problem No. 1.3

Calculate the number of molecule of O_2 produced by thermal decomposition of 490 grams of $KClO_3$.

Solution:

The given mass of $KClO_3$ = 490 g
Formula mass of $KClO_3$ = 122.5 g mole^{-1}
No. of moles of $KClO_3$ = 490 / 122.5
= 4 moles

According to reaction, $2KClO_3 \longrightarrow 2KCl + 3O_2$

Stoichiometrically, 2 moles of $KClO_3$ = 3 moles of O_2
4 moles of $KClO_3$ = 3/2 × 4
= 6 moles of O_2
1 mole = 6.02×10^{23} molecules of O_2
6 moles = $6 \times 6.02 \times 10^{23}$ molecules of O_2
= 3.612×10^{24} molecules of O_2

Q7. Explain Gay Lussac's law of combining volume of gases?

Ans: Gay Lussac's law of combining volume:

According to the Gay Lussacs' Law of combining volumes, **the gases react in simple whole number ratios to produce products.**

For example in the reaction: $H_{2\,(g)} + Cl_{2\,(g)} \longrightarrow 2\,HCl_{(g)}$
is telling that one volume of hydrogen gas reacts with one volume of chlorine gas to produce two volumes of hydrogen chloride gas.

Q8. How will you explain volume of gases at STP?

Ans: Volume of gases at STP:

In stoichiometric calculations the problem can be solved easily if reactants and products are used correctly.

22.414 dm^3 of any gas at STP = 1 mole = 6.02×10^{23} molecules.
22.414 dm^3 of H_2 gas at STP = 2g = 6.02×10^{23} molecules.
22.414 dm^3 of NH_3 gas at STP = 17g = 6.02×10^{23} molecules.

Molar Volume:

A mole and volume relationship exists between reactants and products provided the gases are at S.T.P. This volume of 22.4 dm^3 is called Molar Volume.

Sample Problem No. 1.4

Determine the volume that 2.5 moles of chlorine molecules occupy at STP.

Solution: We know that

22.4 dm^3 of Cl_2 (Chlorine) at S.T.P. = 1 mole

Or 1 mole of Cl_2 occupies a volume of 22.4 dm^3 at S.T.P.

2.5 mole of Cl_2 occupy a volume of $22.4 dm^3 \times 2.5 = 56 dm^3$

Self Check Exercise 1.3

(a) How many moles of oxygen molecule are there in 50.0 dm^3 of oxygen gas at S.T.P?

(b) What volume does 0.80 mole of N_2 gas occupy at S.T.P?

(Ans: (a) 2.23 moles, (b) 17.93 dm^3)

Solution:

(a) *We know that*

1 mole of O_2 occupies volume of 22.414 dm^3 at STP

$$22.414\ dm^3 of\ O_2\ at\ STP = 1\ mole\ of\ O_2$$

$$1\ dm^3\ of\ O_2\ at\ STP = \frac{1}{22.414}$$

$$50\ dm^3\ of\ O_2\ at\ STP = \frac{1}{22.414} \times 50 = 2.23\ moles$$

(b) *We know that*

1 mole of N_2 occupies volume of 22.414 dm^3 at STP

$$1\ mole\ of\ N_2 = 22.414\ dm^3 of\ N_2\ at\ STP$$

$$0.8\ mole\ of\ N_2 = 22.414 \times 0.8 = 18\ dm^3$$

Q9. Define limiting reactant.

Ans: Limiting Reactants:

The reactant that is consumed completely in a chemical reaction is called limiting reactant.

Also it can be defined as the reactant which produces the least number of moles of products in a chemical reaction.

The amount left un-used or un-reacted after completion of reaction is called "*Reactant in excess*".

Q10. How will you identify limiting reactant in a reaction?

Ans: Identification of a Limiting Reactant in a Reaction:

A limiting reactant can be recognized by calculating the number of moles of products formed from data of the given amounts of the reactants using a balanced chemical equation. The reactant, which produces the least amount of products, is the limiting reactant.

Example: For example, 10 moles of H_2 and 7 moles of O_2 were reacted to produce H_2O. Which one of the reactant is the limiting reactant? We can calculate as follows:

The reaction is, $2H_2 + O_2 \longrightarrow 2H_2O$

Stoichiometrically,

(i)	$2H_2$	=	$2H_2O$
	i.e. 2 Moles of H_2	=	2 Moles of H_2O
	10 Moles of H_2	=	10 Moles of H_2O
(ii)	O_2	=	$2H_2O$
	i.e. 1 Mole of O_2	=	2 Moles of H_2O
	so, 7 Moles of O_2	=	$2 \times 7 = 14$ Moles of H_2O

Since H_2 gives the least number of moles of H_2O, i.e. 10 moles, so H_2 is the limiting reactant.

Sample Problem No. 1.5

200 g of $K_2Cr_2O_7$ was reacted with 200g conc. H_2SO_4.Calculate

(a) Mass of atomic oxygen produced

(b) Mass of reactant left unreacted

Solution:

(a) Mass of $K_2Cr_2O_7$ = 200g

Formula Mass of $K_2Cr_2O_7$ = 294g $mole^{-1}$

No of moles of $K_2Cr_2O_7$ = $\frac{200}{294}$

= 0.68 moles

Mass of H_2SO_4 = 200g $mole^{-1}$

Formula Mass of H_2SO_4 = 98g

No of moles of H_2SO_4 = $\frac{200}{98}$

= 2.04 moles

$$K_2Cr_2O_7 + 4H_2SO_4 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + 3(O)$$

0.68 moles   2.04 moles

1 mole of $K_2Cr_2O_7$ = 3 moles of 'O'

0.68 mole of $K_2Cr_2O_7$ = 3×0.68 moles of 'O'

= 2.04 moles of 'O'

4 moles of H_2SO_4 = 3 moles of 'O'

2.04 moles of H_2SO_4 = $\frac{3}{4} \times 2.04$

= 1.53 moles

As H_2SO_4 is producing small amount so, H_2SO_4 is the limiting reactant and produced oxygen = 1.53 moles.

Mass in gram = Number of moles × Molecular mass

= $16 \times 1.53 = 24.48$ g

b) In this problem H_2SO_4 is the limiting reactant and $K_2Cr_2O_7$ is the reactant in the excess

We have 0.68 moles of $K_2Cr_2O_7$ and 2.04 moles of H_2SO_4

According to the reaction,

4 moles of H_2SO_4 = 1 mole of $K_2Cr_2O_7$

2.04 moles of H_2SO_4 = $\frac{1}{4} \times 2.04$

= 0.51 moles of $K_2Cr_2O_7$

No of moles of $K_2Cr_2O_7$ left unreacted = 0.68 – 0.51

= 0.17 moles

Mass of $K_2Cr_2O_7$ = No of moles × Formula Mass of $K_2Cr_2O_7$

= $0.17 \times 294 = 49.98$

So Mass of $K_2Cr_2O_7$ left unreacted = 49.98g

Sample Problem No. 1.6

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

(a) No of H_2SO_4 molecules (b) No of H^+ and SO_4^{2-}

(c) Mass of individual ion

Solution:

a. Mass of H_2SO_4 = 20g

Molar Mass of H_2SO_4 = 98.016g

No of molecules of H_2SO_4 $= \frac{\textit{Mass of } H_2SO_4}{\textit{Molar Mass of } H_2SO_4} \times 6.02\times10^{23}$

$= \frac{20}{98.016} \times 6.02 \times 10^{23}$

$= 1.228 \times 10^{23}$

b. H2SO4 dissolves in water as follows:

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

According to equation

1molecule of H_2SO_4 = 2 H^+ ions

1.228×10^{23} molecules of H_2SO_4 $= 2 \times 1.228 \times 10^{23}$ H^+ ions

As 1molecule of H_2SO_4 = 1 SO_4^{2-} ions

So, 1.228×10^{23} molecule of H_2SO_4 = 1.228×10^{23} SO_4^{2-} ions

c. Mass of individual ions

Mass of H^+ $= \frac{1.008}{6.02\times 10^{23}} \times 2.456\times10^{23}$

$= 0.411g$

Mass of SO_4^{2-} $= \frac{96}{6.02\times 10^{23}} \times 1.228\times10^{23}$ = 19.g

Sample Problem No. 1.7

Magnesium metal reacts with Sulphur to produce MgS. How many grams of magnesium sulphide (MgS) can be made from 1.50g of Mg and 1.50g of sulphur by the reaction

$$Mg + S \longrightarrow MgS$$

Solution: Mass of Mg $= 1.50g$

No. of moles of Mg $= \frac{1.50}{24} = 0.0625$ moles

Mass of S $= 1.50g$

No. of moles of S $= \frac{1.50}{32} = 0.0467$ moles

$Mg + S \longrightarrow MgS$

i.e. 1 mole of Mg = 1 mole of MgS

so, 0.0625 moles of Mg = 0.0625 moles of MgS

also, 1 mole of Mg = 1 mole of MgS

so, 0.0647 moles of Mg = 0.0467 Moles of MgS

Since S gives the least No of moles of products as compared to Mg, so it is the limiting reactant.

Now we calculate the mass of MgS in grams.

Mass of 1 Mole of MgS $= 24 + 32 = 56g$

Mass of 0.0467 Mole of MgS $= 56 \times 0.0467g = 2.6152g$

Self Check Exercise 1.4

(1) Zinc and Sulphur react to form Zinc Sulphide according to the following balanced chemical equation $Zn + S \longrightarrow ZnS$

If 6.00g of Zinc and 4.00g of Sulphur are available for reaction, then determine

(a) The limiting reactant.

(b) The mass of Zinc Sulphide produced.

(Ans. (a) Zinc is the limiting reacted since the whole is consumed.

(b) Mass of Zinc Sulphide produced = 8.94g)

Solution: $Mass\ of\ Zn = 6g$

$Atomic\ mass\ of\ Zn = 65.41$

$$Number\ of\ moles\ of\ Zn = \frac{Mass\ in\ gram}{Atomic\ mass} = \frac{6}{65.41} = 0.0917\ moles$$

$Mass\ of\ S = 4g$

$Atomic\ mass\ of\ S = 32\ g$

$$Number\ of\ moles\ of\ S = \frac{Mass\ in\ gram}{Atomic\ mass} = \frac{4}{32} = 0.125\ moles$$

$Zn + S \longrightarrow ZnS$

$1\ mole\ of\ Zn = 1\ mole\ of\ ZnS$

$0.0917\ moles\ of\ Zn = 0.0917\ moles\ of\ ZnS$

$Also,\ 1\ mole\ of\ S = 1\ mole\ of\ ZnS$

$So,\ 0.125\ moles\ of\ S = 0.125\ moles\ of\ ZnS$

Since Zn gives the least number of moles of products as compared to S, so it is the limiting reactant. Now we calculate the mass of ZnS in grams.

$$Mass\ of\ 1\ mole\ of\ ZnS = 65.41 + 32 = 97.41g$$

$Mass\ of\ 0.0917\ moles\ of\ ZnS = 97.41 \times 0.0917 = 8.94\ g$

Ans. **(a)** Zinc is the limiting reacted since the whole is consumed.

(b) Mass of Zinc Sulphide produced = 8.94 g

(2) **Aluminium reacts with bromine to form Aluminium bromide, as shown by the balanced chemical equation,** $2Al + 3Br_2 \longrightarrow 2AlBr_3$

If 15.8g of Al and 55.6g of Br_2 are available for reaction, then determine

(a) The limiting reactant **(b) The mass of $AlBr_3$ produced**

Ans: **(a) Bromine is the limiting reactant.**

(b) Mass of $AlBr_3$ formed = 61.9g

Solution: $Mass\ of\ Al = 15.8\ g$

$Atomic\ mass\ of\ Al = 27$

$$Number\ of\ moles\ of\ Al = \frac{Mass\ in\ gram}{Atomic\ mass} = \frac{15.8}{27.98} = 0.585\ moles$$

$Mass\ of\ Br_2 = 55.6\ g$

$Molar\ mass\ of\ Br_2 = 79.9 \times 2 = 159.8\ g/mole$

$$Number\ of\ moles\ of\ Br_2 = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{55.6}{159.8} = 0.348\ moles$$

$Now, \quad 2Al + 3Br_2 \longrightarrow 2AlBr_3$

$2\ mole\ of\ Al = 2\ moles\ of\ AlBr_3$

$1\ mole\ of\ Al = \frac{2}{2} = 1\ mole\ of\ AlBr_3$

$0.585\ mole\ of\ Al = 0.585\ mole\ of\ AlBr_3$

$Also, \quad 3\ moles\ of\ Br_2 = 2\ moles\ of\ AlBr_3$

$1\ mole\ of\ Br_2 = \frac{2}{3}$

$\mathbf{1\ mole\ of\ Br_2 = \frac{2}{3} \times 0.348 = 0.232\ moles\ of\ AlBr_3}$

Since Br_2 gives the least number of moles of products as compared to Al, so it is the limiting reactant. Now we calculate the mass of $AlBr_3$ in grams.

$Mass\ of\ 1\ mole\ of\ AlBr_3 = 27 + 79.9 \times 3 = 27 + 239.7 = 266.7\ g$

$Mass\ of\ 0.232\ moles\ of\ AlBr_3 = 0.232 \times 266.7 = 61.9\ g$

Q11. Differentiate between limiting reactant and reactant in excess in a reaction?

Ans: **Difference between limiting reactant and reactant in excess:**

Limiting reactant	Reactant in excess
i. The reactant which controls the quantity of product or which is lesser quantity is called limiting reactant.	**i.** the reactant which remain un-reacted after the completion of a reaction is called reactant in excess.
ii. It is taken in lesser quantity.	**ii.** it is in excess.
iii. It is usually expensive.	**iii.** It is usually cheaper.
iv. It is consumed completely in a chemical reaction.	**iv.** It is not consumed completely in a chemical reaction.

OR

During a reaction in which 2 reactants are reacted sometimes one component is consumed completely and some amount of other reactant is left unreacted. The reactant which is consumed completely during the reaction is called Limiting Reactant and the reactant whose some amount is left unconsumed is called "Reactant in Excess".

Sample Problem No. 1.8

Suppose 1.87 moles of ammonium chloride were reacted with 1.35 mole of calcium hydroxide. How many grams of calcium hydroxide are left unreacted in this reaction?

Solution: According to reaction,

$$Ca(OH)_2 + 2NH_4Cl \longrightarrow CaCl_2 + 2NH_4OH$$

Let us calculate the no. of moles of $Ca(OH)_2$ in above example that reacts completely with 1.87 moles of NH_4Cl.

2 moles of NH_4Cl = 1 mole of $Ca(OH)_2$

1.87 moles of NH_4Cl = $\frac{1}{2} \times 1.87$

= 0.935 moles of $Ca(OH)_2$

So, no. of moles of $Ca(OH)_2$ consumed = 0.935 moles

And no. of moles of $Ca(OH)_2$ initially present = 1.35 moles

So, no. of moles of $Ca(OH)_2$ unconsumed = 1.35 - 0.935

= 0.415 moles

As the molecular mass of $Ca(OH)_2$ = 74

So the mass of 0.415 moles of $Ca(OH)_2$ = 74×0.415

= 30.71g

Result: The excess amount of $Ca(OH)_2$, which is left unreacted is 30.71g. This is also called reactant in excess.

Q12. Differentiate between theoretical yield and actual yield.

Ans:

Theoretical yield	Actual yield.
I. "The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction".	**I.** The quantity of product that is actually produced in a chemical reaction is called the actual yield.
ii. It is calculated from balanced chemical equation	**ii.** It is calculated from experiments.
iii. Theoretical yield is always greater than actual yield.	**iii.** Actual yield is always lesser than theoretical yield

Q13. Define percent yield and write its formula.

Ans: Percent Yield:

Percent yield is a measure of the efficiency of a chemical reaction. Percent yield is calculated to be the experimental yield divided by theoretical yield multiplied by 100%.

$$\text{Percent yield} = \frac{\text{Actual yeild}}{\text{Theoratical yeild}} \times 100$$

Q14. Define Quantitative.

Ans: Quantitative Reaction:

There are many reactions for which the actual yield is almost actually equal to the theoretical yield. Such reactions are quantitative, i.e., they can be used in chemical analysis.

Sample Problem No. 1.9

In an industry Copper metal was prepared by the following reaction,

$$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$$

1.274g $CuSO_4$ when reacted with excess of Zn metal a yield of 0.392g Cu metal was obtained. Calculate the percentage yield.

Solution:

According to reaction, $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

Mass of $CuSO_4$ given = 1.274g

Now we convert the no of grams of $CuSO_4$ into no. of moles.

Molecular mass of $CuSO_4$ = 63.5 + 32 + 64 = 159.6 g/mole

159.6g of $CuSO_4$ = 1 mole

1.274 g of $CuSO_4$ $= \frac{1}{159.6} \times 1.274$

$= 7.982 \times 10^{-3}$ moles.

Stoichiometrically:

1 mole of $CuSO_4$ = 1 mole of Cu

7.98×10^{-3} moles of $CuSO_4$ = 7.982×10^{-3} moles of Cu.

as, 1 mole of Cu = 63.5 g

so, 7.982×10^{-3} moles of Cu = $63.5 \times 7.982 \times 10^{-3}$

= 0.5072 g of Cu.

Hence, Theoretical yield = 0.5072 g

But Actual yield = 0.392 g

So, $\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

$$= \frac{0.392}{0.5072} \times 100 = 77.3\%$$

Sample Problem No. 1.10

In a reaction, 2.00 moles of CH_4 was reacted with an excess of Cl_2. As a result, 177.0 g of CCl_4 is obtained. What is the

(a) theoretical yield (b) actual yield (c) % yield of this reaction?

Solution:

Reaction is, $CH_{4(l)} + 4Cl_{2(g)} \longrightarrow CCl_{4(l)} + 4HCl_{(g)}$

Stiochiometrically,

From 2.0 moles of CH_4, we would expected to obtained 2.0 moles of CCl_4

(a) Theoretical yield = 2.0 moles of CCl_4

2.0 moles of CCl_4 means = 2 x 154 = 308g

(b) Actual yield = 177.0g of CCl_4

(c) Percent Yield:

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

$$\% \text{ yield} = \frac{177}{308} \times 100 = 57.46\ \%$$

Science Titbit

1. The overall balanced equation for the production of ethanol (C_2H_5OH) from sugar is as follows:

$$C_6H_{12}O_{6(aq)} \longrightarrow 2C_2H_5OH_{(aq)} + 2CO_{2(g)}$$

(a) What is the theoretical yield of ethanol available from 10.0 g of sugar.

(b) If in a particular experiment, 10.0 g of sugar produces 0.664 g of ethanol, what is the percentage yield?

Ans: (a) Theoretical yield of ethanol = 5.125g

(b) Percentage yield = 12.89 %)

Solution: **(a)** $Mass\ of\ C_6H_{12}O_6 = 10\ g$

$$Molar\ mass\ of\ C_6H_{12}O_6 = 12 \times 6 + 1 \times 12 + 16 \times 6 = 180\ g/mole$$

$$Number\ of\ moles\ of\ C_6H_{12}O_6 = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{10}{180} = 0.056\ mole$$

$$1\ mole\ of\ C_6H_{12}O_6 = 2\ moles\ of\ C_2H_5OH$$

$$0.056\ moles\ of\ C_6H_{12}O_6 = 2 \times 0.056\ moles\ of\ C_2H_5OH$$

$$= 0.112\ moles\ of\ C_2H_5OH$$

$$Molar\ mass\ of\ C_2H_5OH = 12 \times 2 + 1 \times 6 + 16 \times 1 = 46\ g/mole$$

$$Mass\ of\ C_2H_5OH = 0.112 \times 46 = 5.125\ g$$

(b) $$Percentage\ yield = \frac{Actual\ yield}{Theoretical\ yield} \times 100$$

$$Percentage\ yield = \frac{0.664\ g}{5.11\ g} \times 100 = 12.89\ \%$$

2. Solid carbon dioxide (dry ice) may be used for refrigeration. Some of this carbon dioxide is obtained as a by-product when hydrogen is produced from methane in the following reaction.

$$CH_{4(g)} + 2H_2O_{(g)} \longrightarrow CO_{2(g)} + 4H_{2(g)}$$

(a) What mass of CO_2 should be obtained from the complete reaction of 1250 g of methane.

(b) If the actual yield obtained is 3000 g then what is the percentage yield?

(Ans: a = 3438 g b = 87.3 %)

Solution: (a)

$$Given\ mass\ of\ methane = 1250\ g$$

$$Molar\ mass\ of\ CH_4 = 12 + 4 = 16\frac{g}{mole}$$

$$Number\ of\ moles\ of\ CH_4 = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{1250}{16} = 78.125\ moles$$

Stiochiometrically,

$$1\ mole\ of\ CH_4 = 1\ mole\ of\ CO_2$$

$$78.125\ moles\ of\ CH_4 = 78.125\ moles\ of\ CO_2$$

$$Molar\ mass\ of\ CO_2 = 12 + 32 = 44\ g/mole$$

$$Mass\ of\ CO_2\ obtained = number\ of\ moles \times molar\ mass$$

$$= 78.125 \times 44 = 3437.5g$$

(b)

$$Actual\ yield = 3000\ g$$

$$Theoretical\ yield = 3437.5\ g$$

$$Percentage\ yield = \frac{Actual\ yield}{Theoretical\ yield} \times 100$$

$$Percentage\ yield = \frac{3000}{3437.5} \times 100$$

$$Percentage\ yield = 87.3\ \%$$

EXERCISE

MULTIPLE CHOICE QUESTIONS

1. Select the most suitable answer in the following MCQs.

i. How many molecules are there in one mole of H_2O?

(a) 6.023×10^{19} (b) 6.023×10^{23}

(c) 1.084×10^{18} (d) none of these

ii. A flask contains 500 cm^3 of SO_2 at STP. The flask contains

(a) 40 g (b) 100 g

(c) 50 g (d) none of these

iii. A necklace has 6g of diamond in it. What are the numbers of atoms in it?

(a) 6.02×10^{23} (b) 12.04×10^{23}

(c) 1.003×10^{23} (d) 3.01×10^{23}

(a) What mass of CO_2 should be obtained from the complete reaction of 1250 g of methane.

(b) If the actual yield obtained is 3000 g then what is the percentage yield?

(Ans: a = 3438 g b = 87.3 %)

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$$Molar\ mass\ of\ CO_2 = 12 + 32 = 44\ g/mole$$

$$Mass\ of\ CO_2\ obtained = number\ of\ moles \times molar\ mass$$

$$= 78.125 \times 44 = 3437.5g$$

(b)

$$Actual\ yield = 3000\ g$$

$$Theoretical\ yield = 3437.5\ g$$

$$Percentage\ yield = \frac{Actual\ yield}{Theoretical\ yield} \times 100$$

$$Percentage\ yield = \frac{3000}{3437.5} \times 100$$

$$Percentage\ yield = 87.3\ \%$$

EXERCISE

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(c) 1.003×10^{23} (d) 3.01×10^{23}

iv. **What is the mass of aluminium in 204 g of the aluminium oxide, Al_2O_3**

(a) 26 g (b) 27 g (c) 54 g (d) 108 g

v. **The reactant which is consumed earlier and gives least quantity of product is called.**

(a) Reactant (b) Stoichiometry
(c) Limiting reactant (d) Stoichiometric amount

vi. **Which one of the following compounds contains the highest percentage by mass of nitrogen**

(a) NH_3 (b) N_2H_4 (c) NO (d) NH_4OH

vii. **Vitamin-A has a molecular formula $C_{20}H_{30}O$. The number of vitamin - A molecules in 500 mg of its capsule will be**

(a) 6.02×10^{23} (b) 1.05×10^{23}
(c) 3.01×10^{22} (d) 3.01×10^{23}
(e) none of these

viii. **When one mole of each of the following is completely burnt in oxygen, which will give the largest mass of CO_2?**

(a) Carbon Monoxide (b) Diamond
(c) Ethane (d) Methane

ix. **One mole of ethanol and one mole of ethane have an equal**

(a) Mass (b) Number of Atoms
(c) Number of electron (d) Number of molecules

x. **Methane reacts with steam to form H_2 and CO as shown below,**

$$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$$

What volume of H_2 can be obtained from 100 cm^3 of methane at the standard temperature and pressure?

(a) 300 cm^3 (b) 200 cm^3 (c) 150 cm^3 (d) 100 cm^3

xl. **The Avogadro's constant is the number of**

(a) Atoms in 1g of Helium
(b) Molecules in 35.5g of Chloride
(c) Electrons needed to deposit 24g of Mg
(d) Atoms in 24g of Mg

xii. **How many moles of oxygen ate needed for the complete combustion of two moles of butane?**

(a) 2 (b) 8 (c) 10 (d) 13

xiii. **If four moles of SO_2 are oxidised to SO_3, how many moles of oxygen molecules are required.**

(a) 0.5 (b) 1.0 (c) 1.5 (d) 2.0

xiv. **The relative atomic mass of Chlorine is 35.5. What is the mass of 2 moles of Chlorine gas?**

(a) 142g (b) 71g (c) 35.5g (d) 18.75g

xv. **Which of the following statements is incorrect?**

(a) 12g of Carbon gas contains one mole of atoms
(b) 28g of Nitrogen gas contains one mole of molecules of N2

(c) 1 dm^3 of a 1.0 Mole dm^{-3} solution of NaCl contains one mole of Chloride ions

(d) None of above

xvi. One mole of propane has the same__________

(a) Number of molecules as one mole of methane (CH_4)

(b) Number of C-atoms as in one mole of butane (C_4H_{10})

(c) Mass as half a mole of hexane (C_6H_{14})

(d) Number of molecules as in one mole of ethane (C_2H_6)

xvii. What is the mass of one mole of Iodine Molecules?

(a) 254 g (b) 74 g (c) 106 g (d) 127 g

xviii. What volume of SO_2 at room temperature and pressure is produced on heating 9.7g of Zinc Sulphide (ZnS) if reaction takes place as follows

$$2ZnS_{(s)} + 3O_{2(g)} \longrightarrow 2ZnO_{(s)} + 2SO_{2(g)}$$

(a) 1.2 dm^3 (b) 2.4 dm^3 (c) 3.6 dm^3

(d) 4.8 dm^3 (e) none of these

Answers

i. b	**ii. d**	**iii. d**	**iv. d**	**v. c**	**vi. b**	**vii. e**
viii. c	**ix. d**	**x. a**	**xi. d**	**xii. d**	**xiii. d**	**xiv. a**
xv. d	**xvi. a, d**	**xvii. a**	**xviii. e**			

2. Answer the following questions briefly:

i. 58.5 amu are termed as formula mass and not molecular mass of NaCl. Why?

Ans: Because NaCl is an ionic substance and molecules of ionic substances are termed as formula units but not as molecular mass. That is why 58.5 amu is the formula mass of NaCl and not molecular mass.

ii. Concept of limiting reactant is not applicable to the reversible reactions. Explain.

Ans: "A reactant which consumes earlier due to its smaller amount and produces least amount of product is called limiting reactant."

During a reversible reaction, reactants are converted into products and products convert back into reactants. So 'reactants are not completely consumed. As a result a limiting reactant cannot be identified in a reversible reaction.

OR (Second Answer)

Limiting reactant is the reactant which is completely consumed in a chemical reaction whereas in a reversible reaction, there is no reactant which is 100% consumed but is regenerated due to reversible reaction.

Therefore concept a limiting reactant is not applicable to a reversible reaction.

iii. How many covalent bonds are present in 9g of H_2O?

Ans: $\textit{Mass of } H_2O = 9g$

$\textit{Molecular mass of } H_2O = 16 + 2 = 18\ g/mole$

Number of water molecules $= \frac{\text{Mass of water}}{\text{Molar mass of water}} \times N_A$

$$Number\ of\ molecules\ of\ H_2O = \frac{9}{18} \times 6.02 \times 10^{23} = 3.01 \times 10^{23} molecules$$

$$1\ molecule\ of\ H_2O = 2\ Covalent\ Bonds$$

$$3.01 \times 10^{23} molecules\ of\ H_2O = 2 \times 3.01 \times 10^{23}$$

$$= 6.02 \times 10^{23}\ Covalent\ Bonds$$

iv. Differentiate between limiting and non-limiting reactants.

Ans: Difference between limiting reactant and reactant in excess:

Limiting reactant	Reactant in excess
i. The reactant which controls the quantity of product or which is lesser quantity is called limiting reactant.	**i.** the reactant which remain un-reacted after the completion of a reaction is called reactant in excess.
ii. It is taken in lesser quantity.	**ii.** it is in excess.
iii. It is usually expensive.	**iii.** It is usually cheaper.
iv. It is consumed completely in a chemical reaction.	**iv.** It is not consumed completely in a chemical reaction.

v. How many molecules of water are there in 12 g of ice?

Ans:

$$Mass\ of\ ice = 12\ g$$

$$Molar\ mass\ of\ water = 16 + 2 = 18\ g/mole$$

$$Number\ of\ molecules\ of\ water = \frac{Mass\ of\ water}{Molar\ mass\ of\ water} \times N_A$$

$$Number\ of\ molecules\ of\ water = \frac{12}{18} \times 6.023 \times 10^{23}$$

$$= 4.01 \times 10^{23}\ molecules$$

vi. One mole of H_2SO_4 should completely react with two moles of NaOH. How does Avogadro's number help to explain it?

Ans: $H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$

In above reaction According to Avogadro's number 1 mole of or 6.02×10^{23} molecules of H_2SO_4 react with 2 moles or $2 \times 6.02 \times 10^{23}$ molecules of $NaOH$ to produce 1 mole of Na_2SO_4 or 6.02×10^{23} molecules of Na_2SO_4 and 2 moles or $2 \times 6.02 \times 10^{23}$ molecules of H_2O.

vii. Give reason that 1 mole of different compounds have different masses but have the same number of molecules.

Ans: According to Avogadro's number 1 mole of different compounds have 6.02×10^{23} molecules but have different masses. As different molecules have different atoms having different atomic masses, so their molecular masses may be different but number of molecules in 1 mole must be same i.e. 6.02×10^{23} molecules.

Example: 18g of water = 1 mole = 6.023×10^{23} molecules.

342g of sucrose = 1 mole = 6.023×10^{23} molecules.

viii. 23g of sodium and 238g of uranium have equal number of atoms in them.

Ans: Atomic mass of number is 23 and that of uranium is 238. It means 23g of Na is equal to 1 mole of **Na** and 238 g of uranium is equal to 1 mole of uranium and according to Avogadro's Law 1 mole of any element has 6.02×10^{23} atoms. It means 1 mole of Na and 1 mole of uranium have equal number of atoms in them that is 6.02×10^{23} **atoms.**

ix. What is percentage composition? Calculate percentage composition of Glucose.

Ans: Percentage Composition:

A compound may contain certain elements. The percentage of each element in a compound is called percentage composition of that compound.

It is calculated as follow:

$$\%\ of\ an\ element = \frac{mass\ of\ element\ in\ compound}{Molar\ mass\ of\ compound} \times 100$$

Percentage Composition of Glucose:

$$Molar\ mass\ of\ glucose\ (C_6H_2O_6) = 12 \times 6 + 1 \times 12 + 16 \times 6 = 180g\ mole^{-1}$$

$$\%\ of\ an\ element = \frac{mass\ of\ element\ in\ compound}{Molar\ mass\ of\ compound} \times 100$$

$$\%\ of\ C = \frac{72}{180} \times 100 = 40\%$$

$$\%\ of\ H = \frac{12}{180} \times 100 = 6.67\%$$

$$\%\ of\ O = \frac{96}{180} \times 100 = 53.33\%$$

x. Calculate the weight of oxygen gas evolved when 5.0 g of $KClO_3$ are completely decomposed thermally.

Ans: $Given\ Mass\ of\ KClO_3 = 5g$

$$Formula\ Mass\ of\ KClO_3 = 39 + 35.5 + 16 \times 3 = 122.5\ g/mole$$

$$Number\ of\ Moles\ of\ KClO_3 = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{5}{122} = 0.04\ moles$$

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

According to the above equation

$$2\ moles\ of\ KClO_3 = 3\ moles\ of\ O_2$$

$$1\ moles\ of\ KClO_3 = \frac{3}{2}\ moles\ of\ O_2$$

$$0.04\ moles\ of\ KClO_3 = \frac{3}{2} \times 0.04 = 0.06\ moles\ of\ O_2$$

$$1\ mole\ of\ O_2 = 32\ g\ of\ O_2$$

$$0.06\ mole\ of\ O_2 = 0.06 \times 32 = 1.92\ g$$

xi. What is limiting reactant? How will you dètermine it?

Ans: Excess and Limiting Reactants:

The reactant that is consumed completely in a chemical reaction is called limiting reactant. Also it can be defined as the reactant which produces the least number of moles of products in a chemical reaction.

Determination of Limiting Reactant:

The following three steps should be followed to find out the limiting reactant

(i) Calculate the number of moles from the given amount of reactant.

(ii) Find out the number of moles of product with the help of a balanced chemical equation.

(iii) Identify the reactant which produces the least amount of product as limiting reactant.

xii. Define Theoretical yield and actual yield.

Ans: Theoretical Yield:

"The amount of the products calculated from a balanced chemical equation is called theoretical yield."

i. It is also known as calculated yield or expected yield.

ii. Theoretical yield of a reaction is always greater than the actual yield of the same reaction.

Actual yield:

"The amount of the products obtained with a given amount of the reactant in an actual experiment is called actual yield."

i. It is also known as experimental yield.

i. The actual yield of a chemical reaction is always lesser than the theoretical yield.

xiii. What is conversion factor?

Ans: Conversion factor or Mole ratio:

Conversion factor means the ratios of number of moles of reactants taking part and the number of moles of products formed.

Example Combustion of propane:

For example, combustion of propane

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

The mole ratios between the reactants and products can be shown as, one mole of C_3H_8 reacts with five moles of oxygen to give three moles of CO_2 and four moles of water. The amount of propane used will not affect these ratios.

OR

The simplest stoichiometry problems are mole-to-mole conversions, and for those you need only a balanced chemical equation. The coefficients in the balanced equation are used to construct a conversion factor between moles of one substance and moles of a different substance in the same reaction.

Mole ratios of reactants and products as given by a balanced chemical equation is called conversion factor e.g. in reaction.

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

1 mole of C_3H_8 (propane) reacts with 5 moles of O_2 to produce 3 moles of CO_2 and 4 moles of H_2O. Conversion factor is not effected by amount of reactants or products.

xiv. What is the relationship between mass and volume of a gas at S.T.P.?

Ans: Molar Volume:.

A mole and volume relationship exists between reactants and products provided the gases are at S.T.P. This volume of 22.414 dm^3 is called Molar Volume.

In stoichiometric calculations the problem can be solved easily if reactants and products are used correctly.

22.414 dm^3 of H_2 gas at STP = 2g = 6.02×10^{23} molecules.

22.414 dm^3 of NH_3 gas at STP = 17g = 6.02×10^{23} molecules.

xv. The actual yield is lesser than the theoretical yield. Give reasons.

Ans: Actual yield of a reaction is less than the theoretical yield because of following reasons

i. The reaction may not go to completion and may reduce the yield of product.

ii. Material may be lost in handling.

iii. Side reactions may produce by-products

iv. Reactions are reversible

v. Mechanical loss takes place due to filtration, distillation, and separation by separating funnel, washing and crystallization etc.

vi. Inexperience worker or method may be faulty.

xvi. What are the representative particles in one mole of a gas at S.T.P.?

Ans: One mole of any gas at STP occupies 22.414 dm^3 and contains 6.02×10^{23} particles.

Examples: **(i)** 2g of H_2 = 1 mole of H_2 = 22.414 dm^3

(ii) 16g of CH_4 = 1 mole of CH_4 = 22.414 dm^3

According to Avogadro's Law

There are 6.02×10^{23} Number of particles present in 22.414dm^3 (1 mole) of a gas.

xvii. What is Stoichiometry and Stoichiometric amounts.

Ans: Stoichiometry:

The study of relative amounts of substances involved in a chemical reaction is called Stoichiometry. Such phenomenon is studied through the knowledge of Stoichiometry (Greek word *Stoicheion* means element and *metry* means measurement).

Stoichiometric Amounts:

The amounts of the reactants or the products as given by the balanced chemical equation are called stoichiometric amounts.

$$\begin{array}{ccccc} 2H_2 & + & O_2 & \longrightarrow & 2H_2O \\ 2\,mole & & 1\,mole & & 2\,moles \\ 4\,g & & 32\,g & & 36\,g \end{array}$$

3. (a) What is Avogadro's number? Give examples. How will be explain moles with the help of Avogadro's Number.

(b) The liquid $CHBr_3$ has a density of 2.89 g dm^{-3}. What volume of this liquid should be measured to contain a total of 2.4×10^{24} molecules of $CHBr_3$. **(Ans: 348.4 cm^3)**

Ans: (a) What is Avogadro's number? Give examples. How will be explain moles with the help of Avogadro's Number.

The number of atoms, ions, molecules or formula units present in 1 mole of a substance in called Avogadro's number and its numerical value is 6.02×10^{23}

e.g. 1 mole of Na(23g) = 6.02×10^{23} atoms of Na

1 mole of NaCl = 6.02×10^{23} formula units of NaCl

1 mole of NaCl = 6.02×10^{23} Na^+ ions

1 mole of NaCl = 6.02×10^{23} Cl^- ions

Definition of Mole in terms of Avogadro's number:

6.02×10^{23} atoms of an element, 6.02×10^{23} molecules of a compound or 6.02×10^{23} formula units of an ionic substance is called a mole.

e.g. 6.02×10^{23} atoms of Na = 1 mole of Na

6.02×10^{23} molecules of H_2O = 1 mole of H_2O

6.02×10^{23} formula units of NaCl = 1 mole of NaCl

(b) The liquid $CHBr_3$ has a density of 2.89 g dm^{-3}. What volume of this liquid should be measured to contain a total of 4.8×10^{24} molecules of $CHBr_3$.

Solution: **(Ans: 696.8 dm^3)**

$$Density\ of\ CHBr_3 = 2.98\ g\ dm^{-3}$$
$$Volume = ?$$
$$Molecules\ of\ CHBr_3 = 4.8 \times 10^{24}$$
$$Molar\ mass\ of\ CHBr_3 = 12 + 1.008 + 239.7 = 252.7$$
As we know that
$$Number\ of\ molecules = moles \times N_A$$
$$Number\ of\ molecules = \frac{Mass\ in\ gram}{Molar\ mass} \times N_A$$
$$Mass\ in\ gram = \frac{252.7 \times 4.8 \times 10^{24}}{6.022 \times 10^{23}}$$
$$Mass\ in\ gram = 2014.2\ gram$$
$$Volume = \frac{Mass\ in\ gram}{Density}$$
$$Volume = \frac{2014.2}{2.89} = 696.8\ dm^3$$

4. (a) Differentiate between actual yield and theoretical yield. How will you explain the percentage yield of the substance with the help of;

(b) The following reaction never goes to completion. Therefore less amount of NH_3 is obtained than expected theoretically,

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

42.0 g of H_2 produces 120.2 g of NH_3. Calculate the percentage yield of NH_3. **(Ans: 50.5 %)**

Ans: (a) Differentiate between actual yield and theoretical yield.

Actual Yield:

The quantity of product that is actually produced in a chemical reaction is called the actual yield.

Theoretical Yield:

"The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction".

The theoretical yield is not an estimate, but the calculated amount of the yield based on the best of conditions for the reaction being carried to completion.

The actual yield is the measured amount from the actual experiment. This is often less than the ideal theoretical yield because of other factors that affect the reaction, mainly that the reagent used in the reaction is consumed so that you have less material to progress the reaction to its full theoretical yield.

(b) The following reaction never goes to completion. Therefore less amount of NH_3 is obtained than expected theoretically,

$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$ **42.0 g of H_2 produces 120.2 g of NH_3.**

Calculate the percentage yield of NH_3. **(Ans: 50.5 %)**

Ans. Expected theoretical yield can be calculated as:

Mass of H_2 = 42g

Number of moles of H_2 $= \frac{Mass\ in\ gram}{Molar\ mass} = \frac{42}{2} = 21$ moles

According to reaction equation

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

3 moles of H_2 = 2 moles of NH_3

1 mole of H_2 $= \frac{2}{3}$ moles of NH_3

21 moles of H_2 $= \frac{2}{3} \times 21 =$ 14 moles of NH_3

Mass of NH_3 in grams = 14 × 17 = 238 g

But actual yield = 120.2 g

$$Percentage\ yield = \frac{Actual\ yield}{Theoretical\ yield} \times 100$$

$\%\ yield = \frac{120.2}{238} \times 100 = 50.5\ \%$

5. (a) What do you know about percentage composition? How will you determine the percentage of each element in the substance?

(b) Glucose ($C_6H_{12}O_6$) is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass percentage of each element in glucose.

(Ans: C = 40 %, H = 6.66 %, O = 53.33 %)

Ans: (a) What do you know about percentage composition? How will you determine the percentage of each element in the substance?

Method to determine the percentage composition of a known compound:

i) calculate the molar mass of compound

ii) calculate the percentage of each element in one mole of the compound. This is done by dividing the mass of each element in one mole of the compound by the molar mass multiplied by 100.

$$\%\ of\ an\ element = \frac{mass\ of\ element\ in\ compound}{Molar\ mass\ of\ compound} \times 100$$

(b) Glucose ($C_6H_{12}O_6$) is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass percentage of each element in glucose. (Ans: C = 40 %, H = 6.66 %, O = 53.33 %)

Solution: $Molar\ Mass\ of\ Glucose\ (C_6H_{12}O_6) = 72 + 12 + 96 = 180\ g/mole$

$$\%\ of\ an\ element = \frac{mass\ of\ element\ in\ compound}{Molar\ mass\ of\ compound} \times 100$$

$$\% \, of \, C = \frac{72}{180} \times 100 = 40\%$$
$$\% \, of \, H = \frac{12}{180} \times 100 = 6.66\%$$
$$\% \, of \, O = \frac{96}{180} \times 100 = 53.33\%$$

6. (a) How will you calculate the theoretical yield and percentage yield in a balanced chemical equation?

(b) A small piece of pure Al Metal having a volume of 2.50 cm^3 is reacted with excess of HCl. What is the weight of H_2 liberated? The density of Al is 2.70 g cm^{-3}. (Ans: 0.752g)

Ans: (a) How will you calculate the theoretical yield and percentage yield in a balanced chemical equation?

Theoretical Yield:

"The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction".

Method of finding theoretical yield:

i. find the number of moles of reactant which is not an excess.

ii. Find the number of moles of required product from the balanced chemical equation.

iii. Find the mass of product by using formula: Number of moles = $\frac{Mass \; in \; gram}{Molar \; mass}$

iv. The mass of the product is the theoretical yield.

$$Theoretical \; yield = \frac{Actual \; yield}{\% \; yield} \times 100$$

Percent Yield: Percent yield is a measure of the efficiency of a chemical reaction.

Percent yield can be calculated from the balanced chemical equation it is done by .divided the theoretical yield with actual yield multiplying it by 100.

$$\% \; yield = \frac{Actual \; yield}{Theoretical \; yield} \times 100$$

(b) A small piece of pure Al Metal having a volume of 2.50 cm^3 is reacted with excess of HCl. What is the weight of H_2 liberated? The density of Al is 2.70 g cm^{-3}. (Ans: 0.752g)

$Volume \; of \; Al = 2.50 \; cm^3$, $Density \; of \; Al = 2.70 \; g \; cm^{-3}$

$$Density = \frac{mass}{volume}$$

$$Mass \; of \; Al = Density \times Volume$$

$$= 2.70 \times 2.50 = 6.75 \; g$$

$$Moles \; of \; Al = \frac{Mass \; in \; gram}{Molar \; mass} = \frac{6.75}{27} = 0.25 \; moles$$

According to Reaction;

$$2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$$

$$2 \; moles \; of \; Al = 3 \; moles \; of \; H_2$$

$$1 \; mole \; of \; Al = \frac{3}{2}$$

$$0.125\ moles\ of\ Al = \frac{3}{2} \times 0.25 = 0.375\ moles$$

$molecular\ mass\ of\ H_2 = 2$

$Mass\ of\ H_2\ in\ grams = 0.375 \times 2 = 0.752\ g$

7. How much Silver Chloride will be formed by mixing 120.0 g of Silver Nitrate with a solution of 52.0 g of NaCl.

$(AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3)$ **(Ans: 101.24 g)**

Solution: $Given\ mass\ of\ AgNO_3 = 120\ g$

$Molar\ mass\ of\ AgNO_3 = 107.87 \times 1 + 14 \times 1 + 16 \times 3$

$= 169.87\ g/mole$

$$Number\ of\ moles\ AgNO_3 = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{120}{169.868} = 0.71\ moles$$

$Mass\ of\ NaCl = 52\ g$

$Molar\ mass\ of\ NaCl = 23 + 35.5 = 58.5\ g/mole$

$$Number\ of\ moles\ NaCl = \frac{Mass\ in\ gram}{Molar\ mass} = \frac{52}{58.5} = 0.89\ moles$$

$According\ to\ the\ balanced\ chemical\ equation.$

$1\ mole\ of\ AgNO_3 = 1\ mole\ of\ AgCl$

$0.71\ mole\ of\ AgNO_3 = 0.71\ mole\ of\ AgCl$

$Also$

$1\ mole\ of\ NaCl = 1\ mole\ of\ AgCl$

$0.89\ mole\ of\ AgNO_3 = 0.89\ mole\ of\ AgCl$

Since, $AgNO_3$ produces least amount of product therefore, it is the limiting reactant.

$Number\ of\ moles\ AgCl\ produces = 0.71\ moles\ of\ AgCl$

$Molar\ mass\ of\ AgCl = 107.87 + 35.5 = 143.37\ g/mole$

$Mass\ of\ AgCl\ produced = 0.71 \times 143.37 = 101.7\ g$

8. Which contains more atoms, 1 mole of Fe or 1 mole of H_2? Justify your stand. **(Ans: H_2)**

Solution: 1 mole of Fe atoms $= 6.02 \times 10^{24}$ Fe atoms

1 mole of H_2 atoms $= 2 \times 6.02 \times 10^{24}$ H atoms

$= 12.04 \times 10^{24}$ H atoms

Therefore 1 mole of H_2 contains more atoms than 1 mole of Fe.