CHAPTER # 1

STOICHIOMETRY

QI. 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.

$$\underset{(g)}{2H_2} + \underset{(g)}{O_2} \longrightarrow \underset{(g)}{2H_2O}$$

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 .mass In grams

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

Example: One mole of O = 16 g

One mole of O_2= 32g

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

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

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 $2 \times 6.02 \times 10^{23}$ molecules of hydrogen react with 6.02×10^{23} molecules of oxygen to produce $2 \times 6.02 \times 10^{23}$ molecules of water.

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

$$\text{Number of moles of a substance} = \frac{\text{Number of molecules of substances}}{6.02 \times 10^{23}}$$

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 = 6.02×10^{23} atoms

1 mole of O_2(gas) = 6.02×10^{23} molecules

= 22.414 dm^3 (volume occupied by 1 mole of gas at S.T.P)

1 mole of H20 (1) = 6.02×10^{23} molecules

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

For example,

$$\text{(i)} \quad NaCl \xrightarrow{H_2O} Na^{+1}_{(aq)} + Cl^{-1}_{(aq)}$$

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 x 10^{23}formula units of NaCl are dissolved in water there are produced Na^{1+} and 6.023x Cl^{-1} ions.

$$\text{(ii)} \quad C + O2 \longrightarrow CO_2$$

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

$$C3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

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 = $\dfrac{3\text{moles } O_2}{2\text{moles } CH_30H}$

Desired quantity (of O_2) = Given quantity x 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_30H = 3.5 moles

Desired quantity of H_2O = ?

Conversion factor (or mole ratio) = $\dfrac{4 \text{ moles } H_20}{2 \text{ moles } CH_30H}$.

Desired quantity (i.e. No of moles of H_20 formed)

$$= \text{Given quantity of } CH_30H \times \frac{4 \text{ moles } H_2O}{2 \text{ moles } CH_3OH}$$

$$=3.50 \text{ moles of } \cancel{CH_30H} \times \frac{4 \text{ moles } H_2O}{2 \text{ moles } \cancel{CH_3OH}}$$

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

Quantity of H_2O = 7.00 moles of H_2O

Self Check Exercise 1.1

NH, is an important raw material in the manufacture of fertilizers. It is obtained by the combination of Na and Ha 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: Stoidliometric Calculation:

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

2 moles of NH_3 = 1 mole of N_2

1 moles of NH_3= ½ moles of N_2

5 moles of NH_3= 5 x ½ moles of N_2 = 2.5 moles of N_2

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

2 moles of NH_3 = 3 mole of H_2

1 moles of NH_3 = 3/2 moles of H_2

***5 moles of NH_3*= 5 x 3/2 moles of H_2= 7.5 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)

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

$$= 2.66 \text{ moles of } Fe_2O_3$$

Number of moles of iron produced:

i.e Desired quantity = Given quantity x conversion factor

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

$$= 2.66 \cancel{\text{Moles } Fe_2O_3} \times \frac{\text{No of moles of Fe}}{1 \cancel{\text{moles } Fe_2O_3}}$$

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

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

Desired quantity = Given quantity x conversion factor

$$(\text{Mass of Fe produced}) = 5.\cancel{32 \text{ Mole}} \text{ Fe} \times \frac{55.9g}{1 \cancel{\text{Mole}} \text{ Fe}}$$

$$= 5.32 \times 55.9 \text{ g}$$

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 x IO^5 kg of liquid hydrogen is carried on a particular launch, what mass of liquid oxygen is necessary for all the hydrogen to bum, The equation for the reaction is,

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

(Ans: 8.16 × IO^5 kg oxygen)

Solution:

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

$$= 1.02 \times 10^8 \text{ g}$$

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

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

= 5.1×10^7 moles

From given balanced equation it is clear that:

2 moles of H_2= 1 mole of O_2

1 mole of H_2 = 2 mole of O_2

5.1×10^7 moles of H_2 = ½ x 2 5.1×10^7 moles of O_2

= 2.55×10^7 moles of O_2

Now, Mass of Oxygen = Number of moles x Molar mass

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

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

Sample Problem No. 1.3

Calculate the number of molecules 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.
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 x 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

$= 3.612 \times 10^{24}$ 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 drn ^{3} of any gas at STP = 1 mole = 6.02×10^{23} molecules.

22.414 dm ^{3} of H2 gas at STP = 2g = 6.02×10^{23} molecules.

22 414 dm ^{3} of NH3 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 Cl2 (Chlorine) at S.T.P. = 1 mole

Or 1 mole of Cl2 occupies a volume of 22.4 dm^3 at S.T.P.

2.5 mole of Cl2 occupy a volume of 22.4dm^3 x 2.5 = 56dm^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 N2 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 0f O_2 at STP = 1 mole of O_2

1 dm^3 of O_2at STP = $\frac{1}{22.414}$

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

(b) We know that

1 mole of N2 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 ofN_2 = 22.414 x 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 motes of products formed from data of the given amounts of the reactants using a balanced chemical equation. The reactant, which produces the least number 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 reactants 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

$O_2 = 2H_2O$

i.e. 1 Mole of O_2 = 2 Moles of H_2O

so, 7 Moles of O_2 = 2 x 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.

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

= 0.68 moles

Mass of H_2SO_4 =200g mole.

Formula Mass of H_2SO_4 = 98g

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

= 2.04 moles

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

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 x Molecular mass

= 16 x 1.53 = 24.48 g

b) In this problem H2S04 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 motes 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 = ¼ x 2.04

= 0.51 moles of $K_2Cr_2O_7$

No of moles of K2Cr207 left unreacted= 0.68 - 0.51

= 0.17 moles

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

= 0.17 x 294 = 49.98

So Mass of $K_2Cr_2O_7$ left unreacted = 49.989

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 $= \dfrac{\text{Mass of } H_2SO_4}{\text{Molar Mass of } H_2SO_4} \times 6.02 \times 10^{23}$

$= \dfrac{20}{98.016} \times 6.02 \times 10_{23}$

$= 1.228 \times 10^{23}$

b. H_2SO_4 dissolves in water as follows

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

According to equation

1 molecule of H_2SO_4= $2H^+$ions

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

As 1 molecule 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^+= $1.008 \quad \times 2.456 \times 10^{23}$ ______

6.02

= 0.411g

Mass of SO_4^{2-} = 96 $\times$ $\dfrac{1.228 \times 10^{23}}{6.02 \times 10^{23}}$ = 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 $= \dfrac{1.50}{24} = 0.0625$ 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 ofMgS

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 x 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. (c) Zinc Is the limiting since the whole Is consumed.

(d) Mass of Zinc Sulphide produced 8.94g)

Solution: Mass of Zn = 6g

Atomic mass of Zn = 65.41

Number of moles of Zn = $\frac{\text{Mass in gram}}{}$ = 6 =0.0917 moles

Atomic mass 65.41

Mass of S = 4g

Atomic mass of S = 32 g

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

$Zn + S \longrightarrow ZnS$

1 mole of Zn = 1 mole of ZnS

O. 0917 moles. of Zn = O. 0917 moles of ZnS

Also, 1 mole of S = 1 mole of ZnS

so, O. 125 moles of S O. 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 O. 0917 moles of ZnS = 97.41 x 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, $2Al + 3Br_2 \longrightarrow 2AlBr$

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**

(a) Bromine is the limiting reactant.

(b)Mass of AlBr3 formed = 61.99

Solution: Mass of Al = 15.8 g

Atomic mass of Al = 27

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

Mass of Br_2 = 55.6 g

Molar mass of Br2 = 79.9 x 2 = 159.8 g/mole

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

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

2 mole of Al = 2 moles of $AlBr_3$

1 mole of Al = 2/2 = 1 mole of $AlBr_3$

0.585 mole of Al = 0.585 mole of $AlBr_3$

Also, 3 moles of Br_2 = 2 moles of AlBr3

1 mole of Br_2 = 2/3

1 mole of Br_2 = 2/3 x 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 x 3 = 27+ 239.7 = 266.7 g

Mass of 0.232 moles of AlBr3 = 0.232 x 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 unreacted 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 who's 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 = ½ x 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

Sot 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 x 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.
ii. 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%.

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

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,

$$Zn\,(g) + CuSO_{4\,(aq)} \longrightarrow ZnSO_{4\,(aq)} + Cu(s)$$

1.274g CuS04 when reacted with excess of Zn metal a yield of 0.392g Cu metal was obtained. Calculate the percentage yield.

Solution:

According to reaction $\quad Zn + CuSO_4 \qquad ZnSO_4 + Cu$

Mass of $CuSO_4$ given $\quad = 1\,.274g$

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

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

159.6g of $CuSO_4$ = 1 mole

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

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

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 = 63.5g

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, $CH_4 + 4Cl_2 \qquad CCl_4 + 4HCl_{(g)}$

Stoichiometrically,

From 2.0 moles of CH_4 we would expect to obtained 2.0 moles of CCl_4

(a)Theoretical yield = 2.0 moles of CCl_4

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

(b)Actual yield = 177.09 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 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 x 6 + 1 x 12 + 16 x 6 = 180 g/mole

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

1 mole of $C_6H_{12}O_6$= 2m01es of C_2H_5OH

O. 056 moles of $C_6H_{12}O_6$ = 2 x 0.056 moles of C_2H_5OH

= 0.112 moles of C_2H_5OH

Molar mass of C_2H_5OH = 12 x 2 + 1 x 6 + 16 x 1 = 46 g/mole

Mass of C_2H_5OH = 0.112 x 46 = 5.125 g

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

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.

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

What mass of CO_2 should be obtained from the complete reaction of 1250 g of methane?

If the actual yield obtained is 3000g then what Is the percentage yield?

(Ans: a = 3438 9 b = 87.3 %)

Solution: (a)

Given mass of methane = 1250 g

Molar mass of CH_4= 12 + 14 $\frac{16\ g}{\text{mole}}$

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

Stoichiometrically,

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 x molar mass

= 78.125 x 44 = 3437.5g

(b) Actual yield = 3000g

Theoretical yield = 3437.5g

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

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

Percentage yield = 87.3 %