

CHAPTER # 01

STOICHIOMETRY

This word was derived from *stoicheion* means *element* and *metry* means *measurement*.

STOICHIOMETRY: Study of relative amounts of substances in a chemical reactions is called stoichiometry.

The branch of chemistry which deals with the study of quantitative relationship between reactants and products in balanced chemical equations is called stoichiometry.

STOICHIOMETRIC AMOUNT:

Amounts of reactants and products in balanced chemical equation is called stoichiometric amount.

What is stoichiometry and stoichiometric amount? [Q No 2 (xiii)]

MOLE:

When atomic mass, molecular mass, formula mass or ionic mass of a substance is expressed in gram is called mole.



FORMULA:

$$\text{Mole} = \frac{\text{Mass in gram}}{\text{Molar mass}}$$

MOLE CALCULATION:

$$1 \text{ mole of O} = 16 \text{ g}$$

$$1 \text{ mole of O}_2 = 32 \text{ g}$$

$$1 \text{ mole of NaCl} = 58.5 \text{ g}$$

$$1 \text{ mole of OH} = 17 \text{ g}$$

NOTE: Mole is also called gram atom, gram molecule, gram formula or gram ion.

GRAM ATOM

When atomic mass of an element is expressed in gram is called gram atom or gram mole.

$$\text{Gram atom} = \frac{\text{Mass in gram}}{\text{Atomic mass}}$$

GRAM MOLECULE:

When molecular mass of a molecule or covalent compound is expressed in gram is called gram molecule.

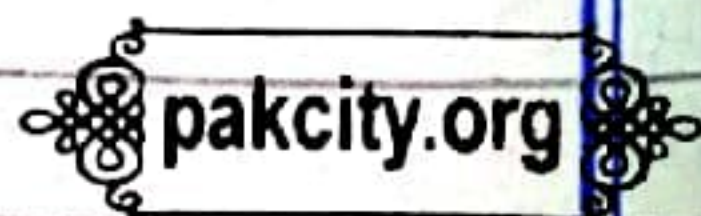
$$\text{Gram molecule} = \frac{\text{Mass in gram}}{\text{Molecular mass}}$$

GRAM FORMULA:

When formula mass of ionic compound is expressed in gram is called gram formula.

$$\text{Gram formula} = \frac{\text{Mass in gram}}{\text{Formula mass}}$$

GRAM ION:



When ionic mass of ionic substance is expressed in gram is called gram ions.

$$\text{Gram ions} = \frac{\text{Mass in gram}}{\text{Ionic mass}}$$

AVOGADRO'S NUMBER (N_A) OR REPRESENTATIVE PARTICLES

It is the number of atoms, ions, molecules or formula units in one mole of a substance is called Avogadro's Number.

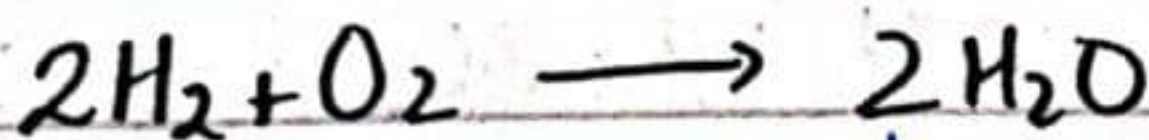


It is denoted by N_A .

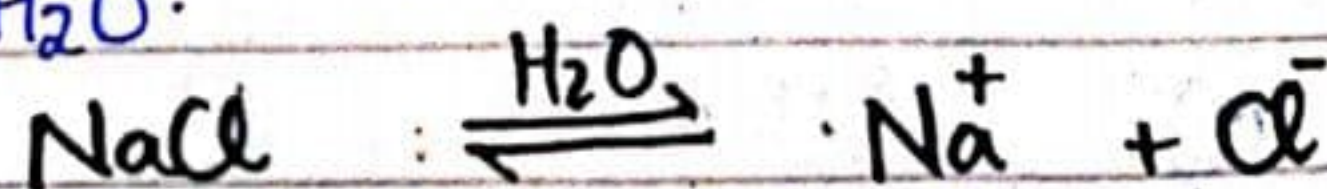
$$N_A = 6.022 \times 10^{23} \text{ Particles.}$$

Examples:

- 1 mole of O = 6.02×10^{23} atoms
- 1 mole of H_2O = 6.02×10^{23} molecules
- 1 mole of NaCl = 6.02×10^{23} formula units
- 1 mole of OH^- = 6.02×10^{23} ions.



According to Avogadro's Number, $2 \times 6.02 \times 10^{23}$ molecules of H_2 react with 6.02×10^{23} molecules of O_2 to produce $2 \times 6.02 \times 10^{23}$ molecules of H_2O .

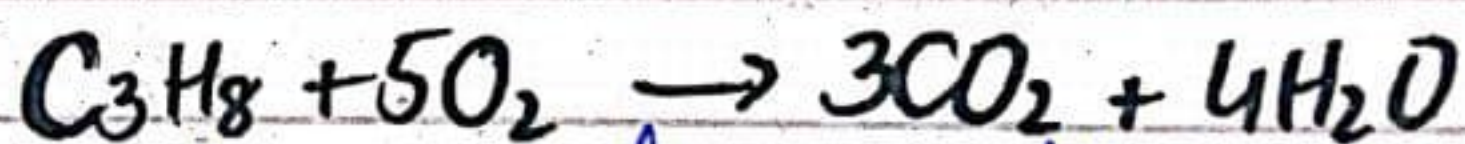


According to Avogadro's Number, 6.02×10^{23} formula units of NaCl dissolve in water produce 6.02×10^{23} ions of sodium and 6.02×10^{23} ions of chlorine.

MOLE RATIO OR CONVERSION FACTOR

It is the ratio between number of moles of reactants taking part in chemical reaction and no. of moles of products produce in a chemical reaction is called mole ratio.

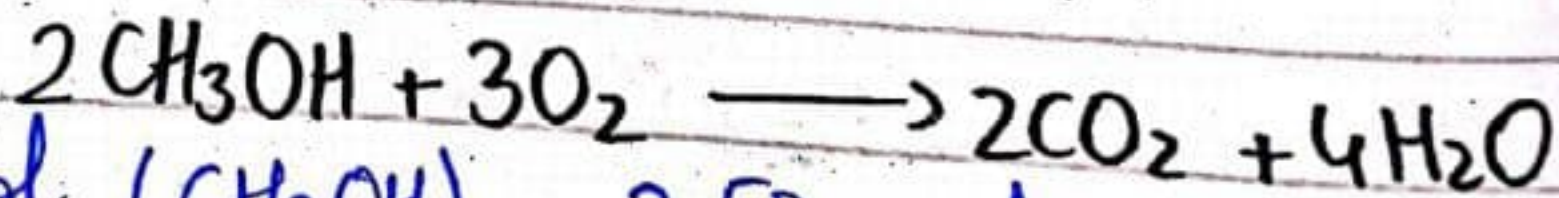
e.g.



According to mole ratio, 1 mole of propane react with 5 mole of oxygen to produce 3 mole of CO_2 and 4 mole of steam.

NUMERICAL

e.g 1-1



a) Methanol (CH_3OH) = 3.50 moles
moles of O_2 used = ?

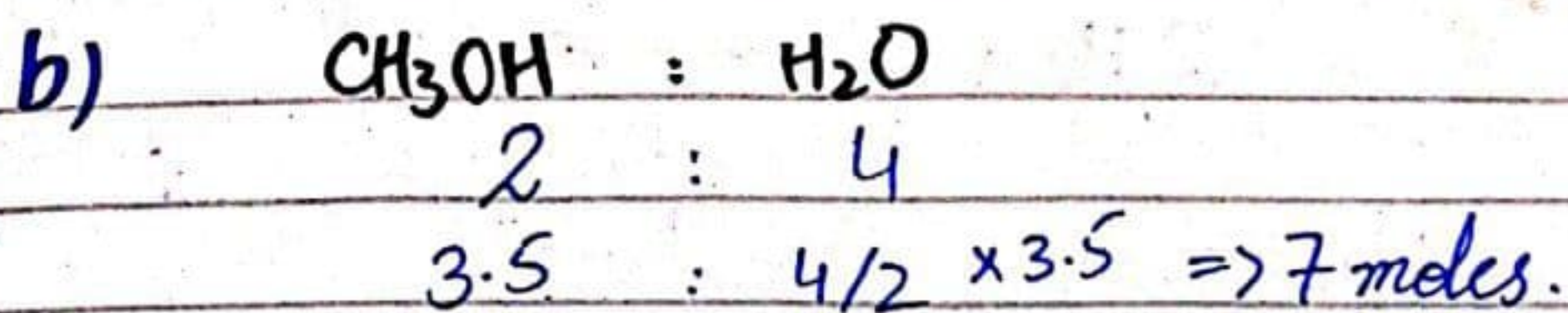
b) moles of H_2O produce = ?

a) $\text{CH}_3\text{OH} : \text{O}_2$

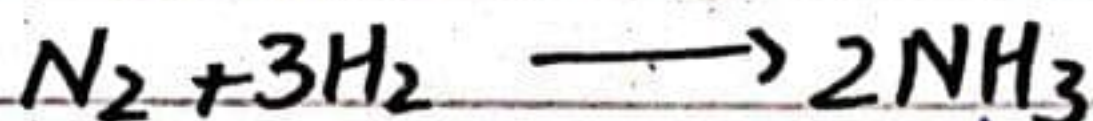
2 : 3

1 : $\frac{3}{2}$

3.5 : $\frac{3}{2} \times 3.5 \Rightarrow 5.25$ moles



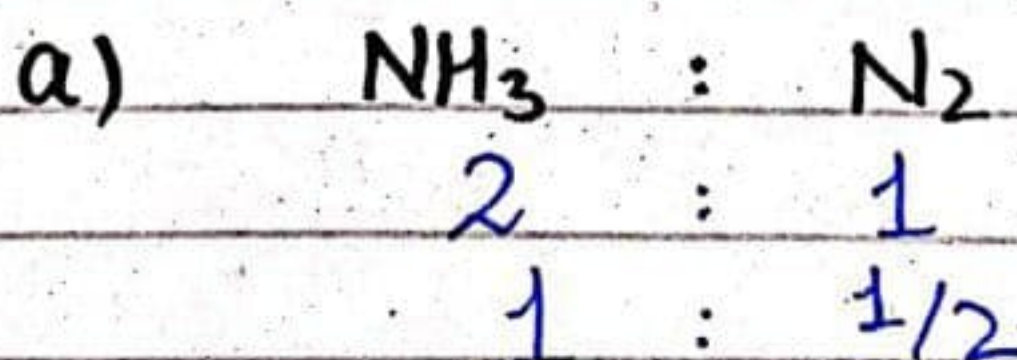
Self check 1-1



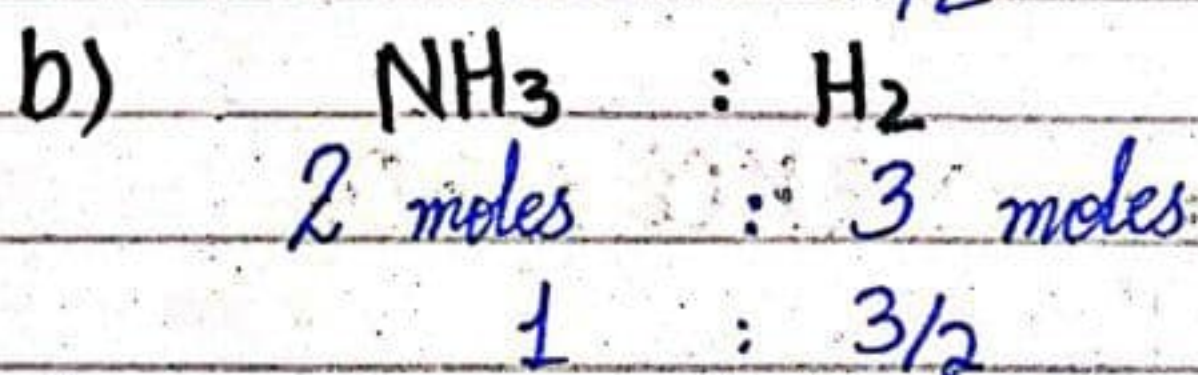
Moles of $\text{NH}_3 = 5.0 \text{ moles}$

a) Mole of $\text{N}_2 = ?$

b) Mole of $\text{H}_2 = ?$

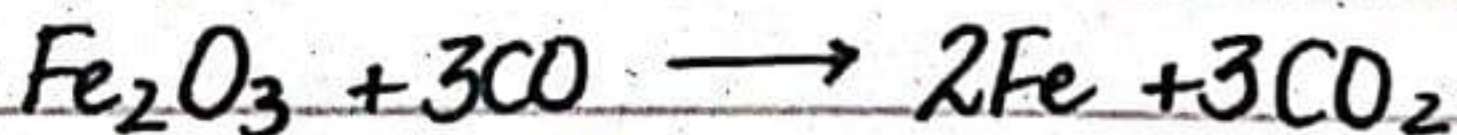


$$5 \times 1 : 1/2 \times 5 \Rightarrow 2.5 \text{ moles}$$



$$1 \times 5 : 3/2 \times 5 \Rightarrow 7.5 \text{ moles}$$

Example 1.2:

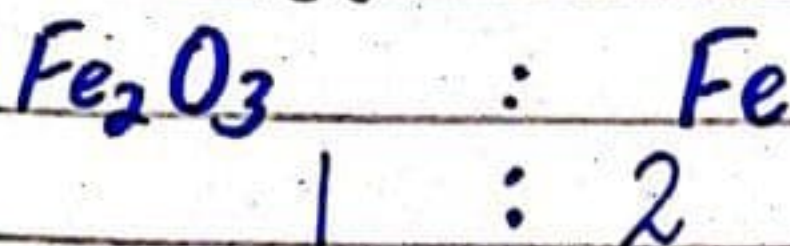


Mass of $\text{Fe}_2\text{O}_3 = 425 \text{ g}$

Mass of Fe formed = ?

Molar mass of $\text{Fe}_2\text{O}_3 = 2 \times 56 + 16 \times 3 = 160 \text{ g}$

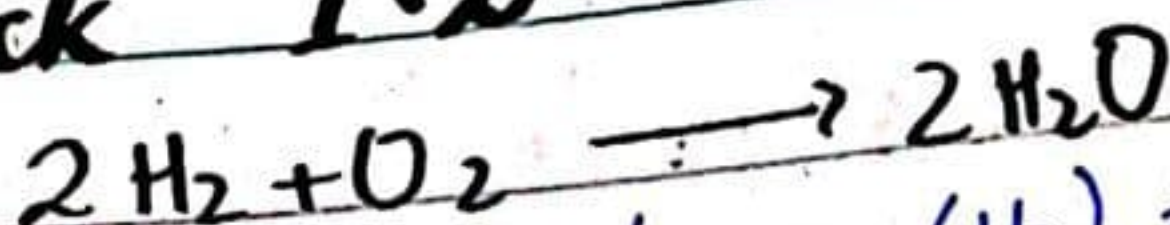
$$\text{Mole} = \frac{425}{160} = 2.66 \text{ moles}$$



$$2.66 : 2 \times 2.66 \Rightarrow 5.32 \text{ moles.}$$

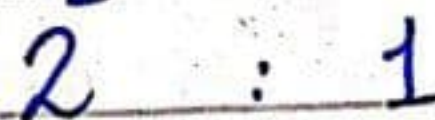
$$\begin{aligned} \text{Mass in gram of Fe} &= \text{Mole} \times \text{Molar mass} \\ &= 5.32 \times 56 \Rightarrow 297.92 \text{ g} \end{aligned}$$

Self-check 1.2



Mass of hydrogen (H_2) = $1.02 \times 10^5 \text{ kg}$
 Mass of oxygen (O_2) = ?

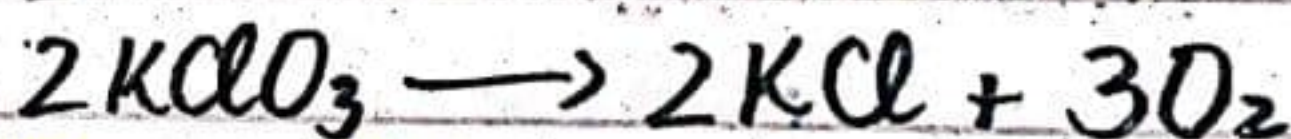
$$\text{Mole} = \frac{1.02 \times 10^5}{2.016} \Rightarrow 0.5059 \times 10^8 \text{ moles}$$



$$0.5059 \times 10^8 : \frac{1}{2} \times 0.5059 \times 10^8 \Rightarrow 0.253 \times 10^8 \text{ moles}$$

$$\begin{aligned} \text{Mass in gram of } (\text{O}_2) &= 0.253 \times 10^8 \times 32 \\ &= 8.09 \times 10^8 \text{ g} \\ &= 8.09 \times 10^5 \text{ kg} \end{aligned}$$

Example 1.3

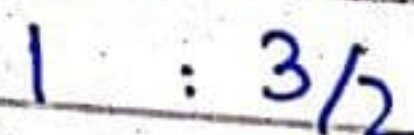
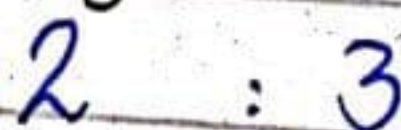
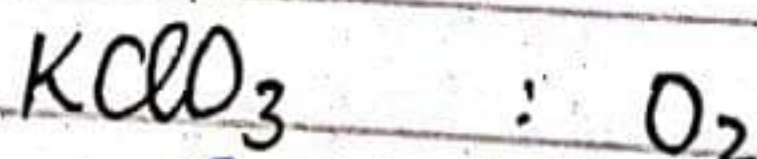


$$\text{KClO}_3 = 490 \text{ g}$$

No: of molecules of O_2 = ?

$$\text{Molar mass of KClO}_3 = 39 + 35.5 \times 16 \times 3 = 122.5 \text{ g/mol}$$

$$\text{Moles of KClO}_3 = \frac{490}{122.5} \Rightarrow 4 \text{ moles}$$



$$4 : \frac{3}{2} \times 4 \Rightarrow 6 \text{ moles}$$

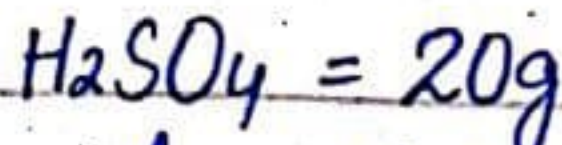
$$\text{No: of molecules} = \text{Mole} \times 6.02 \times 10^{23}$$

$$= 6 \times 6.02 \times 10^{23} \Rightarrow 36.12 \times 10^{23}$$

$$\Rightarrow 3.612 \times 10^{24} \text{ molecules}$$

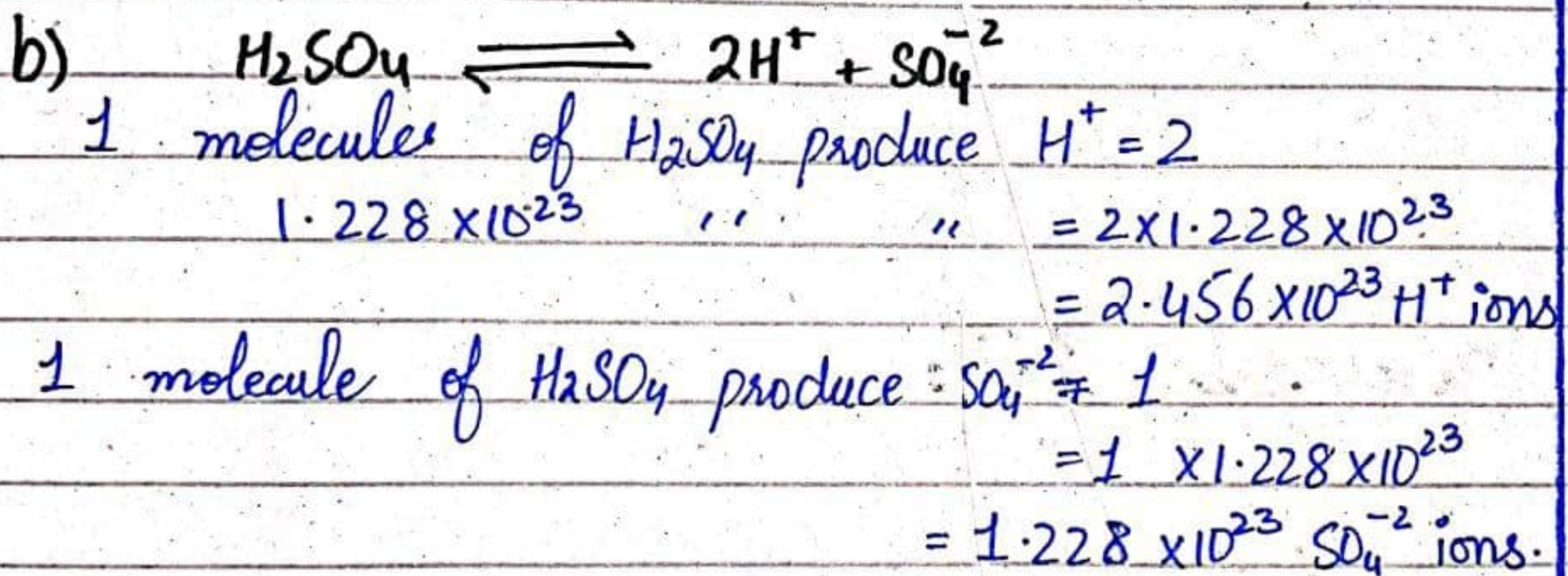
Example 1.4:

Pg 08:

a) No: of H_2SO_4 molecules = ?b) No: of H^+ & SO_4^{-2} ions

c) Mass of individual ions

$$\begin{aligned} \text{a) No: of molecules} &= \frac{\text{Mass in gram}}{\text{Molar mass}} \times 6.02 \times 10^{23} \\ &= \frac{20}{98} \times 6.02 \times 10^{23} \\ &\Rightarrow 1.228 \times 10^{23} \text{ molecules} \end{aligned}$$



$$\text{No: of ions} = \frac{\text{Mass in gram}}{\text{ionic mass}} \times N_A$$

$$2.456 \times 10^{23} = \frac{\text{Mass in gram H}^+}{1.008} \times 6.02 \times 10^{23}$$

$$\text{Mass of H}^+ = 0.411 \text{ g}$$

$$\text{No: of SO}_4^{-2} \text{ ions} = \frac{\text{Mass in gram}}{\text{ionic mass}} \times N_A$$

$$1.228 \times 10^{23} = \frac{\text{Mass in gram of SO}_4^{-2}}{96} \times 6.02 \times 10^{23}$$

$$\text{Mass in gram of SO}_4^{-2} = 19.58 \text{ g}$$

$$22.414 \text{ dm}^3 = 22.414 \text{ L}$$

$$22.414 \text{ cm}^3 = 22.414 \text{ mL}$$

$$\therefore 1 \text{ dm}^3 = 1 \text{ L}$$

$$1 \text{ cm}^3 = 1 \text{ mL} \quad | 1 \text{ dm}^3 = 1000 \text{ cm}^3$$

Pg 09:

MOLAR

VOLUME (V_m)

1 mole of any gas at S.T.P occupy volume 22.414 dm^3 . This volume is called molar volume.

Examples:

1 mole of $\text{H}_2 = 2.016 \text{ g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ molecules
 1 mole of $\text{CH}_4 = 16 \text{ g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ molecules.
 1 mole of $\text{N}_2 = 28 \text{ g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ molecules

Example 1.5:

Moles of $\text{Cl}_2 = 2.5$ moles
 $\therefore$ vol. of Cl_2 at S.T.P = ?
 1 mole of Cl_2 has volume at S.T.P = 22.414 dm^3
 2.5 moles " " = 22.414×2.5
 $= 56.035 \text{ dm}^3$

Self-check 1.3:

a) Moles of oxygen molecules (O_2) =
 vol. of O_2 at STP = 50 dm^3
 22.414 dm^3 of O_2 at STP has no. of mole = 1 mole
 1 dm^3 " " " = 1

$50 \times 1 \text{ dm}^3$ " " = $\frac{50}{22.414}$ moles
 $= 2.23$ moles

b) Mole of N_2 at STP = 0.80 mole
 volume of N_2 at STP = ?

Pg 10:

1 mole of N_2 at STP has volume = 22.414 dm^3
0.80 moles " " = $\frac{22.414}{1} \times 0.80$
 $= 17.93 \text{ dm}^3$

PERCENTAGE



COMPOSITION

Relative amount of each element present in a compound is called percentage composition.

Determination of percentage composition:

There are 2 steps for determination of percentage composition.

1. Calculate molar mass of compound.
2. Calculate percentage of each element by using this formula.

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

Example: Calculate percentage composition of MgO ?
Molar mass of MgO = $24 + 16 = 40 \text{ g}$

Pg 11:

$$\% \text{ of Mg} = \frac{24}{40} \times 100 \Rightarrow 60\%$$

$$\% \text{ of O} = \frac{16}{40} \times 100 \Rightarrow 40\%$$

Example 2: Calculate percentage composition of glucose? Q No 05: b)

$$\text{Molar mass of } C_6H_{12}O_6 = 12 \times 6 + 1 \times 12 + 16 \times 6 \Rightarrow 180 \text{ g mol}^{-1}$$

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

$$\% \text{ of H} = \frac{12}{180} \times 100 \Rightarrow 6.66\%$$

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

Example 3: Calculate percentage composition of sucrose?

$$\text{Molar mass of } C_{12}H_{22}O_{11} = 12 \times 12 + 1 \times 22 + 16 \times 11 \Rightarrow 342 \text{ g mol}^{-1}$$

$$\% \text{ of C} = \frac{144}{342} \times 100 \Rightarrow 42.1\%$$

$$\% \text{ of H} = \frac{22}{342} \times 100 \Rightarrow 6.4\%$$

$$\% \text{ of O} = \frac{176}{342} \times 100 \Rightarrow 51.46\%$$

LIMITING AND NON-LIMITING REACTANT

Q No 02: iv

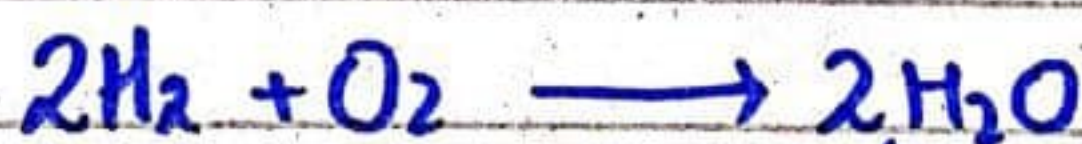
Limiting reactant

* Reactants which consume earlier during the chemical reaction is called limiting reactant.

* Reactant which produce least mole of product during the chemical reaction is called limiting reactant.

* It controls the amount of product.

Example:



Moles of $\text{H}_2 = 10$ moles

Moles of $\text{O}_2 = 7$ moles.

According to balanced chemical equation:

$\text{H}_2 : \text{H}_2\text{O}$

2 : 2

1 : 2/2

10 : $\frac{2}{2} \times 10 \Rightarrow 10$ moles.

According to balanced chemical equation:

$\text{O}_2 : \text{H}_2\text{O}$

1 : 2

7 : $2 \times 7 \Rightarrow 14$ moles.

Pg 13:

1. H_2 is a limiting reactant in this reaction.
2. O_2 is a non-limiting reactant in this reaction.

Identification of limiting reactant

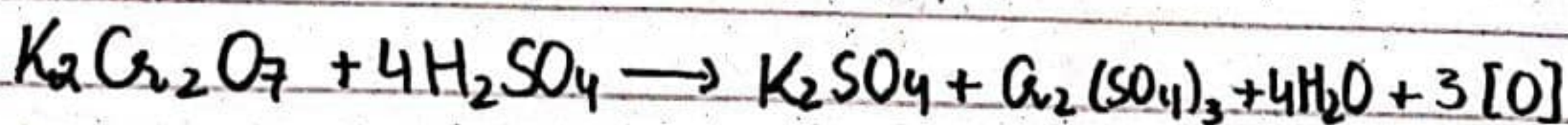
There are three steps for identification of limiting reactant.

- * Calculate no. of moles of reactant from given amount of reactant.
- * Calculate mole of product with respect to each reactant by using balanced chemical equation.
- * Identify the reactant which produce least mole of product.

NOTE:



- * There is no concept of limiting reactant in reversible reaction. Because in reversible reaction, reactant are not completely converted into product.
- * If reactant are mixed with stoichiometric ratio, there is no limiting reactant in this reaction.

Example No 1.6:Mass of $K_2Cr_2O_7 = 200g$ Mass of $H_2SO_4 = 200g$ **A)** Calculate mass of atomic oxygen produce = ?**B)** Mass of reactant left unreactant = ?Molar mass of $K_2Cr_2O_7 = 2 \times 39 + 2 \times 52 + 7 \times 16 \Rightarrow 294 g mol^{-1}$ Molar mass of $H_2SO_4 = 2 \times 1 + 1 \times 32 + 4 \times 16 = 98 g mol^{-1}$ Moles of $K_2Cr_2O_7 = \frac{200}{294} = 0.68 \text{ moles}$ Moles of $H_2SO_4 = \frac{200}{98} \Rightarrow 2.04 \text{ moles}$ $K_2Cr_2O_7 : [O]$

1 : 3

0.68 : $3 \times 0.68 = 2.04 \text{ moles}$ $H_2SO_4 : [O]$

4 : 3

1 : $3/4$ 2.04 : $(3/4) \times 2.04 \Rightarrow 1.53 \text{ moles}$ H_2SO_4 is a limiting reactant.Mole of $[O] = 1.53 \text{ moles}$ Molar mass $\times$ Mole = Mass in gram of $[O]$ Mass of $[O] = 1.53 \times 16$ $\Rightarrow 24.48 g$ $H_2SO_4 : K_2Cr_2O_7$

4 : 1

1 : $1/4$ 2.04 : $1/4 \times 2.04 \Rightarrow 0.51 \text{ moles}$

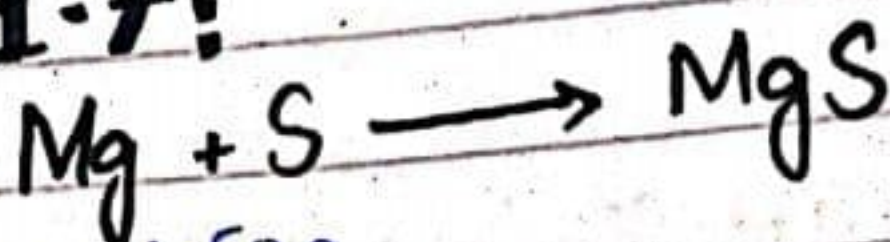
Q 15:

Consumed moles of $K_2Cr_2O_7$ = 0.51 moles.

Excess mole of $K_2Cr_2O_7$ = $0.68 - 0.51 = 0.17$ moles.

Mass of $K_2Cr_2O_7$ = $0.17 \times 294 = 49.98$ g.

Example 1.7:



Mass of Mg = 1.50 g

Mass of S = 1.50 g

Mass of MgS = ?



Moles of Mg = $\frac{1.50}{24} = 0.06$ mol.

Moles of S = $\frac{1.50}{32} = 0.04$ mole.

Mg : MgS
1 : 1

0.06 : 1 $\times$ 0.06 $\Rightarrow$ 0.06 moles.

S : MgS
1 : 1

0.04 : 1 $\times$ 0.04 $\Rightarrow$ 0.04 moles.

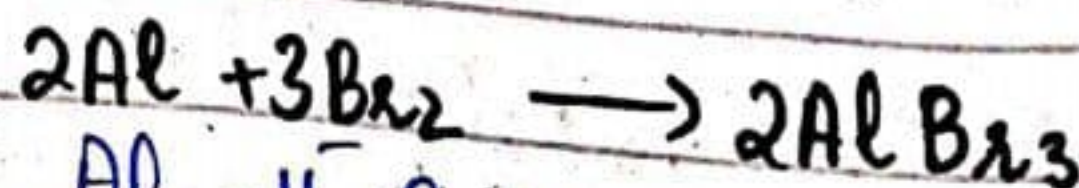
Since S is a limiting reactant.

So moles of MgS = 0.04 moles.

Mass of MgS = 0.04×56
 $= 2.24$ g $\Rightarrow$ 2.576 g

Self-check 1.4:

2)



Mass of Al = 15.8 g

Mass of Br_2 = 55.6 g

limiting reactant = ?

Pg 16:

Mass of AlBr_3 produced = ?

$$\text{Mole of Al} = \frac{15.8}{27} \Rightarrow 0.5851 \text{ moles.}$$

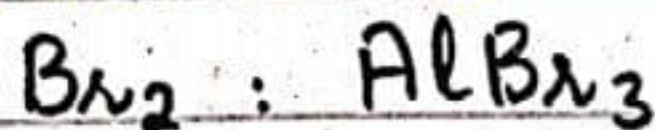
$$\text{Mole of Br}_2 = \frac{55.6}{160} \Rightarrow 0.3475 \text{ moles}$$



$$2 : 2$$

$$1 : 2/2$$

$$0.5851 : 2/2 \times 0.5851 \Rightarrow 0.5851 \text{ moles.}$$



$$3 : 2$$

$$1 : 2/3$$

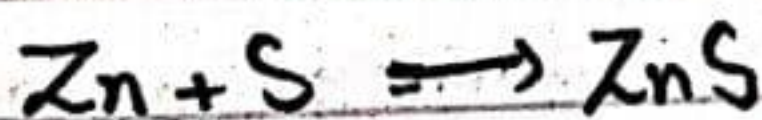
$$1 \times 0.3475 : 2/3 \times 0.3475 \Rightarrow 0.2317 \text{ moles.}$$

Since Br_2 is a limiting reactant.

So Moles of $\text{AlBr}_3 = 0.2317$ moles.

$$\begin{aligned} \text{Mass of AlBr}_3 &= 267 \times 0.2317 \\ &= 61.86 \text{ g} \end{aligned}$$

1)



$$\text{Mass of Zn} = 6 \text{ g}$$

$$\text{Mass of S} = 4 \text{ g}$$

limiting reactant = ?

$$\text{Mass of ZnS} = ?$$

$$\text{Mole of Zn} = \frac{6}{65} = 0.0923 \text{ moles}$$

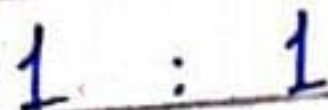
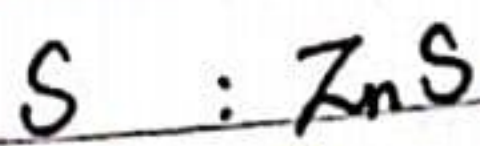
$$\text{Mole of S} = \frac{4}{32} = 0.125 \text{ moles}$$



$$1 : 1$$

$$0.0923 : 1 \times 0.0923 \Rightarrow 0.0923 \text{ moles.}$$

Pg 17:

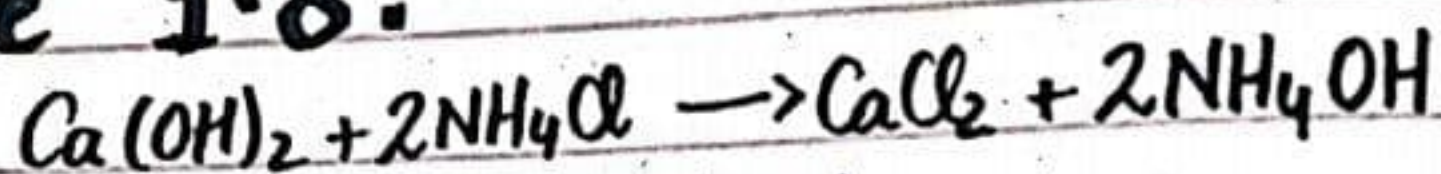


$$0.125 : 1 \times 0.125 \Rightarrow 0.125 \text{ moles.}$$

Since Zn is a limiting reactant.
So moles of ZnS = 0.0923 moles.

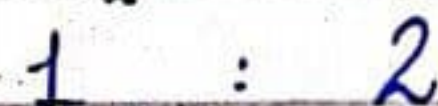
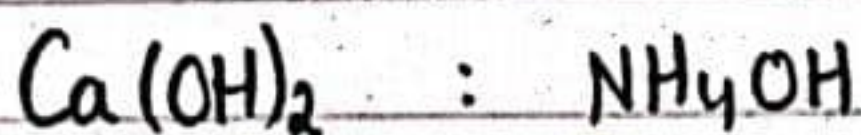
$$\text{Mass of ZnS} = 97 \times 0.0923 \Rightarrow 8.9531 \text{ g}$$

Example 1.8:

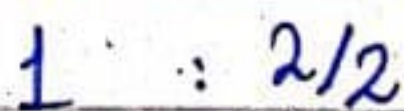
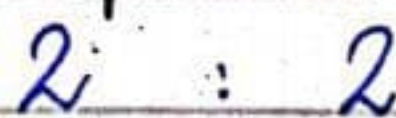
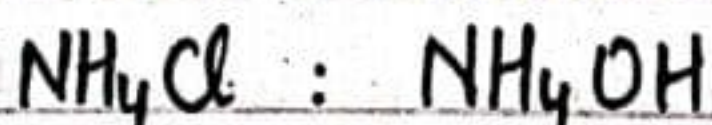


$$\text{Moles of } NH_4Cl = 1.87 \text{ moles}$$

$$\text{Moles of } Ca(OH)_2 = 1.35 \text{ moles}$$

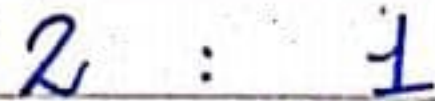
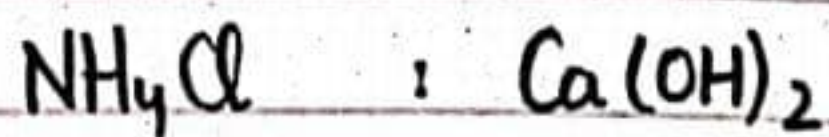


$$1.35 : 2 \times 1.35 \Rightarrow 2.7 \text{ moles}$$



$$1.87 : 2/2 \times 1.87 \Rightarrow 1.87 \text{ moles}$$

Since NH_4Cl is a limiting reactant.



$$1.87 : 1/2 \times 1.87 \Rightarrow 0.935 \text{ moles}$$

$$\text{Consumed moles of } Ca(OH)_2 = 0.935 \text{ moles}$$

$$\text{Excess moles of } Ca(OH)_2 = 1.35 - 0.935$$

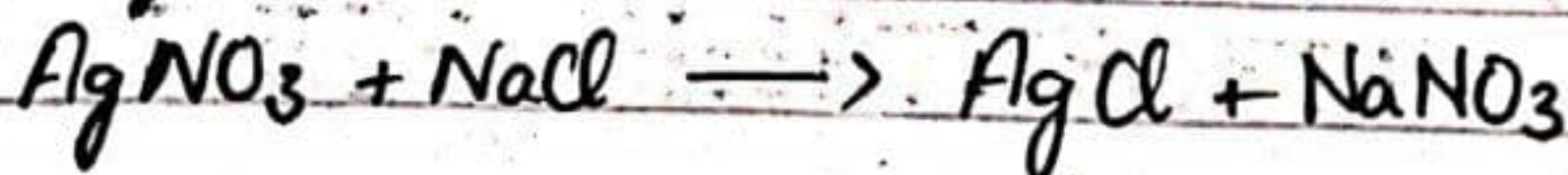
$$= 0.415 \text{ moles}$$

$$\text{Mass of } Ca(OH)_2 = 74 \times 0.415$$

$$\Rightarrow 30.71 \text{ g}$$

Pg 18:

Exercise; Q No 07:



Mass of $\text{AgCl} = ?$

Mass of $\text{AgNO}_3 = 120\text{g}$

Mass of $\text{NaCl} = 52\text{g}$

Moles of $\text{AgNO}_3 = \frac{120\text{g}}{170\text{g mol}^{-1}} \Rightarrow 0.7059 \text{ moles}$

Moles of $\text{NaCl} = \frac{52}{58.5} \Rightarrow 0.89 \text{ moles}$

$\text{AgNO}_3 : \text{AgCl}$
1 : 1

$0.7059 : 1 \times 0.7059 \Rightarrow 0.7059 \text{ moles}$

$\text{NaCl} : \text{AgCl}$
1 : 1

$0.89 : 1 \times 0.89 \Rightarrow 0.89 \text{ moles}$

Since AgNO_3 is a limiting reactant.

So, moles of $\text{AgCl} = 0.7059 \text{ moles}$

Mass of $\text{AgCl} = 0.7059 \times 143.5$
 $= 101.29 \text{ g}$

THEORETICAL YIELD, ACTUAL YIELD AND PERCENTAGE YIELD

Theoretical yield, Numerical yield, Calculated yield or stoichiometric yield:

Amount of product calculated with the help of balanced chemical equation is called theoretical yield.

Actual, experimental or practical yield:

Amount of product actually obtained during the chemical reaction is called actual yield.



Percentage yield:

Efficiency of the reaction can be checked with the help of percentage yield. Higher will be the percentage yield, higher will efficiency will be of the reaction, and vice versa.

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

Q No 02: xi.

Why is actual yield less than theoretical yield?

Due to following reasons, actual yield is lesser than theoretical yield.

- * There may be impurities in reactant.
- * There may be side reaction which produce by product.
- * Reaction may be reversible.
- * There may be mechanical loss during filtration, distillation, crystallization and drying of crystal.
- * Practically inexperienced workers produce low yield due to many short-comings.

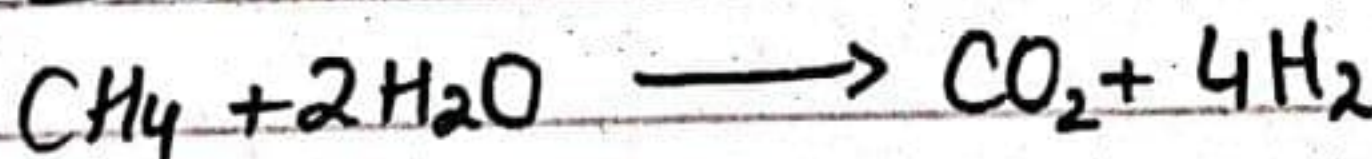


NOTE:

Reaction in which actual yield is almost equal to theoretical yield such reactions are called quantitative reactions.

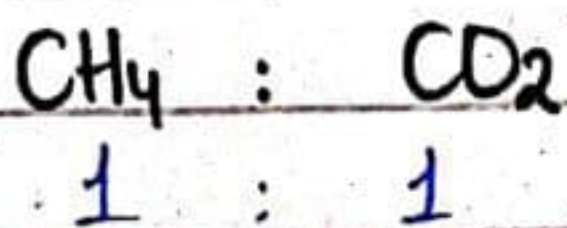
Self-check 1.5

2)



a) Mass of $\text{CH}_4 = 1250\text{g}$
Mass of $\text{CO}_2 = ?$

$$\text{Moles of } \text{CH}_4 = \frac{1250}{16} \Rightarrow 78.125 \text{ moles.}$$



$$78.125 : 1 \times 78.125 \Rightarrow 78.125 \text{ moles.}$$

$$\text{Moles of } \text{CO}_2 = 78.125 \text{ moles.}$$

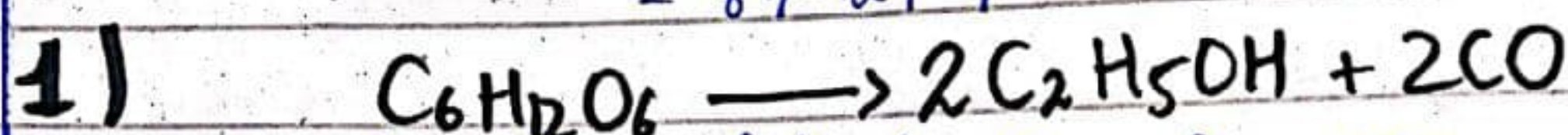
Pg 21:

$$\text{Mass of } \text{CO}_2 = 78.125 \times 44 \Rightarrow 3437.5 \text{ g}$$

$$\text{Theoretical yield of } \text{CO}_2 = 3437.5 \text{ g}$$

$$\text{b. Actual yield of } \text{CO}_2 = 3000 \text{ g}$$

$$\begin{aligned} \% \text{ yield} &= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \\ &= \frac{3000}{3437.5} \times 100 \\ &= 87.27 \% \end{aligned}$$

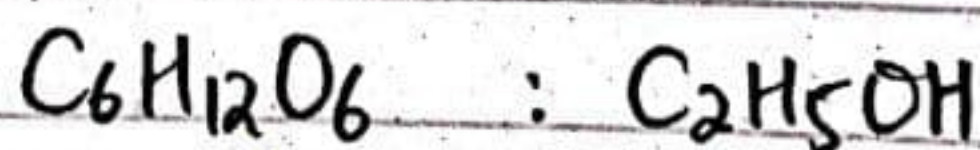


$$\text{a) Actual yield of } \text{C}_6\text{H}_{12}\text{O}_6 = 10 \text{ g}$$

$$\text{C}_2\text{H}_5\text{OH theoretical yield} = ?$$

$$\text{Mass of } \text{C}_6\text{H}_{12}\text{O}_6 = 10 \text{ g}$$

$$\text{Moles of } \text{C}_6\text{H}_{12}\text{O}_6 = \frac{10}{148} \Rightarrow 0.0675 \text{ moles}$$



$$\begin{array}{ccc} 1 & : & 2 \\ 0.0675 & : & 2 \times 0.0675 \Rightarrow 0.135 \text{ moles} \end{array}$$

$$\text{Moles of } \text{C}_2\text{H}_5\text{OH} = 0.135 \text{ moles}$$

$$\text{Mass of } \text{C}_2\text{H}_5\text{OH} = 0.135 \times 46 \Rightarrow 5.11 \text{ g}$$

b)

$$\text{Mass of } \text{C}_6\text{H}_{12}\text{O}_6 = 10 \text{ g}$$

$$\text{actual yield of } \text{C}_2\text{H}_5\text{OH} = 0.664 \text{ g}$$

$$\% \text{ yield of } \text{C}_2\text{H}_5\text{OH} = ?$$

$$\text{theoretical yield of } \text{C}_2\text{H}_5\text{OH} = 5.11 \text{ g}$$

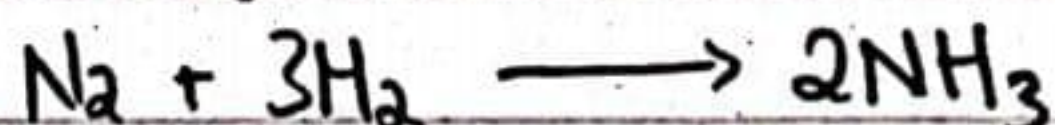
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Pg 22:

$$\% \text{ yield} = \frac{0.664}{5.11} \times 100$$

$$\% \text{ yield of } C_2H_5OH = 12.99\%$$

Exercise Q No 04. b



Mass of $H_2 = 42g$
actual yield of $NH_3 = 120.2g$
 $\% \text{ yield of } NH_3 = ?$

$$\text{Moles of } H_2 = \frac{42}{2} \Rightarrow 21 \text{ moles}$$



$$3 : 2$$

$$1 : \frac{2}{3}$$

$$21 : \frac{2}{3} \times 21 \Rightarrow 14 \text{ moles}$$



$$\text{Moles of } NH_3 = 14 \text{ moles}$$

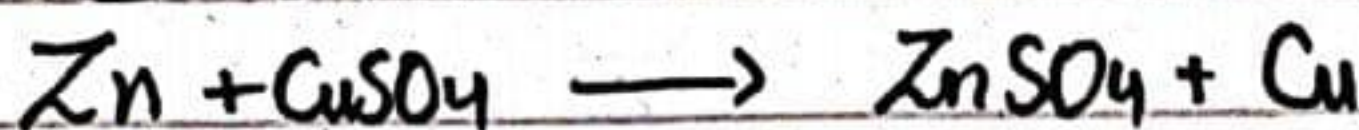
$$\text{Mass of } NH_3 = 238g$$

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

$$\% \text{ yield of } NH_3 = \frac{120.2}{238} \times 100$$

$$= 50.5\%$$

Example 1.9:



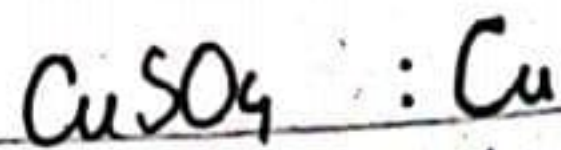
$$\text{Mass of } CuSO_4 = 1.274g$$

$$\text{actual yield of } Cu = 0.392g$$

$$\% \text{ yield of } Cu = ?$$

$$\text{Moles of } CuSO_4 = \frac{1.274}{159.5} = 7.987 \times 10^{-3} \text{ moles}$$

Pg 23:



$$1 : 1$$

$$7.987 \times 10^{-3} : 1 \times 7.987 \times 10^{-3} \Rightarrow 7.987 \times 10^{-3}$$

$$\text{Moles of Cu} = 7.987 \times 10^{-3} \text{ moles}$$

$$\text{Mass of Cu} = 7.987 \times 10^{-3} \times 63.5$$

$$= 0.507 \text{ g}$$

$$\text{Theoretical yield of Cu} = 0.507 \text{ g}$$

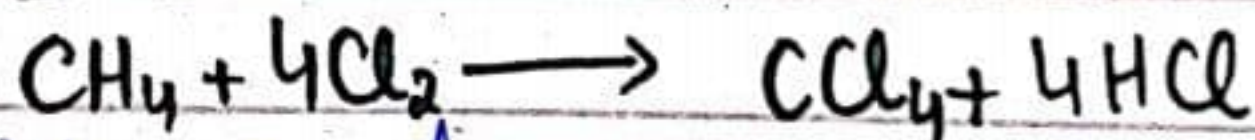
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$\% \text{ yield} = \frac{0.392}{0.507} \times 100$$

$$\% \text{ yield} = 77.32\%$$



Example 1.10



$$\text{Moles of CH}_4 = 2 \text{ moles}$$

$$\text{Mass of CCl}_4 = 177 \text{ g}$$



$$1 : 1$$

$$2 : 1 \times 2 \Rightarrow 2 \text{ moles}$$

$$\text{Moles of CCl}_4 = 2 \text{ moles}$$

$$\text{Mass of CCl}_4 = \text{Moles} \times \text{Molar mass}$$

$$= 2 \times 154$$

$$= 308 \text{ g}$$

$$\text{Theoretical yield of CCl}_4 = 308 \text{ g}$$

$$\text{Actual yield of CCl}_4 = 177 \text{ g}$$

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

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

$$\% \text{ yield of } \text{CCl}_4 = 57.46\%$$

CHAPTER # 01

EXERCISE

Q No 02: Short Q/As

1-

NaCl is ionic compound. It can't exist as individual and independent molecule. Therefore, mass of NaCl is called formula mass not molecular mass.

ii • Concept of limiting reactant is not applicable in reversible reaction. Because in reversible reaction, reactant are not completely converted into product. However, limiting reactant is a reactant which is completely consumed in a chemical reaction.

iii •

Mass of $\text{H}_2\text{O} = 9\text{g}$

No: of covalent bonds = ?

$$\text{No: of molecules} = \frac{9}{18} \times 6.02 \times 10^{23}$$

$$= 3.01 \times 10^{23} \text{ molecules}$$

1 molecule of H_2O has bonds = 2

$$3.01 \times 10^{23} \times 2 = 2 \times 3.01 \times 10^{23}$$

$$= 6.02 \times 10^{23} \text{ No. of covalent bonds.}$$

Mass of ice = 12g

No. of molecules =

$$\begin{aligned} \text{No. of molecules} &= \frac{12}{18} \times 6.02 \times 10^{23} \\ &= 4.01 \times 10^{23} \text{ molecules} \end{aligned}$$

According to Avogadro's No,

1 mole of $H_2 = 6.02 \times 10^{23}$ molecules.

1 mole of Fe = 6.02×10^{23} atoms

1 molecule of H_2 has atoms = 2

$$= 2 \times 6.02 \times 10^{23}$$

$$= 1.204 \times 10^{24} \text{ atoms.}$$

viii.

1 mole of CO_2

1 mole of $N_2 = 28g$ $\therefore$ 1 mole of $CO_2 = 44g = 6.02 \times 10^{23}$ molecules

1 mole of $H_2 = 2.016$

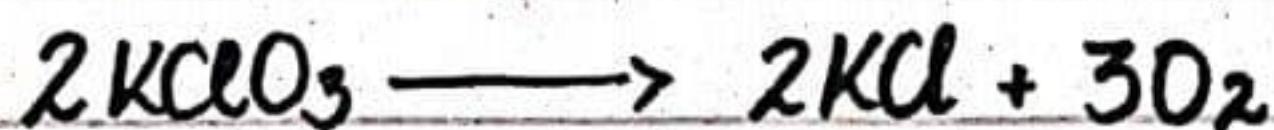
1 mole of $H_2 = 2.016 \text{ g} =$

1 mole of different masses but compound have no. of molecules. This is because same individual molecule of different masses. different compound

VIII.

$$\begin{array}{ll}
 \text{Na} = 23 \text{ g} & ; \quad \text{U} = 238 \text{ g} \\
 \text{No. of atoms} = ? & ; \quad \text{No. of atoms} \\
 \text{No. of atoms} = \frac{23}{23} \times 6.02 \times 10^{23} & ; \quad \text{No. of atoms} = \frac{238}{238} \times 6.02 \times 10^{23} \\
 = 6.02 \times 10^{23} \text{ atoms} & ; \quad = 6.02 \times 10^{23} \text{ atoms}
 \end{array}$$

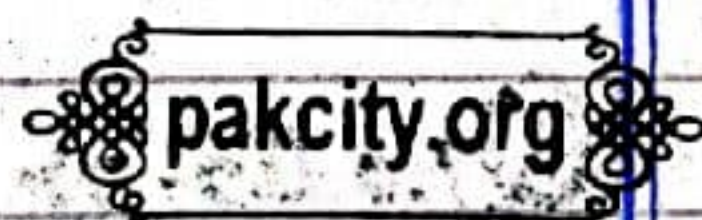
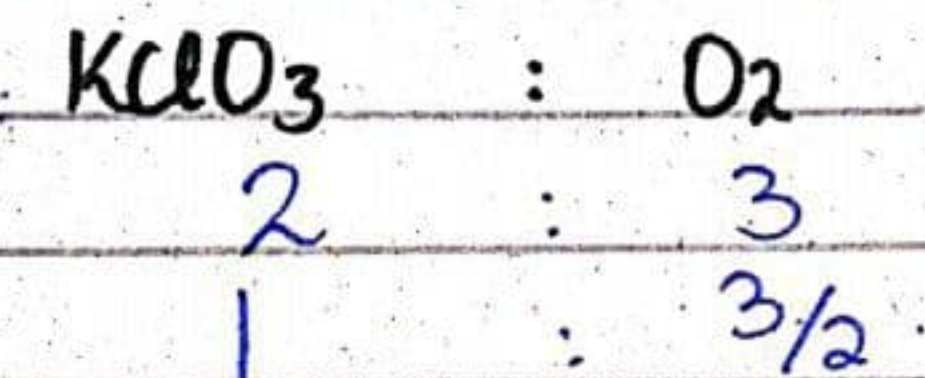
ix.



$\text{KClO}_3 = 5 \text{ g}$
weight of oxygen gas produce = ?

$$\text{Molar mass of } \text{KClO}_3 = 39 + 35.5 + 16 \times 3 = 122.5 \text{ g mol}^{-1}$$

$$\text{Moles of } \text{KClO}_3 = \frac{5}{122.5} \Rightarrow 0.0408 \text{ moles}$$



$$0.0408 : 3/2 \times 0.0408 \Rightarrow 0.0612 \text{ moles}$$

$$\text{Moles of } \text{O}_2 = 0.0612 \text{ moles}$$

$$\begin{aligned}
 \text{Mass of } \text{O}_2 \text{ produce} &= 0.0612 \times 32 \\
 &= 1.9584 \text{ g}
 \end{aligned}$$

x.

$$1 \text{ mole of } \text{CO}_2 = 44 \text{ g} = 22.414 \text{ dm}^3$$

$$1 \text{ mole of } \text{O}_2 = 32 \text{ g} = 22.414 \text{ dm}^3$$

$$1 \text{ mole of } \text{H}_2 = 2.016 \text{ g} = 22.414 \text{ dm}^3$$

Volume occupied by 1 mole of a gas at **S.T.P** is 22.414 dm^3 .
This volume is called **molar volume**.

Pg 27:

Mass of 1 mole of a gas is called molar mass.
Molar mass of a gas is equal to molar volume.

xii.

1 mole of any gas at S.T.P. occupied a volume of 22.414 dm^3 .
This volume is called molar volume. It contains 6.02×10^{23} representative particles.

e.g.

1 mole of $\text{CO}_2 = 44\text{g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ particles.

1 mole of $\text{O}_2 = 32\text{g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ particles.

1 mole of $\text{H}_2 = 2.016\text{g} = 22.414 \text{ dm}^3 = 6.02 \times 10^{23}$ particles.

Q No 03-b

pakcity.org

density of $\text{CHBr}_3 = 2.89 \text{ g cm}^{-3}$

No. of molecules of $\text{CHBr}_3 = 4.8 \times 10^{24}$ molecules

vol. of $\text{CHBr}_3 = ?$

Solution:

Molar mass of $\text{CHBr}_3 = 12 + 1 + 80 \times 3 = 253 \text{ g mol}^{-1}$
No. of molecules = $\frac{\text{Mass in gram}}{\text{Molar mass}} \times 6.02 \times 10^{23}$

Mass in gram = $\frac{\text{No. of molecules} \times \text{Molar mass}}{6.02 \times 10^{23}}$

$$= \frac{4.8 \times 10^{24} \times 253}{6.02 \times 10^{23}}$$

$$= 2017.27 \text{ g}$$

$$d = m/V$$

$$V = m/d$$

$$V = 2017.27 / 2.89 \Rightarrow 698.02 \text{ cm}^3$$



$$\text{Volume of Al} = 2.50 \text{ cm}^3$$

$$\text{Density of Al} = 2.70 \text{ g cm}^{-3}$$

$$\text{weight of H}_2 = ?$$

Solution:

$$\text{mass of Al} = d \times v$$

$$= 2.70 \times 2.50$$

$$= 6.75 \text{ g}$$

$$\text{Mole of Al} = \frac{6.75}{27}$$

$$\Rightarrow 0.25 \text{ moles.}$$

$$\text{Al} : \text{H}_2$$

$$2 : 3$$

$$1 : 3/2$$

$$0.25 : 3/2 \times 0.25 \Rightarrow 0.375 \text{ moles}$$

$$\text{Moles of H}_2 = 0.375 \text{ moles}$$

$$\text{Mass of H}_2 = 0.375 \times 2.016$$

$$\text{Mass in gram of H}_2 = 0.756 \text{ g}$$