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Chemistry Practice questions

GRADE 9

Mole and Stoichiometry

100% Useful For Entrance examination

Hard Work Towards Exellence

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Mole Concept and Stoichiometry

Q. 1. Calculate the volume occupied by 3.4 g of ammonia at S.T.P.

Sol.

$$\text{Gram molecules of } NH_3 = (1 \times 14) + (3 \times 1) = 14 + 3 = 17 \text{ g}$$

$$\text{Mass of one mole of ammonia} = 17 \text{ g}$$

$$\text{Molar volume} = 22.4 \text{ litre}$$

Thus, 17 grams of NH_3 occupied 22.4 litre

$$\therefore \quad 3.4 \text{ gram} = \frac{3.4 \times 22.4}{17} = 4.48 \text{ litre}$$

Volume occupied by 3.4 g of NH_3 at S.T.P. is 4.48 litre.

Q. 2. A cylinder contains 68 g of ammonia gas at S.T.P.

(i) What is the volume occupied by this gas ?

(ii) How many moles of ammonia are present in the cylinder ?

(iii) How many molecules of ammonia are present in the cylinder ?
[N–14, H–1]

Sol. Molecular weight of $NH_3 = (14 + 3 \times 1) = 17$

$$\text{Number of moles of } NH_3 \text{ in 68 g} = \frac{\text{Wt.}}{\text{Mol. wt.}} = \frac{68}{17} = 4 \text{ moles}$$

At S.T.P. one mole of NH_3 occupies 22.4 lit.

(i) $\therefore$ Volume occupied by 4 moles of the gas

$$= 4 \times 22.4 \text{ lit.} = 89.6 \text{ lit. at S.T.P.}$$

(ii) 4 moles of ammonia are present in the cylinder.

(iii) $$\text{Number of molecules} = \text{No. of moles} \times N_A = 4 \times 6.023 \times 10^{23} = 24.092 \times 10^{23}.$$

Q. 3. What weight of zinc is needed to produce 100 ml of dry hydrogen at S.T.P. from dilute sulphuric acid solution ?

Sol. The chemical equation representing the reaction is

$$\underset{(65\text{ g})}{Zn} + H_2SO_4 \longrightarrow ZnSO_4 + \underset{(22400\text{ ml})}{H_2}$$

$\because$ 22400 ml of hydrogen is liberated at S.T.P. from 65 g of zinc

$\therefore$ 100 ml of hydrogen is liberated at S.T.P. from

$$= \frac{65}{22400} \times 100 = 0.29 \text{ g of zinc.}$$

Q. 4. What weight and volume of oxygen at S.T.P. will be given, when 18 g of water is electrolysed ?

Sol. The decomposition of water can be represented as :

$$\underset{(36\text{ g})}{2H_2O} \longrightarrow \underset{(4\text{ g})}{2H_2} + \underset{(32\text{ g})}{O_2}$$

$\because$ 36 g of water yields = 32 g of oxygen.

$$\therefore \quad \text{18 g of water will yield} = \frac{32}{36} \times 18 = 16 \text{ g oxygen.}$$

Again, 32 g oxygen occupies 22.4 litre volume at S.T.P.

$$\therefore \quad \text{16 g oxygen will occupy} = \frac{22.4}{32} \times 16 = 11.2 \text{ litre.}$$

Q. 5. Calculate the volume of propane burnt for every 200 cm^3 of oxygen used in the reaction.

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

Sol. From the above reaction it is clear that for every 5 volumes of oxygen, 1 volume of propane is burnt.

Hence, volume of propane burnt for every 200 cm^3 of oxygen $= \frac{1}{5} \times 200$

$= 40\ cm^3$.

Q. 6. Calculate the volume of methane gas that must be burnt completely to produce 100 lit of CO_2 at S.T.P.

Sol. Let x be the volume of methane to be burnt to produce 100 lit CO_2

$$\underset{\substack{1\text{ mol}\\22.4\text{ lit}}}{\overset{x}{CH_4(g)}} + 2O_2(g) \rightarrow \underset{\substack{1\text{ mol}\\22.4\text{ lit}}}{\overset{100\text{ lit}}{CO_2(g)}} + 2H_2O(l)$$

$$\text{Ratio proportion} = \frac{x\ CH_4}{22.4\text{ lit } CH_4} = \frac{100\text{ lit } CO_2}{22.4\text{ lit } CO_2}$$

$$\therefore \quad x = \frac{100\text{ lit}}{22.4\text{ lit}} \times 22.4\text{ l} = 100\text{ lit}$$

Volume of methane to be burnt is 100 lit.

Q. 7. Calculate the volume of carbon dioxide formed when 8 g of methane gas burns completely as represented by the equation :

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Sol. Let x be the volume of CO_2 produced when 8 g CH_4 is burnt.

$$\underset{\substack{1\text{ mole}\\16\text{ g}}}{CH_4(g)} + 2O_2(g) \rightarrow \underset{\substack{1\text{ mole}\\44\text{ g} = 22.4\text{ litre.}}}{CO_2(g)} + 2H_2O(l)$$

$$\text{Ratio proportion} = \frac{8\text{ g } CH_4}{16\text{ g } CH_4} = \frac{x\text{ lit } CO_2}{22.4\text{ lit } CO_2}$$

$x = \frac{8\text{ g}}{16\text{ g}} \times 22.4\text{ lit} = 11.2$ lit is the volume of CO_2 produced at S.T.P.

Q. 8. What volume of oxygen would be required for the complete combustion of 100 litre of ethane according to the following equation ?

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

Sol.

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

2 volume of ethane = 7 volume of oxygen

2 volume of ethane = 100 litre

$\therefore$ 1 volume of ethane = 50 litre

7 volume of O_2 = $50 \times 7 = 350$ litre

$\therefore$ 100 litre of ethane require = 350 litre of oxygen.

Q. 9. Ammonia may be oxidised to nitrogen monoxide in the presence of a catalyst according to the following equation.

$$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$$

If 27 litre of reactants are consumed, what volume of nitrogen monoxide is produced at the same temperature and pressure ?

Sol. According to equation,

$$4NH_3 + \underset{5\text{ vol.}}{5O_2} \longrightarrow \underset{4\text{ vol.}}{4NO} + 6H_2O$$

$\because$ 5 volumes of reactants are consumed to give 4 volumes of nitrogen monoxide.

$\therefore$ 27 litre of reactants are consumed to give

$$= \frac{27 \times 4}{5} = \frac{108}{5}$$

$= 21.6$ litre.

Q. 10. Hydrogen sulphide gas burns in oxygen to yield 12.8 g of sulphur dioxide according to the equation :

$$2H_2S + 3O_2 \longrightarrow 2H_2O + 2SO_2$$

Calculate the volume of hydrogen sulphide at S.T.P. Also calculate the volume of oxygen required at S.T.P., which will complete the combustion of hydrogen sulphide.
[S = 32; O = 16; H = 1]

Sol. According to equation,

$$\underset{\substack{2 \times 22.4 \\ = 44.8\ cm^3 \text{ at S.T.P.}}}{2H_2S} + \underset{\substack{3 \times 22.4 \\ = 67.2\ dm^3 \text{ at S.T.P.}}}{3O_2} \longrightarrow 2H_2O + \underset{\substack{2\,[32 + 32] \\ = 128\ g}}{2SO_2}$$

$$128 \text{ g of } SO_2 \text{ is produced from } H_2S = 44.8\ dm^3$$

$$\therefore \quad 12.8 \text{ g of } SO_2 \text{ is produced from } H_2S = \frac{44.8 \times 12.8}{128}$$

$$= 4.48\ dm^3$$

$$128 \text{ g of } SO_2 \text{ is produced from } O_2 = 67.2\ dm^3$$

$$\therefore \quad 12.8 \text{ g of } SO_2 \text{ is produced from } O_2 = \frac{67.2 \times 12.8}{128}$$

$$= 6.72\ dm^3$$

Q. 11. Ammonia burns in oxygen and the combustion, in the presence of a catalyst; may be represented as :

$$2NH_3 + 2\frac{1}{2}O_2 \longrightarrow 2NO + 3H_2O$$

(i) What mass of steam is produced when 1.5 g of nitrogen monoxide is formed ?
(ii) What volume of oxygen at S.T.P. is required to form 10 moles of products ?

Sol. (i)

$$2NH_3 + 2\frac{1}{2}O_2 \longrightarrow 2NO + 3H_2O$$

$$\text{Wt. of } 2NO = 2[14 + 16]$$

$$= 60 \text{ g}$$

and

$$\text{Wt. of } 3H_2O = 3[2\,(1) + 16]$$

$$= 54 \text{ g}$$

$$\text{When 60 g of NO is formed, mass of steam produced} = 54 \text{ g}$$

$$\therefore \quad \text{For 1.5 g of NO mass of steam produced} = \frac{54 \times 1.5}{60} = 1.35 \text{ g}$$

(ii)

$$\text{For 5 moles of products, oxygen required} = 2.5 \times 22.4 \text{ litre}$$

$$\therefore \quad \text{For 10 moles of products, oxygen required} = \frac{2.5 \times 22.4 \times 10}{5}$$

$$= 112 \text{ litre.}$$

Q. 12. $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$
(H = 1, N = 14, O = 16, Cu = 64)
Calculate from the equation :
(i) The mass of copper needed to react with 63 g of HNO_3.
(ii) The volume of nitric oxide that can be collected at S.T.P. (the gram molecular volume of a gas at STP = 22.4 litre).

Sol. (i)

$$\underset{\substack{3 \times 64 \\ = 192\ g}}{3Cu} + \underset{\substack{8(1 \times 14 + 48) \\ = 504\ g}}{8HNO_3} \longrightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$$

$\because$ 504 g of HNO_3 reacts with 192 g of Cu.

$$\therefore \quad 1 \text{ g of } HNO_3 \text{ reacts with} = \frac{192}{504} \text{ g of Cu.}$$

$$63 \text{ g of } HNO_3 \text{ reacts with} = \frac{192}{504} \times 63 = 24 \text{ g copper.}$$

(ii) 192 g of Cu with HNO_3 gives out $= 2 \times 22.4$ litre of NO at S.T.P.

$\therefore$ 24 g of Cu with $HNO_3 = \dfrac{2 \times 22.4}{192} \times 24$

$= 5.6$ litre of NO.

Q. 13. From the equation :

(i) $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$

Calculate :

(a) The volume of nitrogen at S.T.P. evolved when 63 g of ammonium dichromate is heated.

(b) The mass of chromium(II) oxide (Cr_2O_3) formed at the same time.

(N = 14; H = 1; Cr = 52; O = 16)

(ii) The reaction $4N_2O + CH_4 \rightarrow CO_2 + 2H_2O + 4N_2$ takes place in the gaseous state. If all volumes are measured at the same temperature and pressure, calculate the volume of dinitrogen oxide (N_2O) required to give 150 cm^3 of steam.

(N = 14; O = 16; C = 12; H = 1)

Sol. (i) The reaction is represented as :

(a)
$$\underset{252\text{ g}}{(NH_4)_2Cr_2O_7} \longrightarrow Cr_2O_3 + 4H_2O + \underset{22.4\text{ litre at S.T.P.}}{N_2}$$

$\therefore$ 63 g $= \dfrac{22.4 \times 63}{252}$

$= 5.6$ litre at S.T.P.

(b) 252 g $(NH_4)_2Cr_2O_7$ gives 152 g of Cr_2O_3

$\therefore$ 63 g $(NH_4)_2Cr_2O_7$ gives $= \dfrac{152 \times 63}{254} = 37.70$ g

(ii)
$$\underset{\substack{4\text{ vol.}\\300\text{ cm}^3}}{4N_2O} + CH_4 \longrightarrow CO_2 + \underset{\substack{2\text{ vol.}\\150\text{ cm}^3}}{2H_2O} + 4N_2$$

300 cm^3 of N_2O will be required to yield 150 cm^3 of steam.

Q. 14. (i) Rewrite the following equation in terms of moles of reactants and products :

$$2NH_4Cl + Ca(OH)_2 \longrightarrow CaCl_2 + 2H_2O + 2NH_3$$

(ii) Use the given equation to calculate the weight of calcium hydroxide needed to decompose 2.14 g NH_4Cl.

(iii) Find out the volume of ammonia at S.T.P. obtained from 4.28 g of NH_4Cl.

Sol. (i) In term of moles, the equation is as follows :

$$\underset{\substack{2\text{ moles}\\2(14+4+35.5)\\=107\text{ g}}}{2NH_4Cl} + \underset{\substack{1\text{ moles}\\40+32+2\\=74\text{ g}}}{Ca(OH)_2} \longrightarrow \underset{\substack{1\text{ mole}\\40+71\\=111\text{ g}}}{CaCl_2} + \underset{\substack{2\text{ moles}\\2(2+16)\\=36\text{ g}}}{2H_2O} + \underset{\substack{2\text{ moles}\\2(14+3)\\=34\text{ g}}}{2NH_3}$$

So, 2 moles of ammonium chloride combines with 1 mole of calcium hydroxide to form 1 mole of calcium chloride, 2 moles of water and 2 moles of ammonia gas.

(ii) 107 g of NH_4Cl needed to react with 74 g of $Ca(OH)_2$

$\therefore$ 2.14 g of NH_4Cl needed $= \dfrac{74 \times 2.14}{107}$ g of $Ca(OH)_2$

$= 1.48$ gm.

(iii) 107 g of NH_4Cl will liberate with slaked lime

$= 2 \times 22.4$ litre of NH_3.

$\therefore$ 4.28 g of NH_4Cl liberates with slaked lime

$= \dfrac{2 \times 22.4 \times 4.28}{107}$

$= 1.792$ litre.

Q. 15. Solve the following :

(i) What volume of oxygen is required to burn completely 90 dm^3 of butane under similar conditions of temperature and pressure ?

$$2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$$

(ii) The vapour density of a gas is 8. What would be the volume occupied by 24.0 g of the gas at STP ?

(iii) A vessel contains X number of molecules of hydrogen gas at a certain temperature and pressure. How many molecules of nitrogen gas would be present in the same vessel under the same conditions of temperature and pressure ?

Sol. (i)

$$\underset{2\text{ vol.}}{2C_4H_{10}} + \underset{13\text{ vol.}}{13O_2} \longrightarrow 8CO_2 + 10H_2O$$

$\because$ 2 vol. of butane require 13 vol. of oxygen (according to Gay-Lussac's law)

$\therefore$ 90 dm^3 of butane require $= \dfrac{13 \times 90}{2} = 585\ dm^3$

Ans. 585 dm^3 of oxygen is required.

(ii) Given : Vapour Density (VD) = 8

$\therefore$ Mol. wt $= 2 \times VD = 2 \times 8 = 16$

No. of moles in 24.0 g of gas $= \dfrac{\text{Wt.}}{\text{mol. wt.}} = \dfrac{24}{16}$

$= 1.5$ moles

At S.T.P. 1 mole of a gas occupies 22.4 l.

$\therefore$ 1.5 moles (or 24.0 g) of the gas will occupy $= \dfrac{22.4 \times 1.5}{1} = 33.6$ l.

(iii) 'X' number of molecules.

Q. 16. O_2 is evolved by heating $KClO_3$ using MnO_2 as a catalyst

$$2KClO_3 \xrightarrow{MnO_2} 2KCl + 3O_2$$

(i) Calculate the mass of $KClO_3$ required to produce 6.72 litre of O_2 at S.T.P. [atomic masses of K = 39, Cl = 35.5, O = 16]

(ii) Calculate the number of moles of oxygen present in the above volume and also the number of molecules.

(iii) Calculate the volume occupied by 0.01 mole of CO_2 at S.T.P.

Sol. (i) Molecular wt. of $KClO_3$ $= 39 + 35.5 + 16 \times 3$

$= 122.5$

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

2 moles → 3 moles

2×122.5 g. → 3×22.4 lit at S.T.P.

$\because$ 3×22.4 lit of oxygen is produced from 2×122.5 g of $KClO_3$

$\therefore$ 6.72 lit of oxygen is produced from $\dfrac{2 \times 122.5 \times 6.72}{3 \times 22.4} = 24.5$ g

Ans. 24.5 g of $KClO_3$ is required to produce 6.72 lit of O_2 at S.T.P.

(ii) At S.T.P. 22.4 lit of a gas = 1 mole

$\therefore$ 6.72 lit $= \dfrac{1 \times 6.72}{22.4} = 0.3$ moles

One mole contains $= 6 \times 10^{23}$ molecules

$\therefore$ 0.3 mole contains $= 6 \times 10^{23} \times 0.3$

$= 1.8 \times 10^{23}$ molecules.

Ans. 6.72 l of oxygen contains 0.3 moles and hence 1.8×10^{23} molecules.

(iii) At STP one mole of CO_2 occupies 22.4 lit

$\therefore$ 0.01 mole of CO_2 occupies $= 22.4 \times 0.01$

$= 0.224$ lit

Ans. 0.01 mole of CO_2 will occupy 0.224 l at S.T.P.

Q. 17. 4.5 moles of calcium carbonate are reacted with dilute hydrochloric acid.

(i) Write the equation for the reaction.

(ii) What is the mass of 4.5 moles of calcium carbonate ? (Relative molecular mass of calcium carbonate is 100).

(iii) What is the volume of carbon dioxide liberated at S.T.P. ?

(iv) What mass of calcium chloride is formed ? (Relative molecular mass of calcium chloride is 111).

(v) How many moles of HCl are used in this reaction ?

Sol. (i)
$$CaCO_3 + 2HCl \longrightarrow \underset{\text{calcium chloride}}{CaCl_2} + H_2O + CO_2$$

(ii) 4.5 moles of $CaCO_3$ $= 4.5 \times 100 = 450$ g

(iii) $\because$ 1 mole $CaCO_3$ liberates 1 mole of CO_2

$\therefore$ 4.5 moles of $CaCO_3$ liberates 4.5 moles of CO_2

$\therefore$ Volume of CO_2 liberated at STP $= 22.4 \times 4.5 = 100.8$ lit

(iv) $\because$ 1 mole of $CaCO_3$ gives 1 mole of $CaCl_2$.

$\therefore$ 4.5 moles of $CaCO_3$ gives 4.5 moles of $CaCl_2 = 4.5 \times 111 = 499.5$ g

(v) $\because$ 1 mole of $CaCO_3$ requires 2 moles of HCl

$\therefore$ 4.5 moles of $CaCO_3$ requires $= 2 \times 4.5 = 9$ moles of HCl

Q. 18. (i) Oxygen oxidises ethyne to carbon dioxide and water as shown by the equation :

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

What volume of ethyne gas at S.T.P. is required to produce 8.4 dm^3 of carbon dioxide at S.T.P. ?

[H = 1, C = 12, O = 16]

(ii) A compound made up of two element X and Y has an empirical formula X_2Y. If the atomic weight of X is 10 and that of Y is 5 and the compound has a vapour density 25, find its molecular formula.

Sol. (i) Given :

$$\underset{\substack{\text{2vol.}\\ \text{1 vol.}}}{2C_2H_2} + 5O_2 \longrightarrow \underset{\substack{\text{4 vol.}\\ \text{2 vol.}}}{4CO_2} + 2H_2O$$

$\therefore$

According to Gay-Lussac's law :

2 volume of CO_2 is produced from 1 vol. of C_2H_2

$\therefore$ 8.4 dm^3 of CO_2 at S.T.P. produced from

$$= \frac{1 \times 8.4}{2}$$

$$= 4.2 \text{ dm}^3 \text{ of } C_2H_2$$

Ans. At S.T.P. 4.2 dm^3 of ethyne is required.

(ii) Molecular formula $=$ (Empirical formula)$_n$

$$n = \frac{\text{Molecular formula weight}}{\text{Empirical formula weight}}$$

$$= \frac{2 \times \text{V.D.}}{(2 \times 10 + 5)} \quad \text{(V.D. = Vapour Density)}$$

$$= \frac{2 \times 25}{25} = 2$$

$\therefore$ Molecular formula $= (X_2 Y)_2$

$= X_4Y_2$

Q. 19. (i) Calculate the volume of 320 g of SO_2 at S.T.P. (Atomic mass : S = 32 and O = 16).

(ii) Calculate the volume of oxygen required for the complete combustion of 8.8 g of propane (C_3H_8). (Atomic mass : C = 14, O = 16, H = 1, Molar volume = 22.4 dm^3 at S.T.P.)

Sol. (i) Gram molar mass of SO_2 = 32 + (2 × 16) = 64 g.

$\therefore$ No. of moles in 320 g of SO_2 $= \frac{320}{64} = 5$ moles.

At S.T.P. 1 mole of SO_2 occupies 22.4 dm^3.

$\therefore$ 5 moles of SO_2 will occupy 5 × 22.4 = 112 dm^3

(ii) Chemical equation for the complete combustion of propane is :

$$\underset{\substack{1\text{ mole}\\1\text{ mole}\\(12\times3)+(1\times8)=44\text{ g}}}{C_3H_8} + \underset{\substack{5\text{ mole}\\5\times22.4\text{ lit at S.T.P.}}}{5O_2} \rightarrow 3CO_2 + 4H_2O$$

$$44\text{g} \equiv 5 \times 22.4 \text{ lit S.T.P.}$$

$\therefore$ 8.8 g of propane would require $\frac{5 \times 22.4 \times 8.8}{44} =$ 22.4 lit of oxygen at S.T.P.

Q. 20. (i) How many grams of sulphuric acid will be required to completely neutralise 16.0 g of caustic soda ?

(ii) What volume of H_2 be evolved at S.T.P. ?

Sol. (i) The complete reaction between H_2SO_4 and NaOH is represented by :

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2$$

$\therefore$ 2 [23 + 16 + 1] g of NaOH react with [2 × 1 + 32 + 4 × 16] g of H_2SO_4

or 80 g of NaOH react with 98 g of H_2SO_4

$\therefore$ 16 g of NaOH react with $\frac{98 \times 16}{80} = 19.6$ g of H_2SO_4.

(ii) 80 g of NaOH formed at S.T.P. = 2 × 22.4 litre of H_2

$\therefore$ 16 g of NaOH formed at S.T.P. $= \frac{2 \times 22.4 \times 16}{80}$

= 8.96 litre of H_2.

Q. 21. (i) How many molecules are present in (the numerical value of Avogadro's number can be used as 6×10^{23}) ?

(a) 2.2 grams of carbon dioxide

(b) 16 grams of sulphur dioxide

(c) 2 grams of oxygen

(d) 58.5 grams of sodium chloride

(C = 12, O = 16, Na = 23, S = 32, Cl = 35.5)

(ii) How many moles are present in ?

(a) 100 g of calcium carbonate, (b) 2.3 g of sodium

(c) 80 g of sodium hydroxide.

Sol. (i) (a) 44 g of CO_2 contain 6×10^{23} molecules at S.T.P.

$\therefore$ 2.2 g of CO_2 $= \frac{6 \times 10^{23}}{44} \times 2.2$

$= 3 \times 10^{23}$ molecules

(b) 64 g of SO_2 contain 6×10^{23} molecules at S.T.P.

$\therefore$ 16 g of S.T.P. $= \frac{6 \times 10^{23}}{64} \times 16$

$= 1.5 \times 10^{23}$ molecules

(c) 32 g of O_2 contain 6×10^{23} molecules at S.T.P.

$\therefore$ 2 g of O_2 $= \frac{6 \times 10^{23}}{32} \times 2$

$= 3.75 \times 10^{23}$ molecules

(d) 58.5 g of NaCl contain 6×10^{23} molecules at S.T.P.

$$\therefore \quad 58.5 \text{ g of NaCl} = \frac{6 \times 10^{23}}{58.5} \times 8.5$$

$$= 0.871 \times 10^{23} \text{ molecules}$$

(ii) (a)
$$\text{Number of moles} = \frac{\text{Weight of the substance (W)}}{\text{Molecular weight (M)}}$$

$$\text{Molecular weight of } CaCO_3 = 40 + 12 + 48 = 100$$

$$\text{No. of moles} = \frac{100}{100} = 1 \text{ mole}$$

(b)
$$\text{No. of moles of sodium} = \frac{2.3}{23} = \text{molecular wt. of sodium} = 23$$

$$= 0.1 \text{ mole}$$

(c)
$$\text{No. of moles of NaOH} = \frac{80}{40} = 2$$

$$\text{Molecular wt. of NaOH} = 23 + 16 + 1 = 40 = 2 \text{ moles}$$

Q. 22. The gases chlorine, nitrogen, ammonia and sulphur dioxide are collected under the same conditions of temperature and pressure. Copy the following table which gives the volumes of gases collected and the number of molecules (X) in 20 litre of nitrogen. You have to complete the table giving the number of molecules in the other gases in terms of X.

Gas	Volume (litres)	Number of Molecules
Chlorine	10	
Nitrogen	20	X
Ammonia	20	
Sulphur dioxide	5	

Sol. In accordance with Avogadro's law which states that, under similar conditions of temperature and pressure, equal volumes of all the gases contain the same number of molecules.

Hence, the table can be represented as :

Gas	Volume (litres)	Number of Molecules
Chlorine	10	$\frac{10}{20}X = X/2$
Nitrogen	20	$\frac{20}{20}X = X$
Ammonia	20	$\frac{20}{20}X = X$
Sulphur dioxide	5	$\frac{5}{20}X = X/4$

Q. 23. A sample of ammonium nitrate when heated yields 8.96 litres of steam (measured at S.T.P.).

$$NH_4NO_3 \longrightarrow N_2O + 2H_2O$$

(i) What volume of dinitrogen oxide is produced at the same time as 8.96 litres of steam ?

(ii) What mass of ammonium nitrate should be heated to produe 8.96 litres of steam ? (Relative molecular mass of ammonium nitreate is 80)

(iii) Determine the precentage of oxygen in ammonium nitrate (O = 16).

Sol. (i)
$$44.8 \text{ L of } H_2O \equiv 22.4 \text{ L of } N_2O \text{ at S.T.P.}$$

$$8.96 \text{ L of } H_2O \equiv \frac{22.4 \text{ L} \times 8.96 \text{ L}}{44.8 \text{ L}} \text{ at S.T.P.}$$

$$= 4.48 \text{ L of } N_2O \text{ at S.T.P.}$$

(ii) 44.8 L steam is liberated by 80 g NH_4NO_3

$\therefore$ 8.96 L steam will be liberated by $\frac{80 \times 8.96}{44.8} = 16$ L

(iii) 80 g NH_4NO_3 contains 48 g O_2

Percentage of oxygen $= \frac{48}{80} \times 100 = 60\%$

Q. 24. (i) What volume of hydrogen sulphide at S.T.P. will burn in oxygen to yield 12.8 g of sulphur dioxide according to the equation ?

$$2H_2S + 3O_2 \longrightarrow 2H_2O + 2SO_2 \quad (H = 1, O = 16, S = 32)$$

(ii) For the volume of hydrogen sulphide determined in (i) above, what volume of oxygen would be required for complete combustion ?

$$\underset{\substack{2 \times 22.4 \\ \text{litres at S.T.P.}}}{2H_2S} + \underset{\substack{3 \times 22.4 \\ \text{litres at S.T.P.}}}{3O_2} \longrightarrow 2H_2O + \underset{\substack{2\,(32 + 2 \times 16) \\ = 128\text{ g}}}{2SO_2}$$

Sol. (i) Volume of H_2S at S.T.P. required to form 128 g of $SO_2 = 2 \times 22.4$ litres

$\therefore$ Volume of H_2S at S.T.P. required to form 1.28 g of $SO_2 = \frac{2 \times 22.4}{128} \times 12.8$ litres

$= 44.8$ litres

(ii) Volume of O_2 required for complete combustion of 2×22.4 litres of H_2S

$= 3 \times 22.4$ litres at S.T.P.

$\therefore$ Volume of O_2 required for combustion of 4.48 litres of H_2S

$= \frac{3 \times 22.4}{2 \times 22.4} \times 4.48 = 6.72$ litres

Q. 25. Iron pyrites has formula FeS_2. What mass of sulphur is contained in 30 g iron pyrites. When roasted, iron pyrites gives sulphur dioxide according to the equation ?

$$4FeS_2 + 11O_2 \longrightarrow 2Fe_2O_3 + 8SO_2$$

What volume of SO_2 at S.T.P. would be liberated by roasting 30 g of iron pyrites [S = 32, Fe = 56, Molar volume of gas is 22.4 litre at S.T.P.]

Sol. (i) Mass of sulphur in iron pyrites.

Gram molecular mass of FeS_2 = $56 + 2 \times 32 = 120$ g

Mass of sulphur = $2 \times 32 = 64$ g

Sulphur contained in 120 g of FeS_2 = 64 g

$\therefore$ Sulphur contained in 30 g of FeS_2 = $\frac{64}{120} \times 30 = 16$ g.

(ii) Volume of SO_2.

$$4FeS_2 + H_2O \longrightarrow 2Fe_2O_3 + 8SO_2$$

Now 4 moles of FeS_2 give 8 moles of SO_2 or 1 mole of FeS_2 gives 2 moles of SO_2.

[Molecular mass of $FeS_2 = 56 + 32 \times 2 = 120$]

Thus, 120 g of FeS_2 gives 2×22.4 litres of SO_2 at S.T.P.

$\therefore$ Volume of SO_2 obtained from 120 g of FeS_2 = 2×22.4 litres at S.T.P.

$\therefore$ Volume of SO_2 obtained from 30 g of FeS_2 = $\frac{2 \times 22.4}{120} \times 30 = 11.2$ litres at S.T.P.

Q. 26. 40 g of sulphur is taken in an enclosed vessel containing 22.4 litre of oxygen, and ignited. Calculate the volume of sulphur dioxide formed and the mass of uncombined sulphur.

(1 mol. of S = 32 g; molar volume = 22.4 litre)

Sol.

$$\underset{32\text{ g}}{S} + \underset{1\text{ vol.}}{O_2} \longrightarrow \underset{1\text{ vol.}}{SO_2}$$

Oxygen : Sulphur dioxide

1 : 1

22.4 : x

Volume of SO_2 produced = 22.4 litres (Equal to the volume of O_2)

Sulphur : Oxygen

32 gm : 1 vol.

$\therefore$ x = 22.4 lit

x = 32 g

$\therefore$ Mass of uncombined sulphur = 40 – 32 = 8 g.

Q. 27. Solid ammonium dichromate (Relative molecular mass 252) decomposes according to the following equation :

$$(NH_4)\,Cr_2O_7 \longrightarrow N_2 + Cr_2O_3 + 4H_2O$$

(i) What volume of nitrogen at S.T.P., will be evolved when 63 g of ammonium dichromate is decomposed ? (H = 1, N = 14, O = 16, Cr = 52)

(ii) If 63 g of ammonium dichromate is heated above 1000°C, what will be the loss of mass ?

Sol. (i) 252 g of ammonium dichromate evolves 22.4 litre N_2 at S.T.P.

$\therefore$ 1 g of ammonium dichromate evolves $= \dfrac{22.4}{252}$ litre N_2

$\therefore$ 63 g of ammonium dichromate evolves $= \dfrac{22.4 \times 63}{252}$ litre N_2

$= 5.6$ litre of N_2.

(ii) Loss of mass due to evolution of nitrogen according to the reaction.

$\because$ 252 g of ammonium dichromate loses 28 g of N_2 gas on heating

$\therefore$ 63 g of ammonium dichromate loses $= \dfrac{28 \times 63}{252}$

$= 7$ g of N_2 gas

Loss due to water vapours :

$\because$ 252 g of ammonium dichromate loses 72 g of H_2O

$\therefore$ 63 g of ammonium dichromate loses $= \dfrac{72 \times 63}{252}$

$= 18$ g of H_2O

$\therefore$ Total loss of mass = Loss of N_2 + Loss of water vapours

= 7 g + 18 g

= 25 g

Q. 28. Solid ammonium dichromate (relative molecular mass = 252) on heating decomposes as follows :

$$(NH_4)_2\,Cr_2O_7 \xrightarrow{\text{Heat}} N_2 + Cr_2O_7 + 4H_2O$$

(i) Calculate the volume of nitrogen at S.T.P., that will be evolved when 31.5 g of ammonium dichromate is heated and

(ii) The mass of chromium(III) oxide formed at the same time

(H = 1; N = 14; O = 16; Cr = 52)

$$\underset{2(14+4)+(52\times 2)+7\times 16\,=\,252\text{ g}}{(NH_4)_2\,Cr_2O_7} \longrightarrow \underset{\substack{22.4\text{ litres}\\ \text{at S.T.P.}}}{N_2} + \underset{\substack{(52\times 2)+(7\times 16)\\ =216\text{ g}}}{Cr_2O_7} + 4H_2$$

Sol. (i) Volume of N_2 evolved at S.T.P. when 31.5 g of $(NH_4)_2Cr_2O_7$ is decomposed

$= \dfrac{22.4\text{ litre}}{252\text{ g}} \times 31.5$ gm

$= 2.80$ litre

(ii) Amount of Cr_2O_3 formed $= \dfrac{216\text{ g}}{252\text{ g}} \times 31.5$ g

$= 27.0$ g

Q. 29. (i) From the balanced chemical equation $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

Compute the mass of oxygen gas that will combine with 8 g of methane.

(Relative atomic mass : H = 1; O = 16; C = 12)

(ii) (a) Iron pyrites has the formula FeS_2. What mass of sulphur is contained in 30 g of pyrites ?

(b) When roasted, iron pyrite gives sulphur dioxide according to the following equation :

$$4FeS_2 + 11O_2 \longrightarrow 2Fe_2O_3 + 8SO_2$$

What volume of sulphur dioxide (at S.T.P.) would be liberated by roasting 30 g of pyrites ?

(S = 32; Fe 56; molar volume of a gas is 22.4 litres at S.T.P.)

Sol. (i) Let x be the mass of oxygen that will combine with 8 g of methane.

$$\begin{array}{rcll} CH_4 & + & 2O_2 \longrightarrow CO_2 + 2H_2O & \\ \text{1 mol} & & \text{2 mol} & \\ \text{1 molar mass} & & \text{2 molar mass} & \\ 12 + 4 \times 1 & & 2 \times (2 \times 16) & \text{(From atomic masses)} \\ = 16\text{ g} & & = 64\text{ g} & \text{(Mass)} \end{array}$$

$$\text{Ratio proportion} = \frac{8 \text{ g } CH_4}{16 \text{ g } CH_4} = \frac{x \text{ g } O_2}{64 \text{ g } O_2}$$

$$\therefore \quad x = \frac{8 \text{ g}}{16 \text{ g}} \times 64 \text{ g} = 32 \text{ g}$$

Mass of oxygen gas that will combine with 8 g of methane is 32 g.

(ii) (a)

$$\begin{aligned} FeS_2 &= 2S \\ (56 + 64 = 120) &= 64 \\ \text{120 g pyrites contain} &= 64 \text{ g S} \\ \text{30 g pyrites contain} &= \frac{64}{120} \times 30 = 16 \text{ g} \end{aligned}$$

(b)

$$\begin{aligned} 4FeS_2 + 11O_2 &= 2Fe_2O_3 + 8SO_2 \\ 4 \times 120 &= 480 \text{ g} = 8 \times 22.4 \text{ litres of } SO_2 \\ \text{480 g pyrites liberate} &= 8 \times 22.4 \text{ litres of } SO_2 \\ \text{30 g pyrites liberate} &= \frac{8 \times 22.4}{480} \times 30 = 11.2 \text{ litres} \end{aligned}$$

Q. 30. Water can be split into hydrogen and oxygen under suitable conditions. The equation representing the change is :

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$

(i) In a given experiment if 2500 cm^3 of hydrogen is being produced, what volume of oxygen is liberated at the same time, under the same conditions of temperature and pressure ?

(ii) The 2500 cm^3 of hydrogen is subjected to $2\frac{1}{2}$ times increase in pressure (temperature remaining constant). What volume will the hydrogen now occupy ?

(iii) Taking the volume of hydrogen calculated in (ii), what change must be made in the Kelvin temperature to return the volume to 2500 cm^3 (P remaining constant) ?

Sol. (i) The reaction is represented as :

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$

From the above reaction, it is clear that the volume of O_2 evolved is half that of hydrogen.

Hence, $$\text{Volume of } O_2 \text{ evolved} = \frac{2500 \text{ cm}^3}{2} = 1250 \text{ cm}^3$$

(ii)

$$\begin{aligned} \text{Let initial pressure of } H_2 &= P_i \\ \text{Final pressure} &= 2.5 \text{ P} \\ P_i V_i &= P_f V_f \\ P \times 2500 &= 2.5 \text{ P} \times V_f \\ V_f &= \frac{P \times 2500}{2.5 \text{ P}} = 1000 \text{ cm}^3. \end{aligned}$$

(iii) Let initial temperature $= T_i$

and Final temperature $= T_f$

According to Charles law, $\frac{V_i}{T_i} = \frac{V_f}{T_f}$

or $T_f = \frac{V_f}{V_i} \times T_i$

$\therefore$ $= \frac{2500 \times T_i}{1000} = 2.5\ T_i$

Thus, the Kelvin temperature must increase by 2.5 times.

Q. 31. Ammonia and oxygen combine to produce water vapour and nitric oxide as per chemical equation : $4NH_3(g) + 5O_2(g) \rightarrow 6H_2O + 4NO.$

(i) Write the chemical equation in terms of Gay-Lussac's law of combining volumes and the Avogadro's law.

(ii) How many moles of oxygen are required to burn 85 g of ammonia ?

(iii) How many moles of nitric oxide will be produced in reaction (i) ?

(iv) What is the volume of NH_3 at S.T.P. that will combine with oxygen in reaction (ii) ?

Sol. (i) The chemical equation can be written as below :

$4NH_3(g)$	+	$5O_2(g)$	$\rightarrow$	$6H_2O$	+	$4NO(g)$	
4 volumes		5 volumes		6 volumes		4 volumes	(In terms of Gay-Lussac's law)
4 mol		5 mol		6 mol		4 mol	(In terms of Avogadro's law)

(ii) Molar mass of ammonia $= \frac{\text{Mol. wt.}}{\text{No. of moles}}$

Moles wt. $= 4\,(14 + 3 \times 1) = 68$ g

No. of moles $= 4$

$\therefore$ Molar mass of ammonia $= \frac{68}{4} = 17$ g/mol.

Moles of ammonia $= \frac{\text{Mass of ammonia}}{\text{Molar mass of ammonia}}$

$= \frac{85 \text{ g}}{17 \text{ g/mol}} = 5$ mol

Let x be the moles of O_2 required to combine with 5 moles of NH_3

$$\underset{4\text{ mol}}{\overset{5\text{ mol}}{4NH_3(g)}} + \underset{5\text{mol}}{\overset{x\text{ mol}}{5O_2(g)}} \rightarrow 6H_2O + 4NO(g)$$

$$\text{Ratio proportion} = \frac{5 \text{ mol } NH_3}{4 \text{ mol } NH_3} = \frac{x \text{ mol } O_2}{5 \text{ mol } O_2}$$

$$x = \frac{5 \text{ mol} \times 5 \text{ mol}}{4 \text{ mol}} = 6.25 \text{ mol } O_2.$$

(iii) Let y be the moles of nitric oxide formed when 5 moles of ammonia are completely burnt in oxygen.

$$\underset{4\text{ mol}}{\overset{5\text{ mol}}{4NH_3(g)}} + 5O_2(g) \rightarrow 6H_2O + \underset{4\text{ mol}}{\overset{y\text{ mol}}{4NO(g)}}$$

$$\text{Ratio proportion} = \frac{5 \text{ mole } NH_3}{4 \text{ mole } NH_3} = \frac{y \text{ mole NO}}{4 \text{ mole NO}}$$

$$y = \frac{5 \text{ mole } NH_3}{4 \text{ mole } NH_3} \times 4 \text{ mole NO}$$

On burning ammonia completely = 5 mole NO will be produced.

(iv) $\because$ 1 mole NH_3 occupies 22.4 litre at S.T.P.

$\therefore$ 5 mole NH_3 will occupy $5 \times 22.4 = 112$ litre.

Volume of ammonia at S.T.P. is 112 litre.

Q. 32. The equations given below relate to the manufacture of sodium carbonate (Molecular weight of $Na_2CO_3 = 106$).

$$NaCl + NH_3 + CO_2 + H_2O \longrightarrow NaHCO_3 + NH_4Cl$$
$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

Questions (i) and (ii) are based on the production of 21.2 g of sodium carbonate.

(i) What mass of sodium hydrogen carbonate must be heated to give 21.2 g of Sodium carbonate (Molecular weight of $NaHCO_3 = 84$) ?

(ii) To produce the mass of sodium hydrogen carbonate calculated in (a), what volume of carbon dioxide, measured at S. T. P., would be required ?

Sol. (i)

$$\underset{\substack{2(23+1+12+48)\\ = 2\times 84\\ = 168\text{ g}}}{2NaHCO_3} \longrightarrow \underset{\substack{(23\times 2+12+48)\\ = 106\\ = 106\text{ g}}}{Na_2CO_3} + H_2O + CO_2$$

$\because$ 106 g of Na_2CO_3 is obtained from 168 g of $NaHCO_3$

$\therefore$ 21.2 g of Na_2CO_3 is obtained from $= \dfrac{168 \times 21.2}{106}$

$= 33.6$ g $NaHCO_3$.

(ii)
$$NaCl + NH_3 + \underset{\substack{1\text{ mole.}\\ 22.4\text{ lit at S.T.P.}}}{CO_2 + H_2O} \longrightarrow \underset{\substack{1\text{ mole}\\ 84\text{ g}}}{NaHCO_3} + NH_4Cl$$

$\because$ 84 g of $NaHCO_3$ requires 22.4 lit of CO_2

$\therefore$ 33.6 g of $NaHCO_3$ requires $= \dfrac{22.4 \times 33.6}{84}$

$= 8.96$ lit of CO_2.

Q. 33. (i) Concentrated nitric acid oxidizes phosphorus to phosphoric acid according to the following equation :

$$P + 5HNO_3 \longrightarrow H_3PO_4 + H_2O + 5NO_2$$

(a) What mass of phosphoric acid can be prepared from 6.2 g of phosphorus ?

(b) What mass of nitric acid will be consumed at the same time ?

(c) What will be the volume of steam at the same time if measured as 760 mm Hg pressure and 273°C ?

[H = 1; N = 14; O = 16; P = 31]

(ii) Ammonia may be oxidised to nitrogen monoxide in the presence of a catalyst according to the following equation :

$$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$$

If 27 litres of reactants are consumed, what volume of nitrogen monoxide is produced at the same temperature and pressure ?

Sol. (i)
$$\underset{(31\text{ g})}{P} + \underset{(315\text{ g})}{5\,HNO_3} \longrightarrow \underset{(98\text{ g})}{H_3PO_4} + H_2O + \underset{(22.4\text{ lit at S.T.P.})}{5NO_2}$$

(a) 31 g of phosphorus produces 98 phosphoric acid

$(H_3PO_4 = 3 \times 1 + 31 + 4 \times 16 = 98\text{ g})$

6.2 g phosphorus produces $= \dfrac{98}{31} \times 62 = 196$ g

(b) 31 g of P consumes 5×63 gm of HNO_3

Mass of HNO_3 consumed by 6.2 g of phosphorus $= \dfrac{5 \times 63}{31} \times 6.2 = 63$ g

(c) 31 g of phosphorus at S.T.P. produces = 22.4 litres of steam.

Vol. of steam S.T.P. formed by 6.2 g of phosphorus

$= \dfrac{22.4}{31} \times 6.2$

$= 4.48$ litres

As we know, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

At S.T.P. $P_1 = 760$ mm; $V_1 = 4.48$ lit; $T_1 = 273$ K

$P_2 = 760$ mm; $T_2 = 273°C + 273 = 546$ K; $V_2 = ?$

$$V_2 = \frac{760 \times 4.48 \times 546}{760 \times 273} = 8.96 \text{ litres}$$

(ii) The balanced equation for the oxidation of ammonia is

$$\underset{4\text{ vols.}}{4NH_3(g)} + \underset{5\text{ vols.}}{5O_2(g)} \longrightarrow \underset{4\text{ vols.}}{4NO(g)} + 6H_2O(l)$$

Total volume of reactant consumed = 27 litres

Ratio by volume of $NH_3 : O_2$ is 4 : 5

$\therefore$ Volume of $NH_3 = \frac{27 \times 4}{9} = 12$ litres

Volume of $O_2 = \frac{27 \times 5}{9} = 15$ litres

$$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$$

9 litres of reactants give 4 litres of NO

$\therefore$ 27 litres would give $\frac{27 \times 4}{9} = 12$ litres

12 litres, of NH_3 produces 4 litres, of NO

$\therefore$ 12 litres of NH_3 produces of 12 litres, of NO

$\therefore$ Volume of NO produced = 12 litres

Q. 34. What is the mass of Nitrogen in 1000 kg of Urea $[CO(NH_2)_2]$? (Answer correct to the nearest kg) [H = 1; C = 12; N = 14; O = 16]

Sol. Molecular formula of urea is NH_2CONH_2.

Molecular mass of Urea $[CO(NH_2)]_2$ = 12 + 16 + 2 (14 + 1 × 2)
= 28 + 32
= 60 g

60 g of urea contains Nitrogen = 2 × 14
= 28 g

$\therefore$ 1000 kg of urea contains Nitrogen = $\frac{28}{60} \times 1000$
= 467 kg.

Q. 35. Calculate the atomicity of oxygen molecule from the following information :

Vapour density of oxygen = 16

Relative atomic mass of oxygen = 16

Show all the calculations.

Sol. Vapour density of oxygen = 16

$\therefore$ Molecular weight of oxygen = 2 × V.D.
= 2 × 16 = 32

Relative atomic mass of oxygen = 16

$\therefore$ No. of atoms in one molecule of oxygen = $\frac{\text{Molecular wt.}}{\text{Atomic wt.}}$
$= \frac{32}{16} = 2$

$\therefore$ Atomicity of Oxygen = 2

Q. 36. (i) 67.2 litres of hydrogen combines with 44.8 litres of nitrogen of form ammonia under specific conditions as :

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Calculate the volume of ammonia produced. What is the other substance, if any, that remains in the resultant mixture ?

(ii) The mass of 5.6 dm^3 of a certain gas at S.T.P. is 12.0 g. Calculate the relative molecular mass of the gas.

(iii) Find the total percentage of Magnesium in magnesium nitrate crystals, $Mg(NO_3)_2.6H_2O$.
[Mg = 24, N = 14; O = 16 and H = 1]

Sol. (i)

N_2	+	$3H_2$	$\longrightarrow$	$2NH_3$
1 vol.		3 vol.		2 vol.
22.4 lit		67.2 lit		44.8 lit

According to Gay-Lussac's law :

3 volumes of H_2 will give 2 vol. of NH_3.

$\therefore$ 67.2 lit of H_2 will give $= \frac{2 \times 67.2}{3} = 44.8$ lit of NH_3

At the same time

3 vol. of H_2 will react with 1 vol. of N_2

$\therefore$ 67.2 lit of H_2 will react with $= \frac{1}{3} \times 67.2 = 22.4$ lit of N_2

$\therefore$ (44.8 – 22.4) = 22.4 lit of N_2 will be left unreacted in the mixture.

(ii) According to molar volume concept weight of 22.4 lit of a gas at S.T.P. is equal to molecular mass of the gas. Now,

$\because$ 5.6 lit of gas at S.T.P. weighs 12.0 g

$\therefore$ 22.4 lit of gas at S.T.P. weighs $= \frac{12 \times 22.4}{5.6} = 48$ g

Hence, molecular mass of the gas is 48 g.

(iii) Relative molecular mass of $Mg(NO_3)_2.6H_2O$

$= (24) + (14 \times 2) + (16 \times 6) + (12 \times 1) + (6 \times 16)$

$= 24 + 28 + 96 + 12 + 96 = 256$ g

Amount of Mg in 256 g of magnesium nitrate = 24 g

$\therefore$ % of Mg in $Mg(NO_3)_2.6H_2O = \frac{24}{256} \times 100 = 9.37\%$

Q. 37. (i) What is the volume (measured in dm^3 or litres) occupied by one mole of a gas at S.T.P. ?

(ii) 112 cm^3 (at S.T.P.) of a gaseous fluoride of phosphorus has a mass of 0.63 g. Calculate the relative molecular mass of the fluoride. If the molecule of the fluoride contains only one atom of phosphorus, then determine the formula of the phosphorus fluoride. (F = 19; P = 31)

(iii) 2.24 litres of a gas weighs 4.4 g at S.T.P. Calculate the vapour density of the gas.

Sol. (i) The volume occupied by one mole of a gas at S.T.P. is 22.4 litre.

(ii) Mass of 112 cm^3 of phosphorus fluoride = 0.63 g

Mass of 22,400 cm^3 of phosphorus fluoride $= \frac{0.63}{112} \times 22400$

= 126 g

Relative molecular mass of fluoride = 126 g

Mass of fluoride in phosphorus fluoride = 126 – 31 = 95 g.

Hence, Number of fluorine in phosphorus fluoride

$= \frac{95}{19} = 5$

Thus, The formula of phosphorus fluoride = PF_5

(iii) $\because$ 2.24 litres of gas weighs 4.4 g at S.T.P.

$\therefore$ 22.4 litres of gas weighs $= \frac{4.4 \times 22.4}{2.24} = 44$ g

$\therefore$ Vapour density $= \frac{\text{Molecular wt.}}{2} = \frac{44}{2} = 22$

Q. 38. Vapour density of a gas Z is 23. Calculate : (i) number of moles, (ii) weight in grams and (iii) number of molecules in 6.72 dm^3 of gas at S.T.P.

Sol. Molecular weight of gas Z $= 2 \times$ V.D.

$= 2 \times 23 = 46$ a.m.u.

$\therefore$ 1 mole of gas $= 46$ g

(i) 22.4 dm^3 of gas Z at S.T.P. $=$ 1 mole

$\therefore$ 6.72 dm^3 of gas Z at S.T.P. $= \frac{6.72}{22.4} = 0.3$ moles.

(ii) 1 mole of gas Z at S.T.P. weighs $= 46$ g

$\therefore$ 0.3 mole of gas Z at S.T.P. weighs $= 46 \times 0.3 = 13.8$ g

(iii) 1 mole of gas Z at S.T.P. contains $= 6 \times 10^{23}$ molecules

$\therefore$ 0.3 mole of gas Z at S.T.P. contains $= 6 \times 10^{23} \times 0.3$

$= 1.8 \times 10^{23}$ molecules.

Q. 39. The atomic weight of oxygen is 16.0 and the formula of oxygen molecule is O_2. Calculate the weight of :

(i) 1 atom of oxygen, (ii) One molecule of oxygen.

Sol. (i) Molecular weight of oxygen $= 16 \times 2 = 32$ g

$\therefore$ 32 g of oxygen $= 2 \times 6.023 \times 10^{23}$ atoms of oxygen

or $2 \times 6.023 \times 10^{23}$ atoms weighs $= 32$ g

$\therefore$ Weight of one oxygen atom $= \frac{32}{2 \times 6.023 \times 10^{23}} = 2.656 \times 10^{-23}$ g

(ii) Similarly 32 g of oxygen contains $= 6.023 \times 10^{23}$ molecules of oxygen

$\therefore$ 1 molecule of oxygen will weighs $= \frac{32}{6.023 \times 10^{23}}$

$= 5.312 \times 10^{-23}$ g.

Q. 40. Calculate the relative molecular masses (or molecular weights) of the following compounds :

(i) Copper sulphate crystals, $CuSO_4.5H_2O$. (ii) Ammonium sulphate, $(NH_4)_2SO_4$.

(iii) Cane sugar, $C_{12}H_{22}O_{11}$.

(Given atomic masses Cu = 63.5, S = 32, O = 16, N = 14, C = 12)

Sol. (i) $CuSO_4.5H_2O = 63.5 + 32 + (16 \times 4) + 5 \times (2 + 16)$

$= 159.5 + 90 = 249.5$ a.m.u.

(ii) $(NH_4)_2SO_4 = N_2H_8SO_4$

$= 14 \times 2 + 1 \times 8 + 32 + 16 \times 4$

$= 28 + 8 + 32 + 64 = 132$ a.m.u.

(iii) $C_{12}H_{22}O_{11} = 12 \times 12 + 22 \times 1 + 11 \times 16$

$= 144 + 22 + 176 = 342$ a.m.u.

Q. 41. (i) An organic compound with vapour density = 94 contains.

C = 12.67%, H = 2.13%, and Br = 85.11%. Find the molecular formula.

[Atomic mass : C = 12, H = 1, Br = 80]

(ii) Calculate the mass of

(a) 10^{22} atoms of sulphur.

(b) 0.1 mole of carbon dioxide.

[Atomic mass : S = 32, C = 12 and O = 16 and Avogadro's Number = 6×10^{23}]

Sol. (i)

Elements	% Ratio	Atomic mass	Relative no. of atoms	Simplest ratio
C	12.67	12	$12.67/12 = 1.055$	$1.055/1.055 = 1$
H	2.13	1	$2.13/1 = 2.13$	$2.13/1.055 = 2$
Br	85.11	80	$85.11/80 = 1.063$	$1.063/1.055 = 1$

$\therefore$ Empirical formula of the compound is CH_2Br

$$\text{Molecular formula} = (\text{Empirical formula})_n$$

$$n = \frac{\text{M.W.}}{\text{Empirical formula weight}}$$

$$= \frac{2 \times \text{V.D.}}{\text{Empirical formula weight}}$$

$$= \frac{2 \times 94}{(12 + 2 + 80)} = \frac{2 \times 94}{94} = 2$$

$$\therefore \quad \text{Molecular formula} = (CH_2Br)_2 = C_2H_4Br_2$$

(ii) (a) 1 mole of sulphur $= 6 \times 10^{23}$ atoms $= 32$ g of sulphur

$$\therefore \quad 10^{22} \text{ atoms} = \frac{32 \times 10^{22}}{6 \times 10^{23}} = \frac{32}{60} = 0.533 \text{ g}$$

(b) 1 mole of carbon dioxide (CO_2)

$$= 12 + (2 \times 16) = 44 \text{ g}$$

$\therefore$ 0.1 mole of carbondioxide $= 0.1 \times 44 = 4.4$ g

Q. 42. Half-litre of carbon dioxide is passed over red hot carbon. The volume becomes 700 ml. Find the composition of the product, assuming that all the volumes of gases are measured at S.T.P.

Sol. The chemical equation representing the reaction is

$$\underset{1\text{ vol.}}{CO_2 + C} \longrightarrow \underset{2\text{ vol.}}{2CO}$$

Let, x ml of CO_2 react with red hot carbon to form $2x$ ml of CO. Initial volume of carbon dioxide = 500 ml.

Final volume of the reaction mixture = 700 ml.

Thus, we have $(500 - x) = (700 - 2x)$

or $2x - x = 700 - 500$

$x = 200$

Volume of carbon monoxide formed $= 2 \times 200 = 400$ ml

Volume of carbon dioxide remains after the reaction

$= (500 - 200) = 300$ ml.

Q. 43. Iron forms three different forms of oxides :

(i) Ferrous oxide [FeO], (ii) Ferric oxide [Fe_2O_3],

(iii) Tri-ferric tetraoxide [Fe_3O_4].

Calculate the percentage of iron in each of the above oxides.

Sol. (i) Molecular formula of ferrous oxide is FeO.

$\because$ Molecular weight of ferrous oxide $= 56 + 16 = 72$ g

$$\therefore \quad \text{\% of Iron in ferrous oxide (FeO)} = \frac{56}{72} \times 100 = \frac{700}{9} = 77.78\%.$$

(ii) Molecular formula of ferric oxide is Fe_2O_3.

$\because$ Molecular weight of ferric oxide $= (2 \times 56) + (3 \times 16)$

$= 112 + 48 = 160$ g

$$\therefore \quad \text{\% of Iron in ferric oxide } (Fe_2O_3) = \frac{112}{160} \times 100 = 70\%.$$

(iii) Molecular formula of tri-ferric tetraoxide is Fe_3O_4.

$$\text{Molecular weight of tri-ferric tetraoxide} = (3 \times 56) + (4 \times 16) = 168 + 64 = 232 \text{ g}$$

$$\therefore \quad \% \text{ of Iron in tri-ferric tetraoxide } (Fe_3O_4) = \frac{168}{232} \times 100 = \frac{2100}{29} = 72.41\%.$$

Q. 44. (i) If 16.4 gram of calcium nitrate is heated :

(a) Calculate the volume of nitrogen dioxide obtained at S.T.P.

(b) The weight of calcium oxide obtained.

(ii) Paddy crop removes 20 kg of nitrogen from soil per hectare. Calculate the amount of calcium nitrate $[Ca(NO_3)_2]$ which should be added to the soil to provide nitrogen to a farm of 10 hectares ? State your answer in kg (approx.). [Ca = 40, N = 14, O = 16]

Sol. (i) The reaction is :

$$\underset{2[40 + 2(14 + 48)] = 328 \text{ g}}{2Ca(NO_3)_2} \longrightarrow \underset{112 \text{ g}}{2CaO} + \underset{4 \text{ moles}}{4NO_2} + O_2$$

(a) 328 g of $Ca(NO_3)_2$ liberate 4×22.4 NO_2

$$\therefore \quad 16.4 \text{ g will liberate} = \frac{4 \times 22.4}{328} \times 16.4 = 4.48 \text{ litre } NO_2 \text{ at S.T.P.}$$

(b) 328 g of $Ca(NO_3)_2$ yields 112 g CaO

$$\therefore \quad 16.4 \text{ g will yields} = \frac{112}{328} \times 16.4 = 5.6 \text{ g CaO}$$

(ii) Total amount of nitrogen required for 10 hectare farm = $10 \times 20 = 200$ kg

$$\text{Gram molecular wt. of } Ca(NO_3)_2 = 1(Ca) + 2(N) + 6(O) = 1 \times (40) + 2 \times (14) + 6 \times (16) = 40 + 28 + 96 = 164 \text{ kg.}$$

28 kg of nitrogen can be obtained from

$$Ca(NO_3)_2 = 164 \text{ kg.}$$

$\therefore$ 200 kg of nitrogen can be obtained from $Ca(NO_3)_2$

$$= \frac{164 \times 200}{28} = 1171.42 \text{ kg.}$$

Q. 45. When heated, potassium permanganate decomposes according to the following equation :

$$2KMnO_4 \longrightarrow \underbrace{K_2MnO_4 + MnO_2}_{\text{Solid residue}} + O_2$$

(i) Some potassium permanganate was heated in a test tube. After collecting one litre of oxygen at room temperature, it was found that the test tube had undergone a loss in mass of 1.32 g. If one litre of hydrogen under the same conditions of temperature and pressure has a mass of 0.0825 g, calculate the relative molecular mass of oxygen.

(ii) Given that the molecular mass of potassium permanganate is 158, what volume of oxygen (measured at room temperature) would be obtained by the complete decomposition of 15.8 of potassium permanganate (molar volume at room temperature is 24 litres) ?

Sol. (i) Mass of one litre of oxygen = 1.32 g

Mass of one litre of hydrogen under the same condition = 0.0825 g

$$\text{V.D. of oxygen} = \frac{\text{Mass of one litre of oxygen}}{\text{Mass of one litre of hydrogen}} = \frac{1.32 \text{ g}}{0.0825 \text{ g}} = 16$$

Relative molecular mass of oxygen $= 2 \times$ V.D.

$= 2 \times 16 = 32$

(ii) $$\underset{(2 \times 158\text{ g})}{2KMnO_4} \longrightarrow K_2MnO_4 + MnO_2 + \underset{1\text{ mol (24 lit)}}{O_2}$$

(2 × 158) g $KMnO_4$ on heating liberates 24 litre O_2 at room temperature

15.8 g $KMnO_4$ on heating liberates $\dfrac{24\text{ litre}}{(2 \times 158)\text{ g}} \times 15.8\text{ g} = 1.2$ litre O_2

Volume of O_2 liberated at room temperature by heating 15.8 g potassium permanganate = 1.2 litres.

Q. 46. (i) The compound A has the following percentage composition by mass.

Carbon = 26.7%, oxygen = 71.1%, hydrogen = 2.2%. Determine the Empirical formula of compound A [work to one decimal place] [H = 1, C = 12, O = 16].

(ii) If the relative molecular mass of A is 90, what is the molecular formula of A ?

(iii) The compound A is weak acid. What is meant by this statement ?

Sol. (i) Relative number of atoms,

$C = \dfrac{26.7}{12} = 2.2$ $\quad O = \dfrac{71.1}{16} = 4.4$ $\quad H = \dfrac{2.2}{1} = 2.2$

Simplest ratio,

$C = \dfrac{2.2}{2.2} = 1$ $\quad O = \dfrac{4.4}{2.2} = 2$ $\quad H = \dfrac{2.2}{2.2} = 1$

Empirical formula of the compound A is CO_2H.

(ii) The empirical formula weight $= [12 + 32 + 1] = 45$

$$n = \frac{\text{Molecular weight}}{\text{Empirical weight}} = \frac{90}{45}$$

$n = 2$

∴ Molecular formula of the compound A is $[CO_2H]_2$

$= C_2O_4H_2$

$$= \begin{array}{c} COOH \\ | \\ COOH \end{array} \text{ (Oxalic acid)}$$

(iii) It does not completely ionise into $C_2O_4^{2-}$ and H^+ ions.

Q. 47. (i) Find the total percentage of oxygen in magnesium nitrate crystals.

$(MgNO_3)_2.6H_2O$ [H = 1, N = 14, O = 16, Mg = 24].

(ii) Calculate the percentage of nitrogen in aluminium nitride.

(Al = 27, N = 14)

(iii) Find out the percentage, by weight of phosphorus present in calcium phosphate.

Sol. (i) $Mg(NO_3)_2.6H_2O$

The molecular mass $= 24 + (2 \times 62) + 6 \times 18$

$= 24 + 124 + 108 = 256$ g

Mass due to oxygen $= (2 \times 48) + (6 \times 16) = 96 + 96 = 192$ g

∴ % of oxygen $= \dfrac{192}{256} \times 100 = 75\%$.

(ii) Molecular weight of AlN = 27 + 14 = 41 g

$$\% \text{ of Nitrogen} = \frac{14 \times 100}{41} = 34.14\%$$

(iii) Molecular formula of calcium phosphate is $Ca_3(PO_4)_2$.

Molecular weight of calcium phosphate

$= 3 \times 40 + 2 \times [31 + (4 \times 16)]$

$= 120 + 2 \times (31 + 64)$

$= 310$ g

and molecular weight of phosphorus = 62 g

$\therefore$ % of phosphorus in calcium phosphate $= \frac{62}{310} \times 100 = 20\%$.

Q. 48. (i) If 112 cm^3 of hydrogen sulphide is mixed with 120 cm^3 of chlorine at S.T.P., what mass of sulphur is formed ?

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(ii) Washing soda has the formula $Na_2CO_3.10H_2O$. What mass of anhydrous sodium carbonate is left when all the water of crystallization is expelled by heating 57.2 g of washing soda ?

(iii) When excess lead nitrate solution was added to a solution of sodium sulphate, 15.15 g of lead sulphate was precipitated. What mass of sodium sulphate was present in the original solution ?

$$Na_2SO_4 + Pb(NO_3)_2 \rightarrow PbSO_4 + 2NaNO_3$$

(H = 1; C = 12; O = 16; Na = 23; S = 32; Pb = 207)

Sol. (i)
$$\underset{1\text{ vol.}}{H_2S} + \underset{1\text{ vol.}}{Cl_2} \longrightarrow 2HCl + S$$

From the above equation, it is clear that 1 vol. of H_2S reacts with 1 vol. of Cl_2. Thus, when 112 cm^3 of H_2S is mixed with 120 cm^3 of Cl_2 at S.T.P. only 112 cm^3 of Cl_2 will react.

Amount of gas in 22,400 cm^3 at S.T.P. = 1 mole

Amount of gas in 112 cm^3 at S.T.P. $= \frac{112}{22400}$

= 0.005 mole

112 cm^3 of H_2S at S.T.P. = 0.005 mole H_2S

112 cm^3 of Cl_2 at S.T.P. = 0.005 mole Cl_2

$$\underset{0.005}{H_2S} + \underset{0.005}{Cl_2} \longrightarrow 2HCl + \underset{0.005\text{ mol.}}{S}$$

Amount of sulphur formed = (0.005 × 32)

= 0.16 g

(ii)
$$\underset{286\text{ g}}{Na_2CO_3.10H_2O} \longrightarrow \underset{106\text{ g}}{Na_2CO_3} + 10H_2O$$

$\because$ 286 g washing soda on heating, produce 106 g anhydrous sodium carbonate.

$\therefore$ 57.2 g washing soda produce $= \frac{57.2 \times 106}{286}$

= 21.20 g.

(iii)
$$\underset{\substack{2 \times 23 + 32 + 16 \times 4 \\ = 142\text{ g}}}{Na_2SO_4} + Pb(NO_3)_2 \longrightarrow \underset{\substack{207 + 32 + 64 \\ = 303\text{ g}}}{PbSO_4} + 2NaNO_3$$

$\therefore$ 303 g $PbSO_4$ is precipitated by 142 g Na_2SO_4

$\therefore$ 15.15 g $PbSO_4$ will be precipitated by $= \frac{15.15 \times 142}{303}$

= 7.1 g Na_2SO_4

Q. 49. Concentrated nitric acid oxidizes phosphorus to phosphoric acid according to the following equation :

$$P + 5HNO_3 \longrightarrow H_3PO_4 + H_2O + 5NO_2$$

(i) What mass of phosphoric acid can be prepared from 6.2 g of phosphorus ?

(ii) What mass of nitric acid will be consumed at the same time ?

[H = 1; N = 14; O = 16; P = 31]

Sol.
$$\underset{31\text{ g}}{P} + \underset{\substack{5[1 + 14 + 48] \\ 5 \times 63 = 315\text{ g}}}{5\,HNO_3} \longrightarrow \underset{\substack{(3) + (31) + (64) \\ = 98\text{ g}}}{H_3PO_4} + 5NO_2 + \underset{\substack{22.4\text{ dm}^3 \\ \text{at S.T.P.}}}{H_2O}$$

(i) $\because$ 31 g of phosphorus can produce phosphoric acid = 98 g

$\therefore$ 6.2 g of phosphorus can produce phosphoric acid $= \frac{98 \times 6.2}{31} = 19.6$ g

(ii) $\because$ 31 g of phosphorus requires nitric acid $= 315$ g

$\therefore$ 6.2 g of phosphorus requires nitric acid $= \dfrac{6.2 \times 315}{31}$

$= 63.0$ g.

Q. 50. Following chemical equation is given :

$$2NH_4Cl + Ca(OH)_2 \longrightarrow CaCl_2 + 2H_2O + 2NH_3 .$$

(i) Calculate the mass of ammonia obtained from 321 g of ammonium chloride.

(ii) Find the mass of ammonium chloride required to obtain 6 moles of H_2O.

(iii) Find the mass of ammonium chloride required to obtain 4 moles of NH_3.

Sol. The chemical equation is :

$$\underset{\substack{2\,(14+4\,(1)+35.5)\\ =107\text{ g}}}{2NH_4Cl} + Ca(OH)_2 \longrightarrow CaCl_2 + 2H_2O + \underset{\substack{2\text{ moles or }2\,[14+(1\times 3)]\\ =2\,(17)=34\text{ g}}}{2NH_3}$$

(i) $\because$ 107 g of NH_4Cl gives 34 g of ammonia

$\therefore$ 1 g of NH_4Cl gives $= \dfrac{34}{107}$ g of ammonia

$\therefore$ 321 g of NH_4Cl gives $= \dfrac{34 \times 321}{107}$ g $= 102$ g of ammonia.

(ii) $\because$ 2 moles of H_2O is obtained from 107 g NH_4Cl

$\therefore$ 6 moles of H_2O will be obtained from

$= \dfrac{107 \times 6}{2} = 321$ g NH_4Cl.

(iii) $\because$ 2 moles of NH_3 is obtained from 107 g of NH_4Cl

$\therefore$ 1 mole of NH_3 is obtained from $= \dfrac{107}{2}$ g of NH_4Cl

$\therefore$ 4 moles of NH_3 is obtained from $= \dfrac{107 \times 4}{2}$ g $= 214$ g of NH_4Cl.

Q. 51. 10 g of a mixture of sodium chloride and anhydrous sodium sulphate is dissolved in water. An excess of barium chloride solution is added and 6.99 g of barium sulphate is precipitated according to the equation given below :

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

Calculate the percentage of sodium sulphate in the original mixture.

(O = 16, Na = 23, S = 32, Ba = 137)

Sol.

$$\underset{\substack{2\times 23+32+4\times 16\\ =\ 142\text{ g}}}{Na_2SO_4} + BaCl_2 \longrightarrow \underset{\substack{137+32+4\times 16\\ =233\text{ g}}}{BaSO_4} + 2NaCl$$

From the above equation, it is clear that 233 g barium sulphate is precipitated from 142 g sodium sulphate.

$\therefore$ 6.99 g barium sulphate is precipitated from $= \dfrac{142\text{ g}}{233\text{ g}} \times 6.99 = 4.26$ g sodium sulphate

Amount of sodium sulphate in 10 g mixture $= 4.26$ g

Percentage of sodium sulphate in original mixture

$= \dfrac{4.26\text{ g}}{10\text{ g}} \times 100 = 42.6\%$

Q. 52. (i) The reaction of potassium permanganate(VII) with acidified iron(II) sulphate is given below :

$$2KMnO_4 + 10FeSO_4 + 8H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 5Fe_2\,(SO_4)_3 + 8H_2O$$

If 15.8 g of potassium permanganate was used in the reaction, calculate the mass of iron(II) sulphate used in the above reaction.

(K = 39, Mn = 55, Fe = 56, S = 32, O = 16)

(ii) 20% nitric acid reacts with 4.11 g of lead carbonate of 65% purity. Calculate the weight of nitric acid to complete the reaction. [Pb = 207; C = 12; O = 16; N = 14]

Sol. (i) $\underset{2 \times 158\ g}{2KMnO_4} + \underset{10 \times 152\ g}{10FeSO_4} + 8H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 5Fe_2(SO_4)_3 + 8H_2O$

(2×158) g $KMnO_4$ uses (10×152) g $FeSO_4$

15.8 g $KMnO_4$ uses $\dfrac{(10 \times 152)\ g}{(2 \times 158)\ g} \times 15.8\ g = 76$ g $FeSO_4$

Amount of $FeSO_4$ used by 15.8 g $KMnO_4$ = 76 g

(ii) Weight of impure lead carbonate = 4.11 g

$\therefore$ Weight of pure lead carbonate $= \dfrac{4.11 \times 65}{100} = 2.6715$ g

Writing gram-molecular weights

$$\underset{\substack{[207 + 12 + 3\,(16)] \\ = 267\ g}}{PbCO_3} + \underset{\substack{2\,[1 + 14 + 3\,(16)] \\ = 126\ g}}{2HNO_3} \longrightarrow Pb(NO_3)_2 + H_2O + CO_2$$

When pure $PbCO_3$ is 267 g pure nitric acid required

= 126 g

When pure $PbCO_3$ is 2.6715 g pure nitric acid required

$= \dfrac{126 \times 2.6715}{267}$ g

= 1.260 g.

$\therefore$ Wt. of pure nitric acid required = 1.260 g

$\therefore$ Wt. of 20% nitric acid required $= \dfrac{1.26 \times 100}{20} = 6.30$ g.

Q. 53. (i) Determine the empirical formula of the compound whose composition by mass is : 42% nitrogen, 48% oxygen and 9% hydrogen. [H = 1; N = 14; O = 16]

(ii) Determine the empirical formula of a compound containing 47.9% potassium, 5.5% beryllium and 46.6% fluorine by mass.

(Atomic weight of Be = 9; F = 19; K = 39) Work to one decimal place.

Sol. (i)

Element	Rel. At. Mass	% age comp.	Relative No. of atoms	Simple Ratio
N	14	42	$\frac{42}{14} = 3$	$\frac{3}{3} = 1$
O	16	48	$\frac{48}{16} = 3$	$\frac{3}{3} = 1$
H	1	9	$\frac{9}{1} = 9$	$\frac{9}{3} = 3$

Empirical formula = NOH_3

(ii)

	%	Atomic mass	Relative no. of atoms	Simplest Ratio
K	47.9	39	$\frac{47.9}{39} = 1.2$	$\frac{1.2}{0.6} = 2$
Be	5.5	9	$\frac{5.5}{9} = 0.6$	$\frac{0.6}{0.6} = 1$
F	46.6	19	$\frac{46.6}{19} = 2.4$	$\frac{2.4}{0.6} = 4$

Empirical formula = K_2BeF_4.

Q. 54. (i) Calculate the Empirical formula of the compound having 37.6% of sodium, 23.1% of silicon and 39.3% of oxygen. [O = 16, N = 23, Si = 28]

(ii) The Empirical formula of a compound is C_2H_5. It has a vapour density of 29. Determine the relative molecular formula mass of the compound and hence its molecular formula.

Sol. (i)

Element	Atomic Mass	% Comp.	Rel. No. of moles	Simple ratio
Na	23	37.6	$\frac{37.6}{23} = 1.63$	$\frac{1.63}{0.82} = 1.98 = 2$
Si	28	23.4	$\frac{23.1}{28} = 0.82$	$\frac{0.82}{0.82} = 1$
O	16	39.3	$\frac{39.3}{16} = 2.45$	$\frac{2.45}{0.82} = 2.98 = 3$

The Empirical formula is Na_2SiO_3.

(ii) $\because$ The Empirical formula $= C_2H_5$

$\therefore$ Empirical formula wt. $= 2\,(C) + 5\,(H)$

$= 2\,(12) + 5\,(1)$

$= 24 + 5 = 29$

The vapour density of C_2H_5 is 29,

$\therefore$ Molecular weight $= 2 \times$ V.D.

$= 2 \times 29$

$= 58$

No. of Molecules $n = \frac{\text{M.F.W.}}{\text{E.F.W.}}$

$n = \frac{58}{29}$

$n = 2$

Molecular formula $= n \times$ [Empirical formula]

$= 2\,[C_2H_5]$

$= C_4H_{10}$

The molecular formula $= C_4H_{10}$

The relative molecular formula mass of $C_4H_{10} = 4\,(C) + 10\,(H)$

$= 4\,(12) + 10\,(1)$

$= 48 + 10$

$= 58$ g

Q. 55. (i) Calculate the percentage of platinum in ammonium chloroplatinate $(NH_4)_2PtCl_6$ (Give your answer correct to the nearest whole number).

(ii) The percentage composition of sodium phosphate as determined by analysis, is 42.1% sodium, 18.9% phosphorus and 39% oxygen. Find the empirical formula of the compound (work to two decimal places).

(H = 1, N = 14, O = 16, Na = 23, P = 31, Cl = 35.5, Pt = 195)

Sol. (i) Mol. wt. of $(NH_4)_2\,PtCl_6 = [14 + 4 \times 1] \times 2 + 195 + 6 \times 35.5$

$= 18 \times 2 + 195 + 213$

$= 444$

% of platinum $= \frac{195}{444} \times 100 = 43.9\%$

(ii)

Element	Percentage	At. wt.	% / At. wt	Simple ratio
Na	42.1%	23	$\frac{42.1}{23} = 1.8$	$\frac{1.8}{0.6} = 3$
P	18.9	31	$\frac{18.9}{31} = 0.6$	$\frac{0.6}{0.6} = 1$
O	39%	16	$\frac{39}{16} = 2.4$	$\frac{2.4}{0.6} = 4$

Empirical formula $= Na_3PO_4$

Q. 56. A compound contains 87.5% by mass of nitrogen and 12.5% by mass of hydrogen. Determine the empirical formula of this compound.

Sol.

Element	Percentage	Relative No. of Atoms	Simple Ratio
N	87.5	$\frac{87.5}{14} = 6.25$	1
H	12.5	$\frac{12.5}{1} = 12.5$	2

Hence, empirical formula = NH_2.

Q. 57. A compound X consists of 4.8% carbon and 95.2% bromine by mass.

(i) Determine the empirical formula of this compound working correct to one decimal place (C = 12; Br = 80).

(ii) If the vapour density of the compound is 252, what is the molecular formula of the compound ?

(iii) Name the type of chemical reaction by which X can be prepared from ethane.

Sol. (i) Empirical formula :

Elements	% composition	At. mass	Relative number of atoms	Simplest Ratio
Carbon	4.8	12	$\frac{4.8}{12} = 0.4$	$\frac{0.4}{0.4} = 1$
Bromine	95.2	80	$\frac{95.2}{80} = 1.19$	$\frac{1.19}{0.4} = 2.975$

Empirical fromula of substance is CBr_3

$$\text{Empirical mass} = 12 + 80 \times 252$$

$$\text{Molecular mass} = 2 \times \text{V.D} = 2 \times 252 = 504$$

$$\therefore \quad n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{504}{252} = 2$$

or $\quad$ Molecular formula $= n \times$ Empirical formula

$$2 \times CBr_3 = C_2Br_6$$

(iii) This substance can be prepared by substitution method.

Q. 58. A gaseous organic compound contains 3.6 g of carbon and 0.8 g of hydrogen. The vapour density of this compound is 22.

(i) Calculate the Empirical formula.

(ii) Calculate the molecular formula of the compound.

(iii) If 4.4 g of the above compound are completely burnt in oxygen, calculate the volume of carbon dioxide formed at S.T.P. [C = 12; H = 1; O = 16]

Sol. (i)

Element	Wt. of atoms	At. wt	Relative No. of atoms	Simple Ratio
C	3.6	12	$\frac{3.6}{12} = 0.3$	3
H	0.8	1	$\frac{0.8}{1} = 0.8$	8

Empirical formula of the compound $= C_3H_8$

(ii) V.D. of the compound = 22

Empirical formula wt. = 36 + 8 = 44

Molecular wt. of compound = 2 × V.D.

$= 2 \times 22 = 44$

$$\text{No. of Molecules} = \frac{\text{Molecular wt.}}{\text{Empirical wt.}} = \frac{44}{44} = 1$$

Molecular formula $= n \times$ Empirical formula

$= 1 \times C_3H_8$

$= C_3H_8$

(iii) $$\underset{44\ g}{C_3H_8} + 5O_2 \longrightarrow \underset{\substack{3 \times 22.4\ dm^3 \text{ at S.T.P.} \\ = 67.2\ dm^3 \text{ at S.T.P.}}}{3CO_2} + 4H_2O$$

44 g of C_3H_8 yields carbon dioxide at S.T.P.

$= 67.2\ dm^3$

$\therefore$ 4.4 g of C_3H_8 will produce carbon dioxide at S.T.P.

$$= \frac{67.2 \times 4.4}{44} = 6.72\ dm^3 \text{ at S.T.P.}$$

Q. 59. (i) A solid organic compound contained 2.15% of hydrogen, 26.8% of carbon, and the rest of oxygen. Take the molecular weight of the compound as 90. Find the empirical and molecular formula of the compound.

(ii) A metal M forms a volatile chloride containing 65.5% chlorine. If the density of the chloride relative to hydrogen is 162.5, find the molecular formula of the chloride.

(M = 56; Cl = 35.5).

Sol. (i) Percentage of oxygen (by difference) $= 100 - (26.8 + 2.15)$

$= 100 - 28.95 = 71.05\%$

Relative number of atoms	Simplest ratio
$C = \frac{26.8}{12} = 2.23$	$C = \frac{2.23}{2.15} = 1$
$H = \frac{2.15}{1} = 2.15$	$H = \frac{2.15}{2.15} = 1$
$O = \frac{71.05}{16} = 4.44$	$O = \frac{4.44}{2.15} = 2$

Empirical formula of the compound is CHO_2

Empirical formula weight $= 12 + 1 + 2 \times 16$

$= 12 + 1 + 32 = 45$

Given, that the molecular weight $= 90$

$$\therefore \quad n = \frac{\text{Relative molecular weight}}{\text{Empirical formula weight}} = \frac{90}{45} = 2$$

Molecular formula $= (CHO_2) \times 2$

Therefore, molecular formula of the compound is $C_2H_2O_4$.

(ii) % of Metal $= 100 - 65.5 = 34.5$

Chlorine % $= 65.5$

Relative number of atoms	Simplest ratio
$\text{Metal} = \frac{34.5}{56} = 0.616$	$\frac{0.616}{0.616} = 1$
$\text{Chlorine} = \frac{65.5}{35.5} = 1.85$	$\frac{1.85}{0.616} = 3$

$\therefore$ Empirical formula $= MCl_3$

Molecular mass $= 2 \times$ V.D.

$= 2 \times 162.5 = 325.0$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula weight}} = \frac{325}{162.5} = 2$$

$\therefore$ Molecular formula $= (\text{Empirical formula})_n$

$= (MCl_3)_2 = M_2Cl_6$

Q. 60. (a) (i) Propane burns in air according to the following equation :

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

What volume of propane is consumed on using 1000 cm^3 of air, considering only 20% of air contains oxygen ?

(ii) The mass of 11.2 litres of a certain gas at S.T.P. is 24 g. Find the gram molecular mass of the gas.

(b) A gas cylinder can hold 1 kg of hydrogen at room temperature and pressure :

(i) Find the number of moles of hydrogen present.

(ii) What weight of CO_2 can the cylinder hold under similar conditions of temperature and pressure ? (H = 1, C = 12, O = 16)

(iii) If the number of molecules of hydrogen in the cylinder is X, calculate the number of CO_2 molecules in the cylinder under the same conditions of temperature and pressure.

(iv) State the law that helped you to arrive at the above result.

Sol. (a) (i) For every 5 moles of O_2, 1 mole of propane is burnt.

20% of 1000 = 20 × 1000/100 = 200 cm^3 of O_2.

Thus, volume of propane = 40 cm^3

(ii) Mass of gas = 24 g

Volume of gas = 11.6 litres

22.4 L of gas at S.T.P. = 1 mole

11.2 L of gas at S.T.P. = 11.2 / 22.4 = 0.5 moles

Mass of 0.5 moles of gas = 24 g

Mass of 1 mole of gas or molar mass = 24/0.5 = 48 g

(b) (i) 1 kg = 1000 grams

2 g of hydrogen molecules = 1 mole

1 g of hydrogen molecules = 1/2 mole

1000 g of hydrogen molecules = 1/2 × 1000 = 500 moles

(ii) Molecular weight of carbon dioxide = 44 g

$$\text{Vapour density} = \frac{44}{2} = 22$$

Now, $$\text{Vapour density} = \frac{\text{Weight of carbon dioxide at certain temperature}}{\text{Weight of same volume of hydrogen at same temperature and pressure}}$$

$$22 = \text{Weight of carbon dioxide} / 1 \text{ kg}$$

$$\text{Weight of carbon dioxide} = 22 \text{ kg.}$$

(iii) If the number of molecules of hydrogen is X, then number of molecules of carbon dioxide will also be X.

(iv) This is according to the **Avogadro's law** which states that, "Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules".

Q. 61. (a) A gas cylinder contains 12×10^{24} molecules of oxygen gas.

If Avogadro's number is 6×10^{23}. Calculate :

(i) The mass of oxygen present in the cylinder.

(ii) The volume of oxygen at S.T.P. present in the cylinder. [O = 16]

(b) A gaseous hydrocarbon contains 82.76% of carbon. Given that its vapour density is 29, find its molecular formula. [C = 12, H = 1]

(c) The equation $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$, represents the catalytic oxidation of ammonia. If 100 cm^3 of ammonia is used calculate the volume of oxygen required to oxidise the ammonia completely.

Sol. (a) (i) 12×10^{24} molecules of O_2

$$\text{Number of mole} = \frac{12 \times 10^{24}}{6 \times 10^{23}} = 20 \text{ mole}$$

$\because$ 1 mole of oxygen has the atomic weight $\longrightarrow$ 32 g

$\therefore$ 20 moles of oxygen have the atomic weight = 32 × 20 = 640 g

(ii) The volume of one mole gas at S.T.P. = 22.4 lit

$\therefore$ 20 mole of gas at S.T.P. will have the volume of oxygen

$$= 20 \times 22.4 \text{ lit}$$

$$= 448 \text{ lit}$$

(b)

Element	Percentage	Molecules	Simple ratio	Simple whole ratio
C	82.76	$\frac{82.76}{12} = 6.89$	1	2
H	17.24	$\frac{17.24}{1} = 17.24$	2.5	5

$\therefore$ Empirical formula $= C_2H_5$

Empirical formula mass $= (12 \times 2) + (1 \times 5)$

$= 24 + 5 = 29$

Vapour density $= 29$ (Given)

Molecular mass = V.D. $\times$ 2 $= 29 \times 2$

$= 58$ gm

Molecular formula mass $= n \times$ Empirical formula mass

$$\Rightarrow \quad n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$

$$= \frac{58}{29} = 2$$

Molecular formula $= n \times$ Empirical formula

$= 2 \times C_2H_5$

$= C_4H_{10}$

(c) $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$

Given : Ammonia used in the reaction = 100 cm^3

From the equation, 4 vol. of NH_3 requires 5 vol. of O_2 for its oxidation.

$\therefore$ 1 vol. will require $= \frac{5}{4}$

Thus 100 cm^3 of ammonia will require $= \frac{5}{4} \times 100$

$= 125$ cm^3 of oxygen

Q. 62. Consider the following reaction and based on the reaction answer the questions that follow :

$$(NH_4)\,Cr_2O_7 \xrightarrow{\text{Heat}} N_2(g) + 4H_2O(g) + Cr_2O_3$$

Calculate :

(i) The quantity in moles of $(NH_4)_2Cr_2O_7$ if 63 g of $(NH_4)_2Cr_2O_7$ is heated.

(ii) The quantity in moles of nitrogen formed.

(iii) The volume in litres or dm^3 of N_2 evolved at S.T.P.

(iv) The mass in grams of Cr_2O_3 formed at the same time.

[Atomic masses : H = 1, Cr = 52, N = 14]

Sol.

$$\underset{\substack{2\,(14+4)+(52\times 2)+(16\times 7) \\ = 36+104+112 = 252\text{ g}}}{(NH_4)Cr_2O_7} \xrightarrow{\text{heat}} N_2(g) + 4H_2O(g) + \underset{\substack{(2\times 52)+(16\times 3) \\ = 104+48 = 152}}{Cr_2O_3}$$

(i) 252 g $(NH_4)_2Cr_2O_7$ = 1 mole

$\therefore$ 63 g $(NH_4)_2Cr_2O_7 = \frac{63}{252} = 0.25$ mole

Hence, 0.25 mole of $(NH_4)_2Cr_2O_7$ is heated.

(ii) From the chemical equation

1 mole of $(NH_4)_2Cr_2O_7$ liberates 1 mole of N_2

$\therefore$ 0.25 mole of $(NH_4)_2Cr_2O_7$ liberates 0.25 moles of N_2.

(iii) Volume of 1 mole of N_2 at S.T.P. is 22.4 l.

$\therefore$ 0.25 mole of N_2 at S.T.P. has volume = 22.4×0.25 = 5.6 lit

(iv) 252 g $(NH_4)_2Cr_2O_7$ gives 152 g Cr_2O_3

$\therefore$ 63 g $(NH_4)_2Cr_2O_7$ gives $= \frac{152 \times 63}{252}$ g Cr_2O_3 = 38 g Cr_2O_3

Hence, the mass of Cr_2O_3 formed is 38 g.

❐