

HINTS & SOLUTIONS

Some basic concept of Chemistry

[Set-1]

SECTION-A

1. (i): Molality is defined as the number of moles of solute per kilogram (or 1000 g) of solvent

2. (ii): $n = \frac{\text{given volume}}{\text{molar volume}} = \frac{\text{mass in grams}}{\text{molar mass}}$

$$\text{molar mass} = \frac{1.1 \times 22400}{612.5} \text{ g/mol}$$

$$\text{molar mass} = 40.23 \text{ g/mol}$$

3. (ii): $1 \text{ g N}^{3-} = \frac{1}{14} \text{ mol N}^{3-} = \frac{1}{14} \times 10 \text{ mol e}^-$

$$1 \text{ g C}^{4-} = \frac{1}{12} \text{ mol C}^{4-} = \frac{1}{12} \times 10 \text{ mol e}^-$$

$$1 \text{ g Na}^+ = \frac{1}{23} \text{ mol Na}^+ = \frac{1}{23} \times 10 \text{ mol e}^-$$

4. (ii): $n_{\text{H}_2\text{O}} = \frac{90}{18} = 5.0 \text{ mol}$

$$1 \text{ mol H}_2\text{O} = 1 \times 2 + 1 \times 8 = 10 \text{ mol protons}$$

$$= 10 N_A$$

$$5 \text{ mol H}_2\text{O} = 10 \times 5 N_A = 50 N_A$$

5. (iv): $\text{Molarity} = \frac{n}{V} = \text{mol/dm}^3$

$$\text{Molality} = \frac{n}{w_A} = \text{mol/kg}$$

$$\text{Volume percent} = \frac{V_i}{V_{\text{total}}} \times 100 = \text{Unitless}$$

but changes with temperature

$$\text{mass percent} = \frac{w_i}{w_{\text{total}}} \times 100 = \text{Unitless}$$

but unaffected by temperature.

6. (i): Atomic mass of hydrogen = 1 u

Atomic mass of aluminium = 27 u

Aluminium is 27 times heavier than hydrogen

7. (iv): In combustion reaction, oxygen is used in excess.

8. (iv): $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \longrightarrow 2\text{NH}_3(\text{g})$

2 mol NH_3 is produced from 3 mol H_2

$\therefore$ 20 moles of NH_3 required 30 mol H_2 .

9. (iii): $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \longrightarrow 2\text{NH}_3(\text{g})$
 $\begin{matrix} 50\text{ml} & 20\text{ml} \end{matrix}$

1 vol N_2 requires 3 vol H_2 and 50 vol N_2 will require 150 vol H_2 but only 20 vol of H_2 given. Hence H_2 is the limiting reactant.

Volume of N_2 used = $\frac{1}{3} \times 20 = 6.67$ ml

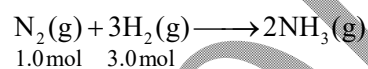
Volume of N_2 left = $50 - 6.67 = 43.33$ ml

Volume of NH_3 produced = $2 \times 6.67 = 13.34$ ml

Total volume of gases left = $43.33 + 13.34 = 56.67$

≈ 57 ml

10. (iii): $n_{\text{H}_2} = \frac{6.02 \times 10^{23}}{6.02 \times 10^{23}} = 1.0 \text{ mol}$, $n_{\text{H}_2} = \frac{6.0}{2} = 3.0 \text{ mol}$



n_{NH_3} produced = 2.0 mol

No. of molecules of NH_3 produced = $2N_A$.

11. Fe = 70 g in 100 g oxide

O = 30 g in 100 g oxide

$$n_{\text{Fe}} : n_{\text{O}} \text{ is } \frac{70}{56} : \frac{30}{16}$$

$$= 1.25 : 1.875$$

$$= \frac{1.25}{1.25} : \frac{1.875}{1.25}$$

$$= 1 : 1.5$$

$$= 2 : 3$$

Empirical formula of oxide is Fe_2O_3 .

12. 100 g chlorophyll contains 18 g Mg

$$\therefore 180 \text{ g chlorophyll contain } \frac{18}{100} \times 180 = \frac{18 \times 18}{10} \text{ g Mg}$$

$$n_{\text{Mg}} = \frac{18 \times 18}{10 \times 24} = 1.35 \text{ mol Mg}$$

$$\begin{aligned} \text{Number of Mg atoms} &= 1.35 \times 6.02 \times 10^{23} \\ &= 8.127 \times 10^{23} \end{aligned}$$

OR

$$V = 0.5 \text{ dm}^3, M = 0.05 \text{ M}$$

$$n_{\text{Na}_2\text{CO}_3} = M \times V = 0.05 \times 0.5 \text{ mol} = \frac{w_{\text{Na}_2\text{CO}_3}}{106}$$

$$\begin{aligned} \text{Mass of Na}_2\text{CO}_3 &= 0.05 \times 0.5 \times 106 \\ &= 2.36 \times 10^{-5} \text{ g} \end{aligned}$$

13. $d = 0.793 \text{ kg/L} = 0.793 \times 10^3 \text{ g/L}$

$$M_{\text{meth}} = \frac{0.793 \times 10^3}{32} = 24.78 \text{ M}$$

$$M_1 = 24.78 \text{ M}$$

$$M_2 = 0.25 \text{ M}$$

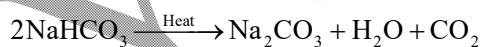
$$V_1 = ?$$

$$V_2 = 2.5 \text{ dm}^3$$

For dilution case $M_1V_1 = M_2V_2$

$$V_1 = 0.25 \times \frac{2.5}{24.78} = 0.025 \text{ dm}^3 = 25 \text{ ml}$$

14. Mass of impure sample = 2.0 g



$$n_{\text{NaHCO}_3} = 2 \times n_{\text{CO}_2} = 2 \times \left(\frac{220}{22400} \right) = 0.0196 \text{ mol}$$

$$w_{\text{NaHCO}_3} = 0.0196 \times 84 = 1.646 \text{ g}$$

$$\% \text{ purity} = \frac{1.646}{2.0} \times 100 = 82.3\%$$

OR

$$C = 24.24\%, H = 4.04\%, Cl = 100 - (24.24 + 4.04) = 71.72\%$$

	C	H	Cl
Mass per 100 g	24.24	4.04	71.72
no. of moles	$\frac{24.24}{12} = 2.02$	$\frac{4.04}{1} = 4.04$	$\frac{71.72}{35.5} = 2.02$
divide by 2.02	$\frac{2.02}{2.02} = 1$	$\frac{4.04}{2.02} = 2$	$\frac{2.02}{2.02} = 1$

Empirical formula = CH_2Cl

Empirical formula mass = $12 + 2 \times 1 + 35.5 = 49.5$

Molecular mass = 98.9

$$\therefore n\text{-factor} = \frac{98.9}{49.5} \approx 2$$

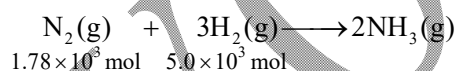
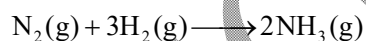
Molecular formula = $(\text{E.F.})_n = (\text{CH}_2\text{Cl})_2$
 $= \text{C}_2\text{H}_4\text{Cl}_2$

15. $n_{\text{N}_2} = \frac{50}{28} \times 10^3$

$$n_{\text{H}_2} = \frac{10}{2} \times 10^3$$

$$n_{\text{N}_2} = 1.78 \times 10^3 \text{ mol}$$

$$n_{\text{H}_2} = 5 \times 10^3$$



1 mol N_2 requires 3 mol H_2 and 1.78×10^3 mol N_2 will require $1.78 \times 10^3 \times 3 = 5.34 \times 10^3$ mol H_2

But only 5.0×10^3 mol H_2 given.

Hence, H_2 is the limiting reactant.

(i) H_2 gas is completely used up in the reaction

(ii) 3 mol H_2 gives 2 mol NH_3 and

$$5.0 \times 10^3 \text{ mol } \text{H}_2 \text{ will liberate } \frac{2 \times 5.0 \times 10^3}{3} \text{ mol } \text{NH}_3$$

$$\text{Volume of NH}_3 \text{ at STP} = n_{\text{N}_2} \times 22.4\text{L}$$

$$= \frac{2 \times 5.0 \times 10^3 \times 22.4}{3} \text{ L}$$

$$= 74.666 \times 10^3 \text{ L}$$

(iii) N_2 is present in excess

$$n_{\text{N}_2} \text{ consumed} = \frac{1}{3} n_{\text{H}_2} = \frac{1}{3} \times 5.0 \times 10^3 = 1.67 \times 10^3 \text{ mol}$$

$$n_{\text{N}_2} \text{ left unused} = 1.78 \times 10^3 - 1.67 \times 10^3 = 0.11 \times 10^3 = 110 \text{ mol}$$

OR

- (i) Law of multiple proportion states that when two elements combine to give two or more compounds, then for a fixed mass of one element, other element bear a simple whole number ratio by mass.
- (ii) Molarity is defined as number of moles of solute per litre of the solution. Molarity of solution changes with change in temperature. Molarity and temperature are inversely related.
- (iii) 0.25 m means 0.25 moles of urea in 1000 g H_2O

$$w_{\text{urea}} = 0.25 \times 60 = 15 \text{ g}$$

$$\text{mass of solution is } 1000 + 15 = 1015 \text{ g}$$

1015 g solution contain 15.0 g urea. Therefore 2.5 kg or 2500 g solution contain

$$\frac{15 \times 2500}{1015} = 36.94 \text{ g urea}$$