

Intext Questions

Question 2.1:

Calculate the mass percentage of benzene (C_6H_6) and carbon tetrachloride (CCl_4) if 22 g of benzene is dissolved in 122 g of carbon tetrachloride.

Answer 2.1:

$$\text{Mass percentage of } C_6H_6 = \frac{\text{Mass of } C_6H_6}{\text{Total mass of the solutions}} \times 100 \%$$

$$= \frac{\text{Mass of } C_6H_6}{\text{Mass of } C_6H_6 + \text{Mass of } CCl_4} \times 100 \% = \frac{22}{22 + 122} \times 100 \% = 15.28 \%$$

$$\text{Mass percentage of } CCl_4 = \frac{\text{Mass of } CCl_4}{\text{Total mass of the solutions}} \times 100 \%$$

$$= \frac{\text{Mass of } CCl_4}{\text{Mass of } C_6H_6 + \text{Mass of } CCl_4} \times 100 \% = \frac{122}{22 + 122} \times 100 \% = 84.72 \%$$

Alternatively, Mass percentage of $CCl_4 = (100 - 15.28) \% = 84.72\%$

Question 2.2:

Calculate the mole fraction of benzene in solution containing 30% by mass in carbon tetrachloride.

Answer 2.2:

Let the total mass of the solution be 100 g and the mass of benzene be 30 g.

\therefore Mass of carbon tetrachloride = $(100 - 30) \text{ g} = 70 \text{ g}$

Molar mass of benzene (C_6H_6) = $(6 \times 12 + 6 \times 1) \text{ g mol}^{-1} = 78 \text{ g mol}^{-1}$

\therefore Number of moles of $C_6H_6 = \frac{30}{78} \text{ mol} = 0.3846 \text{ mol}$

Molar mass of carbon tetrachloride (CCl_4) = $1 \times 12 + 4 \times 35.5 = 154 \text{ g mol}^{-1}$

\therefore Number of moles of $CCl_4 = \frac{70}{154} \text{ mol} = 0.4545 \text{ mol}$

Thus, the mole fraction of C_6H_6 is given as:

$$\frac{\text{Number of moles of } C_6H_6}{\text{Number of moles of } C_6H_6 + \text{Number of moles of } CCl_4} = \frac{0.3846}{0.3846 + 0.4545} = 0.458$$

Question 2.3:

Calculate the molarity of each of the following solutions: (a) 30 g of $Co(NO_3)_2 \cdot 6H_2O$ in 4.3 L of solution (b) 30 mL of 0.5 M H_2SO_4 diluted to 500 mL.

Answer 2.3:

Molarity is given by: $\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litre}}$

(a) Molar mass of $Co(NO_3)_2 \cdot 6H_2O = 59 + 2(14 + 3 \times 16) + 6 \times 18 = 291 \text{ g mol}^{-1}$

\therefore Moles of $Co(NO_3)_2 \cdot 6H_2O = \frac{30}{291} \text{ mol} = 0.103 \text{ mol}$

Therefore, molarity = $\frac{0.103 \text{ mol}}{4.3 \text{ L}} = 0.023 \text{ M}$

(b) Number of moles present in 1000 mL of 0.5 M $H_2SO_4 = 0.5 \text{ mol}$

\therefore Number of moles present in 30 mL of 0.5 M $H_2SO_4 = \frac{0.5 \times 30}{1000} \text{ mol} = 0.015 \text{ mol}$

Therefore, molarity = $\frac{0.015 \text{ mol}}{0.5 \text{ L}} = 0.03 \text{ M}$

Question 2.4:

Calculate the mass of urea (NH_2CONH_2) required in making 2.5 kg of 0.25 molal aqueous solution.

Answer 2.4:

Molar mass of urea (NH_2CONH_2) = $2(1 \times 14 + 2 \times 1) + 1 \times 12 + 1 \times 16 = 60 \text{ g mol}^{-1}$

0.25 molar aqueous solution of urea means: 1000 g of water contains 0.25 mol = $(0.25 \times 60) \text{ g}$ of urea = 15 g of urea

That is,

$(1000 + 15) \text{ g}$ of solution contains 15 g of urea

Therefore, 2.5 kg (2500 g) of solution contains = $\frac{15 \times 2500}{1000 + 15} \text{ g} = 36.95 \text{ g} \approx 37 \text{ g}$ of urea (approximately)

Hence, mass of urea required = 37 g

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(Chapter 2)(Solutions)

(Class 12)

Question 2.5:

Calculate (a) molality (b) molarity and (c) mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g mL⁻¹.

Answer 2.5:

(a) Molar mass of KI = 39 + 127 = 166 g mol⁻¹

20% (mass/mass) aqueous solution of KI means 20 g of KI is present in 100 g of solution.

That is,

20 g of KI is present in (100 - 20) g of water = 80 g of water

Therefore, molality of the solution = $\frac{\text{Moles of KI}}{\text{Mass of water in kg}} = \frac{\frac{20}{166}}{0.08} \text{ m} = 1.506 \text{ m} = 1.51 \text{ m}$ (approximately)

(b) It is given that the density of the solution = 1.202 g mL⁻¹

∴ Volume of 100 g solution = $\frac{\text{Mass}}{\text{Density}} = \frac{100 \text{ g}}{1.202 \text{ mL}^{-1}} = 83.19 \text{ mL} = 83.19 \times 10^{-3} \text{ L}$

Therefore, molarity of the solution = $\frac{\frac{20}{166} \text{ mol}}{83.19 \times 10^{-3} \text{ L}} = 1.45 \text{ M}$

(c) Moles of KI = $\frac{20}{166} = 0.12 \text{ mol}$

Moles of water = $\frac{80}{18} = 4.44 \text{ mol}$

Therefore, mole fraction of KI = $\frac{\text{Moles of KI}}{\text{Moles of KI} + \text{Moles of water}} = \frac{0.12}{0.12 + 4.44} = 0.0263$

Question 2.6:

H₂S, a toxic gas with rotten egg like smell, is used for the qualitative analysis. If the solubility of H₂S in water at STP is 0.195 m, calculate Henry's law constant.

Answer 2.6:

It is given that the solubility of H₂S in water at STP is 0.195 m, i.e., 0.195 mol of H₂S is dissolved in 1000 g of water.

Moles of water = $\frac{1000 \text{ g}}{18 \text{ g mol}^{-1}} = 55.56 \text{ mol}$

∴ Mole fraction of H₂S, $x = \frac{\text{Moles of H}_2\text{S}}{\text{Moles of H}_2\text{S} + \text{Moles of water}} = \frac{0.195}{0.195 + 55.56} = 0.0035$

At STP, pressure (p) = 0.987 bar According to Henry's law: $p = K_H x$

⇒ $K_H = \frac{p}{x} = \frac{0.987}{0.0035} \text{ bar} = 282 \text{ bar}$

Question 2.7:

Henry's law constant for CO₂ in water is 1.67 × 10⁸ Pa at 298 K. Calculate the quantity of CO₂ in 500 mL of soda water when packed under 2.5 atm CO₂ pressure at 298 K.

Answer 2.7:

It is given that: $K_H = 1.67 \times 10^8 \text{ Pa}$

$P_{\text{CO}_2} = 2.5 \text{ atm} = 2.5 \times 1.01325 \times 10^5 \text{ Pa} = 2.533125 \times 10^5 \text{ Pa}$

According to Henry's law: $P_{\text{CO}_2} = K_H x$

⇒ $x = \frac{P_{\text{CO}_2}}{K_H} = \frac{2.533125 \times 10^5}{1.67 \times 10^8} = 0.00152$

We can write, $x = \frac{n_{\text{CO}_2}}{n_{\text{CO}_2} + n_{\text{H}_2\text{O}}} \approx \frac{n_{\text{CO}_2}}{n_{\text{H}_2\text{O}}}$

[Since, n_{CO_2} is negligible as compared to $n_{\text{H}_2\text{O}}$]

In 500 mL of soda water, the volume of water = 500 mL

[Neglecting the amount of soda present]

We can write:

500 mL of water = 500 g of water = $\frac{500}{18}$ mole of water = 27.78 mole of water

Now, $x = \frac{n_{\text{CO}_2}}{n_{\text{H}_2\text{O}}} \Rightarrow 0.00152 = \frac{n_{\text{CO}_2}}{27.78}$

⇒ $n_{\text{CO}_2} = 0.00152 \times 27.78 = 0.042 \text{ mol}$

Hence, quantity of CO₂ in 500 mL of soda water = (0.042 × 44) g = 1.848 g

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Question 2.8:

The vapour pressure of pure liquids A and B are 450 and 700 mm Hg respectively, at 350 K. Find out the composition of the liquid mixture if total vapour pressure is 600 mm Hg. Also find the composition of the vapour phase.

Answer 2.8:

It is given that:

$$P_A^0 = 450 \text{ mm of Hg}$$

$$P_B^0 = 700 \text{ mm of Hg}$$

$$p_{\text{total}} = 600 \text{ mm of Hg}$$

From Raoult's law, we have:

$$P_A = P_A^0 x_A \text{ and } P_B = P_B^0 x_B = P_B^0 (1 - x_A)$$

$$\text{Therefore, total pressure, } P_{\text{total}} = P_A + P_B$$

$$\Rightarrow P_{\text{total}} = P_A^0 x_A + P_B^0 (1 - x_A)$$

$$\Rightarrow P_{\text{total}} = P_A^0 x_A + P_B^0 - P_B^0 x_A$$

$$\Rightarrow P_{\text{total}} = (P_A^0 - P_B^0) x_A + P_B^0$$

$$\Rightarrow 600 = (450 - 700) x_A + 700$$

$$\Rightarrow -100 = -250 x_A$$

$$\Rightarrow x_A = 0.4$$

$$\text{Therefore, } x_B = 1 - x_A = 1 - 0.4 = 0.6$$

$$\text{Now, } P_A = P_A^0 x_A = 450 \times 0.4 = 180 \text{ mm of Hg and}$$

$$P_B = P_B^0 x_B = 700 \times 0.6 = 420 \text{ mm of Hg}$$

$$\text{Now, in the vapour phase: Mole fraction of liquid A} = \frac{P_A}{P_A + P_B}$$

$$= \frac{180}{180 + 420} = \frac{180}{600} = 0.30$$

$$\text{And, mole fraction of liquid B} = 1 - 0.30 = 0.70$$

Question 2.9:

Vapour pressure of pure water at 298 K is 23.8 mm Hg. 50 g of urea (NH_2CONH_2) is dissolved in 850 g of water. Calculate the vapour pressure of water for this solution and its relative lowering.

Answer 2.9:

It is given that vapour pressure of water, $P_1^0 = 23.8$ mm of Hg

Weight of water taken, $w_1 = 850$ g

Weight of urea taken, $w_2 = 50$ g

Molecular weight of water, $M_1 = 18$ g mol⁻¹

Molecular weight of urea, $M_2 = 60$ g mol⁻¹

Now, we have to calculate vapour pressure of water in the solution. We take vapour pressure as p_1 .

Now, from Raoult's law, we have:

$$\frac{P_1^0 - p_1}{P_1^0} = \frac{n_2}{n_1 + n_2}$$

$$\Rightarrow \frac{P_1^0 - p_1}{P_1^0} = \frac{\frac{w_2}{M_2}}{\frac{w_1}{M_1} + \frac{w_2}{M_2}}$$

$$\Rightarrow \frac{23.8 - p_1}{23.8} = \frac{\frac{50}{60}}{\frac{850}{18} + \frac{50}{60}}$$

$$\Rightarrow \frac{23.8 - p_1}{23.8} = \frac{0.83}{47.22 + 0.83}$$

$$\Rightarrow \frac{23.8 - p_1}{23.8} = 0.0173$$

$$\Rightarrow p_1 = 23.4 \text{ mm of Hg}$$

Hence, the vapour pressure of water in the given solution is 23.4 mm of Hg and its relative lowering is 0.0173.

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(Chapter 2)(Solutions)

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Question 2.10:

Boiling point of water at 750 mm Hg is 99.63°C. How much sucrose is to be added to 500 g of water such that it boils at 100°C. Molal elevation constant for water is 0.52 K kg mol⁻¹.

Answer 2.10:

Here, elevation of boiling point $\Delta T_b = (100 + 273) - (99.63 + 273) = 0.37$ K

Mass of water, $w_1 = 500$ g

Molar mass of sucrose (C₁₂H₂₂O₁₁), $M_2 = 11 \times 12 + 22 \times 1 + 11 \times 16 = 342$ g mol⁻¹

Molal elevation constant, $K_b = 0.52$ K kg mol⁻¹

We know that:

$$\Delta T_b = \frac{K_b \times 1000 \times w_2}{M_2 \times w_1}$$

$$\Rightarrow w_2 = \frac{\Delta T_b \times M_2 \times w_1}{K_b \times 1000}$$

$$= \frac{0.37 \times 342 \times 500}{0.52 \times 1000}$$

= 121.67 g (approximately)

Hence, 121.67 g of sucrose is to be added.

Question 2.11:

Calculate the mass of ascorbic acid (Vitamin C, C₆H₈O₆) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5°C. $K_f = 3.9$ K kg mol⁻¹.

Answer 2.11:

Mass of acetic acid, $w_1 = 75$ g

Molar mass of ascorbic acid (C₆H₈O₆), $M_2 = 6 \times 12 + 8 \times 1 + 6 \times 16 = 176$ g mol⁻¹

Lowering of melting point, $\Delta T_f = 1.5$ K

We know that:

$$\Delta T_f = \frac{K_f \times w_2 \times 1000}{M_2 \times w_1}$$

$$\Rightarrow w_2 = \frac{\Delta T_f \times M_2 \times w_1}{K_f \times 1000}$$

$$= \frac{1.5 \times 176 \times 75}{3.9 \times 1000}$$

= 5.08 g (approx)

Hence, 5.08 g of ascorbic acid is needed to be dissolved.

Question 2.12:

Calculate the osmotic pressure in Pascals exerted by a solution prepared by dissolving 1.0 g of polymer of molar mass 185,000 in 450 mL of water at 37°C.

Answer 2.12:

It is given that:

Volume of water, $V = 450$ mL = 0.45 L

Temperature, $T = (37 + 273)$ K = 310 K

Number of moles of the polymer, $n = \frac{1}{185000}$ mol

We know that: $\pi = \frac{n}{V} RT$

Osmotic pressure, $= \frac{1}{185000} \text{ mol} \times \frac{1}{0.45 \text{ L}} \times 8.314 \times 10^3 \text{ Pa L K}^{-1} \times 310 \text{ K}$

= 30.98 Pa

= 31 Pa (approximately)