2. Solutions

❖ Intext Questions

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.

Ans: Mass of solution = Mass of C_6H_6 + Mass of CCl_4
= 22 g+122 g= 144 g
Mass % of benzene = 22/144 x 100 =15.28 %
Mass % of CCl_4 = 122/144 x 100 = 84.72 %

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

Ans: 30% by mass of C_6H_6 in CCl_4 => 30 g C_6H_6 in 100 g solution
.'. no. of moles of C_6H_6,($n_{C_6h_6}$) = 30/78 = 0.385

(molar mass of C_6H_6 = 78g)

no. of moles of

$$CCl_4\,(n_{CCl_4}) = \frac{70}{154} = 0.455$$

$$\therefore x_{C_6H_6} = \frac{n_{C_6H_6}}{n_{C_6H_6} + n_{CCl_4}}$$

$$= \frac{0.385}{0.385+0.455} = \frac{0.385}{0.84} = 0.458$$

$$x_{CCl_4} = 1 - 0.458 = 0.542$$

2.3. Calculate the molarity of each of the following solutions
(a) 30 g of $Co(NO_3)_2 6H_2O$ in 4·3 L of solution
(b) 30 mL of 0-5 M H_2SO_4 diluted to 500 mL.

Ans:

(a) $$\text{Molarity of solution} = \frac{\text{Mass of solute / Molar mass of solute}}{\text{Volume of solution in litres}}$$

Mass of solute, $Co(NO_3)_2.6H_2O$ = 30 g.

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

Volume of solution = 4·3 L

$$\text{Molarity (M)} = \frac{(30\text{g})/(291\text{g mol}^{-1})}{(4{\cdot}3\text{L})} = 0{\cdot}024 \text{ mol L}^{-1} = \mathbf{0{\cdot}024\ M}$$

(b) Volume of undiluted H_2SO_4 solution (V_1) = 30 mL

Molarity of undiluted H_2SO_4 solution (M_1) = 0·5 M

Volume of diluted H_2SO_4 solution (V_2) = 500 mL

Molarity of diluted H_2SO_4 (M_2) can be calculated as :

$$M_1V_1 = M_2V_2$$

or $$M_2 = \frac{M_1V_1}{V_2} = \frac{(30\text{mL}) \times (0{\cdot}5\text{M})}{(500\text{mL})} = \mathbf{0{\cdot}03\ M}$$

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

Ans: 0.25 Molal aqueous solution to urea means that
moles of urea = 0.25 mole
mass of solvent (NH_2CONH_2) = 60 g mol^{-1}
.’. 0.25 mole of urea = 0.25 x 60=15g
Mass of solution = 1000+15 = 1015g = 1.015 kg
1.015 kg of urea solution contains 15g of urea
.’. 2.5 kg of solution contains urea =15/1.015 x 2.5 = 37 g

2.5. Calculate
(a) molality
(b) molarity and
(c) mole fraction of KI if the density of 20% (mass/mass) aqueous KI solution is 1·202 g mL^{-1}.

Ans:

Step I. Calculation of molality of solution

Weight of KI in 100 g of the solution = 20 g
Weight of water in the solution = 100 – 20 = 80 g = 0-08 kg
Molar mass of KI = 39 + 127 = 166 g mol^{-1}.

$$\text{Molality of solution } (m) = \frac{\text{No of gram moles of KI}}{\text{Mass of water in kg}} = \frac{(20\text{g})/(166\text{g mol}^{-1})}{(0{\cdot}08\text{kg})}$$

$$= 1{\cdot}506 \text{ mol kg}^{-1} = \mathbf{1{\cdot}506\ m.}$$

Step II. Calculation of molarity of solution

Weight of solution $= 100$ g ; Density of solution $= 1{\cdot}202$ g mL^{-1}.

$$\text{Volume of solution} = \frac{\text{Weight of solution}}{\text{Density}} = \frac{(100\,\text{g})}{(1{\cdot}202\,\text{g mL}^{-1})} = 83{\cdot}19\text{ mL}$$

$$\text{Molarity of solution (M)} = \frac{\text{No. of gram moles of KI}}{\text{Volume of solution in litres}} = \frac{(20\text{g})/(166\text{g mol}^{-1})}{(0{\cdot}083\text{L})}$$

$$= 1{\cdot}45\text{ mol L}^{-1} = \mathbf{1{\cdot}45\ M}$$

Step III. Calculation of mole fraction of Kl

$$n_{KI} = \frac{\text{Mass of KI}}{\text{Molar mass of KI}} = \frac{(20\text{g})}{(166\text{g mol}^{-1})} = 0{\cdot}12\text{ mol}$$

$$n_{H_2O} = \frac{\text{Mass of water}}{\text{Molar mass of water}} = \frac{(80\text{g})}{(18\text{g mol}^{-1})} = 4{\cdot}44\text{ mol.}$$

$$x_{KI} = \frac{n_{KI}}{n_{KI} + n_{H_2O}} = \frac{(0{\cdot}12\text{mol})}{(0{\cdot}12 + 4{\cdot}44)\text{mol}} = \frac{0{\cdot}12}{4{\cdot}56} = \mathbf{0{\cdot}0263.}$$

2.6. H_2S, a toxic gas with rotten egg like smell, is used for the qualitative analysis. If the solubility of H_2S in water at STP is 0.195 m, calculate Henry's law constant.

Ans: Solubility of H_2S gas = 0.195 m

= 0.195 mole in 1 kg of solvent

1 kg of solvent = 1000g

$$= \frac{1000}{18} = 55.55\text{ moles}$$

$$\therefore\ x_{H_2S} = \frac{0.195}{0.195 + 55.55}$$

$$= \frac{0.195}{55.745} = 0.0035$$

– Pressure at STP = 0.987 bar

Applying Henry's law,

$$P_{H_2S} = K_H \times x_{H_2S}$$

$$\Rightarrow \quad K_H = \frac{P_{H_2S}}{x_{H_2S}} = \frac{0.987}{0.0035} = 282\text{bar}$$

2.7. Henry's law constant for CO_2 in water is 1.67×10^8 Pa at 298 K. Calculate the quantity of CO_2 in 500 mL of soda water when packed under 2.5 atm CO_2 pressure at 298 K.

Ans.:

$$K_H = 1.67 \times 10^8 \text{ Pa}$$
$$P_{CO_2} = 2.5 \text{ atm} = 2.5 \times 101325 \text{ Pa}$$
$$\therefore\ x_{CO_2} = \frac{P_{CO_2}}{K_H} = \frac{2.5 \times 101325}{1.67 \times 10^8} = 1.517 \times 10^{-3}$$

For 500 mL of soda water, water present $\simeq$ 500 mL

$$= 500\text{g} = \frac{500}{18} = 27.78 \text{ moles}$$
$$\therefore\ n_{H_2O} = 27.78 \text{ moles}$$
$$\therefore\ \frac{n_{CO_2}}{27.78} = 1.517 \times 10^{-3}$$
$$n_{CO_2} = 42.14 \times 10^{-3} \text{ mole}$$
$$= 42.14 \text{ m mol}$$
$$= 42.14 \times 10^{-3} \times 44 \text{ g}$$
$$= 1.854 \text{ g}$$

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

Ans:

Vapour pressure of pure liquid A (P°_A) = 450 mm

Vapour pressure of pure liquid B (P°_B) = 700 mm

Total vapour pressure of the solution (P) = 600 mm

According to Raoult's Law, $P = P_A^\circ x_A + P_B^\circ x_B = P_A^\circ x_A + P_B^\circ(1 - x_A)$

$$(600 \text{ mm}) = 450 \text{ mm} \times x_A + 700 \text{ mm}(1 - x_A)$$
$$= 700 \text{ mm} + x_A(450 - 700) \text{ mm}$$
$$= 700 - x_A(250 \text{ mm})$$

or $$x_A = \frac{(600 - 700) \text{ mm}}{-(250 \text{ mm})} = 0{\cdot}40$$

Mole fraction of A (x_A) = 0·40

Mole fraction of B (x_B) = 1 – 0·40 = 0·60

$$P_A = P_A^\circ x_A = (450 \text{ mm}) \times 0{\cdot}40 = 180 \text{ mm}$$

$$P_B = P_B^\circ x_B = (700 \text{ mm}) \times 0{\cdot}60 = 420 \text{ mm}$$

$$\text{Mole fraction of A in the vapour phase} = \frac{P_A}{P_A + P_B} = \frac{(180 \text{ mm})}{(180 + 420) \text{ mm}} = \mathbf{0{\cdot}30}$$

$$\text{Mole fraction of B in the vapour phase} = \frac{P_B}{P_A + P_B} = \frac{(420 \text{ mm})}{(180 + 420) \text{ mm}} = \mathbf{0{\cdot}70}$$

2.9. Vapour pressure of pure water at 298 K is 23.8 m m 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.

Ans:

$P^o = 23.8$ mm

$w_2 = 50$g, M_2 (urea) = 60g mol^{-1}

$w_1 = 850$ g, M_1 (water)= 18g mol^{-1}

To find: P_s and $(P^o - P_s)/P^o$

Solution: Applying Raoult's law,

$$\frac{P^o - P_s}{P^o} = \frac{n_2}{n_1 + n_2} = \frac{w_2 / M_2}{w_1 / M_1 + w_2 / M_2}$$

$$\therefore \quad \frac{P^o - P_s}{P^o} = \frac{50/60}{850/18 + 50/60}$$

$$= \frac{0.83}{47.22 + 0.83} = 0.017$$

Putting $P^o = 23.8$ mm, we have

$$\frac{23.8 - P_s}{P_s} = 0.017$$

$\Rightarrow \quad 23.8 - P_s = 0.017 P_s$

or, $\quad 1.017 P_s = 23.8$

or, $\quad \boldsymbol{P_s = 23.4}$ **mm**

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.

Ans:

Given $\Delta T_b = 100 - 96.63 = 3.37°$

Mass of water, $w_1 = 500$ g

Molar mass of water, $M_1 = 18 \text{g mol}^{-1}$

Molar mass of sucrose, $M_2 = 342 \text{g mol}^{-1}$

To find: Mass of sucrose, $w_2 = ?$

Solution: We know, $\Delta T_b = K_b \times m$

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

$$\Rightarrow w_2 = \frac{M_2 \times w_1 \times \Delta T_b}{1000 \times K_b} = \frac{342 \times 500 \times 3.37}{1000 \times 0.52}$$

$w_2 = 1108.2$ g

$\therefore$ Mass of solute, $w_2 = 1.11$ kg

2.11 Calculate the mass of ascorbic acid (vitamin C, $C_6H_8O_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1·5°C. (K_f for CH_3COOH) = 3·9 K kg mol^{-1})

Ans:

$$W_B = \frac{M_B \times \Delta T_f \times W_A}{K_f}$$

Mass of acetic acid (W_A) = 75 g = 0·075 kg.

Depression in freeing point (ΔT_f) = 1·5°C = 1·5 K

Molar mass of ascorbic acid (M_B) = 6 × 12 + 8 × 1 + 6 × 16 = 176 g mol^{-1}

Molal depression constant (K_f) = 3·9 K kg mol^{-1}

$$W_B = \frac{(176 \text{g mol}^{-1}) \times (1 \cdot 5 \text{K}) \times (0 \cdot 075 \text{kg})}{(3 \cdot 9 \text{K kg mol}^{-1})} = \mathbf{5 \cdot 08 \text{ g}}$$

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.

Ans:

Given: $V = 450\,mL = 0.45\,L$

$T = 37°C = 310\,K$

$R = 8.314\,kPa\,L\,K^{-1}\,mol^{-1}$

To find: $\pi = ?$

Solution: Applying the formula,

$$\pi = CRT = \frac{n}{V}RT$$

$$n = \frac{1.0g}{185{,}000g\ mol^{-1}}$$

$$\therefore P = \frac{1}{185{,}000} \times \frac{1}{0.45} \times 8.314$$

$$\times 10^3\ Pa\,LK^{-1}\ mol^{-1} \times 310K$$

$$= 30.96\ Pa$$

NCERT EXERCISES

2.1. Define the terra solution. How many types of solutions are formed? Write briefly about each type with an example.

Sol: A solution is a homogeneous mixture of two or more chemically non-reacting substances. Types of solutions: There are nine types of solutions.

Types of Solution Examples

Gaseous solutions

(a) Gas in gas Air, mixture of 0_2 and N_2, etc.

(b) Liquid in gas Water vapour

(c) Solid in gas Camphor vapours in N2 gas, smoke etc.

Liquid solutions

(a) Gas in liquid C02 dissolved in water (aerated water), and 02 dissolved in water, etc.

(b) Liquid in liquid Ethanol dissolved in water, etc.

(c) Solid in liquid Sugar dissolved in water, saline water, etc.

Solid solutions

(a) Gas in solid Solution of hydrogen in palladium

(b) Liquid in solid Amalgams, e.g., Na-Hg

(c) Solid in solid Gold ornaments (Cu/Ag with Au)

2.2. Suppose a solid solution is formed between two substances, one whose particles are very large and the other whose particles are very small. What

type of solid solution is this likely to be ?
Sol: The solution likely to be formed is interstitial solid solution.

2.3 Define the following terms:
(i) Mole fraction
(ii) Molality
(iii) Molarity
(iv) Mass percentage
Sol: (i) Mole fraction: It is defined as the ratio of the number of moles of the solute to the total number of moles in the solution. If A is the number of moles of solute dissolved in B moles of solvent, then Mole fraction of solute

$$(X_A) = \frac{n_A}{n_A + n_B} \qquad \ldots (1)$$

Mole fraction of solvent $(X_B) = \frac{n_B}{n_A + n_B}$ ... (2)

Adding the above two equations, we get

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = \frac{n_A + n_B}{n_A + n_B} = 1$$

i.e., $X_A + X_B = 1$

$\therefore X_A = 1 - X_B$ or $X_B = 1 - X_A$

(ii) Molality: It is defined as die number of moles of a solute present in 1000g (1kg) of a solvent.

$$\text{Molality } (m) = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in kg}} = \frac{n}{W}$$

NOTE: Molality is considered better way of expressing concentration of solutions, as compared to molarity because molality does not change with change in temperature since the mass of solvent does not vary with temperature,
(iii) Molarity: It is defined as the number of moles of solute present in one litre of solution.

Molarity (M) =

$$\frac{\text{Number of moles of solute}}{\text{Volume of Solution in litre}} = \frac{n}{V}$$

$$n = \frac{\text{Weight in grams}}{\text{Molecular weight of solute}}$$

$$\therefore \ M = \frac{\text{Weight in grams}}{\text{Volume of solution in litres}} \times \frac{1}{\text{Molecular weight of solute}}$$

Strength : This is weight (in gms) of solute per litre of solution

$$\therefore \ \text{Molarity} = \frac{\text{Strength}}{\text{Molecular weight of solute}}$$

or Strength = Molarity × Molecular weight

NOTE: Molarity is the most common way of expressing concentration of a solution in laboratory. However, it has one disadvantage. It changes with temperature because volume of a solution alters due to expansion and contraction of the liquid with temperature.

(iv) Mass percentage: It is the amount of solute in grams present in 100g of solution.

$$= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

2.4. Concentrated nitric acid used in the laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of acid if the density of the solution is 1·504 g mL^{-1} ?

Sol: Mass of HNO_3 in solution = 68 g

Molar mass of HNO_3 = 63 g mol^{-1}

Mass of solution = 100 g

Density of solution = 1·504 g mL^{-1}

$$\text{Volume of solution} = \frac{\text{Mass of solution}}{\text{Density of solution}}$$

$$= \frac{(100\,g)}{(1\cdot504\ g\ mL^{-1})} = 66\cdot5\ mL = 0\cdot0665\ L$$

$$\text{Molarity of solution (M)} = \frac{\text{Mass of } HNO_3 \text{ / Molar mass of } HNO_3}{\text{Volume of solution in Litres}}$$

$$= \frac{(68g/63g\,mol^{-1})}{(0\cdot0665\,L)} = 16\cdot23\ mol\ L^{-1} = \mathbf{16\cdot23\ M.}$$

2.5. A solution of glucose in water is labelled as 10% w/w, what would be the molality and mole fraction of each component in the solution? If the density of solution is 1 .2 g m L^{-1}, then what shall be the molarity of the solution?

Sol: 10 percent w/w solution of glucose in water means 10g glucose and 90g of water.

Molar mass of glucose = 180g mol^{-1} and molar mass of water = 18g mol^{-1}

$$\therefore \quad 10g \text{ of glucose} = \frac{10}{180} = 0\cdot0555 \text{ moles}$$

$$\text{and } 90g \text{ of } H_2O = \frac{90}{18} = 5 \text{ moles}$$

$\therefore$ Molality of solution

$$= \frac{\text{Moles of solute} \times 1000}{\text{Mass of solution in grams}}$$

$$= \frac{0\cdot0555}{90} \times 1000 = 0\cdot617\,m$$

Mole fraction of glucose

$$= X_g = \frac{\text{No. of moles of glucose}}{\text{No. of moles of glucose} + \text{No. of moles of water}}$$

$$= \frac{0.0555}{5 + 0.0555} = 0.01$$

Mole fraction of water

$$= X_w = \frac{\text{No. of moles of water}}{\text{No. of moles of glucose} + \text{No. of moles of water}}$$

$$= \frac{5}{5 + 0.0555} = 0.99.$$

Volume of 100g of solution

$$= \frac{\text{Mass of solution}}{\text{Density}} = \frac{100}{1.2} = 83.33 \text{ mL}$$

$$\therefore \text{ Molarity of solution} = \frac{0.0555}{83.33} \times 1000$$

$$= 0.67 \text{ M.}$$

2.6. How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na_2C0_3 and $NaHCO_3$ containing equimolar amounts of both?

Sol: Calculation of no. of moles of components in the mixture.

Let x g of Na_2CO_3 is present in the mixture.

$\therefore$ $(1-x)$ g of $NaHCO_3$ is present in the mixture.

Molar mass of Na_2CO_3

$= 2 \times 23 + 12 + 3 \times 16 = 106 \text{ g mol}^{-1}$

and molar mass of $NaHCO_3$

$= 23 \times 1 + 1 + 12 + 3 \times 16 = 84 \text{ g mol}^{-1}$

No. of moles of Na_2CO_3 in x g $= \dfrac{x}{106}$

No. of moles of $NaHCO_3$ in $(1-x)$ g $= (1-x)/84$

As given that the mixture contains equimolar amounts of Na_2CO_3 and $NaHCO_3$, therefore

$$\frac{x}{106} = \frac{1-x}{84}$$

$$106 - 106x = 84x$$

$$106 = 190x$$

$$\therefore \quad x = \frac{106}{190} = 0{\cdot}558\text{g}$$

$\therefore$ No. of moles of Na_2CO_3 present

$$= \frac{0{\cdot}558}{106} = 0{\cdot}00526$$

and no. of moles of $NaHCO_3$ present

$$= \frac{1 - 0.558}{84} = 0.00526$$

Calculation of no. of moles of HCl required

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2$$

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

As can be seen, each mole of Na_2CO_3 needs 2 moles of HCl,

$\therefore$ 0.00526 mole of Na_2CO_3 needs

$= 0.00526 \times 2 = 0.01052$ mole

Each mole of $NaHCO_3$ needs 1 mole of HCl.

$\therefore$ 0.00526 mole of $NaHCO_3$ needs

$= 1 \times 0.00526 = 0.00526$ mole

Total amount of HCl needed will be

$= 0.01052 + 0.00526 = 0.01578$ mole.

0.1 mole of 0.1 M HCl are present in 1000 mL of HCl

$\therefore$ 0.01578 mole of 0.1 M HCl will be present in

$$= \frac{1000}{0.1} \times 0.01578 = 157.8 \text{ mL.}$$

2.7. Calculate the percentage composition in terms of mass of a solution obtained by mixing 300 g of a 25% and 400 g of a 40% solution by mass.

Sol:

Mass of one component in solution $= \frac{(300\text{ g}) \times 25}{100} = 75\text{g}$

Mass of other component in solution $= \frac{(400\text{ g}) \times 40}{100} = 160\text{g}$

Total mass of solute $= (75 + 160)\text{g} = 235\text{ g}$

Total mass of solution $= (300 + 400)\text{g} = 700\text{ g}$

% of solute in the final solution $= \frac{(235\text{g})}{(700\text{g})} \times 100 = \mathbf{33.57}$

% of solvent in the final solution $= 100 - 33.57 = \mathbf{66.43}$

2.8. An antifreeze solution is prepared from 222.6 g of ethylene glycol, ($C_2H_6O_2$) and200 g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g mL^{-1}, then what shall be the molarity of the

solution?

Sol:

Mass of solute = 222·6g

Molar mass of solute, $C_2H_4(OH)_2$

$= 12 \times 2 + 4 + 2(12 + 1) = 62 \text{ g mol}^{-1}$

$$\therefore \quad \text{Moles of solute} = \frac{222 \cdot 6}{62} = 3 \cdot 59$$

Mass of solvent = 200 g

$$\therefore \quad \text{Molality} = \frac{3 \cdot 59}{200} \times 1000 = 17 \cdot 95 \text{ mol kg}^{-1}$$

Total mass of solution = 422·6 g

$$\text{Volume of solution} = \frac{422 \cdot 6}{1 \cdot 072} = 394 \cdot 21 \text{ mL.}$$

$$\therefore \quad \text{Molarity} = \frac{3 \cdot 59}{394 \cdot 2} \times 1000 = 9 \cdot 1 \text{ mol L}^{-1}$$

2.9. A sample of drinking water was found to be severely contaminated with chloroform ($CHCl_3$), supposed to be a carcinogen. The level of contamination was 15 ppm (by mass).

(i) express this in percent by mass.

(ii) determine the molality of chloroform in the water sample.

Sol: 15 ppm means 15 parts in million (10^6) by mass in the solution.

$$\therefore \text{Percentage by mass} = \frac{15}{10^6} \times 100 = 15 \times 10^{-4}\%$$

As only 15g of chloroform is present in 10^6g of the solution, mass of the solvent = 10^6 g

Molar mass of $CHCl_3 = 12 + 1 + 3 \times 35 \cdot 5$

$= 119 \cdot 5 \text{ g mol}^{-1}$

$$\text{Moles of } CHCl_3 = \frac{15}{119 \cdot 5}$$

$$\therefore \text{Molality} = \frac{15/119 \cdot 5 \times 1000}{10^6} = 1 \cdot 25 \times 10^{-4} \text{ m}$$

2.10. What role does the molecular interaction play in solution of alcohol in water?

Sol: In case of alcohol as well as water, the molecules are interlinked by intermolecular hydrogen bonding. However, the hydrogen bonding is also present in the molecules of alcohol and water in the solution but it is comparatively less than both alcohol and water. As a result, the magnitude of attractive forces tends to decrease and the solution shows positive deviation from Raoult's Law. This will lead to increase in vapour pressure of the solution and also decrease in its boiling point.

2.11. Why do gases always tend to be less soluble in liquids as the temperature is raised?

Sol: When gases are dissolved in water, it is accompanied by a release of heat energy, i.e., process is exothermic. When the temperature is increased, according to Lechatlier's Principle, the equilibrium shifts in backward direction, and thus gases becomes less soluble in liquids.

2.12. State Henry's law and mention some of its important applications.

Sol:

Henry's law: The solubility of a gas in a liquid at a particular temperature is directly proportional to the pressure of the gas in equilibrium with the liquid at that temperature.

or

The partial pressure of a gas in vapour phase is proportional to the mole fraction of the gas (x) in the solution. p = KHX

where KH is Henry's law constant.

Applications of Henry's law :

(i) In order to increase the solubility of CO_2 gas in soft drinks and soda water, the bottles are normally sealed under high pressure. Increase in pressure increases the solubility of a gas in a solvent according to Henry's Law. If the bottle is opened by removing the stopper or seal, the pressure on the surface of the gas will suddenly decrease. This will cause a decrease in the solubility of the gas in the liquid i.e. water. As a result, it will rush out of the bottle producing a hissing noise or with a fiz.

(ii) As pointed above, oxygen to be used by deep sea divers is generally diluted with helium inorder to reduce or minimise the painfril effects during decompression.

(iii) As the partial pressure of oxygen in air is high, in lungs it combines with haemoglobin to form oxyhaemoglobin. In tissues, the partial pressure of oxygen is comparatively low. Therefore, oxyhaemoglobin releases oxygen in order to carry out cellular activities.

2.13. The partial pressure of ethane over a solution containing 6.56×10^{-3} g of ethane is 1 bar. If the solution contains 5.00×10^{-2} g of ethane, then what shall be the partial pressure of the gas?

Sol:

We know that, $m = K_H \times P$

$\therefore \quad 6{\cdot}56 \times 10^{-2}\,\text{g} = K_H \times 1\,\text{bar} \qquad \ldots(i)$

$\therefore \quad 5{\cdot}00 \times 10^{-2}\,\text{g} = K_H \times P \qquad \ldots(ii)$

$K_H = 6.56 \times 10^{-2}/1$ bar (from i)

$K_H = 5.00 \times 10^{-2}/p$ bar (from ii),

$$\therefore \quad \frac{6.56 \times 10^{-2}}{1} = \frac{5.00 \times 10^{-2}}{p}$$

$$\therefore \quad P = \frac{5{\cdot}00}{6{\cdot}56} = 0{\cdot}762\ \text{bar.}$$

2.14. According to Raoult's law, what is meant by positive and negative deviaitions and how is the sign of $\Delta_{sol}H$ related to positive and negative deviations from Raoult's law?

Sol: Solutions having vapour pressures more than that expected from Raoult's law are said to exhibit positive deviation. In these solutions solvent – solute interactions are weaker and $\Delta_{sol}H$ is positive because stronger A – A or B – B interactions are replaced by weaker A – B interactions. Breaking of the stronger interactions requires more energy & less energy is released on formation of weaker interactions. So overall $\Delta_{sol}H$ is positive. Similarly $\Delta_{sol}V$ is positive i.e. the volume of solution is some what more than sum of volumes of solvent and solute.
So there is expansion in volume on solution formation.
Similarly in case of solutions exhibiting negative deviations, A – B interactions are stronger than A-A&B-B. So weaker interactions are replaced by stronger interactions so , there is release of energy i.e. $\Delta_{sol}H$ is negative.

2.15. An aqueous solution of 2 percent non-volatile solute exerts a pressure of 1·004 bar at the boiling point of the solvent. What is the molecular mass of the solute ?

Sol:

According to Raoult's Law,

$$\frac{P_A^\circ - P_S}{P_S} = \frac{n_B}{n_A} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

P_A° (for water) $= 1{\cdot}013$ bar ; $P_S = 1{\cdot}004$ bar ; $W_B = 2g$; $W_A = 100 - 2 = 98$ g ;
$M_A = 18 \text{ g mol}^{-1}$.

$$\frac{(1{\cdot}013 - 1{\cdot}004) \text{ bar}}{(1{\cdot}004 \text{ bar})} = \frac{(2 \text{ g}) \times (18 \text{ g mol}^{-1})}{M_B \times (98 \text{ g})}$$

$$\therefore \quad M_B = \frac{(2 \text{ g}) \times (18 \text{ g mol}^{-1}) \times (1{\cdot}004 \text{ bar})}{(0{\cdot}009 \text{ bar}) \times (98 \text{ g})} = \mathbf{41{\cdot}0 \text{ g mol}^{-1}}$$

2.16 Heptane and octane form an ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35.0 g of octane?

Sol.

Molar mass of heptane (C_7H_{16})

$= 7 \times 12 + 16 = 100 \text{ g mol}^{-1}$

Molar mass of octane (C_8H_{18})

$= 8 \times 12 + 18 = 114 \text{ g mol}^{-1}$

Moles of heptane present in mixture

$$= \frac{26{\cdot}0}{100} = 0{\cdot}26 \text{ mol}$$

Moles of octane present in mixture

$$= \frac{35{\cdot}0}{114} = 0{\cdot}307 \text{ mol}$$

Mole fraction of heptane x_H

$$= \frac{0{\cdot}26}{0{\cdot}26 + 0{\cdot}307} = 0{\cdot}458$$

Mole fraction of octane, $x_O = (1 - 0{\cdot}458) = 0{\cdot}542$

Vapour pressure of heptane $= x_H \times P^\circ$

$= 0.458 \times 105.2 \text{ kPa} = 48.18 \text{ kPa}$

Vapour pressure of octane $= x_O \times P^\circ$

$= 0{\cdot}542 \times 46{\cdot}8 \text{ kPa} = 25{\cdot}36 \text{ kPa}$

Vapour pressure of mixture

$= 48{\cdot}18 + 25{\cdot}36 = 73{\cdot}54 \text{ kPa}$

2.17. The vapour pressure of water is 12.3 kPa at 300 K. Calculate vapour pressure of 1 molal solution of a non-volatile solute in it

Sol: 1 molal solution of solute means 1 mole of solute in 1000g of the solvent.

Molar mass of water (solvent) = 18 g mol^{-1}

$$\therefore \text{Moles of water} = \frac{1000}{18} = 55{\cdot}5 \text{ moles.}$$

$$\therefore \text{Mole fraction of solute} = \frac{1}{1 + 55{\cdot}5} = 0{\cdot}0177$$

$$\text{Now, } \frac{P^\circ - P_s}{P^\circ} = x_2$$

$$\frac{12{\cdot}3 - P_s}{12{\cdot}3} = 0{\cdot}0177$$

$$\Rightarrow P_s = 12{\cdot}08 \text{ kPa}$$

2.18. Calculate the mass of a non-volatile solute (molecular mass 40 g mol^{-1}) that should be dissolved in 114 g of octane to reduce its pressure to 80%. (C.B.S.E. Outside Delhi 2008)

Sol: According to Raoult's Law,

$$\frac{P_A^\circ - P_S}{P_S} = \frac{n_B}{n_A} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

Let $P_A^\circ = 1$ atm, $P_S = 0{\cdot}8$ atm ; $P_A^\circ - P_S = 0{\cdot}2$ atm ; $M_B = 40$ g mol^{-1}; $W_A = 114$ g ; M_A (C_8H_{18}) = 114 g mol^{-1}.

$$W_B = \frac{(P_A^\circ - P_S)}{P_S} \times \frac{M_B \times W_A}{M_A}$$

$$= \frac{(0\cdot 2 \text{ atm})}{(0\cdot 8 \text{ atm})} \times \frac{(40 \text{g mol}^{-1}) \times (114 \text{g})}{(114 \text{g mol}^{-1})} = \mathbf{10{\cdot}0\ g.}$$

2.19. A solution containing 30g of non-volatile solute exactly in 90 g of water has a vapour pressure of 2.8 kPa at 298 K. Further, 18g of water is then added to the solution and the new of vapour pressure becomes 2.9 kPa at 298 K. Calculate

(i) molar mass of the solute.

(ii) vapour pressure of water at 298 K.

Sol: Let the molar mass of solute = Mg mol^{-1}

$\therefore$ **Moles of solute present**

$$= \frac{30g}{M\,g\,mol^{-1}} = \frac{30}{M}\text{ mol}$$

Moles of solvent present, $(n_1) = \frac{90}{18} = 5$ moles.

$$\therefore \quad \frac{P^\circ - P_s}{P^\circ} = \frac{n_2}{n_1 + n_2}$$

$$\frac{P^\circ - 2{\cdot}8}{P^\circ} = \frac{30/M}{5 + 30/M}$$

$$1 - \frac{2{\cdot}8}{P^\circ} = \frac{30}{(5M + 30)}$$

$$1 - \frac{30}{5M + 30} = \frac{2{\cdot}8}{P^\circ}$$

$$1 - \frac{6}{M + 6} = \frac{2.8}{P^\circ}$$

$$\frac{M + 6 - 6}{M + 6} = \frac{2{\cdot}8}{P^\circ}$$

$$\frac{M}{M + 6} = \frac{2{\cdot}8}{P^\circ}$$

$$\frac{P^\circ}{2{\cdot}8} = 1 + \frac{6}{M} \qquad ...(i)$$

After adding 18 g of water,
Moles of water becomes

$$= \frac{90 + 18}{18} = \frac{108}{18} = 6 \text{ moles}$$

$$\therefore \frac{P^\circ - P_s}{P^\circ} = \frac{30/M}{6 + 30/M}$$

P_s New vapour pressure = 2·9 kPa

$$\frac{P^\circ - 2{\cdot}9}{P^\circ} = \frac{30\ M}{M\ (6M + 30)} = \frac{5}{M + 5}$$

$$1 - \frac{2{\cdot}9}{P^\circ} = \frac{5}{M + 5}$$

$$1 - \frac{5}{M + 5} = \frac{2{\cdot}9}{P^\circ}$$

$$\frac{M + 5 - 5}{M + 5} = \frac{2{\cdot}9}{P^\circ}$$

$$\frac{P^\circ}{2{\cdot}9} = \frac{M + 5}{M} \Rightarrow = 1 + \frac{5}{M}$$

$$\frac{P^\circ}{2{\cdot}9} = 1 + \frac{5}{M} \qquad ...(ii)$$

Dividing equation (*i*) by (*ii*), we get,

$$\frac{2.9}{2.8}=\frac{1+6/M}{1+5/M}$$

$$2.9\left(1+\frac{5}{M}\right)=2.8\left(1+\frac{6}{M}\right)$$

$$2.9+\frac{2.9\times 5}{M}=2.8+\frac{2.8\times 6}{M}$$

$$2.9+\frac{14.5}{M}=2.8+\frac{16.8}{M}$$

$$0.1=\frac{16.8}{M}-\frac{14.5}{M}=\frac{2.3}{M}$$

$$M=\frac{2.3}{0.1}$$

$M=23\ g\ mol^{-1}$

Putting M = 23, in equation (*i*), we get,

$$\frac{P^\circ}{2.8}=1+\frac{6}{23}=\frac{29}{23}$$

$$P^\circ=\frac{29}{23}\times 2.8=3.53\ kPa.$$

2.20. A 5% solution (by mass) of cane sugar in water has freezing point of 271 K. Calculate the freezing point of 5% glucose in water if freezing point of pure water is 273.15 K.

Sol: Mass of sugar in 5% (by mass) solution means 5gin 100g of solvent (water)

Molar mass of sugar = 342g mol^{-1}

Molality of sugar solution $= \frac{5 \times 1000}{342 \times 100} = 0.146$

$\therefore \Delta T_f$ for sugar solution $= 273.15 - 271 = 2.15°$

$\Delta T_f = K_f \times m$

$\Delta T_f = K_f \times 0.146 \Rightarrow K_f = 2.15/0.146$

Molality of glucose solution

$= \frac{5}{180} \times \frac{1000}{100} = 0.278$

(Molar mass of glucose = 180 g mol^{-1})

$\Delta T_f = K_f \times m = \frac{2.15}{0.146} \times 0.278 = 4.09°$

$\therefore$ Freezing point of glucose solution

$= 273.15 - 4.09 = 269.06$ K.

2.21. Two elements A and B form compounds having formula AB_2 and AB_4. When dissolved in 20g of benzene (C_6H_6), 1 g of AB_2 lowers the freezing point by 2.3 K whereas 1.0 g of AB_4 lowers it by 1.3 K. The molar depression constant for benzene is 5.1 K kg mol^{-1}. Calculate atomic masses of A and B.

Sol:

Using the relation, $M_2 = \frac{1000 \times k_f \times w_2}{w_1 \times \Delta T_f}$

$\therefore \quad M_{AB_2} = \frac{1000 \times 5.1 \times 1}{20 \times 2.3} = 110.87 \text{ g mol}^{-1}$

$M_{AB_4} = \frac{1000 \times 5.1 \times 1}{20 \times 1.3} = 196.15 \text{ g mol}^{-1}$

Let the atomic masses of A and B are 'p' and 'q' respectively.

Then molar mass of

$AB_2 = p + 2q = 110.87 \text{ g mol}^{-1}$ ...(*i*)

And molar mass of

$AB_4 = p + 4q = 196.15 \text{ g mol}^{-1}$ ...(*ii*)

Substracting equation (*ii*) from equation (*i*), we get $2q = 85.28 \Rightarrow q = 42.64$

Putting $q = 42.64$ in equ. (*i*), we get

$p = 110.87 - 85.28$

$p = 25.59$

Thus, atomic mass of $A = 25.59 \text{ g mol}^{-1}$ and atomic mass of $B = 42.64 \text{ g mol}^{-1}$