## 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

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

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

$$= \frac{22}{22+122} \times 100\%$$

$$= 15.28\%$$

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

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

$$= \frac{122}{22+122} \times 100\%$$

$$= 84.72\%$$

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

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

∴Mass of carbon tetrachloride = (100 − 30)g

= 70 g

Molar mass of benzene ($C_6H_6$) = (6 × 12 + 6 × 1) g $mol^{-1}$

= 78 g $mol^{-1}$

∴Number of moles of $C_6H_6 = \frac{30}{78}$ mol

= 0.3846 mol

Molar mass of carbon tetrachloride ($CCl_4$) = 1 × 12 + 4 × 355

= 154 g $mol^{-1}$

∴Number of moles of $CCl_4$ $= \frac{70}{154}$ mol

= 0.4545 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 $Co(NO_3)_2.\ 6H_2O$ in 4.3 L of solution **(b)** 30 mL of 0.5 M $H_2SO_4$ diluted to 500 mL.

Answer

Molarity is given by:

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litre}}$$

**(a)** Molar mass of Co $(NO_3)_2.6H_2O$ = 59 + 2 (14 + 3 × 16) + 6 × 18

= 291 g $mol^{-1}$

∴Moles of Co $(NO_3)_2.6H_2O$ $= \frac{30}{291}$ mol

= 0.103 mol

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

= 0.023 M

**(b)** Number of moles present in 1000 mL of 0.5 M $H_2SO_4$ = 0.5 mol

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

= 0.015 mol

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

= 0.03 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

Molar mass of urea ($NH_2CONH_2$) = 2(1 × 14 + 2 × 1) + 1 × 12 + 1 × 16

= 60 g $mol^{-1}$

0.25 molar aqueous solution of urea means:

1000 g of water contains 0.25 mol = (0.25 × 60)g of urea

= 15 g of urea

That is,

(1000 + 15) 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 g

= 37 g of urea (approximately)

Hence, mass of urea required = 37 g

**Note:** There is a slight variation in this answer and the one given in the NCERT textbook.

**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

**(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 m

= 1.51 m (approximately)

**(b)** It is given that the density of the solution = 1.202 g $mL^{-1}$

∴Volume of 100 g solution $= \frac{\text{Mass}}{\text{Density}}$

$$= \frac{100\ \text{g}}{1.202\ \text{g mL}^{-1}}$$

= 83.19 mL

= 83.19 × $10^{-3}$ L

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

= 1.45 M

**(c)** Moles of KI

Moles of water

Therefore, $$=\frac{20}{166}=0.12\text{ mol}$$ mole fraction of KI

$$=\frac{0.12}{0.12+4.44}$$

= 0.0263 $$=\frac{80}{18}=4.44\text{ mol}$$

$$=\frac{\text{Moles of KI}}{\text{Moles of KI}+\text{Moles of water}}$$

**Question 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.

Answer

It is given that the solubility of $H_2S$ in water at STP is 0.195 m, i.e., 0.195 mol of $H_2S$ is dissolved in 1000 g of water.

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

= 55.56 mol

$$=\frac{\text{Moles of H}_2\text{S}}{\text{Moles of H}_2\text{S}+\text{Moles of water}}$$

∴Mole fraction of $H_2S$, $x$

$$=\frac{0.195}{0.195+55.56}$$

= 0.0035

At STP, pressure ($p$) = 0.987 bar

According to Henry's law: $p = K_Hx$

$$\Rightarrow K_H=\frac{p}{x}$$

$$=\frac{0.987}{0.0035}\text{ bar}$$

= 282 bar

**Question 2.7:**

Henry's law constant for $CO_2$ in water is $1.67 \times 10^8$ 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.

Answer

It is given that:

$K_H = 1.67 \times 10^8$ Pa

$p_{CO_2}$ = 2.5 atm = $2.5 \times 1.01325 \times 10^5$ Pa

= $2.533125 \times 10^5$ Pa

According to Henry's law:

$$p_{CO_2} = K_H x$$

$$\Rightarrow x = \frac{p_{CO_2}}{K_H}$$

$$= \frac{2.533125 \times 10^5}{1.67 \times 10^8}$$

= 0.00152

We $x = \frac{n_{CO_2}}{n_{CO_2} + n_{H_2O}} \approx \frac{n_{CO_2}}{n_{H_2O}}$ can write,

$n_{CO_2}$ is negligible as compared to $n_{H_2O}$ [Since, ]

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

$$= \frac{500}{18} \text{ mol of water}$$

= 27.78 mol of water

Now, $\frac{n_{CO_2}}{n_{H_2O}} = x$

$$\frac{n_{CO_2}}{27.78} = 0.00152$$

$$n_{CO_2} = 0.042 \text{ mol}$$

Hence, quantity of $CO_2$ in 500 mL of soda water = (0.042 × 44)g

= 1.848 g

**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

It is given that:

$p_A^0$ = 450 mm of Hg

$p_B^0$ = 700 mm of Hg

$p_{total}$ = 600 mm of Hg

From Raoult's law, we have:

$$p_A = p_A^0 x_A$$

$$p_B = p_B^0 x_B = p_B^0 (1 - x_A)$$

Therefore, total pressure,

$$p_{total} = p_A + p_B$$

$$\Rightarrow p_{total} = p_A^0 x_A + p_B^0 (1 - x_A)$$

$$\Rightarrow p_{total} = p_A^0 x_A + p_B^0 - p_B^0 x_A$$

$$\Rightarrow p_{total} = (p_A^0 - p_B^0) x_A + p_B^0$$

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

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

$$\Rightarrow x_A = 0.4$$

$$x_B = 1 - x_A$$

Therefore,

= 1 − 0.4

= 0.6

Now, $p_A = p_A^0 x_A$

= 450 × 0.4 = 180 mm of Hg

$p_B = p_B^0 x_B$

= 700 × 0.6

= 420 mm of Hg 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

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

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.

**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

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

$= 0.37$ K

Mass of water, $w_l = 500$ g

Molar mass of sucrose ($C_{12}H_{22}O_{11}$), $M_2 = 11 \times 12 + 22 \times 1 + 11 \times 16$

$= 342$ 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.

**Note:** There is a slight variation in this answer and the one given in the NCERT textbook.

**Question 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. $K_f$ = 3.9 K kg mol$^{-1}$.

Answer

Mass of acetic acid, $w_1$ = 75 g

Molar mass of ascorbic acid ($C_6H_8O_6$), $M_2$ = 6 × 12 + 8 × 1 + 6 × 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.

**Note:** There is a slight variation in this answer and the one given in the NCERT textbook.

**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

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, $n = \frac{1}{185000}\text{ mol}$

We know that:

Osmotic pressure, $\pi = \frac{n}{V}RT$

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

= 30.98 Pa

= 31 Pa (approximately)