INTEX QUESTIONS

Question 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.

Solution 1:

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

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

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

$$= 15.28\%$$

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

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

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

$$= 84.72\%$$

Alternatively,

Mass percentage of CCl_4 =(100-15.28) = 84.72%

Question 2:

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

Solution 2:

30% by mass of C_6H_6 in $CCl4$ $\Rightarrow$ 30 g C_6H_6 in 100 g solution

$\therefore$ Number of moles of C_6H_6 ($n_{C_6H_6}$) = 30/78 = 0.385 mol

(molar mass of C_6H_6 = 78g)

No. of moles of

$$CCl_4\left({}^{n}CCl_4\right)=\frac{70}{154}=0.455$$

$$x_{C_6H_6}=\frac{{}^{n}C_6H_6}{{}^{n}C_6H_6+{}^{n}CCl_4}$$

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

x_{CCl4} = 1 – 0.458 = 0.542

Question 3:

Calculate the molarity of each of the following solution: (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.

Solution 3:

(a) Molar mass of $Co(NO_3)_2.6H_2O$ = 310.7 g mol^{-1}
no. of moles = 30/310.7 = 0.0966
Vol. of solution = 4.3 L
Molarity = 0.0966/4.3 = 0.022 M
(b) 1000 mL of 0.5M H_2SO_4 contain H_2SO_4 = 0.5 mole
30 mL of 0.5 M H_2SO_4 contain H_2SO_4
= 0.5/1000 × 30 = 0.015 mole
Volume of solution = 500mL = 0.5L
Molarity = 0.015/0.5 = 0.03M

Question 4:

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

Solution 4:

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}
$\therefore$ 0.25 mole of urea = 0.25 × 60 = 15g
Mass of solution = 1000 + 15g = 1.015 kg
1.015 kg of urea solution contains 15g of urea
$\therefore$ 2.5 kg of solution contains urea = 15/1.015 × 2.5 = 37 g.

Question 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}

Solution 5:

20% aq. KOH solution $\Rightarrow$ 20g of KI in 100g solution

$\therefore$ Mass of solvent = 100 – 20= 80 g

$$\text{(i) Molality} = \frac{\text{no.of moles of Kl}}{\text{massof solvent (kg)}}$$

$$= \frac{0.120}{0.080} = 1.5 \text{ mol kg}^{-1}$$

(ii) Density of solution = 1.202 g mL^{-1}

$$\text{Volume of solution } = \frac{100}{1.202} = 83.2 \text{ mL}$$

$$= 0.0832 \text{ L}$$

$$\therefore \text{Molarity} = \frac{0.120}{0.0832} = 1.44\ M$$

(iii) No of moles of Kl = 0.120

$$n_{H_2O} = \frac{80}{18} = 4.44$$

$$x_{kl} = \frac{0.120}{0.120 + 4.44}$$

$$= \frac{0.120}{4.560} = 0.0263$$

Question 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.

Solution 6:

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

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

$\Rightarrow K_H = \dfrac{P_{H_2S}}{X_{H_2S}} = \dfrac{0.987}{0.0035} = 282\text{bar}$

Question 7:

Heny's law constant for CO_2 in water is 1.67×10^8Pa 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.

Solution 7:

KH = 1.67×10^8 Pa

P_{CO_2} = 2.5 atm = 2.5 × 101325 Pa

$\therefore X_{CO_2} = \dfrac{p_{CO_2}}{k_H} = \dfrac{2.5 \times 101325}{1.67 \times 10^8} = 1.517 \times 10^{-3}$

For 500 mL of soda water, water $\cong$ 500 mL

$= 500g = \dfrac{500}{18} = 27.78 \text{ moles}$

$\therefore n_{H_2O} = 27.78 \text{ moles}$

$\therefore \dfrac{n_{CO_2}}{27.78} = 27.78\text{moles}$

$n_{CO_2} = 42.14 \times 10^{-3}\text{mole}$

= 42.14m mole

= $42.14 \times 10^{-3} \times 44$g

= 1.854 g

Question 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.

Solution 8:

$P_A^\circ = 450mm,\ P_B^\circ = 700mm,\ P_{\text{total}} = 600mm$

As Ptotal= PA + PB

$= X_A P_A^\circ + (1 - X_A) P_B^\circ$

$= P_B^\circ + \left(P_A^\circ - P_B^\circ\right) X_A$

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

Or x_A= 0.40

$\therefore x_B = 1 - X_A = 1 - 1.40 = 0.60$

$\therefore P_A = x_A P_A^{\circ} = 0.40 \times 450 = 180 mm$

$\therefore P_B = x_B P_A^{\circ} = 0.60 \times 700 = 420 mm$

$\therefore$ Mole fraction of A in vapour phase

$$= \frac{P_A}{P_A + P_B} = \frac{180}{180 + 420} = 0.30$$

And, mole fraction of B in vapour phase

= 1 - 0.30 = 0.70

Question 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.

Solution 9:

P° = 23.8 mm Hg

W2 = 50g, M_2(urea) = 60g mol^{-1}

w1 = 850g, M_1 (water) = 18g mol^{-1}

To find: Ps and (P°- P_s)/ P°

Solution: Applying Raoult's law,

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

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

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

Putting P° = 23.8 mm, we have

23.8 mm, we have

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

$\Rightarrow$ 23.8 - P_s = 0.017 P_s

or 1.017 P_s = 23.8

or P_s = 23.4 mm

Question 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.

Solution 10:

Given $\Delta T_b = 100 - 99.63 = 3.37°$

Mass of water $w_1 = 500g$

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

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

To find Mass of sucrose, $w_2 = ?$

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

$$= K_b \times \frac{w_2}{M_2} \times \frac{100}{w_1}$$

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

$w_2 = 1108.2g$

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

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

Solution 11:

Given: $\Delta T_f = 1.5°$

Mass of CH_3COOH, $w_1 = 75g$

$M_1 = 60g\ mol^{-1}$

$M_2(C_6H_8O) = 176g\ mol^{-1}$

$K_f = 3.9$ kg mol^{-1}

To find: $w_2 = ?$

Solution: Applying $M_2 = \frac{1000 K_f w_2}{w_1 \Delta T_f}$

Or $w2 = \frac{M_2 \times w_1 \times \Delta T_f}{1000 \times K_f}$

$$W_2 = \frac{176 \times 75 \times 1.5}{1000 \times 3.9} = 5.077g$$

Question 12:

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

Solution 12:

Given V = 450mL = 0.45L

T = 37°C = 310K

R = 8.314 kPaL k^{-1} mol^{-1}

To find:

Solution: Applying the formula,

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

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

$$\therefore P = \frac{1}{185{,}000} \times \frac{1}{0.45} \times 8.314 \times 10^{3}\, Pa\ LK^{-1}\ \text{mol}^{-1} \times 310K$$

= 30.96 Pa

NCERT Exercise

Question 1:

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

Solution 1: A Solution is a homogeneous mixture of two or more chemically non-reacting substances. Types of solution: There are nine types of solution.

Types of Solution Examples

Gaseous solutions:

(a) Gas in gas: Air, mixture of O_2 and N_2, etc.

(b) Liquid in gas: Water vapour

(c) Solid in gas: Camphor vapours in N_2 gas, smoke etc.

Liquid solutions

(a) Gas in liquid: CO_2 dissolved in water (aerated water), and O_2 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,eg.Na-Hg.

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

Question 2:

Give an example of a solid solution in which the solute is a gas.

Solution 2:

Solution of hydrogen in palladium and dissolved gases in minerals

Question 3:

Define the following terms:
(i) Mole fraction
(ii) Molality
(iii) Molarity
(iv) Mass percentage

Solution 3:

(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.

$$\left(X_A = \frac{n_A}{n_A + n_B}\right) \text{....(1)}$$

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 the number of moles of a solute present in 1000g (1kg) of a solvent.

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

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.

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

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

$$\therefore M = \frac{\text{Weight in grams}}{\text{Volumeof solutioninlitres}} \times \frac{1}{\text{Molecular weight of solute}}$$

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.

$$= \frac{\text{Massof solute}}{\text{Massof solution}} \times 100$$

Question 4:

Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of the acid if the density of the solution is 1.504 g mL^{-1}?

Solution 4:

68% nitric acid by mass means that 68g mass of nitric acid is dissolved in 100g mass the solution. Molar mass of HNO3 = 63g mol^{-1}

$$\therefore 68\text{g of } HNO_3 = \frac{68}{63} = 1.079\text{mole}$$

Density of solution = 1.504 g mL^{-1} given

$\therefore$ Volume of solution

$$-\frac{\text{mass}}{\text{density}} = \frac{100}{1.504} = 66.5\ mL$$

$\therefore$ Molarity of solution

$$= \frac{\text{moles of slolute} \times 1000}{\text{Volume of solution in mL}}$$

$$= \frac{1.079 \times 1000}{65} = 16.23M$$

Question 5:

A solution of glucose in water is labelled as 10% w/w, that would be the molality and mole fraction of each component in the solution? If the density of solution is 1.2g mL^{-1} then what shall be the molarity of the solution?

Solution 5:

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

$$\therefore 10\text{g of glucose} = \frac{10}{180} = 0.0555 \text{ moles}$$

$$\text{and 90g of H2O} = \frac{90}{18} = 5 \text{ moles}$$

$\therefore$ Molality of solution

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

$$= \frac{0.0555}{90} \times 1000 = 0.617\ m$$

moles fraction of glucose

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

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

Mole fraction of water

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

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

volume of 100g solution

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

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

$$= 0.67\text{M}$$

Question 6:

How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na_2CO_3 and $NaHCO_3$ containing equimolar amounts of both?

Solution 6:

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 × 23 + 12 + 3 × 16 = 84g mol^{-1}

and molar mass of $NaHCO_3$

= 23 × 1 + 1 + 12 + 3 × 16 = 84g mol^{-1}

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

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 = 190 x

$$\therefore x = \frac{106}{190} = 0.558g$$

$\therefore$ No. of moles of Na_2CO_3 present

$$= \frac{0.558}{106} = 0.00526$$

and no. of moles of $NaHCO_3$ present

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

Calculation of no. of moles of HCl required

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

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

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 × 2 = 0.01052 mole

Each mole of NaHCO3 needs 1 mole of HCl

$\therefore$ 0.00526 mole of $NaCHO_3$ needs = 1 × 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

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

Question 7:

A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate the mass percentage of the resulting solution.

Solution 7:

300 g of 25% solution will contain $= \frac{25 \times 300}{100} = 75\text{g of solute.}$

400g of 40% solution will contain $= \frac{40 \times 400}{100} = 160g$ of solute

$\therefore$ Total mass of solute = 160 + 75 = 235g

Total mass of solution = 300 + 400 = 700g

Now, the percentage of solute in solution. $= \frac{235}{700} \times 100 = 33.5\%$

And the percentage of water in solution = 100 - 33.5 = 66.5%

Question 8:

An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200g of water. Calculate the molality of the solution. If the density of the solution is 1.072 mL^{-1}, then what shall be the molarity of the solution?

Solution 8:

Mass of solute = 222.6g

Molar mass of solute $C_2H_6O_2$

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

$\therefore$ Moles of solute $= \frac{222.6}{62} = 3.59$

Mass of solvent = 200 g

$\therefore$ Molality $= \frac{3.59}{200} \times 1000 = 17.95\ mol\ kg^{-1}$

Total mass of solution = 422.6g

Volume of solution $= \frac{422.6}{1.072} = 394.21\ mL$

$\therefore$ Molarity $= \frac{3.59}{394.2} \times 1000 = 9.1\ mol\ L^{-1}$

Question 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 water sample.

Solution 9:

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

$\therefore$ 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 = 10^6g

Molar mass of $CHCl_3$= 12 + 1 + 3 × 35.5 = 119.5 mol^{-1}

Moles of $CHCl_3$= $\frac{15}{119.5}$

$$\therefore \text{ Molality} = \frac{15/119.5\times 1000}{10^{6}} = 1.25\times 10^{-4}\, m$$

Question 10:

What role does the molecular interaction play in a solution of alcohol and water?

Solution 10:

Alcohol and water both have strong tendency to form intermolecular hydrogen bonding. On mixing the two, a solution is formed as a result of formation of H-bonds between alcohol and H_2O molecules but these interactions are weaker and less extensive than those in pure H_2O. Thus they show a positive deviation from ideal behavior. As a result of this, the solution of alcohol and water will have higher vapour pressure and lower boiling point than that of water and alcohol.

Question 11:

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

Solution 11:

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.

Question 12:

State Henry's law and mention some important applications.

Solution 12:

The effect of pressure on the solubility of a gas in a liquid is governed by Henry's Law. It states that the solubility of a gas in a liquid at a given temperature is directly proportional to the partial pressure of the gas. Mathematically, P = KHX where P is the partial pressure of the gas; and X is the mole fraction of the gas in the solution and KH is Henry's Law constant.

Applications of Henry's law:

(i) In the production of carbonated beverages (as solubility of CO_2 increases at high pressure)

(ii) In the deep sea diving.

(iii) For climbers or people living at high altitudes, where low blood O_2 causes climbers to become weak and make them unable to think clearly

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Question 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?

Solution 13:

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

$\therefore 6.56 \times 10^{-2} g = KH \times 1$ bar ………(i)

$\therefore\ 5.00 \times 10^{-2} g = KH \times P$ ………(ii)

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

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

$$\therefore \frac{6.56\times10^{-2}}{1} = \frac{5.00\times10^{-2}}{P}$$

$$\therefore P = \frac{5.00}{6.56} = 0.762 \text{ bar}$$

Question 14:

What is meant by positive and negative deviations from Raoult’s law and how is the sign of $\Delta_{mix}H$ related to positive and negative deviations from Raoult’s law?

Solution 14:

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 $\Delta_{mix}H$ 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 $\Delta_{mix}H$ is positive. Similarly $\Delta_{mix}V$ 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 interaction are stronger than A-A and B-B. So weaker interactions are replaced by stronger interactions, so there is release of enenergy i.e. $\Delta_{mix}H$ is negative.

Question 15:

An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute?

Solution 15:

Vapour pressure of pure water at the boiling point (P°) = 1.013 bar

Vapour pressure of solution (P_s) = 1.004 bar

Mass of solute (w_2) = 2g

Molar mass of solvent, water (M_1) = 18g

Mass of solvent (w_1) = 98g

Mass of solution = 100g

Applying Raoult’s Law for dilute solutions,

$$\frac{P^\circ - P_s}{P^\circ} = \frac{n_2}{n_1 + n_2} \cong \frac{n_2}{n_1} \text{ [Dilute solution being 2\%]}$$

$$\frac{P^\circ - P_s}{P^\circ} = \frac{n_2}{n_1} = \frac{W_2 / W_1}{W_1 / W_1}$$

$$\frac{(1.013 - 1.004)}{(1.013)} = \frac{2 \times 18}{M_2 \times 98}$$

$$\therefore M_2 = \frac{2 \times 18}{98 \times 0.009} \times 1.013 = 41.35 g \text{ mol}^{-1}$$

Question 16:

Heptane and octane form an ideal solution. At 373 K, the vapour pressure of the two liquid components 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?

Solution 16:

Molar mass heptane (C_7H_{16}) = 7 × 12 + 16 = 100g mol^{-1}

Molar mass of octane (C_8H_{18}) = 8 × 12 + 18 = 114 g mol^{-1}

Moles of heptane present in mixture = $\frac{26.0}{100} = 0.26$ mol

Moles of octane present in mixture = $\frac{35.0}{114} = 0.307$ mol

Mole fraction of heptane $x_H = \frac{0.26}{0.26 + 0.307} = 0.458$

Mole fraction of octane, x_O – (1 - 0.458) = 0.542

Vapour pressure of heptane = $x_H \times P^\circ$ = 0.458 × 105.2 kPa = 48.18 kPa

Vapour pressure of octane = $x_O \times P^\circ$ = 0.542 × 46.8 kPa = 25.36 kPa

Vapour pressure of mixture = 48.18 + 25.36 = 73.54 kPa

Question 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

Solution 17:

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

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

$\therefore$ Moles of water = $\frac{1000}{8} = 55.5$ moles

$\therefore$ Mole fraction of solute $= \dfrac{1}{1+55.5} = 0.0177$

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

$\dfrac{12.3 - P_s}{12.3} = 0.0177$

$\Rightarrow P_s = 12.08 \text{ kPa}$

Question 18:

Calculate the mass of a non-volatile solute (molar mass 40 g mol^{-1}) which should be dissolved in 114 g octane to reduce its vapour pressure to 80%

Solution 18:

P_x = 80% of P°

$= \dfrac{80}{100} p^\circ = 0.8 p^\circ$

Let Wg of solute is present in mixture.

Moles of solute present $= \dfrac{W}{40}$ moles

Molar mass of octane, C8H18 = 8 × 12 + 8 = 114 mol^{-1}

$\therefore$ Moles of octane $= \dfrac{114}{114} = 1 \text{ mol}$

Now $\dfrac{P^\circ - P_s}{P^\circ} = x_2 = \dfrac{W/40}{\dfrac{W}{40}+1}$

$\dfrac{P^\circ - 0.80P^\circ}{P^\circ} = \dfrac{W/40}{W/40+1}$

$1 - 0.80 = \dfrac{W \times 40}{40(W+40)} = \dfrac{W}{W+40}$

$0.20 = \dfrac{W}{W+40}$

0.20 W + 8 = W

8 = W (1 - 0.2)

8 = 0.8 W

$\therefore W = \dfrac{8}{0.8} = 10g$

Question 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.

Solution 19:

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

$$\therefore \text{Moles of solute present} = \frac{30g}{\text{Mgmol}^{-1}} = \frac{30}{M}\text{mol}$$

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

$$\frac{P° - P_s}{P°} = \frac{n_2}{n_1 + n_2}$$

$$\frac{P° - 2.8}{P°} = \frac{30/W}{5 + 30/W}$$

$$1 - \frac{2.8}{P°} = \frac{30}{(5M + 30)}$$

$$1 - \frac{30}{5M + 30} = \frac{2.8}{P°}$$

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

$$\frac{M + 6 - 6}{M + 6} = \frac{2.8}{P°}$$

$$\frac{M}{M + 6} = \frac{2.8}{P°}$$

$$\frac{P°}{2.8} = 1 + \frac{6}{M} \quad \text{........(i)}$$

After adding 18 g of water, moles of water becomes

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

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

Ps New vapour pressure = 2.9 kPa

$$\frac{P° - 2.9}{P°} = \frac{30M}{M(6M + 30)} = \frac{5}{M + 5}$$

$$1 - \frac{2.9}{P} = \frac{5}{M + 5}$$

$$1-\frac{5}{M+5}=\frac{2.9}{P^\circ}$$

$$\frac{M+5-5}{M+5}=\frac{2.9}{p^\circ}$$

$$\frac{P^\circ}{2.9}=\frac{M+5}{M}\Rightarrow=1+\frac{5}{M}$$

$$\frac{P^\circ}{2.9}=1+\frac{5}{M} \quad ....(ii)$$

Dividing equation (i) by (ii) we get

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

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

$$2.9+\frac{2.9\times5}{M}=2.8+\frac{2.8\times6}{M}$$

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

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

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

M = 23 g mol^{-1}

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

$$\frac{P^\circ}{.8}=1+\frac{6}{23}=\frac{29}{23}$$

$$P^\circ=\frac{29}{23}\times2.8=3.53\text{ kPa}$$

Question 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 273.15K.

Solution 20:

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

Molality of sugar solution = $\frac{5\times1000}{342\times100}=0.146$

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

$\therefore \Delta T_f = \text{K} \times \text{m}$

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

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 = 269.06K.

Question 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 AB2 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.

Solution 21:

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

$\therefore M_{AB_2} = \frac{1000 \times 5.1 \times 1}{20 \times 2.3} = 110.8 \text{ g mol}^{-1}$

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

Let the atomic masses of and B are ‘p’ and ‘q’ respectively.
Then molar mass of $AB_2 = p + 2q = 110.8$ g mol^{-1}….(i)
And molar mass of $AB_4 = p + 4q = 196.15$ g mol^{-1}……(ii)
Subtracting equation (ii) from equation (i),
we get 2q = 85.28 $\Rightarrow$ q = 42.64
Putting q= 42.64 in eqn. (i) wed get
P = 110.87 - 85.28
P = 25.59
Thus, atomic mass of A = 25.59 u and
atomic mass of B = 42.64 u

Question 22:

At 300 K, 36g of glucose present in a liter of its solution has an osmotic pressure of 4.08 bar. If the osmotic pressure of the solution is 1.52 bars at the same temperature, what would be its concentration?

Solution 22:

$\pi = \text{CRT}$

$$4.98 = \frac{w_1}{M_1} \times R \times 300$$

$$4.98 = \frac{36}{180} \times R \times 300$$

4.98 = 60 R …..(i)

In second case 1.52 = C × R × 300 …..(ii)

Diving equation (ii) by equation (i), we get

$$C = \frac{60 \times 1.52}{300 \times 4.98} = 0.06\text{M}$$

Question 23:

Suggest the most important type of intermolecular attractive interaction in the following pairs:

(i) n-hexane and n-octane

(ii) I_2 and CCl_4.

(iii) $NaClO_4$ and water

(iv) methanol and acetone

(v) acetonitrile (CH_3CN) and acetone (C_3H_6O)

Solution 23:

(i) Both n-hexane and n-octane are non-polar. Thus, the intermolecular interactions will be London dispersion forces.

(ii) Both I_2 and CCl_4 are non-polar. Thus the intermolecular interactions will be London dispersion forces.

(iii)$NaClO_4$ is an ionic compound and gives Na^+ and ClO_4^- ions in the Solution. Water is a polar molecule. Thus, the intermolecular interactions will be ion-dipole interactions.

(iv) Both methanol and acetone are polar molecules. Thus, intermolecular interactions will be dipole-dipole interactions.

(v) Both CH_3CN and C_3H_6O are polar molecules. Thus, intermolecular interactions will be dipole-dipole interactions.

Question 24:

Based on solute-solvent interactions, arrange the following in order of increasing solubility in n-octane and explain.

Cyclohexane, KCl, CH_3OH, CH_3CN.

Solution 24:

(a) Cyclohexane and n-octane both are non-polar. They mix completely in all proportions.

(b) KCl is an ionic compound,KCl will not dissolve in n-octane.
(c) CH_3OH is polar. CH_3OH will dissolve in n-octane.
(d) CH_3CN is polar but lesser than CH_3OH. Therefore, it will dissolve in n-octane but to a greater extent as compared to CH_3OH Hence, the order is KCl< CH_3OH < CH_3CN < Cyclohexane.

Question 25:

Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water
(i) phenol
(ii) toluene
(iii) formic acid
(iv) ethylene glycol
(v) chloroform
(vi) Pentanol

Solution 25:

(i) Phenol (having polar – OH group) – Partially soluble.
(ii) Toluene (non-polar) – insoluble.
(iii) Formic acid (form hydrogen bonds with water molecules) – Highly soluble
(iv) Ethylene glycol (form hydrogen bonds with water molecules) Highly soluble
(v) Chloroform (non-polar) insoluble.
(vi) Pentanol (having polar – OH) – Partially soluble.

Question 26:

If the density of some lake water is 1.25 g mL^{-1} and contains 92g f Na^+ ions per kg of water, calculate the molality of Na^+ ions in the lake.

Solution 26:

Molar mass of Na = 23g mol^{-1}

$$\therefore \text{No. of moles of Na}^+ \text{ ions present} = \frac{92}{23} = 4 \text{ moles}$$

$$\therefore \text{Molality} = \frac{4 \times 1000}{1000} = 4 \text{ m.}$$

Question 27:

If the solubility product of CuS is 6×10^{-16}, calculate the maximum molarity of CuS aqueous solution.

Solution 27:

$CuS \rightleftharpoons Cu \quad - S^{2-}, K_{sp} = 6 \times 10^{-16}$

Maximum molarity of CuS in aqueous solution means solubility of CuS.
Let the solubility of CuS be S mol L^{-1}

$\therefore K_{sp} = [Cu^{2+}][S]^{2-} 6 \times 10^{-16} = S \times S = S^2$

$\therefore S = \sqrt{6 \times 10^{-16}} = 2.45 \times 10^{-8}$ mol L^{-1}

Question 28:

Calculate the mass percentage of aspirin ($C_9H_8O_4$) in acetonitrile (CH_3CN) when 6.5 g of $C_9H_8O_4$ is dissolved in 450 g of CH_3CN.

Solution 28:

6.5 g of $C_9H_8O_4$ is dissolved in 450 g of CH_3CN.
Then, total mass of the solution = (6.5 + 450) g = 456.5 g

Therefore, mass percentage of $C_9H_8O_4$

$= \frac{6.5}{456.5} \times 100\%$

= 1.424%

Question 29:

Nalorphene ($C_{19}H_{21}NO_3$), similar to morphine, is used to combat withdrawal symptoms in narcotic users. Dose of nalorphene generally given is 1.5 mg.

Calculate the mass of 1.5×10^{-3}m aqueous solution required for the above dose.

Solution 29:

The molar mass of nalorphene ($C_{19}H_{21}NO_3$),
$19 \times 12 + 21 \times 1 + 1 \times 14 + 3 \times 16 = 311\, g\, mol^{-1}$

In 1.5×10^{-3}m aqueous solution of nalorphene,
$1.5 \times 10^{-3} \times 311\, g$

1 kg (1000 g) of water contains 1.5×10^{-3} mol
= 0.4665 g
Therefore, total mass of the solution is given as
= (1000 + 0.4665) g
= 1000.4665 g

This implies that the mass of the solution containing 0.4665 g of nalorphene is 1000.4665g.
Therefore, mass of the solution containing 1.5 mg of nalorphene is:

$$\frac{1000.4665\times 1.5\times 10^{-3}}{0.4665}g$$
$$=3.22\,g$$

Hence, the mass of aqueous solution required is 3.22 g.

Question 30:
Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing 250 mL of 0.15M solution in methanol.

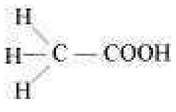
Solution 30:
0.15 M solution of benzoic acid in methanol means,1000 mL of solution contains 0.15 mol of benzoic acid.

$$\frac{0.15\times 250}{1000}$$

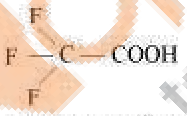
Therefore, 250 mL of solution contains = mol of benzoic acid = 0.0375 mol of benzoic acid.
Molar mass of benzoic acid (C_6H_5COOH) = 7 × 12 + 6 × 1 + 2 × 16 = 122 g mol^{-1}
Hence, required benzoic acid = 0.0375 mol × 122 g mol^{-1}
= 4.575 g

Question 31:
The depression in freezing point of water observed for the same amount of acetic acid, trichloroacetic acid and trifluoroacetic acid increases in the order given above. Explain briefly.

Solution 31:


Acetic acid


Trichloroacetic acid


Trifluoroacetic acid

Among H, Cl, and F, H is least electronegative while F is most electronegative. Then, F can withdraw electrons towards itself more than Cl and H. Thus, trifluoroacetic acid can easily lose H +ions i.e., trifluoroacetic acid ionizes to the largest extent. Now, the more ions produced, the greater is the depression of the freezing point. Hence, the depression in the freezing point increases in the order:
Acetic acid < trichloroacetic acid < trifluoroacetic acid.

Question 32:
Calculate the depression in the freezing point of water when 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water. $Ka = 1.4 \times 10{-}3$, $Kf = 1.86$ K kg mol−1.

Solution 32:
Molar mass of $CH_3CH_2CHClCOOH$ is= 15 + 14 + 13 + 35.5 + 12 + 16 + 16 + 1
$=122.5\text{g mol}^{-1}$

sent in 10 g of $CH_3CH_2CHClCOOH$ is = $\dfrac{10g}{122.5\text{ g mol}^{-1}}$

=0.0816 mol

It is given that 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water.

Molality of the solution, = $\dfrac{0.0186}{250}\times1000$ =0.3264molkg^{-1}

Let α be the degree of dissociation of $CH_3CH_2CHClCOOH$

$CH_3CH_2CHClCOOH$ undergoes dissociation according to the following equation:

$$CH_3CH_2CHClCOOH \leftrightarrow CH_3CH_2CHClCOO^- + H^+$$

Initial conc.	$C\ mol\ L^{-1}$	0	0
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$

$$\therefore K_a = \frac{C\alpha.C\alpha}{C(1-\alpha)}$$

$$= \frac{C\alpha^2}{1-\alpha}$$

Since α is very small with respect to 1, $1 - \alpha \approx 1$

$$\text{Now } K_a = \frac{C\alpha^2}{1}$$

$$\Rightarrow K_a = C\alpha^2$$

$$\Rightarrow \alpha = \sqrt{\frac{K_a}{C}}$$

$= \sqrt{\dfrac{1.4\times10^{-3}}{0.3264}}(\because K_a$ [illegible] $4\times10^{-3})$

$$= 0.0655$$

Again,

$$CH_3CH_2CHClCOOH \leftrightarrow CH_3CH_2CHClCOO^- + H^+$$

Initial moles	1	0	0
At equilibrium	$1-\alpha$	α	α

Total moles of equilibrium = $1 - \alpha + \alpha + \alpha$

$= 1 + \alpha$

$$\therefore i = \frac{1+\alpha}{1}$$

$$= 1+\alpha$$

$$= 1+0.0655$$

$$= 1.0655$$

Hence, the depression in the freezing point of water is given as:

$$\Delta T_f = i.K_f m$$

$$= 1.0655 \times 1.86 K \text{ kg mol}^{-1} \times 0.3264 \text{ mol kg}^{-1}$$

$$= 0.65\ K$$

Question 33:

19.5 g of CH_2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is 1.0°C. Calculate the van’t Hoff factor and dissociation constant of fluoroacetic acid.

Solution 33:

It is given that:

$w_1 = 500 g$

$w_2 = 19.5 g$

$K_f = 1.86\ K \text{ kg mol}^{-1}$

$\Delta T_f = 1 K$

We know that:

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

$$= \frac{1.86 K\ kg\ mol^{-1} \times 19.5 g \times 1000\ g\ kg^{-1}}{500 g \times 1 K}$$

$$= 72.54\ g \text{ mol}^{-1}$$

Therefore, observed molar mass of $CH_2FCOOH, (M_2)_{obs} = 72.54$ g mol

The calculated molar mass of CH_2FCOOH:

$$(M_2)_{cal} = 14 + 19 + 12 + 16 + 16 + 1$$

$$= 78 \text{ g mol}^{-1}$$

Therefore, van’t Hoff factor

$$i = \frac{(M_2)_{cal}}{(M_2)_{obs}} \text{ is}$$

$$= \frac{78 \text{ g mol}^{-1}}{72.54 \text{ g mol}^{-1}}$$

= 1.0753

Let α be the degree of dissociation of CH_2FCOOH:

$$CH_2FCOOH \leftrightarrow CH_2FCOO^- + H^+$$

Initial conc.	C mol L^{-1}	0	0	
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$	Total = C(1+α)

$$\therefore i = \frac{C(1+\alpha)}{C}$$

$$\Rightarrow i = 1 + \alpha$$

$$\Rightarrow \alpha = i - 1$$

$$= 1.0753 - 1$$

$$= 0.0753$$

Now, the value of K_a is given as:

$$K_a = \frac{[CH_2FCOO^-][H^+]}{[CH^2FCOOH]}$$

$$= \frac{C\alpha . C\alpha}{C(1-\alpha)}$$

$$= \frac{C\alpha^2}{1-\alpha}$$

Taking the volume of the solution as 500 mL, we have the concentration:

$$C = \frac{\frac{19.5}{78}}{500} \times 1000M$$

$$= 0.5M$$

$$K_a = \frac{C\alpha^2}{1-\alpha}$$

Therefore,

$$= \frac{0.5 \times (0.0753)^2}{1 - 0.0753}$$

$$= \frac{0.5 \times 0.00567}{0.9247}$$

$$= 0.00307 \text{ (approximately)}$$

$$= 3.07 \times 10^{-3}$$

Question 34:

Vapour pressure of water at 293 K is 17.535 mm Hg. Calculate the vapour pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water.

Solution 34:

Vapour pressure of water, P_1^0 = 17.535 mm of Hg

Mass of glucose, $w_2 = 25$ g

Mass of water, $w_1 = 450$ g

We know that,

Molar mass of glucose (C6H12O6), $M_2 = 6 \times 12 + 12 \times 1 + 6 \times 16 = 180$ g mol^{-1}

Molar mass of water, M_1 = 18 g mol^{-1}
Then, number of moles of glucose,

$$n_2 = \frac{25}{180 g \text{ mol}^{-1}}$$

= 0.139 mol
And, number of moles of water,

$$n_1 = \frac{450 g}{18\ g \text{ mol}^{-1}}$$

= 25 mol

We know that,

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

$$\Rightarrow \frac{17.535 - p_1}{17.535} = \frac{0.139}{0.139 + 25}$$

$$\Rightarrow 17.535 - p_1 = \frac{0.139 \times 17.535}{25.139}$$

$\Rightarrow$ 17.535 − p_1 = 0.097
$\Rightarrow$ p_1 = 17.44 mm of Hg
Hence, the vapour pressure of water is 17.44 mm of Hg.

Question 35:
Henry’s law constant for the molality of methane in benzene at 298 K is 4.27 × 10^5mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.

Solution 35:
Here, p = 760
mm Hg kH = 4.27 × 10^5
mm Hg According to Henry’s law,
p = kHx

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

$$= \frac{760 \text{ mm Hg}}{4.27 \times 10^5 \text{ mm Hg}}$$

= 177.99 × 10^{-5}
= 178 × 10^{-5} (approximately)
Hence, the mole fraction of methane in benzene is 178 × 10^{-5}

Question 36:
100 g of liquid A (molar mass 140 g mol^{-1}) was dissolved in 1000 g of liquid B (molar mass 180 g mol^{-1}). The vapour pressure of pure liquid B was found to be 500 torr. Calculate the vapour pressure of pure liquid A and its vapour pressure in the solution if the total vapour pressure of the solution is 475 Torr.

Solution 36:

Number of moles of liquid A,

$$n_A = \frac{100}{140}\text{mol} \qquad = 0.714 \text{ mol}$$

Number of moles of liquid B,

$$n_B = \frac{1000}{180}\text{mol} \qquad = 5.556 \text{ mol}$$

Then, mole fraction of A,

$$x_A = \frac{n_A}{n_A + n_B}$$

$$= \frac{0.714}{0.714 + 5.556}$$

= 0.114

And, mole fraction of B, $x_B = 1 - 0.114 = 0.886$

Vapour pressure of pure liquid B, $p_B^0 = 500$ torr

Therefore, vapour pressure of liquid B in the solution,

$$p_B = p_B^0 x_B$$

= 500 × 0.886

= 443 torr

Total vapour pressure of the solution, $p_{\text{total}} = 475$ torr

Vapour pressure of liquid A in the solution, $p_A = p_{\text{total}} - p_B = 475 - 443$

= 32 torr

Now,

$$p_A = p_A^0 x_A$$

$$\Rightarrow p_A^0 = \frac{p_A}{x_A}$$

$$= \frac{32}{0.114}$$

= 280.7 torr

Hence, the vapour pressure of pure liquid A is 280.7 torr.

Question 37:

Vapour pressure of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition,

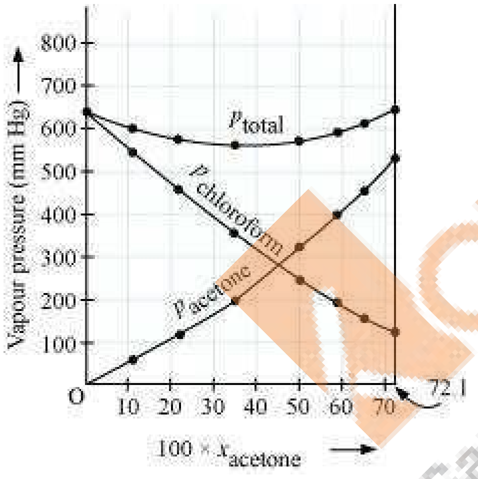
plot p_{total}, $p_{\text{chloroform}}$, and p_{acetone} as a function of x_{acetone}. The experimental data observed for different

compositions of mixture is.

100 × $x_{acetone}$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$p_{acetone}$ /mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{chloroform}$/ mm Hg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7

Solution 37: From the question, we have the following data

100 × $x_{acetone}$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$p_{acetone}$ /mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{chloroform}$/ mm Hg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7
p_{tota}(mm Hg)	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8



It can be observed from the graph that the plot for the *p*total of the solution curves downwards. Therefore, the solution shows negative deviation from the ideal behaviour.

Question 38:

Benzene and toluene form ideal solution over the entire range of composition.
The vapour pressure of pure benzene and naphthalene at 300 K are 50.71 mm Hg and 32.06 mm Hg respectively. Calculate the mole fraction of benzene in vapour phase if 80 g of benzene is mixed with 100 g of toluene.

Solution 38:

Molar mass of benzene

$(C_6H_6) = 6\times12+6\times1 = 78 \text{ g mol}^{-1}$

Molar mass of toluene

$(C_6H_5CH_3) = 7 \times 12 + 8 \times 1 = 92 \text{ g mol}^{-1}$

$= 92 \text{ g mol}^{-1}$

Now, no. of moles present in 80 g of benzene

$= \frac{80}{78} \text{ mol} = 1.026 \text{ mol}$

And, no. of moles present in 100 g of toluene

$= \frac{100}{92} \text{ mol} = 1.087 \text{ mol}$

Mole fraction of benzene, x_b

$= \frac{1.026}{1.026 + 1.087} = 0.486$

And, mole fraction of toluene,

$x_t = 1 - 0.486 = 0.514$

It is given that vapour pressure of pure benzene,

$p_b^0 = 50.71 \text{ mm Hg}$

And, vapour pressure of pure toluene,

$p_t^0 = 32.06 \text{ mm Hg}$

Therefore, partial vapour pressure of benzene,

$P_b = x_b \times p_b$

$= 0.514 \times 32.06$

$= 16.479 \text{ mm Hg}$

Hence, mole fraction of benzene in vapour phase is given by:

$\frac{p_b}{p_b + p_t}$

$= \frac{24.645}{24.645 + 16.479}$

$= \frac{24.645}{41.124}$

$= 0.599$

$= 0.6$

Question 39:

The air is a mixture of a number of gases. The major components are oxygen and nitrogen with approximate proportion of 20% is to 79% by volume at 298 K. The water is in equilibrium with air at a pressure of 10 atm. At 298 K if the Henry's law constants for oxygen and nitrogen are 3.30×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

Solution 39:

Percentage of oxygen (O2) in air = 20 %

Percentage of nitrogen (N2) in air = 79%
Also, it is given that water is in equilibrium with air at a total pressure of 10 atm, that is, (10 × 760) mm Hg = 7600 mm Hg
Therefore, Partial pressure of oxygen,

$$p_{O_2} = \frac{20}{100} \times 7600 \text{ mm Hg} = 1520 \text{ mm Hg}$$

Partial pressure of nitrogen,

$$p_{N_2} = \frac{79}{100} \times 7600 \text{ mm Hg} = 6004 \text{ mmHg}$$

Now, according to Henry's law:
$p = K$H.x
For oxygen:

$$p_{O_2} = k_H . x_{O_2}$$

$$\Rightarrow x_{O_2} = \frac{p_{O_2}}{k_H}$$

$$= \frac{1520 \text{ mm Hg}}{3.30 \times 10^7 \text{ mm Hg}} \quad (\text{Given } \mathrm{k_H} = 3.30 \times 10^7 \text{ mm Hg})$$

$$= 4.61 \times 10^{-5}$$

For nitrogen:

$$p_{N_2} = K_H \cdot x_{N_2}$$

$$\Rightarrow x_{N_2} = \frac{p_{N_2}}{K_H}$$

$$= \frac{6004 \text{ mm Hg}}{6.51 \times 10^7 \text{ mm Hg}}$$

$$= 9.22 \times 10^{-5}$$

Hence, the mole fractions of oxygen and nitrogen in water are 4.61 $\times 10^{-5}$and 9.22 $\times 10^{-5}$ respectively.

Question 40:
Determine the amount of $CaCl_2$ (i = 2.47) dissolved in 2.5 litre of water such that its osmotic pressure is 0.75 atm at 27°C.

Solution 40:
We know that,

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

$$\Rightarrow \pi = i\frac{w}{MV}RT$$

$$\Rightarrow w = \frac{\pi MV}{iRT}$$

$\pi = 0.75$ atm

V = 2.5L

i = 2.47

T = (27 + 273)K = 300K

Here,
R = 0.0821 L atm $K^{-1}mol^{-1}$
M = 1 × 40 + 2 × 35.5
= 111g mol^{-1}
Therefore, w

$$= \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300}$$

= 3.42 g
Hence, the required amount of $CaCl_2$ is 3.42 g.

Question 41:
Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K_2SO_4 in 2 litre of water at 25° C, assuming that it is completely dissociated.

Solution 41:
When K_2SO_4 is dissolved in water,
K^+ and SO_4^{2-} ions are produced.

$$K_2SO_4 \longrightarrow 2K^+ + SO_4^{2-}$$

Total number of ions produced = 3
i =3 Given, w = 25 mg = 0.025g
V = 2 L
T = 250C = (25 + 273) K = 298 K Also,
we know that:
R = 0.0821 L atm $K^{-1}mol^{-1}$
M = (2 × 39) + (1 × 32) + (4 × 16) = 174 g mol^{-1}
Appling the following relation,

$$\pi = i\frac{n}{v}RT$$

$$= i\frac{w}{M}\frac{1}{v}RT$$

$$= 3 \times \frac{0.025}{174} \times \frac{1}{2} \times 0.0821 \times 298$$

$$= 5.27 \times 10^{-3} \text{ atm}$$