Chapter 1 - Solutions
XII

Intext Questions

Question 1.1:

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 1.2:

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 1.3:

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 1.4:

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 1.5:

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 $= \dfrac{\text{Moles of KI}}{\text{Mass of water in kg}}$

$$= \frac{\frac{20}{166}}{0.08}\ \text{m}$$

= 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 $= \dfrac{\text{Mass}}{\text{Density}}$

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

= 83.19 mL

= 83.19 × 10^{-3} L

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

= 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 1.6:

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

∴Mole fraction of H_2S, x $= \frac{\text{Moles of } H_2S}{\text{Moles of } H_2S + \text{Moles of water}}$

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

= 0.0035

At STP, pressure (p) = 0.987 bar

According to Henry's law: $p = K_H x$

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

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

= 282 bar

Question 1.7:

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.

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

= 2.533125×10^5 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 1.8:

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 1.9:

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 1.10:

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.

Answer

Here, elevation of boiling point ΔT_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 × 12 + 22 × 1 + 11 × 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 1.11:

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 ($C_6H_8O_6$), $M_2 = 6 \times 12 + 8 \times 1 + 6 \times 16$

= 176 g mol^{-1}

Lowering of melting point, ΔT_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 1.12:

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

We know that:

Osmotic pressure, $\pi = \frac{n}{V}\mathrm{R}T$

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

= 30.98 Pa

= 31 Pa (approximately)

Exercises

Question 1.1:

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

Answer

A solution is formed when two or more components are mixed together.

The component which is present in a large amount in solution than others is called a solvent and the component which is less in amount in a solution is called the solute.

There are 9 types of solution, which are as follow:

Solution type	Solute	Solvent	Example
Gas in Gas	Gas	Gas	Air
Gas in Liquid	Gas	Liquid	Soda water
Gas in Solid	Gas	Solid	H_2 in Pt
Liquid in gas	Liquid	Gas	Water vapor in the air
Liquid in Liquid	Liquid	Liquid	Alcoholic beverages.
Liquid in Solid	Liquid	Solid	Water into ice
Solid in Gas	Solid	Gas	Camphor in air
Solid in Liquid	Solid	Liquid	Salt in water
Solid in Solid	Solid	Solid	Metal alloys, Bronze

Question 1.2:

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

Answer

In case a solid solution is formed between two substances (one having very large particles and the other having very small particles), an interstitial solid solution will be formed. For example, a solution of hydrogen in palladium is a solid solution in which the solute is a gas.

Question 1.3:
Define the following terms:

(i) Mole fraction (ii) Molality (iii) Molarity (iv) Mass percentage.

Answer

(i) Mole fraction

It is the ratio of the number of moles of a particular component to the total number of moles of all the components in the mixture. It is denoted by symbol χ.

(ii) Molality

It is the number of moles of the solute per kilogram of the solvent. It is the ratio of the number of moles of solute to the mass of solvent in kg.

(iii) Molarity

It is the number of moles of solute per litre of the solution. It is the ratio of the number of moles of solute to the volume of solution in litre.

(iv) Mass percentage

The mass percentage is one way of representing the concentration of an element in a compound or a component in a mixture. The mass percentage is calculated as the mass of a component divided by the total mass of the mixture, multiplied by 100%.

Question 1.4:
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}?

Answer

The molar mass of nitric acid is 63 g/mol.

It is given that the nitric acid used in the laboratory is 68% by mass in aqueous solution.

100 g of solution contains 68 g of nitric acid or $\frac{68}{63} = 1.079$ moles of nitric acid.

Density of solution is 1.504 g/mL.

100 g of solution corresponds to $\frac{100}{1.504} = 66.5$ mL or 0.0665 L

Molarity of the solution is the number of moles of nitric acid present in 1 L of solution.

It is $\frac{1.079}{0.0665} = 16.22$ M

Hence, the molarity should be 16.22 g/mL.

Question 1.5:

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 mL^{-1}, then what shall be the molarity of the solution?

Answer

Step 1:

10% w/w glucose solution means 10 g glucose present in 100 g of solution which contains 90 g of water.

Molality is the number of moles of glucose present in 1 kg of water.

Molar mass of glucose ($C_6H_{12}O_6$) = $6\times12+12\times1+6\times16=180 gmol^{-1}$

$$\text{Number of moles} = \frac{\text{Given mass in grams}}{\text{Molecular mass}}$$

$$\text{Number of moles of glucose} = \frac{10}{180}\text{mol} = 0.056 \text{ mol}$$

$$\text{Molality of solution} = \frac{\text{Number of moles}}{\text{weight of solvent}}$$

$$\therefore \text{Molality of solution} = \frac{0.056 \text{ mol}}{0.09 \text{kg}} = 0.62 \text{ M}$$

$$\text{Number of moles of water} = \frac{90 \text{ g}}{18 \text{ gmol}^{-1}} = 5 \text{ mol}$$

Step 2:

$$\text{Mole fraction} = \frac{\text{Number of moles of solute}}{\text{Number of moles of solute + number of moles of solvent}}$$

$$\text{Mole fraction of glucose, } x_g = \frac{0.056}{0.056+5} = 0.011$$

$$\text{Mole fraction of water, } x_w = 1 - x_g = 1 - 0.011 = 0.989$$

Step 3:

If density of solution is 1.20 g/mL, then the volume of the 100 g solution can be given as:

$$\frac{100 \text{g}}{1.2 \text{ gmL}^{-1}} = 83.33 \text{ mL} = 83.33 \times 10^{-3} \text{ L}$$

$$\therefore \text{Molarity of the solution} = \frac{0.056 \text{ mol}}{83.33 \times 10^{-3} \text{L}} = 0.67 \text{ M}$$

Question 1.6:

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?

Answer

Step 1:

Let the mixture contains x g of sodium carbonate and 1−x g of sodium bicarbonate.

The molar masses of sodium carbonate and sodium bicarbonate are 106 g/mol and 84 g/mol respectively.

The number of moles of sodium carbonate and sodium bicarbonates are

$\frac{x}{106}$ and $\frac{1-x}{84}$ respectively.

Step 2:

Since, it is an equimolar mixture.

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

$84x = 106 - 106x$

$x = 0.5579$

Step 3:

Number of moles of sodium carbonate = $\frac{0.5579}{106}$ = 0.005263

Number of moles of sodium hydrogen carbonate = $\frac{1-0.5579}{84}$ = 0.005263

One mole of sodium carbonate will react with 2 moles of HCl and 1 mole of sodium bicarbonate will react with 1 mole of HCl.

Step 4:

Total number of moles of HCl that will completely neutralize the mixture

= 2 × 0.005263 + 0.005263 = 0.01578 moles

Volume of 0.1 M HCl required = $\frac{0.01578}{0.1}$ = 0.158 L = 158 mL

Question 1.7:

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.

Answer

Given that a solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass.

Total mass of solution =300 + 400 = 700 g

Mass of solute =0.25 × 300 + 0.40 × 400 = 75 + 160 = 235 g

Mass percentage of solution = $\frac{235}{700} \times 100 = 33.6\%$

Hence, the mass percentage of the resulting solution is 33.6%

Question 1.8:

An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200 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?

Answer

Molar mass of ethylene glycol [$C_2H_4(OH)2$]= 2 × 12 + 6 × 1 + 2 ×16

= 62 $gmol^{-1}$

Number of moles of ethylene glycol = $\frac{222.6g}{62\text{gmol}^{-1}}$

= 3.59 mol

Therefore, molality of the solution = $\frac{3.59mol}{0.200kg}$

= 17.95 m

Total mass of the solution = (222.6 + 200) g

= 422.6 g

Given,

Density of the solution = 1.072 g mL^{-1}

∴ Volume of the solution = $\frac{422.6g}{1.072\ gmL^{-1}}$

= 394.22 mL

= 0.3942×10^{-3} L

=>Molarity of the solution = $\frac{3.59mol}{0.39422 \times 10^{-3}L}$

=9.11 M

Question 1.9:

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.

Answer

15 ppm corresponds to 15 g chloroform in 1000,000 g of solution.

$$\text{(i) Percent by mass} = \frac{\text{Mass of chloroform}}{\text{Mass of solution}} \times 100$$

$$\text{Percent by mass} = \frac{15}{1000,000} \times 100 = 1.5 \times 10^{-3}\%$$

$$\text{(ii) Molality} =$$

$$\frac{\text{Mass of chloroform}}{\text{Molar mass of chloroform} \times (\text{Mass of solution} - \text{mass of chloroform})} \times 1000 \text{ (all masses in g)}$$

$$\text{Molality} = \frac{15}{119.5 \times (1000,000 - 15)} \times 1000 = 1.255 \times 10^{-4} \text{ m.}$$

Question 1.10:
What role does the molecular interaction play in a solution of alcohol and water?

Answer

In pure alcohol and water, the molecules are held tightly by a strong hydrogen bonding. The interaction between the molecules of alcohol and water is weaker than alcohol−alcohol and water−water interactions. As a result, when alcohol and water are mixed, the intermolecular interactions become weaker and the molecules can easily escape. This increases the vapour pressure of the solution, which in turn lowers the boiling point of the resulting solution.

Question 1.11:
Why do gases always tend to be less soluble in liquids as the temperature is raised?

Answer
Gas + Liquid ⇌ Dissolving gas; $\Delta H = -\text{ve}$. Dissolution of gas is an exothermic process. As the temperature is raised, the equilibrium shifts in reverse direction (Le-Chatelier's principle).
It results in decrease of solubility of gases in liquid.

Question 1.12:
State Henry’s law and mention some important applications.

Answer
According to Henry's law, the solubility of a gas in a liquid is directly proportional to the pressure of the gas.

$X = K_h \times P$

X is the mole fraction of gas, K_h is Henry's law constant and P is the partial pressure of the gas.

Important applications of Henry's law:

1) In packing of soda cans: Soda water bottles are always packed under higher pressure to increase the solubility of CO_2 gas.

2) In deep-sea diving: Nitrogen is more soluble than Helium in our blood. In the deep sea, the pressure is higher than at the surface of the water. When diver tries to come rapidly towards the surface of the water, pressure decreases and dissolved nitrogen comes back from blood and makes bubbles in veins. Hence, divers use oxygen diluted with helium.

Question 1.13:

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?

Answer

Molar mass of ethane (C_2H_6) $= 2 \times 12 + 6 \times 1$

$= 30$ g mol^{-1}

∴ Number of moles present in 6.56×10^{-2} g of ethane $= \frac{6.56\times10^{-2}}{30}$

$= 2.187 \times 10^{-3}$ mol

Let the number of moles of the solvent be x. According to Henry's law,

$p = K_H x$

$$\Rightarrow 1\ \text{bar} = K_H \cdot \frac{2.187\times10^{-3}}{2.187\times10^{-3}+x}$$

$$\Rightarrow 1\ \text{bar} = K_H \frac{2.187\times10^{-3}}{x}$$

$$\Rightarrow K_H = \frac{x}{2.187\times10^{-3}}\ \text{bar}$$

(Since $x >> 2.187 \times 10^{-3}$)

Number of moles present in 5.00×10^{-2} g of ethane $= \frac{5.00\times10^{-2}}{30}$ mol

$= 1.67 \times 10^{-3}$ mol According to Henry's law,

$p = K_H x$

$$= \frac{x}{2.187\times10^{-3}} \times \frac{1.67\times10^{-3}}{\left(1.67\times10^{-3}\right)+x}$$

$$= \frac{x}{2.187\times10^{-3}} \times \frac{1.67\times10^{-3}}{x}$$

(Since, $x >> 1.67 \times 10^{-3}$)

= 0.764 bar

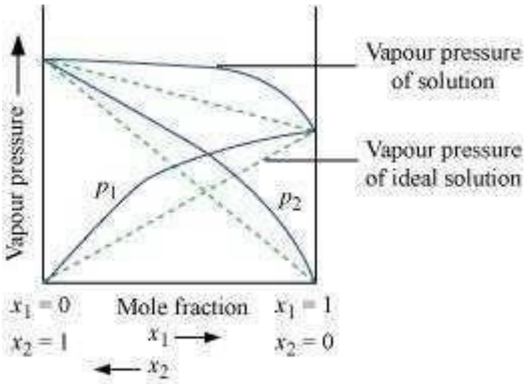
Hence, partial pressure of the gas shall be 0.764 bar.

Question 1.14:

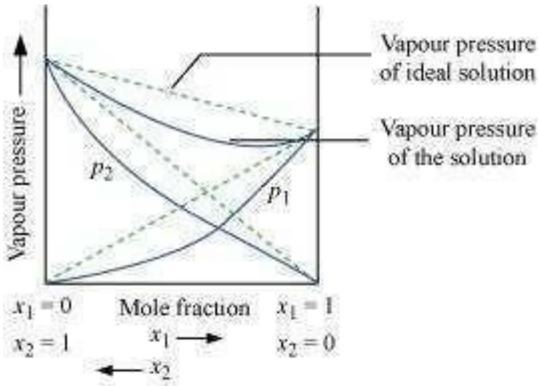
What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{sol}H$ related to positive and negative deviations from Raoult's law?

Answer

According to Raoult's law, the partial vapour pressure of each volatile component in any solution is directly proportional to its mole fraction. The solutions which obey Raoult's law over the entire range of concentration are known as ideal solutions. The solutions that do not obey Raoult's law (non-ideal solutions) have vapour pressures either higher or lower than that predicted by Raoult's law. If the vapour pressure is higher, then the solution is said to exhibit positive deviation, and if it is lower, then the solution is said to exhibit negative deviation from Raoult's law.



Vapour pressure of a two-component solution showing positive deviation from Raoult's law



Vapour pressure of a two-component solution showing negative deviation from Raoult's law

In the case of an ideal solution, the enthalpy of the mixing of the pure components for forming the solution is zero.

$\Delta_{sol}H = 0$

In the case of solutions showing positive deviations, absorption of heat takes place.

$\therefore \Delta_{sol}H = \text{Positive}$

In the case of solutions showing negative deviations, evolution of heat takes place.

$\therefore \Delta_{sol}H$ = Negative

Question 1.15:

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?

Answer

Here,

Vapour pressure of the solution at normal boiling point (p_1) = 1.004 bar

Vapour pressure of pure water at normal boiling point $(p_1^0) = 1.013 \text{ bar}$

Mass of solute, (w_2) = 2 g

Mass of solvent (water), (w_1) = 98 g

Molar mass of solvent (water), (M_1) = 18 g mol^{-1} According to Raoult's law,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

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

$$\Rightarrow \frac{0.009}{1.013} = \frac{2 \times 18}{M_2 \times 98}$$

$$\Rightarrow M_2 = \frac{1.013 \times 2 \times 18}{0.009 \times 98}$$

= 41.35 g mol^{-1}

Hence, the molar mass of the solute is 41.35 g mol^{-1}.

Question 1.16:

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 g of octane?

Answer

Vapour pressure of $(p_1^0) = 105.2 \text{ kPa}$ heptane

Vapour pressure of octane (p_2^0) = 46.8 kPa

We know that,

Molar mass of heptane (C_7H_{16}) = 7 × 12 + 16 × 1

= 100 g mol^{-1}

∴ Number of moles of heptane $= \frac{26}{100}$ mol

= 0.26 mol

Molar mass of octane (C_8H_{18}) = 8 × 12 + 18 × 1

= 114 g mol^{-1}

∴ Number of moles of $= \frac{35}{114}$ mol

= 0.31 mol

Mole fraction of heptane, $x_1 = \frac{0.26}{0.26+0.31}$

= 0.456

And, mole fraction of octane, $x_2 = 1 - 0.456$

= 0.544

Now, partial pressure of heptane, $p_1 = x_1 p_1^0$

= 0.456 × 105.2

= 47.97 kPa

Partial pressure of octane, $p_2 = x_2 p_2^0$

= 0.544 × 46.8

= 25.46 kPa

Hence, vapour pressure of solution, $p_{total} = p_1 + p_2$

= 47.97 + 25.46

= 73.43 kPa

Question 1.17:

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.

Answer

1 molal solution means 1 mol of the solute is present in 100 g of the solvent (water).

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

∴ Number of moles present in 1000 g of water $= \frac{1000}{18}$

= 55.56 mol

Therefore, mole fraction of the solute in the solution is

$x_2 = \frac{1}{1+55.56} = 0.0177$.

It is given that,

Vapour pressure of water, p_1^0 = 12.3 kPa

Applying the relation, $\frac{p_1^0 - p_1}{p_1^0} = x_2$

$$\Rightarrow \frac{12.3 - p_1}{12.3} = 0.0177$$

$\Rightarrow 12.3 - p_1 = 0.2177$

$\Rightarrow p_1 = 12.0823$

= 12.08 kPa (approximately)

Hence, the vapour pressure of the solution is 12.08 kPa.

Question 1.18:

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%.

Answer

Let the vapour pressure of pure octane p_1^0.

Then, the vapour pressure of the octane after dissolving the non-volatile solute is $\frac{80}{100} p_1^0 = 0.8 p_1^0$.

Molar mass of solute, M_2 = 40 g mol^{-1}

Mass of octane, w_1 = 114 g

Molar mass of octane, (C_8H_{18}), $M_1 = 8 \times 12 + 18 \times 1$

= 114 g mol^{-1}

Applying the relation,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

$$\Rightarrow \frac{p_1^0 - 0.8 p_1^0}{p_1^0} = \frac{w_2 \times 114}{40 \times 114}$$

$$\Rightarrow \frac{0.2 p_1^0}{p_1^0} = \frac{w_2}{40}$$

$$\Rightarrow 0.2 = \frac{w_2}{40}$$

$$\Rightarrow w_2 = 8 \text{ g}$$

Hence, the required mass of the solute is 8 g.

Question 1.19:

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

i. molar mass of the solute

ii. vapour pressure of water at 298 K.

Answer

(i) Let, the molar mass of the solute be M g mol^{-1}

Now, the no. of moles of solvent (water), $n_1 = \frac{90 \text{ g}}{18 \text{ g mol}^{-1}} = 5 \text{ mol}$

And, the no. of moles of solute, $n_2 = \frac{30 \text{ g}}{\text{M mol}^{-1}} = \frac{30}{\text{M}} \text{mol}$

$p_1 = 2.8 \text{ kPa}$

Applying the relation:

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

$$\Rightarrow \frac{p_1^0 - 2.8}{p_1^0} = \frac{\frac{30}{\text{M}}}{5 + \frac{30}{\text{M}}}$$

$$\Rightarrow 1 - \frac{2.8}{p_1^0} = \frac{\frac{30}{\text{M}}}{\frac{5\text{M} + 30}{\text{M}}}$$

$$\Rightarrow 1 - \frac{2.8}{p_1^0} = \frac{30}{5\text{M} + 30}$$

$$\Rightarrow \frac{2.8}{p_1^0} = 1 - \frac{30}{5\text{M} + 30}$$

$$\Rightarrow \frac{2.8}{p_1^0} = \frac{5\text{M} + 30 - 30}{5\text{M} + 30}$$

$$\Rightarrow \frac{2.8}{p_1^0} = \frac{5\text{M}}{5\text{M} + 30}$$

$$\Rightarrow \frac{p_1^0}{2.8} = \frac{5\text{M} + 30}{5\text{M}} \qquad (i)$$

After the addition of 18 g of water:

$n_1 = \frac{90 + 18\text{g}}{18} = 6 \text{ mol}$

$p_1 = 2.9 \text{ kPa}$

Again, applying the relation:

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

$$\Rightarrow \frac{p_1^0 - 2.9}{p_1^0} = \frac{\frac{30}{M}}{6 + \frac{30}{M}}$$

$$\Rightarrow 1 - \frac{2.9}{p_1^0} = \frac{\frac{30}{M}}{\frac{6M + 30}{M}}$$

$$\Rightarrow 1 - \frac{2.9}{p_1^0} = \frac{30}{6M + 30}$$

$$\Rightarrow \frac{2.9}{p_1^0} = 1 - \frac{30}{6M + 30}$$

$$\Rightarrow \frac{2.9}{p_1^0} = \frac{6M + 30 - 30}{6M + 30}$$

$$\Rightarrow \frac{2.9}{p_1^0} = \frac{6M}{6M + 30}$$

$$\Rightarrow \frac{p_1^0}{2.9} = \frac{6M + 30}{6M} \qquad (ii)$$

Dividing equation ***(i)*** by ***(ii)***, we have:

$$\frac{2.9}{2.8} = \frac{\frac{5M + 30}{5M}}{\frac{6M + 30}{6M}}$$

$$\Rightarrow \frac{2.9}{2.8} \times \frac{6M + 30}{6} = \frac{5M + 30}{5}$$

$$\Rightarrow 2.9 \times 5 \times (6M + 30) = 2.8 \times 6 \times (5M + 30)$$
$$\Rightarrow 87M + 435 = 84M + 504$$
$$\Rightarrow 3M = 69$$
$$\Rightarrow M = 23\,u$$

Therefore, the molar mass of the solute is 23 g mol^{-1}.

(ii) Putting the value of 'M' in equation **(*i*)**, we have:

$$\frac{p_1^0}{2.8} = \frac{5 \times 23 + 30}{5 \times 23}$$

$$\Rightarrow \frac{p_1^0}{2.8} = \frac{145}{115}$$

$$\Rightarrow p_1^0 = 3.53$$

Hence, the vapour pressure of water at 298 K is 3.53 kPa.

Question 1.20:

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.

Answer

Here, ΔT_f = (273.15 − 271) K

= 2.15 K

Molar mass of sugar ($C_{12}H_{22}O_{11}$) = 12 × 12 + 22 × 1 + 11 × 16

= 342 g mol^{-1}

5% solution (by mass) of cane sugar in water means 5 g of cane sugar is present in (100 − 5)g = 95 g of water.

Now, number of moles of cane $= \frac{5}{342}$ mol

= 0.0146 mol

Therefore, molality of the solution, $m = \frac{0.0146 \text{ mol}}{0.095 \text{ kg}}$

= 0.1537 mol kg^{-1}

Applying the relation,

$\Delta T_f = K_f \times m$

$$\Rightarrow K_f = \frac{\Delta T_f}{m}$$

$$= \frac{2.15 \text{ K}}{0.1537 \text{ mol kg}^{-1}}$$

= 13.99 K kg mol^{-1}

Molar of glucose ($C_6H_{12}O_6$) = 6 × 12 + 12 × 1 + 6 × 16 = 180 g mol^{-1}

5% glucose in water means 5 g of glucose is present in (100 − 5) g = 95 g of water.

∴ Number of moles of glucose $= \frac{5}{180}$ mol

= 0.0278 mol

Therefore, molality of the solution, $m = \frac{0.0278 \text{ mol}}{0.095 \text{ kg}}$

= 0.2926 mol kg^{-1}

Applying the relation,

$\Delta T_f = K_f \times m$

= 13.99 K kg mol^{-1} × 0.2926 mol kg^{-1}

= 4.09 K (approximately)

Hence, the freezing point of 5% glucose solution is (273.15 − 4.09) K= 269.06 K.

Question 1.21:

Two elements A and B form compounds having formula AB_2 and AB_4. When dissolved in 20 g 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 Kkg mol^{-1}.

Calculate atomic masses of A and B.

Answer

We know that,

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

Then, $M_{AB_2} = \frac{1000 \times 1 \times 5.1}{2.3 \times 20}$

= 110.87 g mol^{-1}

$$M_{AB_4} = \frac{1000 \times 1 \times 5.1}{1.3 \times 20}$$

= 196.15 g mol^{-1}

Now, we have the molar masses of AB_2 and AB_4 as 110.87 g mol^{-1} and 196.15 g mol^{-1} respectively.

Let the atomic masses of A and B be x and y respectively.

Now, we can write:

$$x + 2y = 110.87 \qquad \text{(i)}$$
$$x + 4y = 196.15 \qquad \text{(ii)}$$

Subtracting equation (i) from (ii), we have

$2y$ = 85.28

⇒ y = 42.64

Putting the value of 'y' in equation (1), we have x + 2 × 42.64 = 110.87

⇒ x = 25.59

Hence, the atomic masses of A and B are 25.59 u and 42.64 u respectively.

Question 1.22:

At 300 K, 36 g of glucose present in a litre of its solution has an osmotic pressure of 4.98 bar. If the osmotic pressure of the solution is 1.52 bars at the same temperature, what would be its concentration?

Answer

Here,

T = 300 K π

= 1.52 bar

R = 0.083 bar L K^{-1} mol^{-1}

Applying the relation, π =

CRT

$$\Rightarrow C = \frac{\pi}{RT}$$

$$= \frac{1.52 \text{ bar}}{0.083 \text{ bar L K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}$$

= 0.061 mol

Since the volume of the solution is 1 L, the concentration of the solution would be 0.061 M.

Question 1.23:

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).

Answer

(i) Van der Wall's forces of attraction.

(ii) Van der Wall's forces of attraction.

(iii) Ion-diople interaction.

(iv) Dipole-dipole interaction.

(v) Dipole-dipole interaction.

Question 1.24:

Based on solute-solvent interactions, arrange the following in order of increasing solubility in n-octane and explain. Cyclohexane, KCl, CH_3OH, CH_3CN.

Answer

n-octane is a non-polar solvent. Therefore, the solubility of a non-polar solute is more than that of a polar solute in the *n*-octane.

The order of increasing polarity is:

Cyclohexane < CH_3CN < CH_3OH < KCl

Therefore, the order of increasing solubility is:

KCl < CH_3OH < CH_3CN < Cyclohexane

Question 1.25:

Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

(i) phenol **(ii)** toluene **(iii)** formic acid

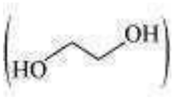
(iv) ethylene glycol **(v)** chloroform **(vi)** pentanol.

Answer

(i) Phenol (C_6H_5OH) has the polar group −OH and non-polar group $-C_6H_5$. Thus, phenol is partially soluble in water.

(ii) Toluene ($C_6H_5-CH_3$) has no polar groups. Thus, toluene is insoluble in water.

(iii) Formic acid (HCOOH) has the polar group −OH and can form H-bond with water. Thus, formic acid is highly soluble in water.

(iv) Ethylene glycol  has polar −OH group and can form H−bond. Thus, it is highly soluble in water.

(v) Chloroform is insoluble in water.

(vi) Pentanol ($C_5H_{11}OH$) has polar −OH group, but it also contains a very bulky nonpolar $-C_5H_{11}$ group. Thus, pentanol is partially soluble in water.

Question 1.26:

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

Answer

Number of moles present in 92 g of Na^+ ions = $\frac{92\text{ g}}{23\text{ g mol}^{-1}}$

= 4 mol

Therefore, molality of Na^+ ions in the lake $= \frac{4\text{ mol}}{1\text{ kg}}$

= 4 m

Question 1.27:

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

Answer

Solubility product of CuS, $K_{sp} = 6 \times 10^{-16}$ Let *s* be the solubility of CuS in mol L^{-1}.

$$CuS \leftrightarrow \underset{s}{Cu^{2+}} + \underset{s}{S^{2-}}$$

Now, $K_{sp} = \left[Cu^{2+}\right]\left[S^{2-}\right]$

$= s \times s$

$= s^2$

Then, we have, $K_{sp} = s^2 = 6 \times 10^{-16}$

$\Rightarrow s = \sqrt{6 \times 10^{-16}}$

$= 2.45 \times 10^{-8}$ mol L^{-1}

Hence, the maximum molarity of CuS in an aqueous solution is 2.45×10^{-8} mol L^{-1}.

Question 1.28:

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.

Answer

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 1.29:

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.

Answer

The molar mass of nalorphene $(C_{19}H_{21}NO_3)$ is given as:

$19 \times 12 + 21 \times 1 + 1 \times 14 + 3 \times 16 = 311 \text{ g mol}^{-1}$

In 1.5×10^{-3}m aqueous solution of nalorphene,

1 kg (1000 g) of water contains 1.5×10^{-3} mol $= 1.5 \times 10^{-3} \times 311$ g

$= 0.4665$ g

Therefore, total mass of the solution $= (1000 + 0.4665)$ g

$= 1000.4665$ g

This implies that the mass of the solution containing 0.4665 g of nalorphene is 1000.4665 g.

Therefore, mass of the solution containing 1.5 mg of nalorphene is:

$$\frac{1000.4665 \times 1.5 \times 10^{-3}}{0.4665} \text{ g}$$

$$= 3.22 \text{ g}$$

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

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

Question 1.30:

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

Answer

0.15 M solution of benzoic acid in methanol means,

1000 mL of solution contains 0.15 mol of benzoic acid

Therefore, 250 mL of solution contains = $\frac{0.15 x 250}{1000}$ 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}

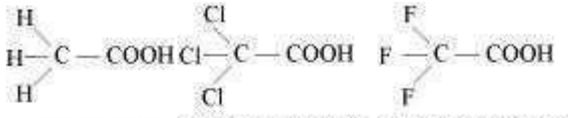
Hence, required benzoic acid = 0.0375 mol × 122 g mol^{-1}

= 4.575 g

Question 1.31:

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.

Answer



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 1.32:

Calculate the depression in the freezing point of water when 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water. $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86$ K kg mol^{-1}.

Answer

Molar mass of $CH_3CH_2CHClCOOH = 15+14+13+35.5+12+16+16+1$

$= 122.5 \text{ g mol}^{-1}$

∴No. of moles present in 10 g of $CH_3CH_2CHClCOOH = \frac{10 \text{ g}}{122.5 \text{ g mol}^{-1}}$

$= 0.0816 \text{ mol}$

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

∴Molality of the solution, $= \frac{0.0186}{250} \times 1000$

$= 0.3264 \text{ mol kg}^{-1}$

Let *α* be the degree of dissociation of $CH_3CH_2CHClCOOH.$

$CH_3CH_2CHClCOOH$ undergoes dissociation according to the following equation:

$$\begin{array}{llll} & CH_3CH_2CHClCOOH & \leftrightarrow CH_3CH_2CHClCOO^- & + H^+ \\ \text{Initial conc.} & C \text{ mol L}^{-1} & 0 & 0 \\ \text{At equilibrium} & C(1-\alpha) & C\alpha & C\alpha \end{array}$$

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

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

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

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

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

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

$$= \sqrt{\frac{1.4 \times 10^{-3}}{0.3264}} \quad \left(\because K_a = 1.4 \times 10^{-3}\right)$$

$$= 0.0655$$

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:

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

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

$$= 0.65 \text{ K}$$

Question 1.33:

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.

Answer

It is given that:

$w_1 = 500 \text{ g}$

$w_2 = 19.5 \text{ g}$

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

$\Delta T_f = 1 \text{K}$

We know that:

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

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

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

Therefore, observed molar mass of $\text{CH}_2\text{FCOOH}, (M_2)_{\text{obs}} = 72.54\,\text{g mol}$

The calculated molar mass of CH_2FCOOH

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

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

Therefore, van’t Hoff factor, $i = \frac{(M_2)_{\text{cal}}}{(M_2)_{\text{obs}}}$ 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

$$\text{CH}_2\text{FCOOH} \leftrightarrow \text{CH}_2\text{FCOO}^- + \text{H}^+$$

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

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

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

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

$$= 1.0753 - 1$$

$$= 0.0753$$

Now, the value of K_a is given as:

$$K_a = \frac{[\text{CH}_2\text{FCOO}^-][\text{H}^+]}{[\text{CH}_2\text{FCOOH}]}$$

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

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

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

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

$$= 0.5\,\text{M}$$

Therefore, $K_a = \dfrac{C\alpha^2}{1-\alpha}$

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

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

$$= 0.00307 \text{ (approximately)}$$
$$= 3.07 \times 10^{-3}$$

Question 1.34:

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.

Answer

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 ($C_6H_{12}O_6$), M_2 = 6 × 12 + 12 × 1 + 6 × 16 = 180 g mol^{-1}

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

Then, number of moles of glucose, $n_2 = \dfrac{25}{180\text{ g mol}^{-1}}$

= 0.139 mol

And, number of moles of water, $n_1 = \frac{450\text{ g}}{18\text{ g mol}^{-1}}$

= 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 1.35:

Henry's law constant for the molality of methane in benzene at 298 K is 4.27×10^5 mm Hg. Calculate the solubility of methane in benzene at 298 Kunder 760 mm Hg.

Answer

Here, p = 760 mm Hg

$k_H = 4.27 \times 10^5$

mm Hg According to

Henry's law,

$p = k_H x$

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

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

$= 177.99 \times 10^{-5}$

$= 178 \times 10^{-5}$ (approximately)

Hence, the mole fraction of methane in benzene is 178×10^{-5}.

Question 1.36:

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.

Answer

Number of moles of liquid A, $n_A = \frac{100}{140}$ mol

= 0.714 mol

Number of moles of liquid B, $n_B = \frac{1000}{180}$ mol

= 5.556 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_{total} = 475 torr

∴ Vapour pressure of liquid A in the solution,

$p_A = p_{total} - p_B = 475 - 443$

= 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 1.37:

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.

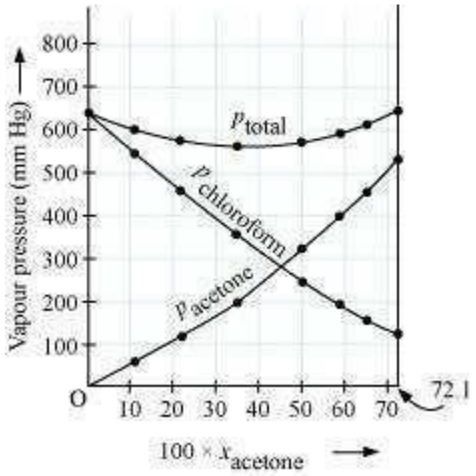
100 ×x_{acetone}	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
p_{acetone} /mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{\text{chloroform}}$/mm Hg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7

Plot this data also on the same graph paper. Indicate whether it has positive deviation or negative deviation from the ideal solution

Answer

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
$p_{\text{chloroform}}$/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 1.38:

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.

Answer

Molar mass of benzene $(C_6H_6) = 6 \times 12 + 6 \times 1$

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

Molar mass of toluene $(C_6H_5CH_3) = 7 \times 12 + 8 \times 1$

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

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, $p_b = x_b \times p_b$

$= 0.486 \times 50.71$

$= 24.645 \text{ mm Hg}$

And, partial vapour pressure of $p_t = x_t \times p_t$

$= 0.514 \times 32.06$

$= 16.479 \text{ mm Hg}$

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 1.39:

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 at 298 k are 3.30×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

Answer

Percentage of oxygen (O_2) in air = 20 %

Percentage of nitrogen (N_2) 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 mm Hg

Partial pressure of nitrogen, $p_{N_2} = \frac{79}{100} \times 7600 \text{ mmHg}$

= 6004 mmHg

Now, according to Henry's law:

$p = K_H.x$

For oxygen:

$$p_{O_2} = K_H \cdot x_{O_2}$$

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

$$= \frac{1520 \text{ mm Hg}}{3.30 \times 10^7 \text{ mm Hg}} \qquad \left(\text{Given } K_H = 3.30 \times 10^7 \text{ mm Hg}\right)$$

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

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 4.61×10^{-5} and 9.22×10^{-5} respectively.

Question 1.40:

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.

Answer

We know that,

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

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

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

$\pi = 0.75\,\text{atm}$

$V = 2.5\,\text{L}$

$i = 2.47$

$T = (27+273)\,\text{K} = 300\,\text{K}$

Here,

R = 0.0821 L atm $K^{-1}mol^{-1}$

M = 1 × 40 + 2 × 35.5

= 111g mol^{-1}

Therefore, $w = \frac{0.75\times111\times2.5}{2.47\times0.0821\times300}$

= 3.42 g

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

Question 1.41:

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

Answer

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

$\therefore i = 3$ Given, w = 25 mg = 0.025 g

V = 2 L

T = 25^0C = (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,

$$
\begin{aligned}
\pi &= i\frac{n}{v}\mathrm{R}T \\
&= i\frac{w}{M}\frac{1}{v}\mathrm{R}T \\
&= 3\times\frac{0.025}{174}\times\frac{1}{2}\times 0.0821\times 298 \\
&= 5.27\times 10^{-3}\,\mathrm{atm}
\end{aligned}
$$