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Question 2.1:

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

Answer:

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

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

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

$$= 15.28\%$$

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

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

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

$$= 84.72\%$$

Alternatively,

Mass percentage of CCl_4 = (100 – 15.28)%

= 84.72%

Question 2.2:

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

Answer:

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

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

= 70 g

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

= 78 g mol^{-1}

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

= 0.3846 mol

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

= 154 g mol^{-1}

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

= 0.4545 mol

Thus, the mole fraction of C_6H_6 is given as:

$$\frac{\text{Number of moles of } C_6H_6}{\text{Number of moles of } C_6H_6 + \text{ Number of moles of } CCl_4}$$

$$= \frac{0.3846}{0.3846 + 0.4545}$$

= 0.458

Question 2.3:

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

Answer:

Molarity is given by:

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

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

= 291 g mol^{-1}

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

= 0.103 mol

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

= 0.023 M

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

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

= 0.015 mol

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

= 0.03 M

Question 2.4:

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

Answer:

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

= 60 g mol^{-1}

0.25 molar aqueous solution of urea means:

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

= 15 g of urea

That is,

(1000 + 15) g of solution contains 15 g of urea

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

= 36.95 g

= 37 g of urea (approximately)

Hence, mass of urea required = 37 g

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

Question 2.5:

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

Answer:

(a) Molar mass of KI = 39 + 127 = 166 g mol^{-1}

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

That is,

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

Therefore, molality of the solution $= \frac{\text{Moles of KI}}{\text{Mass of water in kg}}$

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

= 1.506 m

= 1.51 m (approximately)

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

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

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

= 83.19 mL

= 83.19 × 10^{-3} L

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

= 1.45 M

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

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

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

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

= 0.0263

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Question 2.6:

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

Answer:

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

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

= 55.56 mol

∴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 2.7:

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

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

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

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

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

= 27.78 mol of water

Now,

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

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

$n_{CO_2} = 0.042$ mol

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

= 1.848 g

Question 2.8:

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

Answer:

It is given that:

p_A^0 = 450 mm of Hg

p_B^0 = 700 mm of Hg

p_{total} = 600 mm of Hg

From Raoult's law, we have:

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

Therefore, total pressure, $p_{total} = p_A + p_B$

$$\Rightarrow p_{total} = p_A^0 x_A + p_B^0 (1 - x_A)$$
$$\Rightarrow p_{total} = p_A^0 x_A + p_B^0 - p_B^0 x_A$$
$$\Rightarrow p_{total} = (p_A^0 - p_B^0) x_A + p_B^0$$
$$\Rightarrow 600 = (450 - 700) x_A + 700$$
$$\Rightarrow -100 = -250 x_A$$
$$\Rightarrow x_A = 0.4$$

Therefore, $x_B = 1 - x_A$

= 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

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Question 2.9:

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

Answer:

It is given that vapour pressure of water, p_1^0 = 23.8 mm of Hg

Weight of water taken, w_1 = 850 g

Weight of urea taken, w_2 = 50 g

Molecular weight of water, M_1 = 18 g mol^{-1}

Molecular weight of urea, M_2 = 60 g mol^{-1}

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

Now, from Raoult's law, we have:

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

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

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

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

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

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

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

Question 2.10:

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

Answer:

Here, elevation of boiling point 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 2.11:

Calculate the mass of ascorbic acid (Vitamin C, $C_6H_8O_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5°C. K_f = 3.9 K kg mol^{-1}.

Answer:

Mass of acetic acid, w_1 = 75 g

Molar mass of ascorbic acid ($C_6H_8O_6$), M_2 = 6 × 12 + 8 × 1 + 6 × 16

= 176 g mol^{-1}

Lowering of melting point, 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 2.12:

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

Answer:

It is given that:

Volume of water, V = 450 mL = 0.45 L

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

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

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Question 2.1:

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

Answer:

Homogeneous mixtures of two or more than two components are known as solutions.

There are three types of solutions.

(i) Gaseous solution:

The solution in which the solvent is a gas is called a gaseous solution. In these solutions, the solute may be liquid, solid, or gas. For example, a mixture of oxygen and nitrogen gas is a gaseous solution.

(ii) Liquid solution:

The solution in which the solvent is a liquid is known as a liquid solution. The solute in these solutions may be gas, liquid, or solid.

For example, a solution of ethanol in water is a liquid solution.

(iii) Solid solution:

The solution in which the solvent is a solid is known as a solid solution. The solute may be gas, liquid or solid. For example, a solution of copper in gold is a solid solution.

Question 2.2:

Give an example of 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 2.3:

Define the following terms:

(i) Mole fraction

(ii) Molality

(iii) Molarity

(iv) Mass percentage.

Answer:

(i) Mole fraction:

The mole fraction of a component in a mixture is defined as the ratio of the number of moles of the component to the total number of moles of all the components in the mixture.

i.e.,

Mole fraction of a component $= \frac{\text{Number of moles of the component}}{\text{Total number of moles of all components}}$

Mole fraction is denoted by ‘x’.

If in a binary solution, the number of moles of the solute and the solvent are n_A and n_B respectively, then the mole fraction of the solute in the solution is given by,

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

Similarly, the mole fraction of the solvent in the solution is given as:

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

(ii) Molality

Molality (m) is defined as the number of moles of the solute per kilogram of the solvent. It is expressed as:

Molality (m) $= \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$

(iii) Molarity

Molarity (M) is defined as the number of moles of the solute dissolved in one Litre of the solution.

It is expressed as:

Molarity (M) $= \frac{\text{Moles of solute}}{\text{Volume of solution in Litre}}$

(iv) Mass percentage:

The mass percentage of a component of a solution is defined as the mass of the solute in grams present in 100 g of the solution. It is expressed as:

Mass % of a component $= \frac{\text{Mass of component in solution}}{\text{Total mass of solution}} \times 100$

Question 2.4:

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:

Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in an aqueous solution. This means that 68 g of nitric acid is dissolved in 100 g of the solution.

Molar mass of nitric acid (HNO_3) = 1 × 1 + 1 × 14 + 3 × 16 = 63 g mol^{-1}

Then, number of moles of HNO_3 $= \frac{68}{63}$ mol

$= 1.079$ mol

Given,

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

∴ Volume of 100 g solution = $\frac{100}{1.504}$ mL

$= 66.49$ mL
$= 66.49 \times 10^{-3}$ L

Molarity of solution $= \frac{1.079 \text{ mol}}{66.49 \times 10^{-3} \text{ L}}$

$= 16.23$ M

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Question 2.5:

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

Answer:

10% w/w solution of glucose in water means that 10 g of glucose in present in 100 g of the solution i.e., 10 g of glucose is present in (100 − 10) g = 90 g of water.

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

Then, number of moles of glucose $= \frac{10}{180}$ mol

= 0.056 mol

∴ Molality of solution $= \frac{0.056 \text{ mol}}{0.09 \text{ kg}}$ = 0.62 m

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

= 5 mol

⇒ Mole fraction of glucose $(x_g) = \frac{0.056}{0.056 + 5}$
$= 0.011$

And, mole fraction of water $x_w = 1 - x_g$

= 1 − 0.011

= 0.989

If the density of the solution is $1.2\ \text{g mL}^{-1}$, then the volume of the 100 g solution can be given as:

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

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

= 0.67 M

Question 2.6:

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:

Let the amount of Na_2CO_3 in the mixture be x g.

Then, the amount of $NaHCO_3$ in the mixture is (1 – x) g.

Molar mass of $Na_2CO_3 = 2 \times 23 + 1 \times 12 + 3 \times 16$

$= 106\ \text{g mol}^{-1}$

∴ Number of moles Na_2CO_3 $= \frac{x}{106}\ \text{mol}$

Molar mass of $NaHCO_3 = 1 \times 23 + 1 \times 1 \times 12 + 3 \times 16$

$= 84\ \text{g mol}^{-1}$

∴ Number of moles of $NaHCO_3$ $= \frac{1-x}{84}\ \text{mol}$

According to the question,

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

$\Rightarrow 84x = 106 - 106x$

$\Rightarrow 190x = 106$

$\Rightarrow x = 0.5579$

Therefore, number of moles of Na_2CO_3 $= \frac{0.5579}{106}\ \text{mol}$

= 0.0053 mol

And, number of moles of $NaHCO_3$ $= \frac{1-0.5579}{84}$

= 0.0053 mol

HCl reacts with Na_2CO_3 and $NaHCO_3$ according to the following equation.

$$\begin{array}{ll} 2HCl + Na_2CO_3 \longrightarrow 2NaCl + H_2O + CO_2 \\ \text{2 mol} \quad \text{1 mol} \end{array}$$

$$\begin{array}{ll} HCl + NaHCO_3 \longrightarrow NaCl + H_2O + CO_2 \\ \text{1 mol} \quad \text{1 mol} \end{array}$$

1 mol of Na_2CO_3 reacts with 2 mol of HCl.

Therefore, 0.0053 mol of Na_2CO_3 reacts with 2 × 0.0053 mol = 0.0106 mol.

Similarly, 1 mol of $NaHCO_3$ reacts with 1 mol of HCl.

Therefore, 0.0053 mol of $NaHCO_3$ reacts with 0.0053 mol of HCl.

Total moles of HCl required = (0.0106 + 0.0053) mol

= 0.0159 mol

In 0.1 M of HCl,

0.1 mol of HCl is preset in 1000 mL of the solution.

Therefore, 0.0159 mol of HCl is present in $\frac{1000 \times 0.0159}{0.1}$ mol

= 159 mL of the solution

Hence, 159 mL of 0.1 M of HCl is required to react completely with 1 g mixture of $Na_2CO_3$and $NaHCO_3$, containing equimolar amounts of both.

Question 2.7:

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:

Total amount of solute present in the mixture is given by,

$$300 \times \frac{25}{100} + 400 \times \frac{40}{100}$$

= 75 + 160

= 235 g

Total amount of solution = 300 + 400 = 700 g

Therefore, mass percentage (w/w) of the solute in the resulting solution, $= \frac{235}{700} \times 100\%$

= 33.57%

And, mass percentage (w/w) of the solvent in the resulting solution,

= (100 – 33.57)%

= 66.43%

Question 2.8:

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 $\left[C_2H_4(OH)_2\right]$ = 2 × 12 + 6 × 1 + 2 ×16

= 62 $gmol^{-1}$

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

= 3.59 mol

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

= 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 $= \frac{422.6\text{ g}}{1.072\text{ g mL}^{-1}}$

= 394.22 mL

= 0.3942×10^{-3} L

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

= 9.11 M

Question 2.9:

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

(i) express this in percent by mass

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

Answer:

(i) 15 ppm (by mass) means 15 parts per million (10^6) of the solution.

Therefore, percent by mass $= \frac{15}{10^6} \times 100\%$

$= 1.5 \times 10^{-3}$ %

(ii) Molar mass of chloroform ($CHCl_3$) = 1 × 12 + 1 × 1 + 3 × 35.5

$= 119.5 \text{ g mol}^{-1}$

Now, according to the question,

15 g of chloroform is present in 10^6 g of the solution.

i.e., 15 g of chloroform is present in $(10^6 - 15) \approx 10^6$ g of water.

∴ Molality of the solution $= \dfrac{\frac{15}{119.5} \text{ mol}}{10^6 \times 10^{-3} \text{ kg}}$

$= 1.26 \times 10^{-4}$ m

Question 2.10:

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 2.11:

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

Answer:

Solubility of gases in liquids decreases with an increase in temperature. This is because dissolution of gases in liquids is an exothermic process.

$$\text{Gas} + \text{Liquid} \longrightarrow \text{Solution} + \text{Heat}$$

Therefore, when the temperature is increased, heat is supplied and the equilibrium shifts backwards, thereby decreasing the solubility of gases.

Question 2.12:

State Henry’s law and mention some important applications?

Answer:

Henry’s law states that partial pressure of a gas in the vapour phase is proportional to the mole fraction of the gas in the solution. If p is the partial pressure of the gas in the vapour phase and x is the mole fraction of the gas, then Henry’s law can be expressed as:

$p = K_H x$

Where,

K_H is Henry’s law constant

Some important applications of Henry’s law are mentioned below.

(i) Bottles are sealed under high pressure to increase the solubility of CO_2 in soft drinks and soda water.

(ii) Henry’s law states that the solubility of gases increases with an increase in pressure. Therefore, when a scuba diver dives deep into the sea, the increased sea pressure causes the nitrogen present in air to dissolve in his blood in great amounts. As a result, when he comes back to the surface, the solubility of nitrogen again decreases and the dissolved gas is released, leading to the formation of nitrogen bubbles in the blood. This results in the blockage of capillaries and leads to a medical condition known as ‘bends’ or ‘decompression sickness’.

Hence, the oxygen tanks used by scuba divers are filled with air and diluted with helium to avoid bends.

(iii) The concentration of oxygen is low in the blood and tissues of people living at high altitudes such as climbers. This is because at high altitudes, partial pressure of oxygen is less than that at ground level. Low-blood oxygen causes climbers to become weak and disables them from thinking clearly. These are symptoms of anoxia.

Question 2.13:

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

Answer:

Molar mass of ethane (C_2H_6) = 2 × 12 + 6 × 1

= 30 g mol^{-1}

∴Number of moles present in 6.56×10^{-3} g of ethane $= \frac{6.56 \times 10^{-3}}{30}$

= 2.187×10^{-4} mol

Let the number of moles of the solvent be *x.*

According to Henry's law,

$p = K_H x$

$$\Rightarrow 1\,bar = K_H \cdot \frac{2.187 \times 10^{-4}}{2.187 \times 10^{-4} + x}$$

$$\Rightarrow 1\,bar = K_H \cdot \frac{2.187 \times 10^{-4}}{x} \quad (Since\ x >> 2.187 \times 10^{-4})$$

$$\Rightarrow K_H = \frac{x}{2.187 \times 10^{-4}} bar$$

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

= 1.67×10^{-3} mol

According to Henry's law,

$p = K_H x$

$$= \frac{x}{2.187 \times 10^{-4}} \times \frac{1.67 \times 10^{-3}}{(1.67 \times 10^{-3}) + x}$$

$$= \frac{x}{2.187 \times 10^{-4}} \times \frac{1.67 \times 10^{-3}}{x} \qquad (\sin ce\ x >> 1.67 \times 10^{-3})$$

= 7.636 bar

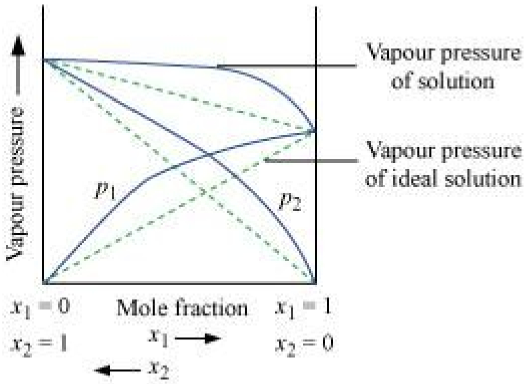
Hence, partial pressure of the gas shall be 7.636 bar.

Question 2.14:

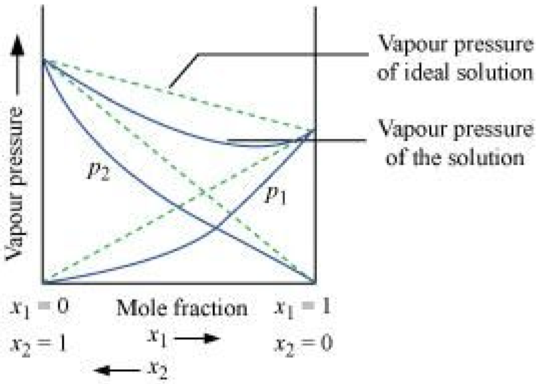
What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{sol}H$ related to positive and negative deviations from Raoult's law?

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.

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

Question 2.15:

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 $\left(p_1^0\right) = 1.013 \text{ bar}$

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 2.16:

Heptane and octane form an ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35 g of octane?

Answer:

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

Vapour pressure of octane $\left(p_2^0\right)$ = 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 $= \frac{26}{100} \text{ mol}$

= 0.26 mol

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

= 114 g mol^{-1}

∴Number of moles of octane $= \frac{35}{114}$ 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 2.17:

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

Answer:

1 molal solution means 1 mol of the solute is present in 1000 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 2.18:

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 be 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 × 12 + 18 × 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 2.19:

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:

1. molar mass of the solute
2. 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:

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

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

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

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

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

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

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

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

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

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

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

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

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 2.20:

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

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 sugar $=\frac{5}{342}\text{mol}$

= 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 2.21:

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 Kwhereas 1.0 g of AB_4lowers 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.

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Question 2.22:

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 2.23:

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 2.24:

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 2.25:

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

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

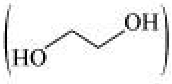
(iv) ethylene glycol (v) chloroform (vi) pentanol.

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 non-polar --C_5H_{11} group. Thus, pentanol is partially soluble in water.

Question 2.26:

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 2.27:

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}.

$$\begin{array}{llll} CuS & \leftrightarrow Cu^{2+} & + & S^{2-} \\ & \quad s & & s \end{array}$$

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

$= s \times s$

$= s^2$

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 2.28:

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 2.29:

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 $= 1.5 \times 10^{-3} \times 311 \text{ g}$

$= 0.4665 \text{ g}$

Therefore, total mass of the solution $= (1000 + 0.4665) \text{ g}$

$= 1000.4665 \text{ g}$

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 2.30:

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

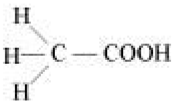
Hence, required benzoic acid = 0.0375 mol × 122 g mol^{-1}

= 4.575 g

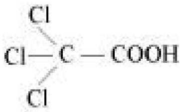
Question 2.31:

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.

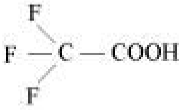
Answer:



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 2.32:

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 $CH_3CH_2CHClCOOH = 15+14+13+35.5+12+16+16+1$

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

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

$= 0.0816 \text{ mol}$

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

$\therefore$Molality of the solution, $= \dfrac{0.0186}{250} \times 1000$

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

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

$CH_3CH_2CHClCOOH$ undergoes dissociation according to the following equation:

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

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

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

Since α is very small with respect to 1, 1 − α ≈ 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,

$$\begin{array}{llll} & \mathrm{CH_3CH_2CHClCOOH} & \leftrightarrow \mathrm{CH_3CH_2CHClCOO^-} & + \mathrm{H^+} \\ \text{Initial moles} & 1 & 0 & 0 \\ \text{At equilibrium} & 1-\alpha & \alpha & \alpha \end{array}$$

Total moles of equilibrium = 1 − α + α + α

= 1 + α

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

$$\begin{aligned} &= 1+\alpha \\ &= 1+0.0655 \\ &= 1.0655 \end{aligned}$$

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

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

$$\begin{aligned} &= 1.0655 \times 1.86\,\mathrm{K\,kg\,mol^{-1}} \times 0.3264\,\mathrm{mol\,kg^{-1}} \\ &= 0.65\ \mathrm{K} \end{aligned}$$

Question 2.33:

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:

$$\begin{aligned} w_1 &= 500\text{ g} \\ w_2 &= 19.5\text{ g} \\ K_f &= 1.86\text{ K kg mol}^{-1} \\ \Delta T_f &= 1\text{K} \end{aligned}$$

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 $\mathrm{CH_2FCOOH}, (M_2)_{\text{obs}} = 72.54\text{ g mol}$

The calculated molar mass of $\mathrm{CH_2FCOOH}$ is:

$$\begin{aligned} (M_2)_{\text{cal}} &= 14 + 19 + 12 + 16 + 16 + 1 \\ &= 78\text{ g mol}^{-1} \end{aligned}$$

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

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

$$= 1.0753$$

Let α be the degree of dissociation of $\mathrm{CH_2FCOOH}$

	$\mathrm{CH_2FCOOH} \leftrightarrow$	$\mathrm{CH_2FCOO^-}$ +	$\mathrm{H^+}$	
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:

$$K_a = \frac{[\mathrm{CH_2FCOO^-}][\mathrm{H^+}]}{[\mathrm{CH_2FCOOH}]}$$

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

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

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

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

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

Therefore,

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

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

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

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

Question 2.34:

Vapour pressure of water at 293 Kis 17.535 mm Hg. Calculate the vapour pressure of water at 293 Kwhen 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, $n_2 = \frac{25}{180 \text{ g mol}^{-1}}$

= 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,

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

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

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

⇒ 17.535 − p_1 = 0.097

⇒ p_1 = 17.44 mm of Hg

Hence, the vapour pressure of water is 17.44 mm of Hg.

Question 2.35:

Henry's law constant for the molality of methane in benzene at 298 Kis 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 × 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 × 10^{-5}

= 178 × 10^{-5} (approximately)

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

Question 2.36:

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.

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Question 2.37:

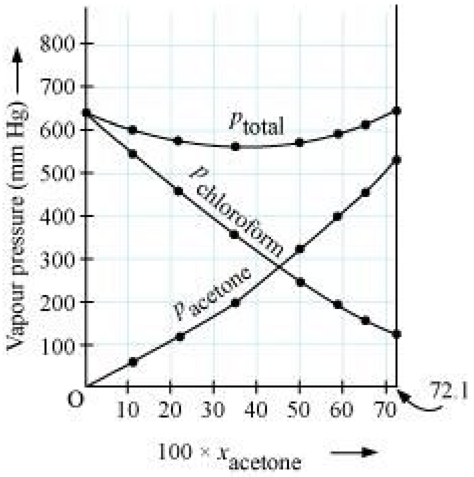
Vapour pressure of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot p_{total}’ $p_{chloroform}$’ and $p_{acetone}$ as a function of $x_{acetone}$. The experimental data observed for different 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

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
$p_{chloroform}$/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 2.38:

Benzene and toluene form ideal solution over the entire range of composition. The vapour pressure of pure benzene and toluene 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 toluene, $p_t = x_t \times p_t$

$= 0.514 \times 32.06$

$= 16.479 \text{ mm Hg}$

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 2.39:

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 Kif 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.

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:

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

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

$$= \frac{1520\,\text{mm Hg}}{3.30\times10^7\,\text{mm Hg}} \qquad \left(\text{Given } K_H = 3.30\times10^7\,\text{mm Hg}\right)$$

$$= 4.61\times10^{-5}$$

For nitrogen:

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

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

$$= \frac{6004\,\text{mm Hg}}{6.51\times10^7\,\text{mm Hg}}$$

$$= 9.22\times10^{-5}$$

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

Question 2.40:

Determine the amount of CaCl2 (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,

$$\pi = i\frac{n}{V}\text{R}T$$

$$\Rightarrow \pi = i\frac{w}{MV}\text{R}T$$

$$\Rightarrow w = \frac{\pi MV}{i\text{R}T}$$

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

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

$i = 2.47$

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

Here,

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

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 2.41:

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, $K^+ \text{ and } SO_4^{2-}$ 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,

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