NCERT SOLUTIONS

CLASS - 12th



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Class : 12th
Subject : Chemistry
Chapter : 2
Chapter Name : Solutions

Q2.1 Calculate the mass percentage of benzene ($\left(C_6H_6\right)$) and carbon tetrachloride $\left(CC1_4\right)$) 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 %

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Q2.2 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.
$\therefore$ Mass of carbon tetrachloride = (100 — 30) g
= 70 g
Molar mass of benzene (C_6H_6) = $(6 \times 12 + 6 \times 1) g\text{mol}^{-1}$

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

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

$= 154\ g\ mol^{-1}$

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

$= 0.4545 mol$

Thus, the mole fraction of C_6H_6 is given as:

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

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

$= 0.458$

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Q2.3 Calculate the molarity of each of the following solutions: (a) 30 g of $Co\left[NO_3\right]_2 . 6H_2O$ 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) $Co\left(NO_3\right)_2 \cdot 6H_2O = 59 + 2(14 + 3 \times 16) + 6 \times 18$

$= 291 g mol^{-1}$

$\therefore$ Moles of $Co\left(NO_3\right)_2 \cdot 6H_2O = \frac{30}{291} mol$

$= 0.103$ mol

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

$= 0.023$ M

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

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

$= 0.015$ mol

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

$= 0.03$ M

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Q2.4 Calculate the mass of urea $\left(NH_2CONH_2\right)$ required in making 2.5 kg of 0.25 molal aqueous solution.

Answer.

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

$= 60 gmol^{-1}$

0.25 molar aqueous solution of urea means:

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

= 15g of urea

That is,

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

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

= 36.95 g

= 37 g of urea (approximately)

Hence, mass of urea required = 37 g

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Q2.5 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 = $166 gmol^{-1}$

(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 - 200) g of water = 80 g of water

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

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

= 1.506 m

= 1.51 m (approximately)

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

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

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

$= 83.19 mL$

$= 83.19 \times 10^{-3} L$

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

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

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

$= 0.0263$

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Q2.6 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 = $= \frac{1000\text{g}}{18\text{gmol}^{-1}}$

= 55.56 mol

$\therefore$ 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:

$p = K_H x$

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

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

$= 282\text{bar}$

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Q2.7 Henry's law constant for CO_2 in water is 1.67×10^8 Pa at 298 K. Calculate the quantity of in 500 mL of soda water when packed under 2.5 atm CO_2 pressure at 298 K.

Answer. given that

Pressure of $CO_2 = 2.5\text{atm}$

$1\text{ atm} = 1.01325 \times 10^5$ Pa So that

Pressure of $CO_2 = 2.5 \times 1.01325 \times 10^5\text{Pa} = 2.533125 \times 10^5\text{Pa}$

$K_H = 1.67 \times 10^8\text{Pa}$

According to Henry's law:

$p = K_H \times X$

$X = p/K_H = 2.533125 \times 10^5/1.67 \times 10^8 = 1.52 \times 10^{-3}$

But

We have 500 mL of soda water so that

Volume of water = 500 mL

[Neglecting the amount of soda present]

Density of water = 1 g/ ml

Use formula mass = volume × density we get

500 mL of water = 500 g of water

$\text{Molar mass of water } \left(H_2O\right) = 18\text{gmol}^{-1}$

Use formula

Number of mole of water = 500 / 18 = 27.78 mol of water

Now use the formula of mole fraction

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

Value of mole fraction is very small so it is negligible as compared to 1

We get

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

$1.52 \times 10^{-3} = n_{CO2}/27.78$

After calculation we get

$n_{CO2} = 0.042\text{mol}$

molar mass of $CO_2 = 12 + 2 \times 16 = 44\text{gmol}_{-1}$

use formula

mass = molar mass \times number of moles

= 1.848g

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Q2.8 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:

$P_A = P_A^0 x_A$

$p_B = p_B^0 x_B = p_B^0\left(1 - x_A\right)$ Therefore, total pressure, $P_{total} = P_A + P_B$

$\Rightarrow P_{\text{total}} = P_A^0 x_A + P_B^0 \left(1 - x_A\right)$

$\Rightarrow P_{\text{total}} = P_A^0 x_A + P_B^0 - P_B^0 x_A$

$\Rightarrow P_{\text{total}} = \left(P_A^0 - P_B^0\right) x_A + P_B^0$

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

$\Rightarrow -100 = -250 x_A$

$\Rightarrow x_A = 0.4$

Therefore, $x_B = 1 - x_A$

= 1- 0.4

= 0.6

Now, $p_A = p_A^0 x_A$

= 450 x 0.4

= 180 mm of Hg

$p_B = p_B^0 x_B$

$= 700 \times 0.6$

= 420mm of Hg

Now, in the vapour phase :

Mole of fraction of liquid A = $= \frac{P_A}{P_A + P_B}$

$= \frac{180}{180 + 420}$

$= \frac{180}{600}$

= 0.30

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Q2.9 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 = 5O g

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

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

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:

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

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

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

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

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

$\Rightarrow p_1 = 23.4$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.

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Q2.10 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_1 = 500 g
Molar mass of sucrose $\left(C_{12}H_{22}O_{11}\right)$, M_2 = 11 x 12 + 22 x 1 + 11 x 16
= 342 g mol^{-1}
Molal elevation constant, Kb = 0.52 K kg mol-
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 s to be added.

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Q2.11 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 $\left(C_6H_8O_6\right)$, M_2 = 6 × 12 + 8 × 1 + 6 × 16

= 176 g mol^{-1}

Lowering of melting point, ΔT_f = 1.5 K

We know that:

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

$$\Rightarrow w_2 = \frac{\Delta T_f \times M_2 \times w_1}{K_f \times 1000} = \frac{1.5 \times 176 \times 75}{3.9 \times 1000}$$

$$= 5.08\text{g(approx)}$$

Hence, 5.08 g of ascorbic acid is needed to be dissolved.

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Q2.12 Calculate the osmotic pressure in pascals exerted by a solution prepared by dissolving 1.0 g of polymer of molar mass 185,000 in 450 mL of water at 37°C.

Answer. It is given that:
Volume of water, V = 450 mL = 0.45 L
Temperature, T = (37 + 273) K = 310 K

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

We know that:

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

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

= 30.98Pa

= 31pa(approximately)

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Q2.1 Define the term solution. How many types of solutions are formed? Write briefly about each type with an example.

Answer. A solution is a homogenous mixture of two or more than two substances whose composition can change within a certain limits. A solution of two substances is called binary solution.

In binary solution, the component which is present in smaller amount is called solute and other one is called solvent.
Types of solutions and there example:-
first component is solute and second one solvent
1.Solid in solid - stainless steel (C and Fe)
2.Liquid in solid - water and sodium carbonate (10H2O+Na2CO3)

3.Gas in solid - H2 in Pd
4.Solid in liquid - salt water
5.Liquid in liquid - alcohol in water
6.Gas in liquid - O2 in water
7.Solid in gas - Camphor in N2 gas
8.Liquid in gas - moisture
9.Gas in gas - O2 and N2 gas in air

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Q2.2 Give an example of a solid solution in which the solute is a gas.

Answer. Solution of hydrogen and palladium. In this hydrogen which is a gas is solute and palladium is solvent.

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Q2.3 Define the following terms:
(i) Mole fraction
(ii) Molality
(iii) Molarity
(iv) Mass percentage

Answer. (i)Mole fraction –Mole fraction is a way to express concentration of a solution.
Mole fraction of a constituent (either of solute or solvent) is the ratio of number of moles of one component to the total number of moles of all component present in a solution.
Mole fraction is denoted by letter, 'x'.
(ii) Molality: Molality is a way to express concentration of solution.

$$= \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$$

(iii) Molarity:

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

(iv) Mass percentage

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

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Q2.4 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 $\left(HNO_3\right) = 1 \times 1 + 1 \times 14 + 3 \times 16 = 63 g mol^{-1}$

Then, number of moles of $HNO_3 = 68/63 mol$

= 1.08 mol

$= 1.504 g/mL^{-1}$

Therefore from the formula density = mass / volume,we get

Volume of solution = 1000/1.504 = 66.49 mL

Therefore molarity of nitric acid = (1.08/66.49) x 1000 = 16.24

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Q2.5 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 $\left(C_6H_{12}O_6\right) = 6 \times 12 + 12 \times 1 + 6 \times 16 = 180 g mol^{-}$

1

Then, number of moles of glucose $= 10/180 mol$

$= 0.056 mol$

$\therefore$ Molality of solution $= 0.056 mol/0.09 kg = 0.62 m$

Number of moles of water $= 90 g/18 g$ mol 1 = 5mol

Mole fraction of glucose $\left(x_g\right) = 0.056/(0.056 + 5) = 0.011$

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

$= 1 - 0.011 = 0.989$

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

$= 100 g/1.2 gmL^{-1}$

$= 83.33 mL$

$= 83.33 \times 10^{-3} L$

$\therefore$ Molarity of the solution $= 0.056 mol/83.33 \times 10^{-3} L$

$= 0.67 M$

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Q2.6 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 be x

& that of $NaHCO_3$ be $1 - x$

Now moles of Na_2CO_3 = $\times$ /106

& moles of $NaHCO_3$ = $1 - x/84$

Now according to question, number of moles of Na_2Co_3 = number of

moles of $NaHCO_3$

Therefore $\times 106 = 1 - x/84$

$84x = 106 - 106x$

$84x + 106x = 106$

$190x = 106$

&

moles of $NaHCO_3 = 1 - 0.558/84 = 0.0053$

Now Hcl reacts with Na_2Co_3&$NaHCO_3$ as follows:

$Na_2Co_3 + 2Hcl \longrightarrow 2Nacl + H_2o + CO_2$

$NaHCO_3 + Hcl \longrightarrow Nacl + H_2o + CO_2$

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Q2.7 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,

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

$= 75 + 160$

$= 235g$

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%

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Q2.8 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 $1.072gmL^{-1}$, then what shall be the molarity of the solution?

Answer.

Calculation of Molality:

Mass of ethylene glycol = 222.6 (Given)

Molar mass of ethylene glycol $\left[C_2H_4(OH)_2\right]$

$= 2 \times 12 + 6 \times 1 + 2 \times 16$

$= 62$

Therefore moles of ethylene glycol

$= 222.6g/62gmol^{-1}$

$= 3.59mol$

Mass of water $= 200g$ (Given)

Therefore molality of the solution is = (moles of ethylene glycol / mass of water) x 1000

$= (3.59/200) \times 1000$

$= 17.95m$

Calculation of Molarity:

Moles of ethylene glycol = 3.59 mol (already calculated)

Total Mass of solution = 200 + 222.6

= 422.6g

Volume of solution = mass / density volume

= 422.6 / 1.072

= 394.22 ml

now molarity of the solution is = (moles of ethylene glycol / volume of solution) x 1000

= (3.59 / 394.22) x 1000

= 9.11 M

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Q2.9 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.

1) 15 ppm means : 15 parts per million $\left(10^6\right)$ of the solutions

So, Percent by mass $=$ (mass of chloroform / total mass) $\times 100$

$= \left(15/10^6\right) \times 100$

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

2) Molality Mass of chloroform = 15g

Molar mass of chloroform $\left(CHCl_3\right)$ = 1 × 12 + 1 × 1 + 3 × 35.5

= 119.5gmol^{-1}

Moles of chloroform = 15/119.5 = 0.1255mol

Mass of water = 10^6

Therefore molality = (moles of chloroform / mass of water) × 1000 = $\left(0.1255/10^6\right)$ × 1000 = 1.255 × 10^{-4}m

$\text{Therefore molality} = (\text{moles of chloroform} / \text{mass of water}) \times 1000$

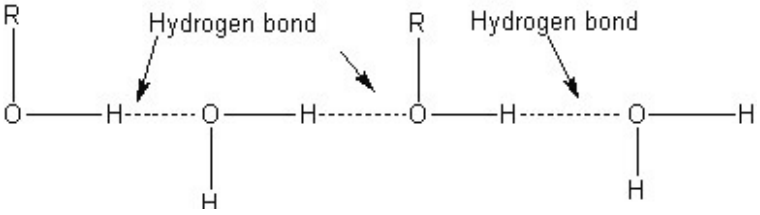
$= \left(0.1255/10^6\right) \times 1000$

$= 1.255 \times 10^{-4}\text{m}$

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Q2.10 What role does the molecular interaction play in a solution of alcohol and water?

Answer. The lower members of alcohols are highly soluble in water but the solubility decreases with increase in the molecular weight. The solubility of lower alcohols in water is due to formation of hydrogen bonds(Hydrogen bonding) between alcohols & water molecules.



However, as the size of alcohol molecule increases, the alkyl groups becomes larger & prevents the formation of hydrogen bonds with water, & hence the solubility goes on decreasing with increase in the length of carbon chain.Also 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.

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Q2.11 Why do gases always tend to be less soluble in liquids as the temperature is raised?

Answer. The dissolution of a gas in a liquid is exothermic process.Therefore according to Le Chatelier principle,with the increase in temperature,the equilibrium shifts in the backward direction.
Gas + Liquid --> Solution + Heat
Therefore the solubility of gas in solution decreases with rise in temperature.

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Q2.12 State Henry's law and mention some important applications.

Answer. Henry's law states that the mass of a gas dissolved per unit volume of the solvent at a given temperature is proportional to the pressure of the gas in equilibrium with the solution.
Or
It also states that the pressure of a gas over a solution in which the gas is dissolved is proportional to the mole fraction of the gas dissolved in the solution.
The important applications of Henry's law are as follows:
1) In the production of carbonated beverages-in order to increase the solubility of CO2 in cold drinks,beer etc,the bottle are sealed under high pressure.When the bottle is opened under normal atmospheric pressure,the pressure inside the bottle falls to atmospheric pressure & the excess CO2 bubbles out of the bottle causing effervescence.
2) At high altitudes-the partial pressure of oxygen at high altitudes is less than the ground level.this results in low concentration of oxygen in the blood & tissues of the peoples
3) In scuba diving- during scuba diving,when the diver breaths in compressed air from the supply tank,more nitrogen dissolves in the blood & other body fluids because the pressure at that depth is far greater than the surface atmospheric pressure.

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Q2.13 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 $\left(C_2H_6\right) = 2 \times 12 + 6 \times 1$

$= 30 gmol^{-1}$

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

$= 2.187 \text{ x } 10^{-3}$mol

Let the number of moles of the solvent be x.

According to Henry's law,

$p = K_H x$

$\Rightarrow 1 \text{ bar } = K_H \cdot \frac{2.187 \times 10^{-3}}{2.187 \times 10^{-3} + x}$

$\Rightarrow 1 \text{ bar } = K_H \frac{2.187 \times 10^{-3}}{x}$

$\Rightarrow K_H = \frac{x}{2.187 \times 10^{-3}} \text{ bar}$

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

$= = 1.67 \times 10^{-3}$mol

According to Henry's law,

$p = K_H x$

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

$$= \frac{x}{2.187 \times 10^{-3}} \times \frac{1.67 \times 10^{-3}}{x}$$

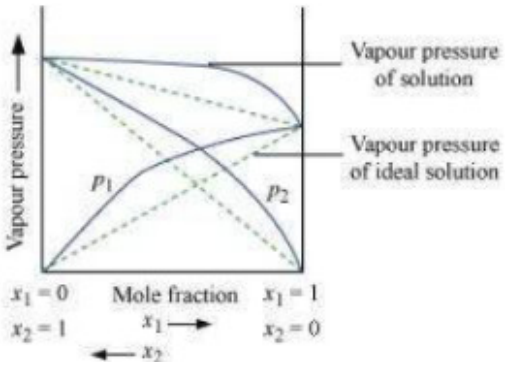
= 0.764 bar

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

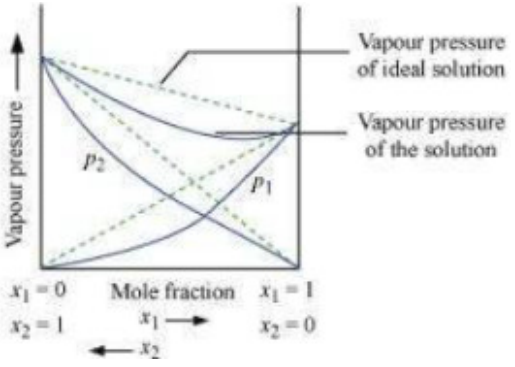
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Q2.14 What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{mix}H$ related to positive and negative deviations from Raoult's law?

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.

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

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

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

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Q2.15 An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 barat 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 $\left(p_1\right)$ = 1.004bar

Vapour pressure of pure water at normal boiling point $\left(p_1^0\right)$ = 1.013bar

Mass of solute, $\left(w_2\right)$ = 2 g

Mass of solvent (water), $\left(w_1\right)$ = 98 g

Molar mass of solvent (water), $\left(M_1\right)$ = $18gmol^{-1}$

According to Raoult's law,

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

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

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

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

$= 41.35gmol^{-1}$

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

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Q2.16 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. kPa

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

Vapour pressure of octane $\left(p_2^0\right)_{=46.8kPa}$

We know that,

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

$= 100 \text{gmol}^{-1}$

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

$= 0.26$ mol

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

$= 114 \text{gmol}^{-1}$

$\therefore$ Number of moles of octane $= \frac{35}{114} \text{mol}$

$= 0.31$ 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, $p_1 = x_1 p_1^0$

$= 0.456$ x 105.2

$= 47.97$ kPa

Now, partial pressure of octane, $p_2 = x_2 p_2^0$

$= 0.544$ x 46.8

$= 25.46$ kPa

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

$= 47.97 + 25.46$

$= 73.43$ kPa

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Q2.17 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 $= 18 \text{gmol}^{-1}$

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 the water, $p_1^0 = 12.3 \text{kPa}$

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

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

$\Rightarrow 12.3 - p_1 = 0.2177$

$\Rightarrow p_1 = 12.0823$

$= 12.08\text{kPa}$ (approximately)

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

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Q2.18 Calculate the mass of a non-volatile solute (molar mass 40 g $mo1^{-1}$) 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 80

$\frac{80}{100}p_1^0 = 0.8p_1^0$

Molar mass of solute, $M_2 = 40\text{gmol}^{-1}$

Mass of octane, $w_1 = 114\text{g}$

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

$= 114\text{gmol}^{-1}$

Applying the relation,

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

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

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

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

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

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

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Q2.19 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 $\mathrm{Mgmol^{-1}}$

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

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

$p_1 = 2.8\mathrm{kPa}$

Applying the relation:

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

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

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

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

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

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

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

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

After the addition of 18 g of water:

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

$p_1 = 2.9\mathrm{kPa}$

Again, applying the relation:

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

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

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

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

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

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

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

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

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

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

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

$$\Rightarrow 2.9 \times 5 \times (6M + 30) = 2.8 \times 6 \times (5M + 30)$$

$$\Rightarrow 87M + 435 = 84M + 504$$

$$\Rightarrow 3M = 69$$

$$\Rightarrow M = 23u$$

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.

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Q2.20 A 5% solution (by mass) of cane sugar in water has freezing point of 271K.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

= 2.15K

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

$= 342 \text{gmol}^{-1}$

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, $m = \frac{0.0146\text{mol}}{0.095\text{kg}}$

$= 0.1537 \text{molkg}^{-1}$

Applying the relation,

$$\Delta T_f = K_f \times m$$

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

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

$$= 13.99\text{Kkgmol}^{-1}$$

Molar of glucose $\left(C_6H_{12}O_6\right) = 6 \times 12 + 12 \times 1 + 6 \times 16$

$= 180 \text{gmol}^{-1}$

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

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

= 0.0278 mol

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

$= 0.2926 \text{molkg}^{-1}$

Applying the relation,

$\Delta T_f = K_f \times m$

$= 13.99\text{Kkgmol}^{-1} \times 0.2926\text{molkg}^{-1}$

= 4.09K (approximately)

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

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Q2.21 Two elements A and B form compounds having formula AB_2 and AB_4. When dissolved in 20 g of benzene $\left(C_6H_6\right)$, 1 g of AB_2 lowers the freezing point by 2.3 K whereas 1.0 g of AB_4 lowers it by 1.3 K. The molar depression constant for benzene is 5.1 K kg mol^{-1}. Calculate atomic masses of A and B.

Answer. We know that,

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

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

$= 110.87 \text{gmol}^{-1}$

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

Now, we have the molar masses of AB_2 and AB_4 as 110.87gmol^{-1} and 196.15gmol^{-1} respectively.
Let the atomic masses of A and B be x and y respectively.
Now, we can write:
x + 3y = 110.87 (i)
x + 4y = 196.15 (ii)
Subtracting equation (i) from (ii), we have
2y = 85.28
$\Rightarrow y = 42.64$
Putting the value of 'y' in equation (i), we have
x +2 x 42.64 = 110.87
$\Rightarrow x = 25.59$
Hence, the atomic masses of A and 3 are 25.59 u and 42.64 u respectively.

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Q2.22 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
n = 1.52 bar
R = 0.083 bar $LK^{-1}mol^{-1}$
Applying the relation,
n = CRT
$\Rightarrow C = \frac{\pi}{RT}$
$= \frac{1.52\text{bar}}{0.083 \text{ bar LK}^{-1}\text{mol}^{-1} \times 300\text{K}}$
= 0.061 mol
Since the volume of the solution is 1 L, the concentration of the solution would be 0.061 M.

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Q2.23 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 $\left(CH_3CN\right)$ and acetone $\left\{C_3H_6O\right\}$

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.

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Q2.24 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 < CH3OH < CH3CN < Cyclohexane

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Q2.25 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.

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Q2.26 If the density of some lake water is $1.25 g mL^{-1}$ and contains 92g 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 = $\frac{92g}{23gmol^{-1}}$

= 4 mol

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

= 4 m

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Q2.27 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 $molL^{-1}$

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

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

= s x s

$= s^2$

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

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

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

Hence, the maximum molarity of Cus in an aqueous solution is 2.45 x $10^{-8} molL^{-1}$.

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Q2.28 Calculate the mass percentage of aspirin $\left(C_9H_8O_4\right)$ in acetonitrile $\left(CH_3CN\right)$ when 6.5 g of $C_2H_8O_4$ 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%

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Q.2.29 Nalorphene $\left(C_{19}H_{21}NO_3\right)$, 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 $\left(C_{19}H_{21}NO_3\right)$ is given as:

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

In 1.5×10^{-3}m aqueous solution of nalorphene,
1kg(1000g) 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:

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

= 3.22 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.

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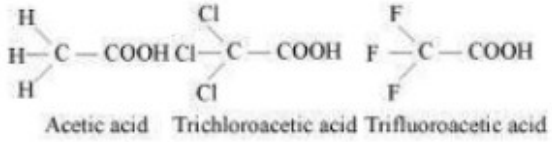
Q2.30 Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing 250 mL of 0.15 M solution in methanol.

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 $\left(C_6H_5COOH\right)$ = 7 x 12 + 6 x 1 + 2 x 16
= 122 g mol^{-1}
Hence, required benzoic acid = 0.0375 mol x 122gmol^{-1}
= 4.575 g

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Q2.31 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.



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 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

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Q2.32 Calculate the depression in the freezing point of water when 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water. $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86\ \text{Kkgmol}^{-1}$.

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 10g of $CH_3CH_2CHCICOOH = \frac{10\text{g}}{122.5\text{gmol}^{-1}}$
$= 0.0816$ mol

It is given that 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water.
$\therefore$ Molality of the solution, $= \frac{0.0186}{250} \times 1000$
$= 0.3264\ \text{mol molkg}^{-1}$

Let α be the degree of dissociation of $CH_3CH_2CHClCOOH$
$CH_3CH_2CHClCOOH$ undergoes dissociation according to the following equation:

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

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

$\therefore K_\alpha = \frac{C\alpha \cdot C\alpha}{C(1-\alpha)}$
$= \frac{C\alpha^2}{1-\alpha}$
Since a is very small with respect to 1, $1 - \alpha \approx 1$
Now, $K_\alpha = \frac{C\alpha^2}{1}$

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

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

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

$= 0.0655$
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:

$\Delta T_f = i.K_f m$

$= 1.0655 \times 1.86 \text{Kkgmol}^{-1} \times 0.3264 \text{molkg}^{-1}$

$= 0.65\text{K}$

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Q2.33 19.5 g of CH_2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is $1.0\ ^\circ C$. Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.

Answer. It is given that:

$w_1 = 500$ g

$w_2 = 19.5\text{g}$

$K_f = 1.86 \text{Kkgmol}^{-1}$

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

We know that:

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

$= \frac{1.86 \text{Kkgmol}^{-1} \times 19.5\text{g} \times 1000 \text{gkg}^{-1}}{500\text{g} \times 1\text{K}}$

$= 72.54 \text{gmol}^{-1}$

Therefore, observed molar mass of CH_2FCOOH - 72.54 g mol

The calculated molar mass of CH_2FCOOH is:

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

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

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

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

$$= 1.0753$$

Let α be the degree of dissociation of CH_2FCOOH

	$CH_2FCOOH \leftrightarrow CH_2FCOO^- + H^+$			
Initial conc.	C mol L^{-1}	0	0	
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$	Total $= C(1+\alpha)$

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

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

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

$$= 1.0753 - 1$$

$$= 0.0753$$

Now, the value of K_α is given as:

$$K_\alpha = \frac{\left[CH_2FCOO^-\right]\left[H^+\right]}{\left[CH_2FCOOH\right]}$$

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

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

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

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

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

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

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

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Q2.34 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 $\left(C_6H_{12}O_6\right)$, M , = 6 x 12 + 12 x 1+6 x 16
$= 180 gmol^{-1}$
Molar mass of water, $M_1 = 18 gmol^{-1}$

Then, number of moles of glucose, $n_2 = \frac{25}{180 gmol^{-1}}$
= 0.139 mol

And, number of moles of water, $n_1 = \frac{450g}{18 gmol^{-1}}$
= 25 mol
We know that,

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

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

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

$$\Rightarrow 17.535 - p_1 = 0.097$$

$$\Rightarrow p_1 = 17.44 mm \text{ of } Hg$$

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

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Q2.35 Henry's law constant for the molality of methane in benzene at 298 K is 4.27×10^5mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.

Answer. p = 760 mm Hg
k_H = = 4.27 x 10^3 mm Hg

According to Henry's law,
$p = k_H x$

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

$$= \frac{760 mmHg}{4.27 \times 10^5 mmHg}$$

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

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

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

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Q2.36 100 g of liquid A (molar mass 140gmol^{-1}) was dissolved in 1000 g of liquid B (molar mass 180gmol^{-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, $n_{\text{A}} = \frac{100}{140} \text{mol}$

= 0.714 mol

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

= 5.556 mol

Then, mole fraction of A, $x_{\text{A}} = \frac{n_{\text{A}}}{n_{\text{A}} + n_{\text{B}}}$

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

= 0.114

And, mole fraction of B, $x_{\text{B}} = 1 - 0.114$

= 0.886

Vapour pressure of pure liquid B, $p_{\text{B}}^{0} = 500$ torr

Therefore, vapour pressure of liquid B in the solution,

$p_{\text{B}} = p_{\text{B}}^{0} x_{\text{B}}$

= 500 x 0.886

= 443 torr

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

Vapour pressure of liquid A in the solution,

$p_{\text{A}} = p_{\text{total}} - p_{\text{B}}$

= 475 − 443

= 32 torr

Now,

$p_A = p_A^0 x_A$

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

$= \frac{32}{0.114}$

= 280.7 torr

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

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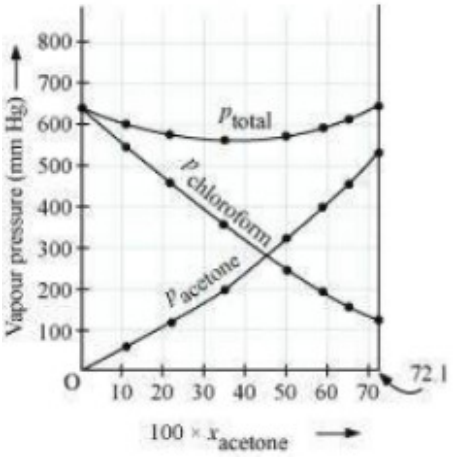
Q2.37 Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot P_{total}, $P_{\text{chloroform}}$, and P_{acetone} as a function of x_{acetone}. The experimental data observed for different compositions of mixture is:

$100 \times x_{\text{acetone}}$	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

$100 \times x_{\text{acetone}}$	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.

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Q2.38 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 $\left(C_6H_6\right)$ = 6 × 12 + 6 × 1

= 78gmol^{-1}

Molar mass of toluene $\left(C_6H_5CH_3\right)$ = 7 × 12 + 8 × 1

= 92gmol^{-1}

Now, no. of moles present in 80 g of benzene = = $\frac{80}{78}$mol = 1.026mol

And, no. of moles present in 100 g of benzene = = $\frac{100}{92}$mol = 1.087mol

∴ 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° = 50.71mmHg

And, vapour pressure of pure toluene, p_t° = 32.06mmHg

Therefore, partial vapour pressure of benzene, $p_b = x_b \times p_b$

= 0.486 × 50.71

= 24.645mmHg

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

= 0.514 × 32.06

= 16.479mmHg

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

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Q2.39 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 $\left(O_2\right)$ in air = 20%

Percentage of nitrogen $\left(\mathrm{N}_2\right)$ 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)mmHg = 7600mmHg

Therefore,

Partial pressure of oxygen, $p_{\mathrm{O}_2} = \frac{20}{100} \times 7600\text{mmHg}$

= 1520 mm Hg

Partial pressure of nitrogen, $p_{\mathrm{N}_2} = \frac{79}{100} \times 7600\text{mmHg}$

= 6004 mm Hg

Now, according to Henry's law:

$p = K_{\mathrm{H}} \cdot x$

For oxygen:

$p_{\mathrm{O}_2} = K_{\mathrm{H}} \cdot x_{\mathrm{O}_2}$

$\Rightarrow x_{\mathrm{O}_2} = \frac{p_{\mathrm{O}_2}}{K_{\mathrm{H}}}$

$= \frac{1520\text{mmHg}}{3.30 \times 10^7\text{mmHg}}$ (Given $K_{\mathrm{H}} = 3.30 \times 10^7\text{mmHg}$)

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

For nitrogen:

$p_{\mathrm{N}_2} = K_{\mathrm{H}} \cdot x_{\mathrm{N}_2}$

$\Rightarrow x_{\mathrm{N}_2} = \frac{p_{\mathrm{N}_2}}{K_{\mathrm{H}}} = \frac{6004\text{mmHg}}{6.51 \times 10^7\text{mmHg}}$

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

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

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Q2.40 Determine the amount of $\mathrm{CaCl}_2 (i = 2.47)$ 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,

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

$M = 1 \times 40 + 2 \times 35.5$

$= 111\text{gmol}^{-1}$

Therefore, $w = \frac{0.75 \times 11 \times 2.5}{2.47 \times 0.0821 \times 300}$

$= 3.42\text{g}$

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

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Q2.41 Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K_2SO_4 in 2 litre of water at 25° C, assuming that it is completely dissociated.

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,

$w = 25\text{mg} = 0.025\text{g}$

$V = 2\text{L}$

$T = 25^\circ\text{C} = (25 + 273)\text{K} = 298\text{K}$

Also, we know that:

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

$M = (2 \times 39) + (1 \times 32) + (4 \times 16) = 174\text{gmol}^{-1}$

Applying the following relation,

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

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

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

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

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