2. SOLUTION

2.36 100g of liquid A (molar mass 140 g mol^{-1}) was dissolved in 1000g of liquid B (molar mass 180g mol^{-1}). The vapour pressure of pure liquid B was found to be 500 torr. Calculate the vapour pressure of pure liquid A and its vapour pressure in the solution if the total vapour pressure of the solution is 475 torr.

Solution:

No. of moles of solute, $n_2 = \frac{100}{140} = \frac{5}{7}$ mole

No. of moles of solvent, $n_1 = \frac{1000}{180} = \frac{50}{9}$ mole

Mole fraction of solute,

$$x_2 = \frac{n_2}{n_1 + n_2} = \frac{5/7}{5/7 + 50/9} = 0{\cdot}114$$

Mole fraction of solvent, $x_1 = (1 - x_2) = (1 - 0{\cdot}114)$
$= 0{\cdot}886$

According to Raoult's law

$$P_A = x_A P_A^\circ = 0{\cdot}114 \times P_A^\circ$$

$$P_B = x_B P_B^\circ = 0{\cdot}886 \times 500 = 443 \text{ torr}$$

$$P_{\text{Total}} = P_A + P_B$$

$$475 = 0{\cdot}114\ P_A^\circ + 443$$

$$P_A^\circ = \frac{475 - 443}{0{\cdot}114} = 280{\cdot}7 \text{ torr}$$

$$\therefore P_A = 0{\cdot}114 \times 280{\cdot}7 = 32 \text{ torr.}$$

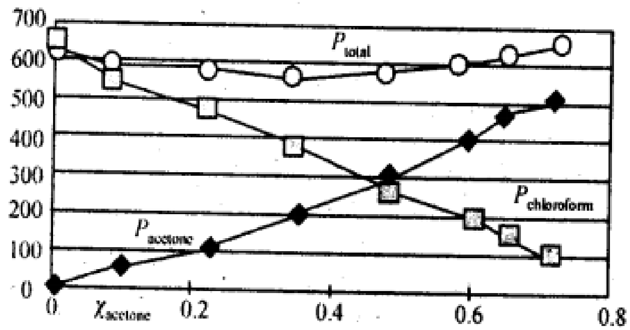
2.37 Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot P_{total}, $P_{chlroform}$ and $P_{acetone}$ as a function of $\chi_{acetone}$. The experimental data observed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

$100 \times \chi_{acetone}$	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

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

Solution:

$\chi_{acetone}$	0.0	0.118	0.234	0.360	0.508	0.582	0.645	0.721
$P_{acetone}$/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_{total}	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8



As the plot for P_{total} dips downwards, hence the solution shows negative deviation from the ideal behaviour.

Mole fraction of $C_6H_5CH_3$,

$$x_T = \frac{1.087}{1.026 + 1.087} = 0.514$$

According to Raoult's Law,

$$P_B = x_B \times P_B^{\circ} = 0.486 \times 50.71 = 24.65 \text{mm}$$

$$P_T = x_T \times P_T^{\circ} = 0.514 \times 32.06 = 16.48 \text{ mm}$$

Mole fraction of C_6H_6 in vapour phase

$$= \frac{P_B}{P_B + P_T} = \frac{24.65}{24.65 + 16.48} = 0.599.$$

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 80g of benzene is mixed with 100g of toluene.

Solution:

Molar mass of $C_6H_6 = 78 \text{ g mol}^{-1}$

Molar mass of $C_6H_5CH_3 = 92 \text{ gmol}^{-1}$

$$\text{No. of moles of } C_6H_6 = \frac{80}{78} = 1.026 \text{ mole}$$

$$\text{No. of moles of } C_6H_5CH_3 = \frac{100}{92} = 1.087 \text{ mole}$$

Mole fraction of C_6H_6,

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

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 K if the Henry's law constants for oxygen and nitrogen are 3.30 x 10^7 mm and 6.51 x 10^7 mm respectively, calculate the composition of these gases in water.

Solution:

Air containing 20% oxygen and 79% nitrogen by volume means

Partial pressure of O_2 (P_{O_2}) = $\frac{20}{100} \times 10 = 2$ atm

$= 2 \times 760 \text{ mm} = 1520 \text{ mm}$

Partial pressure of N_2 (P_{N_2}) = $\frac{79 \times 10}{100} = 7{\cdot}9$ atm

$= 7{\cdot}9 \times 760 \text{ mm} = 6004 \text{ mm}$

According to Henry's Law,

$$P_{O_2} = K_H \, x_{O_2}$$

$$x_{O_2} = \frac{P_{O_2}}{K_H} = \frac{1520}{3{\cdot}30 \times 10^7} = 4{\cdot}61 \times 10^{-5}$$

$$P_{N_2} = K_H \, x_{N_2}$$

$$x_{N_2} = \frac{P_{N_2}}{K_H} = \frac{6004}{6{\cdot}51 \times 10^7} = 9{\cdot}22 \times 10^{-5}.$$

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

Solution:

Using relation, $\pi = iCRT = i \frac{n}{V} RT$

$$n = \frac{\pi V}{i RT} = \frac{0.75 \times 2.5}{2.47 \times 0.0821 \times 300} = 0.0308 \text{ mole}$$

Molar mass of $CaCl_2 = 40 + 2 \times 35.5 = 111 \text{ g mol}^{-1}$

$\therefore$ Amount of $CaCl_2$ dissolved

$= 0.0308 \times 111 = 3.42$ g.