NCERT Solutions for Class 12 Science

Chapter 1- Solutions

Intext Questions

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.

Ans: Mass percentage is the ratio of mass of solute divided by the total mass of solution multiplied by 100.

$$\text{Mass\%} = \frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{solution}}} \times 100$$

Here,

Mass percentage of both the compounds are given as;

$$\text{Mass percentage of } C_6H_6 = \frac{\text{mass}_{C_6H_6}}{\text{mass}_{C_6H_6} + \text{mass}_{CCl_4}} \times 100$$

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

$$= 15.28\ \%$$

Now, as the solution consists of only two components i.e., C_6H_6 and CCl_4.

Thus,

$$\text{Mass percentage of } CCl_4 = 100 - 15.28$$

$$= 84.72\%$$

$\therefore$ Mass percentages of C_6H_6 and CCl_4 are 15.28% and 84.72% respectively.

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

Ans: Mole fraction is the ratio of moles of solute to the total moles of solution.

$$\text{Mole fraction} = \frac{n_{\text{solute}}}{n_{\text{solution}}}$$

Here,

We have given a solution containing 30% by mass of benzene in carbon tetrachloride; which clearly states that 30 g of benzene is present in 100 g of solution.

Thus, carbon tetrachloride is 70 g in 100 g of solution.

We know that,

Molar mass of benzene = 78 g/mol

Molar mass of carbon tetrachloride = 154 g/mol

Thus,

Number of moles of benzene = $\frac{30}{78}=0.385\text{mol}$

Number of moles of carbon tetrachloride = $\frac{70}{154}=0.455\text{mol}$

Now,

Mole fraction of benzene = $\frac{n_{C_6H_6}}{n_{C_6H_6}+n_{CCl_4}}$

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

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

$\therefore$ Mole fraction of benzene is 0.458 in the solution.

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

(a) 30 g of $Co(NO_3)_2.6H_2O$ in 4.3 L of solution.

Ans: Molarity is the number of moles of solute per liter of solution.

Molarity = $\frac{n_{solute}}{V_{solution}}$ M

Here,

Molarity of 30 g of $Co(NO_3)_2.6H_2O$ in 4.3 L of solution is given as;

Molar mass of $Co(NO_3)_2.6H_2O$ = 310.7 g/mol.

Number of moles = $\frac{30}{310.7}=0.0966$

Thus, molarity = $\frac{0.0966}{4.3}=0.022\text{M}$

(b) 30 mL of 0.5 M H_2SO_4 diluted to 500 ml.

Ans: Molarity of 30 mL of 0.5 M H_2SO_4 diluted to 500 ml.

1000 ml of 0.5 M H_2SO_4 contains 0.5 moles of H_2SO_4.

Thus,

30 ml of 0.5 M H_2SO_4 contains;

$=\frac{0.5}{1000}\times30=0.015\text{moles } H_2SO_4$.

Also, volume of solution = 500 ml = 0.5 L.

Now,

$\therefore$ Molarity = $\frac{0.015}{0.5} = 0.03\text{M}$

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

Ans: Molality is given as ratio of moles of solute per gm of mass of solvent.

Molality = $\frac{n_{solute}}{m_{solvent}}$ m

Here,

0.25 molal solution of urea states that 0.25 moles of urea is present in 1000 gm of solvent.

Molar mass of solute i.e., urea = 60 g/mol

Thus,

Mass of urea in the solution = $0.25 \times 60 = 15\text{g}$

Total mass of solution = 1000 gm solvent + 15 gm of solute
= 1015 gm = 1.015 Kg

This states that,

1.015 Kg solution has 15 gm urea.

Thus,

2.5 Kg solution will have;

$\frac{15}{1.015} \times 2.5 = 37\text{g}$ urea.

$\therefore$ Mass of urea required will be 37 g.

5. Calculate

(a) Molality

Ans: Given that,

20% of aqueous KI solution has 20 gm of KI in 100 gm of solution.

Mass of water (solvent) = 100 – 20 = 80 gm.

Volume of solution is given as,

Volume = $\frac{\text{mass}}{\text{density}} = \frac{100}{1.202} = 83.194\text{ml}$

Molar mass of KI = 166.0028 g/mol

Thus,

Number of moles of KI = $\frac{20}{166.0028} = 0.120\text{mol}$

Molar mass of water = 18 g/mol

Thus,

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

Molality of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g / ml;

Molality = $\frac{0.120}{0.08} = 1.50\text{m}$

(b) Molarity and

Ans: Molarity of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g / ml.

$$\text{Molarity} = \frac{0.120}{0.083194} = 1.44\text{M}$$

(c) Mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g/ml.

Ans: Mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g / ml.

Mole fraction = $\frac{0.120}{0.120 + 4.44} = 0.0263$

6. H_2S , a toxic gas with a 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.

Ans: Molality (solubility) of H_2S in water = 0.195 m which states that,

0.195 moles of H_2S in 1Kg of water (1000 g).

Now,

Molar mass of water = 18 g/mol.

Number of moles of water = $\frac{1000}{18} = 55.55\text{mol}$

Mole fraction of H_2S;

$$X_{H_2S} = \frac{0.195}{0.195 + 55.55} = 0.00349$$

Pressure at STP is given as 0.987 bar. Thus, applying Henry's law we get;

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

$$K_H = \frac{P_{H_2S}}{X_{H_2S}} = \frac{0.987}{0.00349} = 282.80\text{bar}$$

$\therefore$ Henry's constant will be 282.80 bar.

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.

Ans: Given that,

Henry's law constant, $K_H = 1.67 \times 10^8 \text{Pa}$

Pressure, $P_{CO_2} = 2.5\text{atm} = 2.5 \times 101325 = 253312.5\text{Pa}$

Now, by Henry's law;

$$P_{CO_2} = K_H \times X_{CO_2}$$

$$X_{CO_2} = \frac{253312.5}{1.67 \times 10^8} = 1.5168 \times 10^{-3}$$

Now,

500 ml soda water is equivalent to 500 ml water which indirectly signifies 500 gm of the same.

Thus,

We know that, molar mass of water = 18 g/mol

Number of moles of water = $\frac{500}{18} = 27.78\text{mol}$

Number of moles of CO_2 is given as,

$$X_{CO_2} = \frac{n_{CO_2}}{n_{CO_2} + n_{water}}$$

$$n_{CO_2} = X_{CO_2}\left(n_{CO_2} + n_{water}\right)$$

$$n_{CO_2} = 42.20 \times 10^{-3}\text{mol}$$

Mass of CO_2 is given as,

Molar mass of CO_2 = 44 g/mol.

Thus,

Mass = $n_{CO_2} \times 44 = 1.857\text{gm}$

8. The vapor 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 vapor pressure is 600 mm Hg. Also find the composition of the vapor phase.

Ans: Vapor pressures of the given pure liquids are;

$P_A^{\ 0} = 450\text{mm Hg}$, $P_B^{\ 0} = 700\text{mm Hg}$ and $P_{Total} = 600\text{mm Hg}$

By Raoult's law;

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

$$P_{Total} = X_A P_A^{\ 0} + X_B P_B^{\ 0} = X_A P_A^{\ 0} + \left(1 - X_A\right) P_B^{\ 0}$$

$$P_{Total} = P_B^{\ 0} + \left(P_A^{\ 0} - P_B^{\ 0}\right) X_A$$

600 = 700 + (450-700) X_A

$\therefore X_A = 0.4$

Thus,

$X_B = 1 - 0.4 = 0.6$

$P_A = X_A P_A{}^0 = 0.4 \times 450 = 180 \text{mm Hg}$

$P_B = X_B P_B{}^0 = 0.6 \times 700 = 420 \text{mm Hg}$

Thus,

Mole fraction of A in vapor phase = $\dfrac{180}{180 + 420} = 0.3$

Mole fraction of B in vapor phase = 1 – 0.3 = 0.7

9. Vapor 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 vapor pressure of water for this solution and its relative lowering.

Ans: Given that,

Vapor pressure of pure water, P^0 = 23.8 mm Hg

Mass of urea, W_2 = 50 g

Molar mass of urea, M_2 = 60 g/mol

Mass of water, W_1 = 850 g

Molar mass of water, M_1 = 18 g/mol

By Raoult's law;

$\Delta P = P^0 - P_s = X_2 \times P^0$

$$\frac{P^0 - P_s}{P^0} = X_2 = \frac{n_2}{n_1 + n_2} = \frac{W_2/M_2}{W_1/M_1 + W_2/M_2}$$

$X_2 = 0.017$

Now,

We have given that P^0 = 23.8 mm Hg;

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

Thus, V. P. of solution, P_s = 23.4 mmHg.

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.

Ans: Given that,

Boiling point of water = 100° C

Boiling point of water at 750 mm Hg = 99.63° C

Elevation in boiling point, $\Delta T_b = 100 - 99.63 = 0.37$

Ebullioscopic constant, $K_b = 0.52$

Mass of water, W_1 = 500 g

Molar mass of water, M_1 = 18 g/mol

Molar mass of sucrose, M_2 = 342 g/mol

Mass of sucrose, $W_2 = ?$

The elevation in boiling point is given as;

$$\Delta T_b = K_b \times m$$

$$\Delta T_b = K_b \times \frac{n_2}{W_1} = K_b \times \frac{W_2}{M_2} \times \frac{1000}{W_1}$$

$$W_2 = \frac{\Delta T_b \times M_2 \times W_1}{W_2 \times 1000} = 121.67g$$

$\therefore$ 121.67 g sucrose to be added in water such that it boils at 100° C.

11. Calculate the mass of ascorbic acid (Vitamin C, $C_6H_8O_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5° C. $K_f = 3.9 Kg/mol$.

Ans: Given that,

Lowering in freezing point, ΔT_f = 1.5° C

Cryoscopic constant, $K_f = 3.9 Kg/mol$

Mass of acetic acid, W_1 = 75 g

Molar mass of acetic acid, M_1 = 60 g/mol

Molar mass of ascorbic acid, M_2 = 176 g/mol

Mass of ascorbic acid, $W_2 = ?$

The lowering in freezing point is given as;

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

$$\Delta T_f = K_f \times \frac{n_2}{W_1} = K_f \times \frac{W_2}{M_2} \times \frac{1000}{W_1}$$

$$W_2 = \frac{\Delta T_f \times M_2 \times W_1}{K_f \times 1000} = 5.077g$$

$\therefore$ Mass of ascorbic acid is 5.077 g.

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

Ans: Given that,

Mass of polymer, W = 1 g

Molar mass of polymer, M = 185000 g/mol

Volume of water, V = 450 ml = 0.45 L

Temperature, T = 37° C = 310 K

Gas constant, R = $8.314\text{kPa.L.K}^{-1}\text{mol}^{-1} = 8.314\times10^{3}\text{Pa.L.K}^{-1}\text{mol}^{-1}$

Now, using formula for osmotic pressure;

$$\pi = CRT = \frac{n}{V}RT = \frac{W}{M}\times\frac{1}{V}\times RT$$

$$\pi = 30.96\text{Pa}$$

∴ Osmotic pressure is 30.96 Pa.

NCERT Exercise

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

Ans: A homogeneous mixture of two or more chemically non-reacting substances is called a solution. There are nine types of solution under the main 3 heads i.e., gaseous solution, liquid solution and solid solution.

Types of Solution with examples;

Gas in gas: Air i.e., mixture of O_2 and N_2 , etc.

Liquid in gas: Water vapor.

Solid in gas: Smoke, Camphor vapors in N_2 gas, etc.

Gas in liquid: Aerated water, O_2 dissolved in water, etc.

Liquid in liquid: Vinegar solution, etc.

Solid in liquid: Glucose dissolved in water, saline water, etc.

Gas in solid: Solution of hydrogen in platinum, etc.

Liquid in solid: Amalgams eg. Mg-Hg.

Solid in solid: Ornaments (Cu/Ag with Au).

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

Ans: It is said that gas is solute in the solid solution i.e., gas – solid solution. The examples are solid carbon dioxide in fire extinguishers, solution of hydrogen in palladium, dissolved gases in underground minerals and many more.

3. Define the following terms:

(i) Mole fraction

Ans: Mole fraction is the ratio of moles of solute to the total moles of solution.

Mole fraction = $\frac{n_{\text{solute}}}{n_{\text{solution}}}$

(ii) Molality

Ans: Molality is given as ratio of moles of solute per gm of mass of solvent.

Molality = $\frac{n_{solute}}{m_{solvent}}$ m

It is better way of expressing concentration of the solute because this does not change with the change in temperature as that of molarity (dependent on volume of solution).

(iii) Molarity

Ans: Molarity is the number of moles of solute per liter of solution.

Molarity = $\frac{n_{solute}}{V_{solution}}$ M

(iv) Mass percentage

Ans: Mass Percentage:

Mass percentage is the ratio of mass of solute divided by the total mass of solution multiplied by 100.

$$\text{Mass\%} = \frac{\text{mass}_{solute}}{\text{mass}_{solution}} \times 100$$

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?

Ans: Given that,

We have 68% nitric acid by mass in aqueous solution which means 68 g nitric acid is present in 100 g of solution.

Molar mass of nitric acid = 63 g/mol

Number of moles of nitric acid = $\frac{68}{63} = 1.079\text{mol}$

Also,

Density of the solution is given as 1.504 g/ml. Thus,

Volume = $\frac{\text{mass}}{\text{density}} = \frac{100}{1.504} = 66.489\text{ml}$

Now,

Molarity is given as;

Thus,

Molarity = $\frac{1.079 \times 1000}{66.489} = 16.22\text{M}$

5. A solution of glucose in water is labelled as 10% w/w, that would be the molality and mole fraction of each component in the solution? If the density

of solution is 1.2 g/ml then what shall be the molarity of the solution?

Ans: We have given that,

10% w/w solution of glucose in water i.e., 10 g of glucose in 90 g of water.

Now, as we know;

Molar mass of glucose = 180 g/mol

Molar mass of water = 18 g/mol

Thus,

Number of moles of glucose in the solution = $\frac{10}{180} = 0.055\text{mol}$

Number of moles of water in the solution = $\frac{90}{18} = 5\text{mol}$

Taking into consideration the above values;

Molality is given as –

Molality = $\frac{0.055 \times 1000}{90} = 0.617\text{m}$

Mole fraction of each component can be given as –

Mole fraction of glucose = $\frac{0.055}{0.055 + 5} = 0.0108$

Mole fraction of water = 1 – 0.0108 = 0.9892

Again,

We have given, density of solution is 1.2 g/ml. Thus,

Volume = $\frac{\text{mass}}{\text{density}} = \frac{100}{1.2} = 83.33\text{ml}$

Thus,

Molarity can be given as;

Molarity = $\frac{0.055 \times 1000}{83.33} = 0.66\text{M}$

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?

Ans: Let us consider that we have x g of Na_2CO_3 in 1g mixture. Thus, we will have (1 - x) g of $NaHCO_3$ in the same.

We know that,

Molar mass of Na_2CO_3 = 106 g/mol

Molar mass of $NaHCO_3$ = 84 g/mol

Thus,

Number of moles of Na_2CO_3 in mixture = $\frac{x}{106}$

Number of moles of $NaHCO_3$ in mixture = $\frac{(1-x)}{84}$

Now, as we have given that they are equimolar; thus,

$$\frac{x}{106}=\frac{(1-x)}{84}$$
$$84x=106-106x$$
$$\therefore x=0.557$$

Hence,

Number of moles of Na_2CO_3 in mixture = $\frac{0.557}{106}=0.00526$

Number of moles of $NaHCO_3$ in mixture = $\frac{(1-0.577)}{84}=0.00503$

To calculate how many ml of 0.1 M HCl is required to react completely with 1 g mixture of Na_2CO_3 and $NaHCO_3$ we need to analyze reactions for both;

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

Here, each mole of Na_2CO_3 requires 2 moles of HCl. Thus,

0.00526 moles of Na_2CO_3 requires = $0.00526\times 2=0.01052$ moles of HCl.

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

Here, each mole of $NaHCO_3$ requires a mole of HCl. Thus,

0.00503 moles of $NaHCO_3$ will require 0.00503 HCl.

Total moles of HCl required = 0.01052 + 0.00503 = 0.01555 moles

Now,

1.1 moles of 0.1 M HCl are present in 1000 ml of solution. Thus,

0.01555 moles will be present in = $\frac{0.01555\times 1000}{0.1}=155.5\text{ml}$ of solution.

$\therefore$ Volume required to react completely with the mixture will be 155.5 ml.

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.

Ans: Mass percentage is defined as the ratio of mass of solute by the mass of solution multiplied by 100.

Mass percentage = $\frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{solution}}}\times 100$

Here,

We have given a mixture of 300 g of 25% solution and 400 g of 40% solution by mass. Thus,

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

400 g of 40% solution will contain = $\frac{40\times400}{100}=160$ g of solute

Now, the resulting solution will have contents as;

Total mass of solute = 75 + 160 = 235 g

Total mass of solution = 300 + 400 = 700 g

Thus,

Mass percentage is given as,

Mass percentage of solute in the solution = $\frac{235}{700}\times100=33.57\,\%$

Similarly,

Mass percentage of water in the solution = 100 – 33.57 = 66.42%

8. An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g/ml, then what shall be the molarity of the solution?

Ans: Given that,

Mass of solute ($C_2H_6O_2$) = 222.6 g

Molar mass of $C_2H_6O_2$ = 62 g/mol

Number of moles of $C_2H_6O_2$ = $\frac{222.6}{62}=3.59\text{mol}$

Mass of solvent (water) = 200 g

Total mass of solution = 422.6 g

Density of solution = 1.072 g/ml

Volume = $\frac{\text{mass}}{\text{density}}=\frac{422.6}{1.072}=394.21\text{ml}$

Thus,

Molality is given as,

Molality = $\frac{3.59}{200}=17.95\text{m}$

And, molarity is given as,

Molarity = $\frac{3.59\times1000}{394.21}=9.106\text{M}$

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.

Ans: We know that ppm is the units which expresses the quantity in parts per million i.e., parts per 10^6 of solution.

Here, 15 ppm means 15 parts per 10^6 of solution.

Thus, mass percentage is given as;

$$\text{Mass percentage} = \frac{15}{10^6}\times 100 = 15\times 10^{-4}$$

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

Ans: 15 ppm states that we have 15 g chloroform in 10^6 g of solution i.e., mass of solvent = 10^6 g

Molar mass of chloroform = 119.5 g/mol

$$\text{Number of moles of chloroform} = \frac{15}{119.5} = 0.125\text{mol}$$

Thus,

Molality is given as,

$$\text{Molality} = \frac{0.125\times 1000}{10^6} = 125\times 10^{-6}\text{m}$$

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

Ans: Alcohol and water both possess a strong tendency to form intermolecular hydrogen bonding. When we mix the two liquids, a solution is formed as a result of formation of H-bonds between alcohol and H_2O molecules. These interactions are weaker and less extensive than those in pure H_2O. Hence, they show a positive deviation from ideal behavior. As a result, the solution of alcohol and water will have higher vapor pressure and lower boiling point than that of pure water and pure alcohol.

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

Ans: When gases dissolve in water, the process is accompanied by release of excess heat energy, i.e., exothermic. According to Le Chatelier's principle, when the temperature of the process is increased further the equilibrium shifts in backward direction. Hence, gases become less soluble in liquids.

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

Ans: Henry's law:

It states that the solubility of a gas in a liquid at a given temperature is directly proportional to the partial pressure of the gas. The effect of pressure on the solubility of a gas in a liquid is governed by this law.

Mathematically;

$$P = K_H x$$

where,
P is the partial pressure of the gas
x is the mole fraction of the gas in the solution
K_H is Henry's Law constant.

Applications of Henry's law:

a. In the sea diving.
b. In the production of carbonated beverages.

13. The partial pressure of ethane over a solution containing $6.56\times10^{-3}g$ of ethane is 1 bar. If the solution contains $5\times10^{-2}g$ of ethane, then what shall be the partial pressure of the gas?

Ans: By Henry's law, we know that the solubility of gas in liquid is directly proportional to the pressure of the gas. The proportionality sign is then replaced by Henry's constant. This is stated as;

$m=K_H\times P$

Also, mole fraction is directly proportional to the mass of the ethane.
Thus,
For case 1,

$6.56\times10^{-3}=K_H\times1$

For case 2,

$5\times10^{-2}=K_H\times x$

where, x is the partial pressure of gas when the solution contains $5\times10^{-2}g$ of ethane.
Now, equating both the above equations;

$$6.56\times10^{-3}=\frac{5\times10^{-2}}{x}$$

Thus,
The partial pressure of gas, $x=7.62$ bar

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?

Ans: Positive deviation from Raoult's law:
Solutions exhibit positive deviation from Raoult's law when they have vapor pressure more than expected from the law.
Here, $\Delta_{mix}H$ is positive as the energy is consumed for breaking the strong interaction and form weaker interactions. In a similar way, $\Delta_{mix}V$ is positive as the expansion of volume takes place.

- Negative deviation from Raoult's law:

Solutions exhibit negative deviation from Raoult's law when they have vapor pressure less than expected from the law.

Here, $\Delta_{mix}H$ is negative as the energy is released due to replacement of weaker interactions by stronger ones.

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?

Ans: Given that,

Vapor pressure of pure water at boiling point, P^0 = 1.103 bar

Vapor pressure of solution, P_s = 1.004 bar

An aqueous solution has 2% non-volatile solute. Thus, this states that the solution is 100 g from which 2 g $\left(W_2\right)$ is the non-volatile solute and the solvent is 98 g$\left(W_1\right)$.

Molar mass of water $\left(M_1\right)$ = 18 g/mol.

Now, by Raoult's law for dilute solutions;

$$\frac{P^0 - P_s}{P^0} = \frac{n_2}{n_1 + n_2} \approx \frac{n_2}{n_1}$$

$$\frac{P^0 - P_s}{P^0} = \frac{W_2/M_2}{W_1/M_1}$$

Thus,

The molar mass of solute, $M_2 = 41.34 g/mol$

16. Heptane and octane form an ideal solution. At 373 K, the vapor pressure of the two liquid components is 105.2 kPa and 46.8 kPa respectively. What will be the vapor pressure of a mixture of 26.0 g of heptane and 35.0 g of octane?

Ans: We have given the vapor pressures of the two liquid components i.e.

V. P. of heptane = 105.2 kPa

V. P. of octane = 46.8 kPa

Mass of heptane = 26 g

Mass of octane = 35 g

Molar mass of heptane = 100 g/mol

Molar mass of octane = 114 g/mol

Thus,

Number of moles of heptane present = $\frac{26}{100} = 0.26 mol$

Number of moles of octane present = $\frac{35}{114} = 0.307 mol$

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

Mole fraction of octane = 1 – 0.458 = 0.541

Now,

By Raoult's law;

Vapor pressure of heptane = $x_H \times P^0$

$= 0.458 \times 105.2 = 48.1816 \text{kPa}$

Vapor pressure of octane = $x_O \times P^0$

$= 0.541 \times 46.8 = 25.3188 \text{kPa}$

Thus,

Vapor pressure of mixture = 48.1816 + 25.3188 = 73.5004kPa

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

Ans: We have given 1 molal solution i.e., 1 mole of non-volatile solute in 1000 g water.

Vapor pressure of water, P^0 = 12.3 kPa

Now,

Molar mass of water = 18 g/mol

Number of moles of water in the solution = $\frac{1000}{18}=55.56 \text{mol}$

Mole fraction of solute = $\frac{1}{1+55.56}=0.0176$

Thus,

By Raoult's law,

$$\frac{P^0 - P_s}{P^0} = x$$

$$\frac{12.3 - P_s}{12.3} = 0.0176$$

Thus,

V. P. of solution, $P_s = 12.082 \text{kPa}$

18. Calculate the mass of a non-volatile solute (molar mass 40 g/mol) which should be dissolved in 114 g octane to reduce its vapor pressure to 80%.

Ans: Given that,

Reduced vapor pressure = 80% of vapor pressure of pure components.

$$\therefore P_s = \frac{80}{100} P^0 = 0.8P^0$$

Now, let us consider the mass of a non-volatile solute as W g.

Molar mass of the same solute = 40 g/mol

So, number of moles of same solute = $\frac{W}{40}$mol

Mass of octane = 144 g
Molar mass of octane = 144 g/mol
So, number of moles of octane = 1 mol
Now,

Mole fraction of non-volatile solute = $\frac{W/40}{W/40+1}$

By Raoult's law,

$$\frac{P^0-P_s}{P^0}=x$$

$$\frac{P^0-0.8P^0}{P^0}=x=\frac{W/40}{W/40+1}$$

Thus,
Mass of solute, W = 10 g

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

(i) molar mass of the solute

Ans: Given that,
Mass of non-volatile solute = 30 g
Let molar mass of the same be M g/mol

Thus, number of moles of solute = $\frac{30}{M}$mol

Case 1
Mass of water = 90 g
Molar mass of water = 18 g/mol

So, number of moles of water involved = $\frac{90}{18}=5$mol

Vapor pressure of solution, P_s = 2.8 kPa

Mole fraction of solute = $\frac{30/M}{30/M+5}$

By Raoult's law,

$$\frac{P^0 - P_s}{P^0} = x$$

$$\frac{P^0 - 2.8}{P^0} = \frac{30/M}{30/M + 5}$$

$$\therefore \frac{P^0}{2.8} = 1 + \frac{6}{M}$$

Case 2

Mass of water = 90 + 18 = 108 g

Molar mass of water = 18 g/mol

So, number of moles of water involved = $\frac{108}{18} = 6\text{mol}$

Vapor pressure of solution, P_s = 2.9 kPa

Mole fraction of solute = $\frac{30/M}{30/M + 6}$

By Raoult's law,

$$\frac{P^0 - P_s}{P^0} = x$$

$$\frac{P^0 - 2.8}{P^0} = \frac{30/M}{30/M + 6}$$

$$\therefore \frac{P^0}{2.9} = 1 + \frac{5}{M}$$

Now,

Considering both the above equations, we get;

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

$\therefore$ Molar mass of solute;

M = 23 g/mol

(ii) vapor pressure of water at 298 K.

Ans: Now, putting this value in equation (1.1)

$$\frac{P^0}{2.8} = 1 + \frac{6}{M} = 1 + \frac{6}{23}$$

$\therefore$ V. P. of water, $P^0 = 3.53\text{kPa}$

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

Ans: Given that,

Freezing point of pure water = 273.15 K

Case 1 –

A solution of 5% of cane sugar is in water i.e., 5 g of cane sugar is present in 100 g of water.

Molar mass of cane sugar = 342 g/mol

Molality is given as,

Molality, m = $\frac{5\times1000}{342\times100}=0.146\text{m}$

Freezing point of solution = 271 K

$\Delta T_f = 273.15-271=2.15\text{K}$

Now,

Lowering in freezing point is given as,

$\Delta T_f = K_f m$

$K_f = \frac{2.15}{0.146}=14.726$

Case 2 –

A solution of 5% glucose is in water i.e. 5 g of glucose in 100 g of water.

Molar mass of glucose = 180 g/mol

Molality is given as,

Molality, m = $\frac{5\times1000}{180\times100}=0.277\text{m}$

$K_f = 14.726$

Lowering in freezing point is given as,

$\Delta T_f = K_f m$

$= 14.726\times0.277=4.079$

Thus,

$\Delta T_f = 273.15-T=4.079$

Thus, the freezing point;

T = 269.07 K

21. Two elements A and B form compounds having formula AB_2 and AB_4. When dissolved in 20g of benzene (C_6H_6). 1 g of 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. Calculate atomic masses of A and B.

Ans: We know that the lowering in freezing point is stated as,

$\Delta T_f = K_f m$

which also can be evaluated as;

$$\Delta T_f = K_f \times \frac{W_2 \times 1000}{M_2 \times W_1}$$

Mass of solvent (benzene), W_1 = 20 g

Molal depression constant, K_f = 5.1 K.Kg / mol

Now,

Case 1 –

Solute is AB_2;

Mass of solute, W_2 = 1 g

ΔT_f = 2.3 K

Thus,

Molar mass for the solute i.e., AB_2 is given as,

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{5.1 \times 1 \times 1000}{2.3 \times 20} = 110.869 g / mol$$

Case 2 –

Solute is AB_4;

Mass of solute, W_2 = 1 g

ΔT_f = 1.3 K

Thus,

Molar mass for the solute i.e., AB_4 is given as,

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{5.1 \times 1 \times 1000}{1.3 \times 20} = 196.153 g / mol$$

Now,

Molar mass for AB_2 is given as;

A + 2 (B) = 110.869 g/mol

Molar mass for AB_4 is given as;

A + 4 (B) = 196.153 g/mol

Therefore,

Solving above 2 equations we get, the atomic masses as;

A = 25.59 u

B = 42.64 u

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

Ans: The formulation for the osmotic pressure is given as;

$\pi = CRT$

Case 1 –

T = 300 K

Mass of glucose = 36 g

Molar mass of glucose = 180 g/mol

Osmotic pressure = 4.08 bar

According to the formula;

$$\pi = \frac{W}{M}RT$$

$$4.08 = \frac{36}{180} \times R \times 300$$

R = 0.068 units

Case 2 –

Osmotic pressure = 1.52 bar

T = 300 K

Thus, by given formula;

$\pi = CRT$

Thus, the concentration will be,

$$C = \frac{1.52}{0.068 \times 300} = 0.0745M$$

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

(i) n-hexane and n-octane

Ans: Both are nonpolar and hence, the intermolecular interactions will be London dispersion forces.

(ii) I_2 and CCl_4 .

Ans: Both are nonpolar and hence, the intermolecular interactions will be London dispersion forces.

(iii) $NaClO_4$ and water

Ans: The intermolecular interactions will be ion-dipole interactions as $NaClO_4$ is an ionic compound and water is a polar molecule.

(iv) methanol and acetone

Ans: Both are polar molecules and hence, intermolecular interactions will be dipole-dipole interactions.

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

Ans: Both are polar molecules and hence, intermolecular interactions will be dipole-dipole interactions.

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.

Ans: The arrangement according to the increasing order of solubility in n-octane is; KCl < CH_3OH < CH_3CN < Cyclohexane.

This is because;

a. KCl is an ionic compound and will not dissolve in n-octane.

b. CH_3OH is a polar molecule and hence, will dissolve in n-octane.

c. CH_3CN is also a polar molecule but less than that of CH_3OH. Thus, it will dissolve in n-octane to a great extent.

d. Cyclohexane is also a polar and simpler molecule which will cause its dissolving in all proportions.

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

(i) Phenol

Ans: Insoluble in water:

(ii) Toluene

Ans: Toluene is not soluble in one another due to their non-polar nature.

(iii) Formic acid

Ans: Partially soluble in water:

(iv) Ethylene glycol

Ans: Partially soluble

(v) Chloroform

Ans: Highly insoluble in water:

(vi) Pentanol

Ans: Partially soluble in water due to the presence of the alcohol functional group.

26. If the density of some lake water is 1.25 g/ml and contains 92g of Na^+ ions per kg of water, calculate the molality of Na^+ ions in the lake.

Ans: Given that,

Mass of Na^+ ions = 92 g

Molar mass of Na^+ ions = 23 g/mol

Mass of water = 1000 g

$\therefore$ Molality is given as;

Molality, m = $\frac{92\times1000}{23\times1000}=4m$

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

Ans: The dissociation reaction is given as;

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

The solubility product is given as; $K_{sp}=6\times10^{-16}$

Let the solubility of each ion be denoted as X mol/L.

Thus,

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

$X^2=6\times10^{-16}$

$X=\sqrt{6\times10^{-16}}$

$\therefore$ Maximum molarity of CuS,

$X=2.44\times10^{-8}mol/L$

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.

Ans: Given that,

Mass of aspirin = 6.5 g

Mass of acetonitrile = 450 g

Total mass of solution = 6.5 + 450 = 456.5 g

Mass percentage is given as,

Mass percentage = $\frac{mass_{solute}}{mass_{solution}}\times100$

$= \frac{6.5}{456.5}\times100=1.423\%$

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.

Ans: We have given,

Molality of solution = 1.5×10^{-3}m; which states that the 1.5×10^{-3}moles are present in 1 kg of solvent.

Molar mass of nalorphene = 311 g/mol

Thus,

Mass of solute = $1.5\times10^{-3}\times311=0.4665\text{g}$

Total mass of the solution = 0.4665 + 1000 = 1000.4665 g

From this we can say that 0.4665 g narlophene requires total solution of 1000.4665 g.

Likewise;

1.5 mg narlophene will have;

Mass of solution = $\dfrac{1.5\times10^{-3}\times1000.4665}{0.4665}=3.216\text{g}$

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

Ans: Given that,

We have a 0.15M solution which states that the solution of 1000 ml has 0.15 moles of solute.

Thus, to prepare 250 ml of solution we will require = $\dfrac{250\times0.15}{1000}=0.0375\text{mol}$

Molar mass of benzoic acid = 122 g/mol

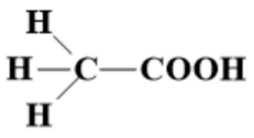
Therefore, mass of benzoic acid is given as,

Mass = $0.0375\times122=4.575\text{g}$

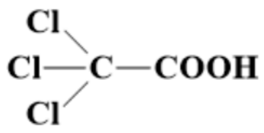
31. The depression in freezing point of water observed for the same amount of acetic acid, trichloroacetic acid and trifluoroacetic acid increases in the order given below. Explain briefly.

Acetic acid < trichloroacetic acid < trifluoroacetic acid.

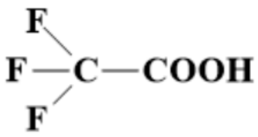
Ans: The structures for the given acids are;



Acetic acid



Trichloroacetic acid



Trifloroacetic acid

We know that, H is least electronegative whereas, F is most electronegative. Thus, F can withdraw electrons towards itself more than Cl and H. This signifies that,

trifluoroacetic acid can easily lose H^+ ions i.e., ionizes to the greatest extent. The more ions produced, the greater the depression of the freezing point. Therefore, the depression in the freezing point increases in the order:
Acetic acid < trichloroacetic acid < trifluoroacetic acid.

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

Ans: Given that,
Mass of $CH_3CH_2CHClCOOH$ = 10 g
Molar mass of $CH_3CH_2CHClCOOH$ = 122.5 g/mol

Number of moles of the same, n = $\frac{10}{122.5} = 0.0816 \text{moles}$

Mass of water (solvent), W_1 = 250 g
Molality is given by;

Molality = $\frac{n}{W_1} = \frac{0.0816 \times 1000}{250} = 3.265 \times 10^{-1} m$

Also, we know that,
$K_a = 1.4 \times 10^{-3}, K_f = 1.86 K.Kg / mol.$
Now,
The dissociation reaction is given as,

$$\underset{\text{at initial conc.}}{} \quad \underset{C\ mol\ L^{-1}}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{0}{CH_3CH_2CHClCOO^-} + \underset{0}{H^+}$$

$$\underset{\text{at equilibrium}}{} \quad \underset{C(1-\alpha)}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{C\alpha}{CH_3CH_2CHClCOO^-} + \underset{C\alpha}{H^+}$$

where, α is the degree of dissociation.
Now,
Dissociation constant is given as;

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

As, $\alpha << 1, (1-\alpha) \approx 1$

$$K_a = C\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}} = 0.06548$$

Again,

$$\underset{\text{at initial conc.}}{} \quad \underset{1}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{0}{CH_3CH_2CHClCOO^-} + \underset{0}{H^+}$$

$$\underset{\text{at equilibrium}}{} \quad \underset{1-\alpha}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{\alpha}{CH_3CH_2CHClCOO^-} + \underset{\alpha}{H^+}$$

Thus,

Total moles at equilibrium = $1-\alpha+2\alpha=1+\alpha$

Then Van’t Hoff factor is given as,

$i=1+\alpha$

Now, putting the values we get,

$i=1+0.06548=1.06548$

Now,

Lowering in freezing point is given as;

$\Delta T_f = iK_f m$

$\Delta T_f = 1.06548\times1.86\times0.3265=0.6470K$

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 fluoro acetic acid.

Ans: Given that,

Mass of CH_2FCOOH = 19.5 g

Molar mass (actual) of CH_2FCOOH = 78 g/mol

Mass of water = 500 g

$K_f = 1.86K.Kg/mol.$

Depression in freezing point, $\Delta T_f = 1°C = 1K$

Now,

Lowering in freezing point is stated as;

$$\Delta T_f = K_f m = K_f \frac{W_2\times1000}{M_2\times W_1}$$

$$M_2 = \frac{K_f\times W_2\times1000}{\Delta T_f\times W_1}$$

The observed molar mass will be,

$M_2 = 72.54g/mol$

Thus,

The Van’t Hoff factor can be calculated as;

$$i = \frac{(M_2)_{actual}}{(M_2)_{observed}}$$

i = 1.0752

Now,

The dissociation reaction is given as,

$$\text{at initial conc.}\quad \underset{\text{C mol L}^{-1}}{CH_2FCOOH} \leftrightarrow \underset{0}{CH_2FCOO^-} + \underset{0}{H^+}$$

$$\text{at equilibrium}\quad \underset{C(1-\alpha)}{CH_2FCOOH} \leftrightarrow \underset{C\alpha}{CH_2FCOO^-} + \underset{C\alpha}{H^+}$$

where, α is the degree of dissociation.
Total resulting concentration is given as,
Total = $C(1-\alpha+2\alpha) = C(1+\alpha)$
From this, Van't Hoff factor is given as;

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

i = 1.0752 = $1+\alpha$
Thus,
α = 0.0752
And,
The concentration (molarity) can be calculated as;
Having volume of 500 ml –

$$\text{Concentration} = \frac{19.5\times1000}{78\times500} = 0.5\text{M}$$

From all these data, dissociation constant can be calculated as;

$$K_a = \frac{\left[CH_2FCOO^-\right]\left[H^+\right]}{\left[CH_2FCOOH\right]} = \frac{C^2\alpha^2}{C(1-\alpha)}$$

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

Putting the specified values, we get;

$$K_a = \frac{0.5\times0.0752^2}{1-0.0752}$$

Thus,
Dissociation constant, $K_a = 2.299\times10^{-4}$.

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

Ans: Given that,
Vapor pressure of water, P^0 = 17.535 mm Hg
Mass of glucose, W_2 = 25 g
Molar mass of glucose, M_2 = 180 g/mol
Mass of water, W_1 = 450 g

Molar mass of water, $M_1 = 18$ g/mol

Now,

Number of moles of glucose in the solution can be stated as;

$$n_2 = \frac{W_2}{M_2} = \frac{25}{180} = 0.138 \text{mol}$$

Number of moles of water in the solution can be stated as;

$$n_1 = \frac{W_1}{M_1} = \frac{450}{18} = 25 \text{mol}$$

Thus, by Raoult's law,

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

$$\frac{17.535 - P_s}{17.535} = \frac{0.138}{0.138 + 25}$$

Thus, V. P. of water,

$$P_s = 17.44 \text{mmHg}$$

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

Ans: Given that,

Henry's law constant, K_H = 4.27×10^5 mm Hg

Pressure = 760 mm Hg

Thus, the molality can be calculated as;

By henry's law,

$$P = K_H m$$

$$m = \frac{P}{K_H} = \frac{760}{4.27 \times 10^5}$$

Molality / solubility = 1.779×10^{-3}m

36. 100 g of liquid A (molar mass 140 g/mol) was dissolved in 1000 g of liquid B (molar mass 180 g/mol). The vapor pressure of pure liquid B was found to be 500 torr. Calculate the vapor pressure of pure liquid A and its vapor pressure in the solution if the total vapor pressure of the solution is 475 Torr.

Ans: Given that,

Total vapor pressure of the solution, P_{Total} = 475 Torr

Liquid A –

Mass = 100 g

Molar mass = 140 g/mol

Number of moles of A, $n_A = \frac{100}{140} = 0.7142 \text{mol}$

Liquid B -
Mass = 1000 g
Molar mass = 180 g/mol

Number of moles of B, $n_B = \frac{1000}{180} = 5.56 \text{mol}$

V. P. of pure liquid B, $P_B{}^0$ = 500 Torr

Thus,

Mole fraction of A, $x_A = \frac{n_A}{n_A + n_B}$

$\therefore x_A = 0.1138$

Mole fraction of B, $x_B = \frac{n_B}{n_B + n_A}$

$\therefore x_B = 0.886$

Now,
By Raoult's law,

$P_B = P_B{}^0 x_B$

$P_B = 500 \times 0.886 = 443.08 \text{Torr}$

Hence,

$P_{Total} = P_A + P_B$

$\Rightarrow P_A = P_{Total} - P_B$

$P_A = 475 - 443.08 = 31.91 \text{Torr}$

Now,
V. P. of pure liquid A can be found as;
By Raoult's law,

$P_A = P_A{}^0 x_A$

Thus,
V. P. of pure liquid A,

$$P_A{}^0 = \frac{P_A}{x_A} = 280.45 \text{Torr}$$

37. Vapor 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

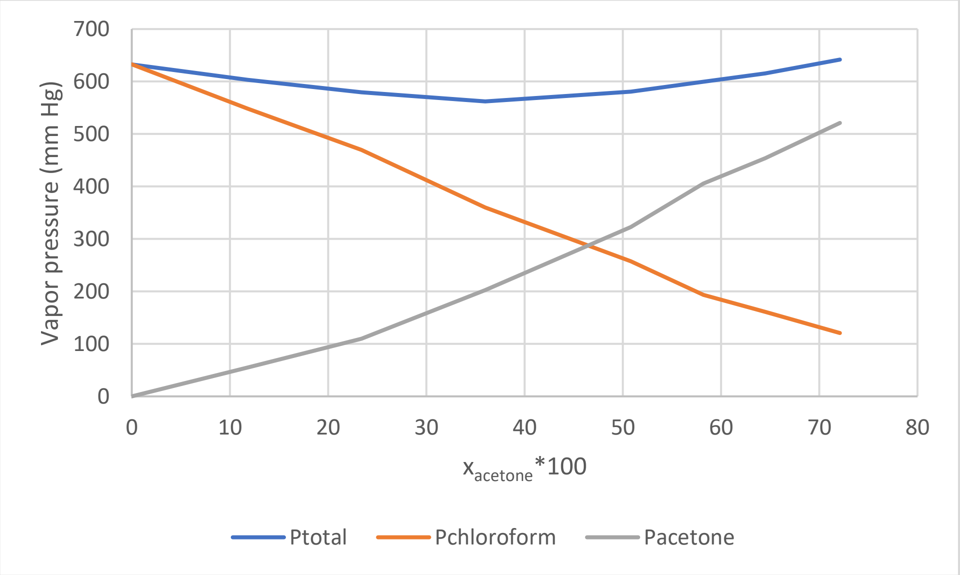
compositions of mixture is.

$x_{Acetone} \times 100$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$P_{Acetone}$/mmHg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$P_{Chloroform}$/mmHg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7

Ans: From the given table, following data can be generated;

$x_{Acetone} \times 100$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$P_{Acetone}$/mmHg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$P_{Chloroform}$/mmHg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7
P_{Total} /mmHg	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8

The graph can be drawn as;



From the above table and the graph, we can say that the solution shows the negative deviation from ideal behavior (graph of P_{Total} is downwards).

38. Benzene and toluene form an ideal solution over the entire range of composition. The vapor 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 the vapor phase if 80 g of benzene is mixed with 100 g of toluene.

Ans: Given that,

Mass of benzene = 80 g

Molar mass of benzene = 78 g/mol

Number of moles of benzene, $n_B = \frac{80}{78} = 1.025 \text{moles}$

Vapor pressure of pure benzene, $P_B{}^0$ = 50.71 mm Hg

Mass of toluene = 100 g

Molar mass of toluene = 92 g/mol

Number of moles of toluene, $n_T = \frac{100}{92} = 1.086$ moles

Vapor pressure of pure toluene, $P_T{}^0$ = 32.06 mm Hg

Now,

Mole fraction of benzene, $x_B = \frac{n_B}{n_B + n_T}$

$\therefore x_B = 0.4855$

Mole fraction of toluene, $x_T = 1 - 0.4855 = 0.5144$

Now,

By raoult's law,

The partial v. p. of benzene, $P_B = P_B{}^0 x_B$

$\therefore P_B = 24.61 \text{mmHg}$

Similarly, the partial v. p. of toluene, $P_T = P_T{}^0 x_T$

$\therefore P_T = 16.491 \text{mmHg}$

Thus, the mole fraction of benzene in vapor phase is given as;

Mole fraction = $\frac{P_B}{P_B + P_T} = 0.598$

39. The air is a mixture of a number of gases. The major components are oxygen and nitrogen with an approximate proportion of 20% 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 3.3×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

Ans: Given that,

Oxygen in air = 20% i.e., 20 g of oxygen is present in 100 g air.

Nitrogen in air = 79% i.e., 79 g of nitrogen is present in 100 g of air.

Henry's law constant;

For oxygen = 3.3×10^7 mm Hg

For nitrogen = 6.51×10^7 mm Hg

The water is in equilibrium with air at a pressure of 10 atm i.e., 7600 mm Hg.

Thus, the partial pressures are given as;

Oxygen;

$$P_{O_2} = \frac{20}{100} \times 7600 = 1520 \text{mmHg}$$

Nitrogen;

$$P_{N_2} = \frac{79}{100} \times 7600 = 6004 \text{mmHg}$$

Therefore, by Henry's law;

Oxygen –

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

$$x_{O_2} = \frac{1520}{3.3 \times 10^7} = 4.606 \times 10^{-5}$$

Nitrogen –

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

$$x_{N_2} = \frac{6004}{6.51 \times 10^7} = 9.222 \times 10^{-5}$$

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

Ans: Given that,

Volume = 2.5 L

Van't Hoff factor = 2.47

Osmotic pressure = 0.75 atm

Gas constant = $0.0821 \text{LatmK}^{-1}\text{mol}^{-1}$

Temperature = 273 + 27 = 300 K

Molar mass of $CaCl_2$ = 111g/mol

Osmotic pressure is given as,

$$\pi = iCRT$$

$$\pi = i\frac{n}{V}RT = i\frac{W}{M \times V}RT$$

$$W = \frac{\pi MV}{iRT} = \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300}$$

Thus, the amount of $CaCl_2$ required,

W = 3.42 g

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

Ans: Given that,

Mass of K_2SO_4 = 25 mg = 0.025 g

Molar mass of K_2SO_4 = 174 g/mol
Volume = 2 L
Temperature = 25 + 273 = 298 K
Gas constant = $0.0821 \mathrm{L atm K^{-1} mol^{-1}}$
Dissociation reaction is given by,

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

Number of ions produced = Van’t Hoff factor = i = 3

Thus,

Osmotic pressure is given as,

$$\pi = iCRT$$

$$\pi = i\frac{n}{V}RT = i\frac{W}{M \times V}RT$$

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

$$\pi = 5.272 \times 10^{-3} \text{atm}$$

$\therefore$ The osmotic pressure of the solution is 5.272×10^{-3} atm .