Solutions

Concentration Term

Molality (m) = $\dfrac{\text{Number of moles of solute}}{\text{Weight of solvent (kg)}}$

Molarity (M) = $\dfrac{\text{Number of moles of solute}}{\text{Volume of solution (L)}}$

Normality (N) = $\dfrac{\text{Number of gram-equivalents of solute}}{\text{Volume of solution (L)}}$

Formality (F) = $\dfrac{\text{Number of gram-formula mass of solute}}{\text{Volume of solution (L)}}$

PPM = mole fraction $\times 10^6$ = mass fraction $\times 10^6$

Henry's Law

The dissolution of a gas in a liquid is governed by Henry's law, according to which, at a given temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas.

$p = K_H.X \begin{cases} p = \text{partial pressure of gas in vapour phase} \\ K_H = \text{Henry's law constant} \\ X = \text{Mole fraction of gas} \end{cases}$

Clausius - Clapeyron Equation

$$\log \frac{P_2}{P_1} = \frac{\Delta H_{vapour}}{2.303R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

where, P_1, P_2 are vapour pressure of liquid at T_1 and T_2 respectively.

Roult's Law

$P_T = P_A^o \chi_A + P_B^o \chi_B$ (Solution)

$Y_A = \dfrac{P_A}{P_T} = \dfrac{P_A^o \chi_A}{P_A^o \chi_A + P_B^o \chi_B}$

{Y_A; Y_B = mole fraction of A and B in vapour phase

$P_T = P_A + P_B$

Colligative Properties

These are the properties of solutions which depends on the number of solute particles and independent of their chemical identity. These are relative lowering of vapour pressure, elevation in boiling point, depression in freezing point and osmotic pressure.

Relative Lowering of Vapour Pressure

The vapour pressure of the solvent is lowered by the presence of a non-volatile solute in the solution and this lowering is governed by Raoult's law, according to which relative lowering of vapour pressure of solvent is equal to the mole fraction of non-volatile solute present in solution.

Relative lowering of vapour pressure = χ_B

$\dfrac{P_A^o - P_S}{P_A^o} = \chi_B$ {χ_B = mole fraction of solute}

Elevation of boiling point

$\Delta T_b = K_b \times m$ {m = molality}

$K_b = \dfrac{RT_b^2}{1000\, L_{vapourisation}} = \dfrac{M_A RT_b^2}{1000\Delta H_{vapourisation}}$

where $L_{vapourisation}$ is latent heat of vapourisation in cal./g

$\Delta H_{vapourisation}$ is enthalpy of vapourisation in cal./mol

Depression in freezing point

$\Delta T_f = K_f \times m$

$K_f = \dfrac{RT_f^2}{1000\, L_{fusion}} = \dfrac{M_A RT_f^2}{1000\Delta H_{fusion}}$

where L_{fusion} is latent heat of fusion in cal./g

ΔH_{fusion} is enthalpy of fusion in cal./mol

Osmotic pressure

$\pi = CRT$

For isotonic solution, $\pi_1 = \pi_2$ thus, $C_1 = C_2$.

Van't Hoff Factor

Van't Hoff Factor 'i' is the extent to which a solute is dissociated or associated. This can be defined as ratio of observed colligative property to calculated colligative property.

$i = \dfrac{\text{Experimental colligative property (observed)}}{\text{Calculated colligative property (Normal)}}$

$i = \dfrac{\text{Theoritical molar mass of solute}}{\text{Actual / observed molar mass of solute}}$

Solutes which dissociate in solution exhibit the molar mass lower than the actual molar mass and those which associate show higher molar mass.

$$\frac{P_A^o - P_S}{P_A^o} = \frac{(i\,n_B)}{(i\,n_B + n_A)}$$

where n_B = moles of solute

n_A = moles of solvent
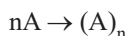
P_A^o is vapour pressure of pure solvent A

P_S is vapour pressure of solution

$\Delta T_b = i \times K_b \times m$
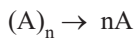
$\Delta T_f = i \times K_f \times m$

$\pi = iCRT$

Association of solute

$nA \rightarrow (A)_n$

$$i = \frac{(1-\alpha) + \dfrac{\alpha}{n}}{1} \quad \{\alpha = \text{degree of association}\}$$

Dissociation of solute

$(A)_n \rightarrow nA$

$$i = \frac{1 + (n-1)\alpha}{1} \quad \{\alpha = \text{degree of dissociation}\}$$

Ideal and Non-Ideal Solutions

Solutions which obey Raoult's law over the entire range of concentration are called ideal solutions. For non-ideal solutions, positive and negative deviations are observed from Raoult's law. Azeotropes arise due to very large deviations from Raoult's law.