Solutions

Molality into molarity $M = \frac{m\ \rho}{1+m\ M_2}$

Molarity into molality $m = \frac{M}{\rho - M\ M_2}$

Where M_1 *and* M_2 are the molar masses of
solvent and solute and ρ is the density of solution.

Volume of solution = n1M1 + n2M2/ p ,n1 and n2 are amount of solvent and solute in solution, p is density of solution.

Morality of solution, M= amount of solute /Volume of solution in dm3

Morality of solution, m= amount of solute /mass of solvent in Kg = n2/n1M1

Mole fraction of solute =x2 = n2 /n1 + n2

A liquid solution is said to be ideal solution if its constituents follow Raoult's law.

An ionic compound that attracts atmospheric water so strongly that a hydrate is formed is called hygroscopic.

Desalinating of sea water is done using Reverse Osmosis.

In the mode of Molality, the expression of concentration of a solution is independent of temperature, while in Molarity, the concentration changes with temperature.

Ways to describe the concentration of a solution quantitatively:

1) Mass percentage: Mass % of a component= Mass of the component in the solution ***/*** Total mass of the solution) * 100
2) Volume percentage: volume % of a component= (volume of the component / total volume of the solution) * 100
3) Mass by volume % = It is the mass of solute dissolved in 100 mL of the solution.
4) Parts per million (ppm) = When a solute is present in **trace** quantities, it is convenient to express concentration in **parts per million (ppm) a**nd is defined as:

(Number of parts of the component)/(total number of parts of all the components in the solution) * 10^6

5) Mole fraction: Commonly used symbol for mole fraction is x and subscript used on the right hand side of x denotes the component. It is defined as:

Number of moles of the component/total number of moles of all the components

6) Molarity: Molarity (*M*) is defined as number of moles of solute dissolved in one litre (or one cubic decimetre) of solution,

Molarity= (moles of solute)/(volume of solution in litre)

7) Molality (m) : defined as the number of moles of the solute per kilogram (kg) of the solvent and is expressed as:

Molality (m)= (moles of solute)/(mass of solvent in Kg)

Solubility:

Solubility of a substance is its maximum amount that can be dissolved in a specified amount of solvent.

A solution in which no more solute can be dissolved at the same temperature and pressure is called a **saturated solution**.

Effect of temperature: In general, if in a *nearly saturated solution,* the dissolution process is endothermic (Δsol H > 0), the solubility should increase with rise in temperature and if it is exothermic (Δsol H > 0) the solubility should decrease.

Effect of pressure: no effect on the solubility of solids in liquids owing to the fact that it’s not compressible. But the effect on the solubility of gases in liquids is profound.

“**the partial pressure of the gas in vapour phase (*p*) is proportional to the mole fraction of the gas (*x*) in the solution”** and is expressed as:

$p = K\text{H}\, x$

The partial vapour pressure of each component in the solution is directly proportional to its mole fraction

$$p_1 \propto x_1$$

$$p_1 = p_1^0 x_1$$

Where p_1^0 is the vapour pressure of pure component and p_1 is the partial vapour pressure of the component in the solution. $x_1$is the mole fraction of the component.

$$p_{total} = p_1^0 + (p_2^0 - p_1^0)x_2$$

This is Raults law.

1) Total vapour pressure can be related to mole fraction of any one component.
2) Total vapour pressure over the solution varies linearly with the mole fraction of the second component.
3) Depending on the vapour pressure of the pure components 1 & 2, total vapour pressure of the solution decreases or increases with the increase of the mole fraction of component 1.

Solutions can be broadly classified in 2 ways:

1) Ideal solutions: they obey Raults law over the entire range of concentration.

Enthalpy of mixing of components to form solution is zero.

Volume of mixing is zero.

The solutions of n-hexane and n-heptane; bromethane and chloroethane; benzene and toluene are all examples of ideal solutions.

2) Non ideal solutions: they do not obey Raults law over the entire range of concentration.

There are again 2 types of non ideal solutions:

i) (+)ve deviation: the actual vapour pressure of the solution is greater than the predicted vapour pressure. This is due to the fact that the intermolecular attraction between the solute and the solvent is very weak.

Eg: ethanol and acetone solution.

ii) (-)ve deviation: the actual vapour pressure of the solution is lower than the predicted vapour pressure. This is due to the fact that the intermolecular attraction between the solute and the solvent is very high.

Eg: phenol and aniline solution.

AZEOTROPES: they are binary mixtures having same composition in liquid and vapour phase and boil at constant temperature. It is NOT possible to separate the mixture by fractional distillation.

They are of 2 types:

1) Minimum boiling type: solutions with large (+)ve deviation form this type of solution.
2) Maximum boiling type: solutions with large (-)ve deviation form this type of solution.

$$\left(\frac{\Delta p_1}{p_1^0}\right) = \left(\frac{p_1^0 - p_1}{p_1^0}\right) = x_2$$

$$\Delta p_1 = reduction\ in\ vapour\ pressure\ of\ solvent$$

$$= \left(\frac{n_2}{n_1 + n_2}\right)$$

$$n_1\ and\ n_2\ are\ the\ number\ of\ moles\ of\ solvent\ and\ solute\ respectively$$

$$\Delta p_1 = \frac{W_2 * M_1}{M_2 * W_1}$$

$$W_1 \; and \; W_2 \; are \; the \; masses \; while \; M_1 \; and \; M_2 \; are \; the \; molar \; masses$$

Due to the decrease in vapour pressure there's an elevation in the boiling point:

$$\Delta T_b = T_b - T_b^0$$

$$\Delta T_b \propto m$$

$$\Delta T_b = K_b m$$

Isotonic solutions are 2 solutions having same Osmosis Pressure at a given temperature.

$$O_p = CRT$$

$$= \left(\frac{n_2}{V}\right) RT$$

$$O_p V = \left(\frac{W_2 RT}{M_2}\right)$$

i = Van’t Hoff factor

which is the extent of association or dissociation in the solution

$$= \left(\frac{Normal\ molar\ mass}{Abnormal\ molar\ mass}\right)$$

$$= \left(\frac{observed\ colligative\ properties}{calculated\ colligative\ property}\right)$$

$$= \left(\frac{total\ number\ of\ moles\ of\ part\ after\ association\ or\ dissociation}{number\ of\ moles\ of\ particles\ before\ association\ or\ dissociation}\right)$$

Sample Examples

Calculate the mole fraction of ethylene glycol ($C_2H_6O_2$) in a solution containing 20% of $C_2H_6O_2$ by mass.

Solution:

Assume that we have 100 g of solution (one can start with any amount of solution because the results obtained will be the same). Solution will contain 20 g of ethylene glycol and 80 g of water. Molar mass of C2H6O2 = 12 × 2 + 1 × 6 + 16 × 2 = 62 g mol^{-1}.

Moles of $C_2H_6O_2$ = (20 g)/(62g/mol) =0.322 mol

Moles of water = (80g)/(18g/mol) =4.444 mol

X(glycol)= (Moles of $C_2H_6O_2$)/((Moles of $C_2H_6O_2$)+ (Moles of water))

=0.068

X(water) = 0.932 (1- 0.068)