Solutions

Parts per million (*ppm*) and parts per billion (*ppb*): It is

convenient to measure the concentration of *trace quantities* of solutes in parts per million and parts per billion. *It is independent of the temperature.*

$$ppm = \frac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^6$$

$$ppb = \frac{\text{mass of solute component}}{\text{Total mass of solution}} \times 10^9$$

Normality (*N*): Normality of solution is equal to the number of gram equivalents of a solute present per *litre* of the solution.

It can be calculated by following formulas;

(i) $\text{Normality (N)} = \frac{\text{Number of gram equivelent of solute}}{\text{Volume of solution (l)}}$

(ii) $\text{Normality}\ (N) = \frac{\text{Weight of solute in g.}}{\text{gram equivelent weight of solute} \times \text{Volume of solution (l)}}$

Molarity (*M*): Molarity of a solution is equal to the *number of moles of the solute present per litre of solution.*

It can be calculated by following formulas,

(i) $M = \frac{\text{No. of moles of solute}(n)}{\text{Vol. of solution in litres}}$,

(ii) If molarity and volume of solution are changed from M_1, V_1 to M_2, V_2. Then,

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$$M_1V_1 = M_2V_2 \text{ (Molarity equation)}$$

(ii) In balanced chemical equation, if n_1 moles of reactant one react with n_2 moles of reactant two. Then,

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

(iii)   Volume of water added to get a solution of molarity M_2 from $V_1 \, ml$ of molarity M_1 is

$$V_2 - V_1 = \left(\frac{M_1 - M_2}{M_2}\right)V_1$$

(iv) If two solutions of the same solute are mixed then molarity (M) of resulting solution.

$$M = \frac{M_1V_1 + M_2V_2}{(V_1 + V_2)}$$

Here, M_1V_1 are molarity and volume of initial solution and M_2V_2 are

molarity and volume of final solution.

Relation between molarity and normality

Normality of solution = molarity $\times \dfrac{\text{Molecular mass}}{\text{Equivalent mass}}$

Normality $\times$ equivalent mass = molarity $\times$ molecular mass

Molality (*m*) : Mathematically molality can be calculated by following formulas,

(i) $m = \dfrac{\text{Number of moles of the solute}}{\text{Weight of the solvent in grams}} \times 1000$

(ii) $m = \dfrac{\text{Strength per 1000 grams of solvent}}{\text{Molecular mass of solute}}$

Relation between molarity (*M*) and molality (*m*)

Molality $(m) = \dfrac{\text{Molarity}}{\text{Density} - \dfrac{\text{Molarity} \times \text{molecular mass}}{1000}}$

Formality (*F*) : Formality of a solution may be defined as the number of gram formula masses of the ionic solute dissolved per litre of the solution. It is represented by F.

Formality (F) $= \dfrac{\text{Number of gram formula masses of solute}}{\text{Volume of solution in litres}}$

Mole fraction (*X*): To find the mole fraction, divide moles of a constituent(either solute or solvent) by the total moles of both the constituents (solute and solvent).It is denoted by the letter X.

Let in a solution A is solute and B is solvent, then their mole fractions are given below;

Mole fraction of Solute A, $X_A = \dfrac{n_A}{n_A + n_B}$

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Mole fraction of solvent B, $X_B = \dfrac{n_B}{n_A + n_B}$

Azeotropic mixture

Azeotropes are defined as the mixtures of liquids that boil like a pure liquid at constant temperature and possess components in same composition, in liquid as well as in vapour phase.

Osmotic pressure (π)

The osmotic pressure of a solution at a particular temperature may be defined as the excess hydrostatic pressure that builds up when the solution is separated from the solvent by a semi-permeable membrane. It is denoted by π.

According to Boyle Van't Hoff law (at conc. temp.)

$\pi \propto C$ .....(i)

According to Gaylussac Van't Hoff law (at conc. temp.)

$\pi \propto T$ .....(ii)

From equation (i) and (ii)

$\pi \propto CT$, $\pi = CRT$ (Van't Hoff equation)

$\pi = \dfrac{n}{V} RT$ $\left(\because C = \dfrac{n}{V} \right)$; $\pi = \dfrac{w}{m} \dfrac{RT}{V}$ $\left(\because n = \dfrac{w}{m} \right)$

Here, C = concentration of solution in *moles* per *litre*

R = gas constant; T = temperature

n = number of *moles* of solute; V = volume of solution

m = molecular weight of solute; w = weight of solute

(iv) ***Relation of osmotic pressure with different colligative properties*:** Osmotic pressure is related to relative lowering of vapour pressure, elevation of boiling point and depression of freezing point according to the following relations,

(a) $\pi = \left(\dfrac{P_A^o - P_A}{P_A^o} \right) \times \dfrac{dRT}{M}$ (b) $\pi = \Delta T_b \times \dfrac{dRT}{1000 \times K_b}$

(c) $\pi = \Delta T_f \times \dfrac{dRT}{1000 \times K_f}$

In the above relations, π = Osmotic pressure; d = Density of solution at temperature T; R = Universal gas constant;

M = Mol. Mass of solute;

K_b = Molal elevation constant of solvent;

K_f = Molal depression constant of solvent

Elevation in b.pt. of the solvent

If T_b is the boiling point of the solvent and T is the boiling point of the solution, the difference in the boiling point (ΔT or ΔT_b) is called the elevation of boiling point.

$T - T_b = \Delta T_b$ or ΔT

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Elevation in boiling point is determined by Landsberger's method and Cottrell's method. Study of elevation in boiling point of a liquid in which a non-volatile solute is dissolved is called as ebullioscopy.

Important relations concerning elevation in boiling point

(1) $\Delta T_b \propto p^0 - p$

Here, $p^0 - p$ is *lowering of vapour pressure.*

(ii) $\Delta T_b = K_b \times m$

Here, $K_b =$ molal elevation constant; $m =$ Molality of the solution; $\Delta T_b =$ Elevation in boiling point

As you know, $m = \dfrac{w \times 1000}{W}$

Thus, $\Delta T_b = \dfrac{1000 \times K_b \times w}{m \times W}$ or $m = \dfrac{1000 \times K_b \times w}{\Delta T_b \times W}$

Here, K_b is *molal elevation constant*

Depression in f.pt. of the solvent

The depression in freezing point $(\Delta T$ or $\Delta T_f)$ of a solvent is the difference in the freezing point of the pure solvent (T_s) and the solution (T_{sol}).

$T_s - T_{sol} = \Delta T_f$ or ΔT

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Important relations concerning depression in freezing point.

(i) $\Delta T_f \propto p^0 - p$

Here, $p^0 - p$ is Depression in freezing point.

(ii) $\Delta T_f = K_f \times m$

Here, $K_f =$ molal depression constant $m =$ Molality of the solution; $\Delta T_f =$ Depression in freezing point.

As you know, $m = \dfrac{w \times 1000}{W}$

Thus,

$$\Delta T_f = \dfrac{1000 \times K_f \times w}{m \times W} \text{ or } m = \dfrac{1000 \times K_f \times w}{\Delta T_f \times W}$$

Where K_f is *molal depression constant* .

Van't Hoff's factor (*i*): In 1886, Van't Hoff introduced a factor '*i*' called Van't Hoff's factor, to express the extent of association or dissociation of solutes in solution. It is ratio of the normal and observed molecular masses of the solute, *i.e.*,

$$i = \dfrac{\text{Normal molecular mass}}{\text{Observed molecular mass}}$$

Relative lowering of vapour pressure $= \dfrac{P_A^o - P_A}{P_A^o} = iX_B$

Elevation of boiling point, $\Delta T_b = ik_b m$

Depression in freezing point, $\Delta T_f = ik_f m$

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Osmotic pressure, $\pi = \dfrac{inRT}{V}$; $\pi = iCRT$

From the value of 'i', it is possible to calculate degree of dissociation or degree of association of substance.

Degree of dissociation (α) : or ions.

The fraction of total molecules that dissociate into simpler mole cules or ions is known as Degree of dissociation.

$\alpha = \dfrac{i-1}{m-1}$; m= number of particles in solution

Degree of association (α) :

The fraction of the total number of molecules which associate or combine together resulting in a bigger molecules being formed.

$\alpha = \dfrac{i-1}{1/m-1}$; m = number of particles in solution.