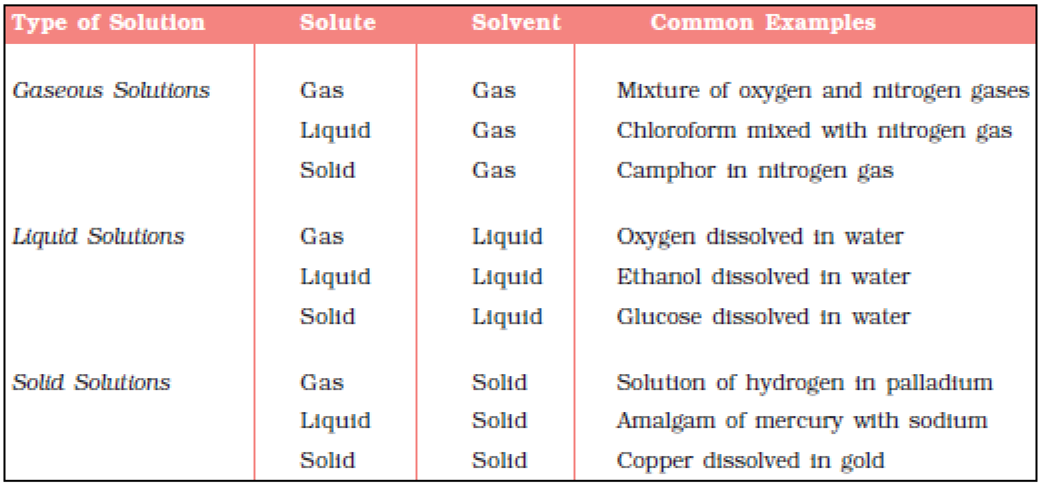
Chapter 2- Solution

1. Solution: Solution is a homogeneous mixture of two or more substances in same or different physical phases. The substances forming the solution are called components of the solution. On the basis of number of components a solution of two components is called binary solution.

Solute and Solvent: In a binary solution, solvent is the component which is present in large quantity while the other component is known as solute

2. Types of Solutions:

(A) Following types of solutions are seen on the basis of physical state of solute and solvent.



Type of Solution	Solute	Solvent	Common Examples
Gaseous Solutions	Gas	Gas	Mixture of oxygen and nitrogen gases
	Liquid	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	Copper dissolved in gold

If water is used as a solvent, the solution is called aqueous solution and if not, the solution is called non-aqueous solution.

(B) Depending upon the amount of solute dissolved in a solvent we have the following types of solutions:

(i) Unsaturated solution: A solution in which more solute can be dissolved without raising temperature is called an unsaturated solution.

(ii) Saturated solution: A solution in which no solute can be dissolved further at a given temperature is called a saturated solution.

(iii) Supersaturated solution: A solution which contains more solute than that would be necessary to saturate it at a given temperature is called a supersaturated solution.

3. Solubility: The maximum amount of a solute that can be dissolved in a given amount of solvent (generally 100 g) at a given temperature is termed as its solubility at that temperature. The solubility of a solute in a liquid depends upon the following factors:

(i) Nature of the solute
(ii) Nature of the solvent
(iii) Temperature of the solution
(iv) Pressure (in case of gases)

4. Henry's Law: It states that at a constant temperature, the solubility of a gas in a liquid is directly proportional to the pressure of the gas. In other words the partial pressure (p) of the gas in vapour phase is proportional to the mole fraction (x) or solubility of the gas in the solution, it expressed as

$p = K_H \cdot x$ Here K_H is the Henry's law constant and its unit is Pa (1 atm = 10^5Pa)

Applications

(i) In manufacture of soft drinks and soda water, CO_2 is passed at high pressure to increase its solubility.

(ii) To minimize the painful effects (bends) accompanying the decompression of deep sea divers. O_2 diluted with less soluble. He gas is used as breathing gas.

(iii) At high altitudes, the partial pressure of O_2 is less then that at the ground level. This leads to low concentrations of O_2 in the blood of climbers which causes anoxia.

5. Methods of Expressing Concentration of Solutions:

(i) Mass Percentage (w/w): It is defined as the amount of solute present in 100 g of solution.

$$\text{Mass \% of a solute} = \frac{\text{Mass of the solute}}{\text{Total mass of the solution}} \times 100$$

$$\text{Mass \% of a solute} = \frac{\text{Mass of the solute}}{\text{Mass of solute (gm)} + \text{Mass of Solvent (gm)}} \times 100$$

(ii) Volume Percentage (V/V): It is defined as the volume of solute present in 100 mL of solution.

$$\text{Volume \% of a solute} = \frac{\text{Volume of the solute}}{\text{Total volume of solution}} \times 100$$

(iii) Mass by Volume Percentage (w/V): The mass of solute present in 100 mL of solution.

$$\text{Mass by volume \%} = \frac{\text{Mass of the solute}}{\text{Total volume of solution}} \times 100$$

(iv) Mole fraction (x):

$$\textbf{Mole fraction } (x_B) = \frac{\text{No. of moles of solute } (n_B)}{\text{No. of mole of solute } (n_B) + \text{No. of moles of Solvent } (n_A)}$$

$$x_B = \frac{n_B}{n_A + n_B} = \frac{\dfrac{W_B}{M_B}}{\dfrac{W_A}{M_A} + \dfrac{W_B}{M_B}}$$

Where: W_B = Mass of solute (B); M_B = Molar mass of solute (B); W_A = Mass of solvent (A); M_A = Molar mass of solvent (A)

(v) Parts per million (ppm):

$$\text{Parts per million} = \frac{\text{Number of parts of the component}}{\text{Total number of parts of all components of the solution}} \times 10^6$$

(vi) Molarity (M): It is the number of moles of solute present in 1 litre (dm3) of the solution. Molarity varies with temperature due to change in volume of solution.

$$\textbf{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (liter)}}$$

$$\textbf{Molarity} = \frac{\text{mass of solute (gm) x 1000}}{\text{Molar mass (m.wt.) of solute (gm /mole) x Vol. of sol}^{\text{n}}\text{ (ml)}}$$

$$\textbf{Molarity} = \frac{\text{Density of solution (gm / liter)}}{\text{Molar mass (m.wt.) of solute (gm / mole)}}$$

$$\textbf{Molarity} = \frac{\text{\% concentration x density x 10}}{\text{Molar mass}}$$

When molarity of a solution is 1 M, it is called a molar solution. 0.1 M solution is called a decimolar solution while 0.5 M solution is known as semi molar solution.

(vii) Molality (m): It is the number of moles of solute per kilogram of the solvent.

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$$

(viii) Dilution law: $M_1V_1 = M_2V_2$ or $M_1V_1/n_1 = M_2V_2/n_2$ (For reaction b/w two reactants)
Where, n_1 and n_2 arc stoichiometric coefficient in balanced equation.

6. Raoult's Law for volatile solute (Vapour pressure of liquid - liquid solution): It states that for a solution of volatile liquids the partial pressure of each component is directly proportional to its mole fraction. The partial vapour pressure of component 1 is

$$p_1 \propto x_1$$
$$p_1 = p_1^0 x_1$$

Similarly for another component 2

$$p_2 = p_2^0 x_2$$

p_1^0 and p_2^0 are constant at a given temperature. It is equal to the vapour pressure of pure solvent or components 1 and 2.

According to Dalton's law of partial pressure, if *p* is total vapour pressure then

$$p_{\text{total}} = p_1 + p_2$$

$p_{total} = x_1 p_1^0 + x_2 p_2^0$ While $x_1 + x_2 = 1$

$= (1 - x_2) p_1^0 + x_2 p_2^0$

$= p_1^0 - x_2 p_1^0 + x_2 p_2^0$

$= p_1^0 + (p_2^0 - p_1^0) x_2$

Following conclusion carried out from above equation-

- Total vapor pressure over the soln can be related to the mole fraction of any one component.
- Total vapor pressure over the soln varies linearly with the mole fraction of any one component (here- x_2).
- Depending on the vapor pressures of the pure components 1 and 2, total vapor pressure over the solution decreases or increases with the increases of the mole fraction of component 1.

The composition of vapour phase in equilibrium with the solution is determined by the partial pressure of components. If y_1 and y_2 are the mole fraction of component 1 and 2 respectively in the vapour phase, then

$p_1 = y_1 p_{total}$

$p_2 = y_2 p_{total}$

In general $p_i = y_i p_{total}$

7. Raoult's Law for a solution containing a non-volatile solute and volatile solvent (Vapour pressure of solution of solids in liquids): It states that the relative lowering of vapour pressure is equal to mole fractions of solute which is non volatile. Mathematically,

$p_{total} = p_1 + p_2$

$p_{total} = p_1$ (Since solute 2 is non volatile)

$p_{total} = x_1 p_1^0$

$p_{total} = (1 - x_2) p_1^0 = p_1^0 - p_1^0 x_2$

$$x_2 = \frac{p_1^0 - p_{total}}{p_1^0}$$

Or Mole fraction of solute = Relative lowering of vapour pressure

8. Ideal Solutions: A solutions which obey Raoult's law over the entire range of concentration at a specified temperature are known as ideal solutions. The ideal solutions have three other important properties:

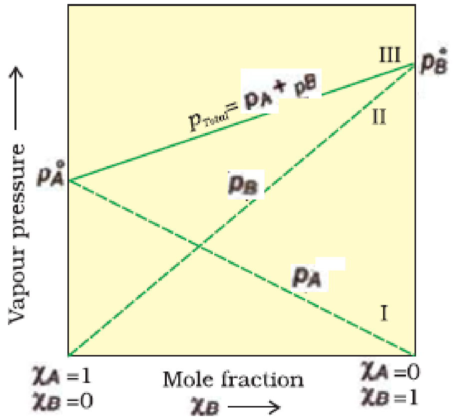
(i) ΔHmix = 0 (No heat is absorbed or evolved when the components are mixed.)

(ii) ΔVmix = 0 (No expansion or contraction when the components are mixed.)

(iii) $p_{total} = p_A^0 \chi_A + p_B^0 \chi_B$

Reason for formation of ideal solution:

- Liquids having similar structure and polarity are likely to form ideal solutions. For example mixture of benzene and toluene, n-hexane and n-heptane, methanol and ethanol etc.
- In pure components, the intermolecular attractive interactions will be of types A-A and B-B, whereas in the binary solutions in addition to these two interactions, A-B type of interactions will also be present. If the intermolecular attractive forces between the A-A and B-B are nearly equal to those between A-B, this leads to the formation of ideal solution. i.e.,
  A-B = A-A and B-B interactions.



9. Non-ideal Solutions: Those solutions which show deviation from Raoult‟s law is called non-ideal solution. For such solutions,

$\Delta H_{mix} \neq 0$

$\Delta V_{mix} \neq 0$

A non ideal solution can show either positive or negative deviation from Raoult’s law.

(i) Non-ideal solutions showing positive deviation:

- The partial vapour pressure of a solution is higher than that predicted by Raoult’s law.

$$p_A > p_A^\circ \chi_A,\ p_B > p_B^\circ \chi_B$$

$$p_{total} > p_A^\circ \chi_A + p_B^\circ \chi_B$$

- The intermolecular attractive forces b/w solute solvent molecules are weaker than those b/w solute-solute and solvent-solvent molecules i.e., A-B< A-A and B-B interactions.

For such solution $\Delta H_{mix} > 0$

$\Delta V_{mix} > 0$

- Some examples of the solution exhibiting positive deviations are ethanol and water, acetone and carbon disulphide, acetone and benzene, carbon tetrachloride and benzene.

(ii) Non-ideal solution showing negative deviation:

- The vapour pressure of a solution is lesser than that predicted by Raoult’s law.

$$p_A < p_A^\circ \chi_A,\ p_B < p_B^\circ \chi_B$$

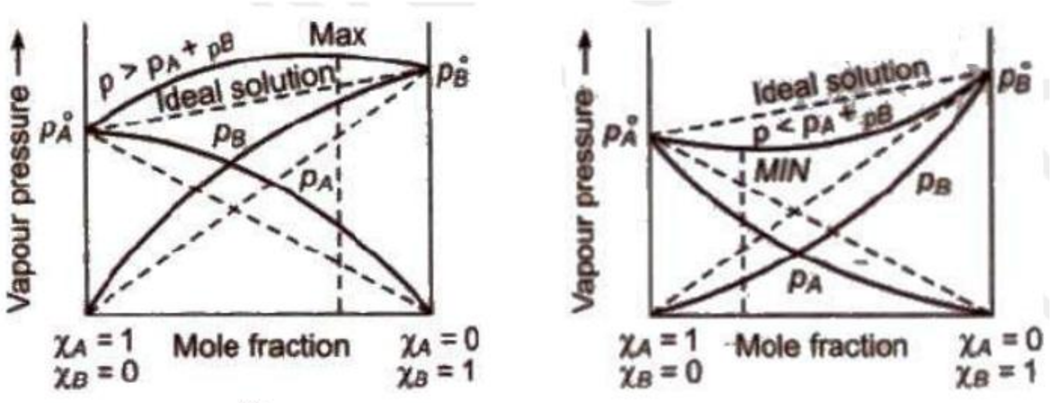
$$p_{total} < p_A^\circ \chi_A + p_B^\circ \chi_B$$

- The intermolecular attractive forces b/w solute solvent molecules are stronger than those b/w solute-solute and solvent-solvent molecules i.e., A-B > A-A and B-B interactions.

For such solution $\Delta H_{mix} < 0$

$\Delta V_{mix} < 0$

- Some examples of the solution exhibiting negative deviations are nitric acid and water, chloroform and acetone, HCl and water, acetic acid and pyridine.



(a) (b)

Vapour pressure diagram showing (a) positive deviation (b) negative deviation

10. Azeotropes or Azeotropes Mixture: These are the binary mixture having the same composition in liquid and vapour phase and boil at a constant temperature. Liquid forming azeotropes cannot be separated by fractional distillation. These are formed by non ideal solutions.

Types of Azeotropes:

(i) Minimum boiling azeotropes: Those liquid pairs which show positive deviation from Raoult's law form minimum boiling azeotropes at a specific composition. Such azeotropes have boiling points lower than either of the two components, e.g., a mixture of ethanol (95.57%) and water (4.43%) by mass.

(ii) Maximum boiling azeotropes: Those liquid pairs which show negative deviation from Raoult's law form maximum boiling azeotropes at a specific composition. Such azeotropes have boiling points higher than either of the two components, e.g., a mixture of water (20.22%) and HCl (79.78%) by mass.

11. Colligative Properties: Those properties which depend only upon the number of solute particles (molecules, atoms or ions) in a solution but not upon their nature are called colligative properties. Following are the main colligative properties-

(i) Relative Lowering of Vapour Pressure
(ii) Elevation in Boiling Point (ΔT_b)
(iii) Depression in Freezing Point (ΔT_f)
(iv) Osmotic Pressure (π)

(i) Relative Lowering of Vapour Pressure:

- The addition of a non volatile solute to a volatile solvent decreases the vapor pressure of a solvent during formation of solution.

- The difference in the vapor pressures of pure solvent p°_A and solution p_A represent the lowering in vapor pressure ($p^\circ_A - p_A$).
- So the ratio of lowering in vapour pressure to vapour pressure of pure solvent is called Relative Lowering of Vapour Pressure ($p^\circ_A - p_A$) / p°_A .
- The relative lowering in vapour pressure is directly proportional to the mole fraction of solute in the solution so it's a colligative property.

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \chi_B$$

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \frac{n_B}{n_A + n_B}$$

For dilute solutions, $n_B << n_A$

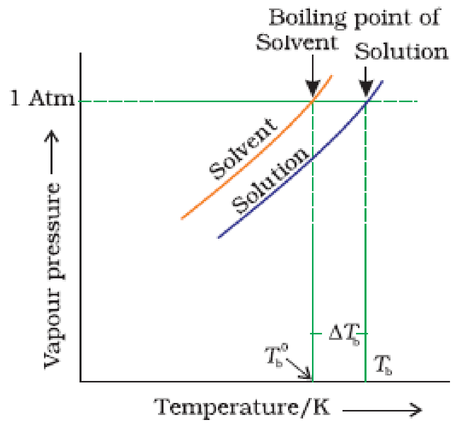
$$\frac{p_A^\circ - p_A}{p_A^\circ} = \frac{n_B}{n_A} = \frac{W_B \times M_A}{M_B \times W_A}$$

$$M_B = \frac{W_B}{W_A} \times M_A \times \frac{p_A}{(p_A^\circ - p_A)}$$

Above expression is used to find the molecular weight of an unknown solute dissolved in a given solvent. Where, W_B and W_A are mass of solute and solvent respectively. M_B and M_A are molecular weight of solute and solvent respectively.

(ii) Elevation in Boiling Point (ΔT_b):

- Boiling point of a liquid is the temperature at which its vapor pressure becomes equal to the atmospheric pressure. As the vapor pressure of solution is containing a non volatile solute is lower than that of pure solvent so it boiling point will be always higher than that of pure solvent in which the solution is prepared as show in figure-



- So the difference in boiling point of solution T_b and pure solvent T°_b is known as elevation in boiling point ΔT_b.

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

- For a dilute solution elevation in boiling point is directly proportional to molal concentration of the solute in solution hence it is a colligative property.
  $\Delta T_b = K_b$ m (where; m = Molality)
  Where K_b is molal elevation constant or ebullioscopic constant and its unit is K kg $mole^{-1}$, for water, the value of $K_b = 0.52$ K kg mol^{-1}.
- Molecular mass of solute can be calculated as

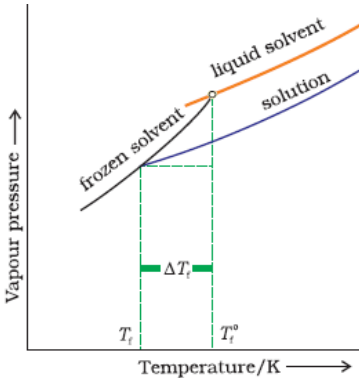
$$\Delta T_b = K_b \times \frac{W_B \times 1000}{M_B \times W_A}$$

$$M_B = \frac{K_b \times W_B \times 1000}{\Delta T_b \times W_A}$$

  Where, W_B and W_A = mass of Solute and solvent respectively, M_B = molecular weight of solute.

(iii) Depression in Freezing Point (ΔT_f):

- Freezing point of a liquid is the temperature at which vapour pressure of the solvent in its liquid and solid phase become equal. As we know that vapour pressure of solution containing non-volatile solute is lower than that of pure solvent, solid form gets separated out at a lower temperature as shown in the figure.



- This decrease in freezing point of a liquid is known as depression in freezing point.
  Depression in freezing point $\Delta T_f = T_f^0 - T_f$
- For a dilute solution depression is a colligative property because it is directly proportional to molal concentration of solute. To find molecular mass of solute,

$$\Delta T_f = K_f \times \frac{W_B \times 1000}{M_B \times W_A}$$

$$M_B = \frac{K_f \times W_B \times 1000}{\Delta T_f \times W_A}$$

  Where K_f is molal depression constant or cryoscopic constant. Unit of K_f is K kg mol^{-1}.
- Ethylene glycol is usually added to water in the radiator to lower its freezing point. It is called antifreeze solution.

- Common salt (NaCl) and anhydrous $CaC1_2$ are used to clear the snow on the roads because they depress the freezing point of water.
- The value of K_b and K_f, depend upon the nature of the solvent and concentration of the solution, which can be calculated as following

$$K_b = \frac{R \times M_A \times (T_b^0)^2}{1000 \times \Delta_{vap}H}$$

$$K_f = \frac{R \times M_A \times (T_f^0)^2}{1000 \times \Delta_{fus}H}$$

Where, R= Universal gas constant
M_A = Molecular mass of solvent
T_b^0 = Boiling point of pure solvent
$\Delta_{vap}H$ = Enthalpy for the vaporization of solvent
T_f^0 = Freezing point of pure solvent
$\Delta_{fus}H$ = Enthalpy for the fusion of solid solvent

(iv) Osmosis and Osmotic Pressure (π):

- **Osmosis:** is the phenomenon of spontaneous flow of the solvent molecules through a semi permeable membrane from pure solvent to solution or from a dilute solution to concentrated solution. Some natural semi permeable membranes are animal bladder, cell membrane etc.
- Copper ferrocyanide $Cu_2[Fe(CN)_6]$ is an artificial semi permeable membrane which does not work in non-aqueous solutions as it dissolves in them.
- **Osmotic pressure (π):** The hydrostatic pressure developed on the solution which just prevents the osmosis of pure solvent into the solution through a semi permeable membrane is called osmotic pressure.

$$\pi = CRT$$

$$\pi = (n_B / V)\, R\, T = \frac{w_B\, R\, T}{M_B\ V}$$

$$M_B = \frac{w_B\, R\, T}{\pi\ \ V}$$

- Osmotic pressure is a colligative property b/c it depends on the number of solute particles and not their identity. On the basis of osmotic pressure, the solution can be

a) **Isotonic solution:** Two solutions are called isotonic if they exert the same osmotic pressure. These solutions have same molar concentration. 0.91% solution of pure NaCl is isotonic with human RBC"s.

b) **Hypotonic solution:** A solution is called hypotonic if its osmotic pressure is lower than that of the solution from which it is separated by a semipermeable membrane.

c) **Hypertonic solution:** A solution is called hypertonic if its osmotic pressure is higher than that of the solution from which it is separated by a semipermeable membrane.

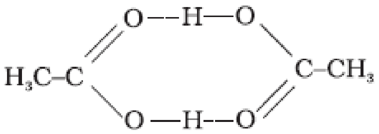
d) **Plasmolysis:** When a plant cell is placed in a hypertonic solution, the fluid from the plant cell comes out and cell shrinks, this phenomenon is called plasmolysis.

e) **Reverse osmosis:** When the external pressure applied on the solution is more than osmotic pressure, the solvent flows from the solution to the pure solvent, I which is called reverse osmosis. Desalination of sea water is done by reverse Osmosis.

12. Abnormal Molar Masses: Molar mass that is either lower or higher than expected or normal moral mass is called as abnormal molar mass. It can occurs due to following two reasons-

(i) Association of molecules: Higher values of molar mass are observed in case of association of molecules. For example CH_3COOH is dissolved in benzene it form dimmer due to associate of molecules hence its show a molar mass 120 (normal molar mass is 60). Hence the number of solute particles becomes nearly half the initial number of solute.

$$2\,CH_3COOH \rightleftharpoons (CH_3COOH)_2 \quad \text{or}$$



(ii) Dissociation of molecules: Lower values of molar mass are observed in case of dissociation of molecules. For example dissociation of KCl in water in to K^+ and Cl^- ions so its observed molar mass is 37.25 (74.5/2).

$$KCl \rightleftharpoons K^+ + Cl^-$$

These observed values are corrected by multiplying with Van't Hoff's Factor (i).

Van't Hoff's Factor (i): The ratio of observed colligative property to the calculated colligative property is known as Van't Hoff factor (i).

$$i = \frac{\text{Normal molar mass (when no association or dissociation)}}{\text{Abnormal molar mass}}$$

$$i = \frac{\text{Observed colligative property}}{\text{Calculated colligative property}}$$

$$i = \frac{\text{Total No. of moles of particles after association / dissociation}}{\text{No. of moles of particles before association / dissociation}}$$

Actually, abnormal molecular masses are experimentally determined when association or dissociation takes place.

Now, there may be three cases

(a) For a solute which doesn't undergo any association or dissociation, i = 1, so observed colligative property is equal to calculated colligative property, e.g. Glucose, Urea, Sucrose etc. (when dissolved in water)

(b) For a solute which undergoes dissociation, number of solute particles in solution increases and therefore, i > 1 means observed colligative property is higher than calculated Colligative Property. For example, if complete dissociation occurs then

i = 2 for KCl, $NaNO_3$

i = 3 for $CaCl_2$, $Mg(NO_3)_2$, Na_2SO_4

(c) For solute which undergoes association, number of solute particles in solution decreases and therefore, i < 1 means observed colligative property is less than calculated colligative properties. In case of complete association i = ½ for CH_3COOH or C_6H_5OH

13. Modified form of colligative properties: To correct the observed value of molar mass, van't Hoff factor must be included in different expressions for colligative properties-

Relative lowering of vapour pressure of solvent, $\frac{p_1^\circ - p_1}{p_1^\circ} = i.\frac{n_B}{n_A}$

Elevation of Boiling point, $\Delta T_b = i\ K_b\ m$

Depression of Freezing point, $\Delta T_f = i\ K_f\ m$

Osmotic pressure of solution, $\pi = i\ n_B R\ T\ /\ V$

14. Degree of dissociation (α): If one molecule of substance A get dissociate in to n particles or molecule nP, then degree of dissociation can be find out as

	A	→	nP
Initially	**1 mol**		**0**
Equillibrium	**1- ά**		**n α**

Total number of moles at equilibrium = 1- ά + n α where n= no. of particles (product)

$$i = \frac{\text{Total No. of moles of particles after dissociation}}{\text{No. of moles of particles before dissociation}}$$

$$i = \frac{1 - ά + n\alpha}{1}$$

$$\therefore \quad \alpha = \frac{i-1}{n-1}$$

Dissociation constant: $ka = \frac{C\alpha^2}{1-\alpha}$

15. Degree of association (α): If n molecule of substance A associate to form a bigger molecule A_n then degree of association can be find out as

	nA	→	A_n
Initially	**1 mol**		**0**
Equillibrium	**1- ά**		$\frac{\alpha}{n}$

Total number of moles at equilibrium $= 1- \acute{\alpha} + \frac{\alpha}{n}$,Where n = number of particles

$$i = \frac{\text{Total No. of moles of particles after association}}{\text{No. of moles of particles before association}}$$

$$i = \frac{1- \acute{\alpha} + \frac{\alpha}{n}}{1}$$

$$\alpha = \frac{i - 1}{\frac{1}{n} - 1}$$

Van't Hoff factor i > 1 then solute undergoing dissociation and it is i < 1 then solutes undergoing association

Some important formulas

1. Mass Percentage (w/w): It is defined as the amount of solute present in 100 g of solution.

$$\text{Mass \% of a solute} = \frac{\text{Mass of the solute}}{\text{Total mass of the solution}} \times 100$$

$$\text{Mass \% of a solute} = \frac{\text{Mass of the solute}}{\text{Mass of solute (gm)} + \text{Mass of Solvent (gm)}} \times 100$$

2. Volume Percentage (V/V): It is defined as the vol. of solute present in 100 mL of solution.

$$\text{Volume \% of a solute} = \frac{\text{Volume of the solute}}{\text{Total volume of solution}} \times 100$$

3. Mass by Volume Percentage (w/V): The mass of solute present in 100 mL of solution.

$$\text{Mass by volume \%} = \frac{\text{Mass of the solute}}{\text{Total volume of solution}} \times 100$$

4. Mole fraction (x):

$$\textbf{Mole fraction } (x_B) = \frac{\text{No. of moles of solute } (n_B)}{\text{No. of mole of solute } (n_B) + \text{No. of moles of Solvent } (n_A)}$$

$$x_B = \frac{n_B}{n_A + n_B} = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

Where: W_B = Mass of solute (B); M_B = Molar mass of solute (B); W_A = Mass of solvent (A) etc.

5. Molarity (M):

$$\textbf{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (liter)}}$$

$$\textbf{Molarity} = \frac{\text{mass of solute (gm) x 1000}}{\text{Molar mass (m.wt.) of solute (gm /mole) x Vol. of sol}^{n}\text{ (ml)}}$$

$$\textbf{Molarity} = \frac{\text{Density of solution (gm / liter)}}{\text{Molar mass (m.wt.) of solute (gm / mole)}}$$

$$\textbf{Molarity} = \frac{\text{\% concentration x density x 10}}{\text{Molar mass}}$$

6. Molality (m): It is the number of moles of solute per kilogram of the solvent.

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}}$$

7. Dilution law: $M_1V_1 = M_2V_2$ or $M_1V_1/n_1 = M_2V_2/n_2$ (For reaction b/w two reactants)
Where, n_1 and n_2 arc stoichiometric coefficient in balanced equation.

8. Henry's Law: Partial pressure $p = K_H . x$; K_H = Henry's law constant and its unit Pa.
1 atm = 10^5Pa; 1 atm = 1.013 bar; 1 atm = 760 mm Hg

9. Raoult's law: $p_{total} = x_1 p_1^0 + x_2 p_2^0$ While $x_1 + x_2 = 1$

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

10. Relative Lowering of Vapour Pressure: $\frac{p_A^\circ - p_A}{p_A^\circ} = \chi_B$

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \frac{n_B}{n_A + n_B}$$

11. Elevation in boiling point: $\Delta T_b = T_b - T_b^0$

$$\Delta T_b = K_b \times \frac{W_B \times 1000}{M_B \times W_A}$$

12. Depression in Freezing Point (ΔT_f): $\Delta T_f = T_f^0 - T_f$

$$\Delta T_f = K_f \times \frac{W_B \times 1000}{M_B \times W_A}$$

13. Osmotic pressure (π): $\pi = CRT$

$$\pi = (n_B / V) R T = \frac{w_B R T}{M_B V}$$

14. Van't Hoff's Factor (i): $= \frac{\text{Observed colligative property}}{\text{Calculated colligative property}}$

$$i = \frac{\text{Total No. of moles of particles after association / dissociation}}{\text{No. of moles of particles before association / dissociation}}$$

15. Modified form of colligative properties: To correct the observed value of molar mass, van't Hoff factor must be included in different expressions for colligative properties.

Relative lowering of vapour pressure of solvent, $\frac{p_1^\circ - p_1}{p_1^\circ} = i . \frac{n_B}{n_A}$

Elevation of Boiling point, $\Delta T_b = i K_b m$

Depression of Freezing point, $\Delta T_f = i K_f m$

Osmotic pressure of solution, $\pi = i n_B R T / V$

14. Degree of dissociation (α): If one molecule of substance A get dissociate in to n particles or molecule nP, then degree of dissociation can be find out as

	A	$\longrightarrow$	nP
Initially	**1 mol**		**0**
Equillibrium	$1-\acute{\alpha}$		$n\alpha$

Total number of moles at equilibrium $= 1-\acute{\alpha} + n\alpha$ where n= no. of particles (product)

$$i = \frac{\text{Total No. of moles of particles after dissociation}}{\text{No. of moles of particles before dissociation}}$$

$$i = \frac{1-\acute{\alpha} + n\alpha}{1}$$

$$\therefore \quad \alpha = \frac{i-1}{n-1}$$

Dissociation constant: $ka = \dfrac{C\alpha^2}{1-\alpha}$

15. Degree of association (α): If n molecule of substance A associate to form a bigger molecule A_n then degree of association can be find out as

	nA	$\longrightarrow$	A_n
Initially	**1 mol**		**0**
Equillibrium	$1-\acute{\alpha}$		$\dfrac{\alpha}{n}$

Total number of moles at equilibrium $= 1-\acute{\alpha} + \dfrac{\alpha}{n}$,Where n = number of particles

$$i = \frac{\text{Total No. of moles of particles after association}}{\text{No. of moles of particles before association}}$$

$$i = \frac{1-\acute{\alpha} + \dfrac{\alpha}{n}}{1}$$

$$\alpha = \frac{i-1}{\dfrac{1}{n}-1}$$

Van't Hoff factor i > 1 then solute undergoing dissociation and it is i < 1 then solutes undergoing association