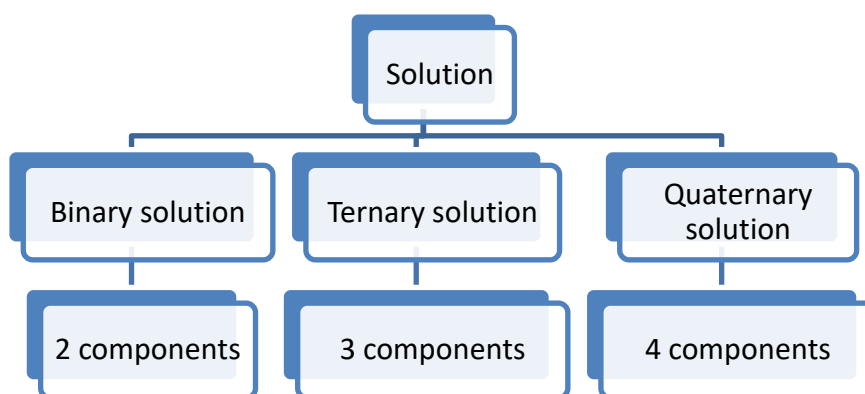# ICSE Class XII Solutions

Solution – The homogeneous mixture of two or more components is called solution. The component present in largest quantity is known as solvent. One or more components present in solution are known as solutes.

Classification of solution

Types of solutions

1. Gaseous Solution

Solute	Solvent	Examples
Gas	Gas	Mixture of O_2 and N_2 gas
Liquid	Gas	$CHCl_3$ mixed with N_2 gas
Solid	Gas	Camphor in N_2 gas

2. Liquid Solution

Solute	Solvent	Examples
Gas	Liquid	O_2 gas dissolved in water
Liquid	Liquid	Ethanol dissolved in water
Solid	Liquid	Glucose dissolved in water

3. Solid solutions

Solute	Solvent	Examples
Gas	Solid	Solution of H_2 gas in Palladium
Liquid	Solid	Amalgam of mercury with sodium
Solid	Solid	Copper dissolved in gold

Concentration of solutions

- Constitution of a solution can be described by expressing its concentration.
- The concentration of solution can be expressed qualitatively and quantitatively.
- There are many ways to describe concentration quantitatively.

1. **Mass percent or weight percent (w/W%)-** the mass percentage of a component of a solution is defined as:

   Mass percent of a component (w/W %)

   $$= \frac{\text{Mass of component in the solution}}{\text{Total mass of solution}} \times 100$$

   Example – If a solution is described by 10% sodium chloride in water by mass, it means 10 g of sodium chloride is dissolved in 90g of water resulting in a 100g solution.

2. **Volume percent (V/V%) -**
   The volume percentage is defined as:

   Volume % of a component (V/V%) = $\frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$

   For example: 10% C_2H_5OH solution in water means 10ml of C_2H_5OH is dissolved in 100 ml water in such a way that total volume of solution is 100ml.

3. **Mass by volume percentage (w/V%)-**
   The mass of solute dissolved in 100ml of the solution.

   Mass by volume percentage (w/V%) = $\frac{W_B}{V_A + V_B} \times 100$

4. **Parts per million-** When a solute is present in very small quantity. It is easier to express its concentration in parts per million (ppm).
   Parts per million (ppm)

   $$= \frac{\text{No. of parts of the component}}{\text{Total no. of parts of all the components of solution}} \times 10^6$$

5. Mole fraction-

➢ It is defined as the number of moles of a component present in the total number of moles of solution. It is denoted by X.

Mole fraction of a component = $\dfrac{\text{No. of moles of the component}}{\text{Total no. of moles of all the components}}$

➢ If Substance A is dissolved in substance B. Their numbers of moles are n_A and n_B respectively.
➢ Then mole fraction of A = X_A

$$= \frac{n_A}{n_A + n_B}$$

➢ Mole fraction of B = X_B

$$= \frac{n_B}{n_A + n_B}$$

➢ For a solution containing i number of components, we have:

$$X_i = \frac{n_i}{n_1 + n_2 + ... + n_i} = \frac{n_i}{\sum n_i}$$

➢ In a given solution sum of all the mole fractions is unity.

$$x_1 + x_2 + ...+ x_i = 1$$

Example: Assume that we have 150g of solution. Solution contains 40 g ethylene glycol and 110g of water.

Molar mass of ethylene glycol $C_2H_6O_2$ = 62g mol^{-1}

Moles of $C_2H_6O_2$ = $\dfrac{40g}{62g\ mol^{-1}}$ = 0.645 mol

Moles of water = $\dfrac{110}{18\ g\ mol^{-1}}$ = 6.11mol

x_{glycol} = $\dfrac{\text{Moles of } C_2H_6O_2}{\text{Moles of } C_2H_6O_2 + \text{moles of } H_2O}$

$= \dfrac{0.645}{0.645 + 6.11\ mol}$ = 0.096

x_{water} = $\dfrac{6.11\ mol}{}$ = 0.904

$$0.645 + 6.11 \text{ mol}$$

Or

$$x_{water} = 1 - 0.096 = 0.904$$

6. Molarity

➤ It is denoted by 'M'.
➤ It is defined as the number of moles of solute dissolved in one litre of the solution.

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

$$M = \frac{W_B}{M_B \times V \text{ (ml)}} \times 1000$$

Here W_B is the weight of solute
        M_B is the molar mass of solute
        V is the volume of solution

➤ Example – 2 g of NaOH are dissolved in water and the solution is made to 500cm^3 in a volumetric flask. Find the molarity of the solution?
Solution-

$$M = \frac{2}{40 \times 500} \times 1000 = 0.1 \text{ M}$$

7. Molality

➤ It is denoted by 'm'.
➤ It is the number of moles of solute present in 1 kilogram of solvent.

$$m = \frac{\text{no. of moles of solute}}{\text{Mass of solvent in Kg}}$$

$$m = \frac{n_B}{W_A \text{ (gm)}} \times 1000$$

$$m = \frac{W_B}{M_B \times W_A \text{ (gm)}} \times 1000$$

Here W_A = weight of solution in gm

W_B = weight of solute in g

M_B = molar mass of solute

➢ Example – Calculate the molality of solution which contains 18 gm of glucose in 250g of water?

Molar mass of $C_6H_{12}O_6$= 180

Solution-

W_A = 250g

W_B = 18g

M_B = 180

$$m = \frac{18}{180 \times 250} \times 1000 = 0.4 \text{ m}$$

Solubility

Maximum amount of a substance that can be dissolved in a specific amount of solvent at a specific temperature is known as solubility of that substance.

Solubility of a gas in a liquid-

➢ Many gases dissolve completely in water. Oxygen dissolves to a small extent in water. On the other hand hydrogen chloride gas is highly soluble in water.

➢ Solubility of gases in liquid is affected by pressure and temperature.
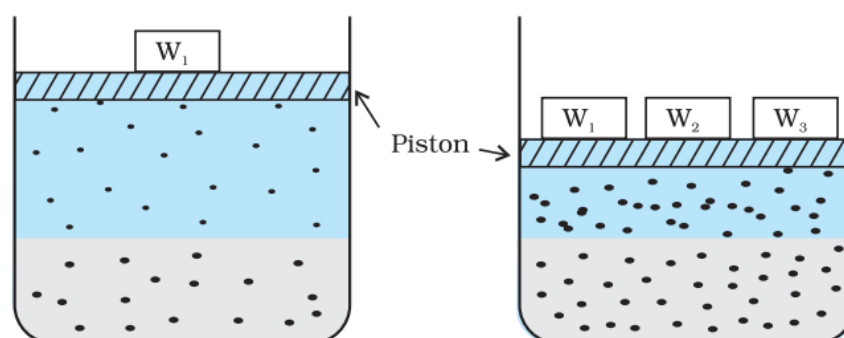
Solubility $\propto$ pressure

Figure: Effect of pressure on the solubility of a gas.

Henry's law:

➢ The solubility of a gas in a liquid is directly proportional to the partial pressure of the gas present above the surface of liquid or solution.

➢ In terms of mole fraction- The mole fraction of a gas in the solution is proportional to the partial pressure of the gas over the solution.

➢ The partial pressure of the gas in vapour phase (p) is proportional to the mole fraction of the gas (x) in the solution.

$$p = K_H \, x$$

Here K_H is the Henry's law constant; it is function of the nature of gases.

➢ The solubility of gases increases with decrease in temperature.

Applications of Henry's Law:

➢ To increase the solubility of carbon dioxide in soft drinks and soda water, the bottle is sealed under high pressure.

➢ The tanks used by scuba divers are filled with air diluted with (11.7% Helium, 56.2% Nitrogen and 32.1% Oxygen) to avoid the bend formation in blood vessels.

➢ Climbing at high altitudes causes low level of oxygen in blood.

Vapour pressure of liquid solutions-

• Liquid solutions are formed when solvent is a liquid. The solute can be a gas, a liquid or a solid. Liquid solutions are generally volatile.

• We will discuss here binary solution, mainly solution of a) liquid in liquid b) solids in liquid.

• A binary solution of two volatile liquids. The component denoted as 1 and 2. When taken in closed vessel, components evaporate and equilibrium is attained between vapour phase and liquid phase. Let the total vapour pressure is p_{total} and p_1, p_2 are the partial pressures of two components 1 and 2 respectively. These partial vapour pressures are related to the mole fractions x_1 and x_2.

• In 1886 Francots Marte Raoult formulated the quantitative relationship between partial pressure and mole fraction.

- **Raoult's law-** For a solution of volatile liquids, the partial vapour pressure of each component of the solution is directly proportional to its mole fraction present in solution.

  For component 1

  $$p_1 \propto x_1$$

  $$p_1 = p°_1 x_1$$

Here $p°_1$ is the vapour pressure of pure component 1 at the same temperature.

For component 2

$$P_2 \propto x_2$$

$$p_2 = p°_2 x_2$$

Here $p°_2$ is the vapour pressure of pure component 2 at the same temperature.

- According to Dalton's law of partial pressure, the total pressure over the solution phase in the container is the sum of the partial pressures of the components of the solution.

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

  By substituting the values of p_1 and p_2 –

  $$p_{total} = x_1 p°_1 + x_2 p°_2$$

  $$= (1-x_2) p°_1 + x_2 p°_2$$

  $$= p°_1 + (p°_2 - p°_1) x_2$$

- Following are the conclusions of this equation:
  1. Total vapour pressure over the solution can be related to the mole fraction of any one component.
  2. Total vapour pressure over the solution varies linearly with the mole fraction of component 2.
  3. The total vapour pressure over the solution decreases or increases with the increase of the mole fraction of component 1.
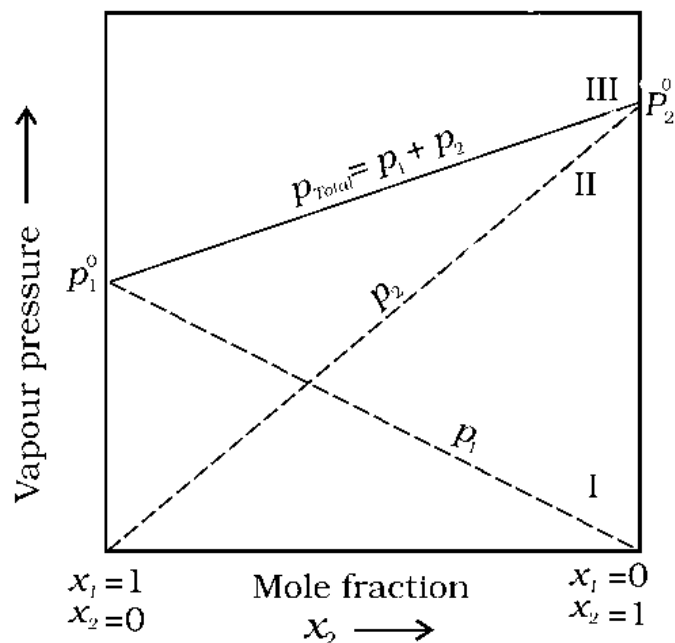
Figure: the plot of vapour pressure and mole fraction of an ideal solution at constant temperature

Raoult's Law as a special case of Henry's Law-

According to Raoult's law

$$p_1 = x_1 \, p°_1$$

According to Henry's Law

$$p = K_H \, x$$

if we compare Raoult's and Henry's Law, the partial pressure of volatile component or gas is directly proportional to its mole fraction. Only the K_H and $P°_1$ are different. Thus Raoult's law becomes a special case of Henry's law.

Vapour pressure of solutions of solids in liquids-

- Another class of solutions in which solid is dissolved in liquid. Example- sodium chloride, glucose, sugar in water and iodine and sulphur dissolved in carbon disulphide.
- At a given temperature liquids vaporise and at equilibrium the pressure exerted by the vapours over the liquid phase is called

vapour pressure. In a pure liquid the entire surface is occupied by the molecules of the liquid.

- If a non-volatile solute is added to a solvent to give a solution, the vapour pressure of the solution is solely from solvent alone. This vapour pressure is lower than the vapour pressure of pure solution at the same temperature.
- As the surface is partially covered by the solvent molecules consequently less number of solvent molecules escape from surface so the vapour pressure gets reduced.
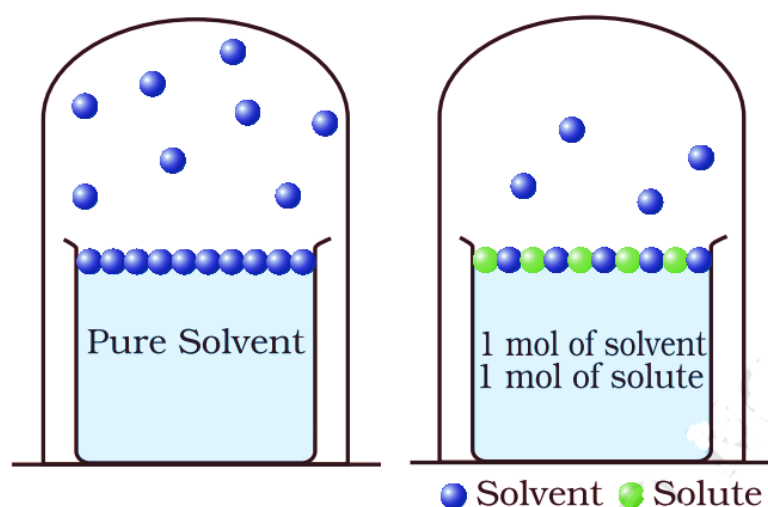

Figure: Decrease in the vapour pressure of the solvent on account of the presence of solute in solvent.

Ideal solution

- Solution which obeys Raoult's law over the entire range of concentration is known as ideal solutions.
- For ideal solutions:

$$\Delta_{mix}H = 0$$
$$\Delta_{mix}V = 0$$

The enthalpy of mixing of pure components to form the solution is zero and the volume of mixing is also zero. Means during the mixing of components no heat is absorbed or evolved.

- Examples – n-hexane and n-heptane
  Bromoethane and chloroethane

Non ideal solution

- Solution which does not obey Raoult's law over the entire range of concentration is known as non-ideal solutions.
- When the vapour pressure is high then the solution shows positive deviation. If it is low then it shows negative deviation from the Raoult's law.
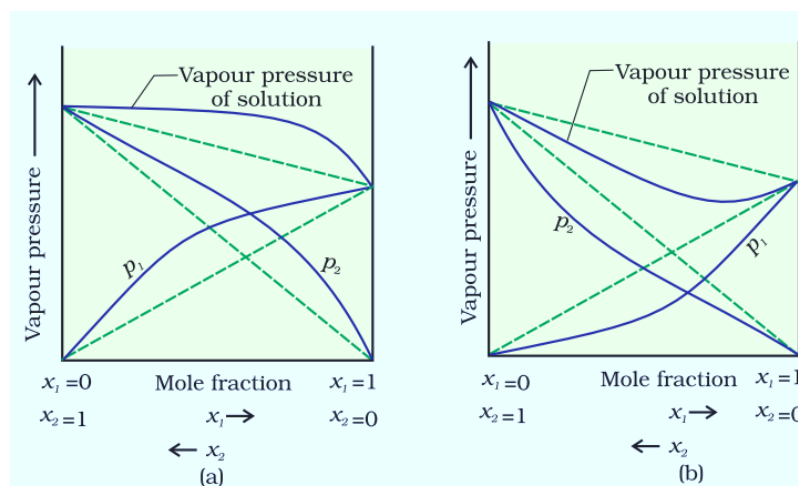

Figure: the vapour pressures of two component systems as a function of composition (a) solution shows positive deviation (b) solution show negative deviation from Raoult's law.

- Example of positive deviation – Carbon di sulphide in acetone
- Example of negative deviation – chloroform and acetone

Azeotropes – A binary mixtures which have same composition in liquid and vapour phase and boil at a constant temperature are known as azeotropes.

- ➤ There are two types of azeotropes:
    1. Minimum boiling azeotropes- They show positive deviation from Raoult's law. Example: ethanol-water
    2. Maximum boiling azeotropes- They show large negative deviation from Raoult's law. Example: nitric acid and water

Colligative properties:

Properties which depend on the concentration of molecules or ions of the solute but not on the identity of the solute are known as Colligative properties.

Example: Lowering of vapour pressure, elevation of boiling point, depression of the freezing point, osmotic pressure.

1. **Relative lowering of vapour pressure-**
   ➢ There is a relation between vapour pressure of solution, mole fraction and partial vapour pressure-

$$p_1 = x_1 \, p^{\circ}_1$$

The reduction in vapour pressure of solvent—

$$\Delta p_1 = p^{\circ}_1 - p_1$$

$$= p^{\circ}_1 - x_1 \, p^{\circ}_1$$

$$= p^{\circ}_1 \, (1 - x_1)$$

As we know $x_1 + x_2 = 1$

$$\Delta p_1 = x_2 \, p^{\circ}_1$$

$$\frac{\Delta p_1}{p^{\circ}_1} = \frac{p^{\circ}_1 - p_1}{p^{\circ}_1} = x_2$$

The relative lowering of vapour pressure and is equal to the mole fraction of the solute. The above equation can be written as:

$$\frac{p^{\circ}_1 - p_1}{p^{\circ}_1} = \frac{n_2}{n_1 + n_2} \qquad (x_2 = n_2/n_1 + n_2)$$

Here n_1 and n_2 are the number of moles of solvent and solute respectively.

In case of dilute solutions $n_2 <<< n_1$

$$\frac{p^{\circ}_1 - p_1}{p^{\circ}_1} = \frac{n_2}{n_1}$$

or

$$\frac{p^{\circ}_1 - p_1}{p^{\circ}_1} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

Here w_1 = mass of the solvent
w_2 = mass of the solvent
M_1 = molar mass of the solvent
M_2 = molar mass of the solute

Example: The vapour pressure of pure benzene at a certain temperature is 0.850 bar. 5g of a non-volatile, non-electrolyte solid added to 39g of benzene. Vapour pressure of the solution is 0.845 bar. What is the molar mass of the solid substance?
Solution:
$p°_1$ = 0.845 bar
p = 0.845 bar
M_1 = 78g mol^{-1}
w_2 = 0.5g
w_1 = 39g

$$\frac{p°_1 - p_1}{p°_1} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

$$\frac{0.845 - 0.845}{0.845} = \frac{0.5 \times 78}{M_2 \times 39}$$

$$M_2 = 170 \text{ g mol-1}$$

2. **Elevation of boiling point**
   ➢ A temperature at which a liquid boil and its vapour pressure is equal to the atmospheric pressure is known as boiling point. Water boils at 100°C, at this temperature the vapour pressure of water is 1.013 bar.
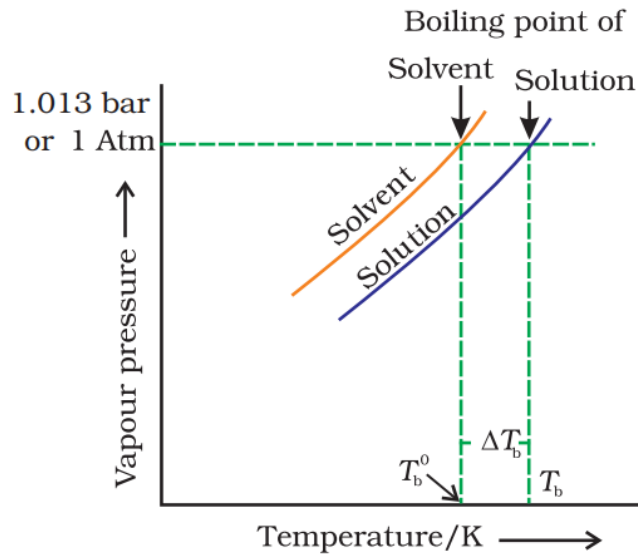
Figure: the vapour pressure curve for solution lies below the curve of pure water. Here ΔT shows the elevation of boiling point of a solvent in solution.

➢ The boiling point of a solution is always higher than that of the boiling point of the pure solvent in which the solution is prepared.

➢ Let T°_b be the boiling point of pure solvent and T_b be the boiling point of solution.

➢ The increase in the boiling point is known as elevation of boiling point.

The increase in the boiling point = $\Delta T_b = T_b - T^{\circ}_b$

➢ In case of dilute solutions, the elevation of boiling point (ΔT_b) is directly proportional to the molal concentration of the solute in a solution.

$$\Delta T_b \propto m$$
$$\Delta Tb = K_b \, m$$

Here m = molality (The number of moles of solute dissolved in 1 Kg of solvent)

K_b = Boiling point elevation constant or Molal elevation constant

Unit of K_b = K kg mol^{-1}

➢ If w_2 g of solute of molar mass M_2 is dissolved in w1 g of solvent, then molality m of the solution is given by the expression:

$$m = \frac{w_2/M_2}{w_1/1000} = \frac{1000 \times w_2}{M_2 \times w_1}$$

Substituting the value of molality in upper equation

$$\Delta Tb = \frac{K_b \times 1000 \times w_2}{M_2 \times w_1}$$

$$M_2 = \frac{1000 \times w_2 \times K_b}{\Delta Tb \times w_1}$$

Example: The boiling point of benzene is 353.23K. When 1.80g of a non-volatile solute was dissolved in 90g of benzene, the boiling point is raised to 354.11K. Calculate the molar mass of the solute. K_b for benzene is 2.53K kg mol^{-1}

Solution: ΔTb = 354.11-353.23 = 0.88K

$$M_2 = \frac{1000 \times w_2 \times K_b}{\Delta Tb \times w_1}$$

$$M_2 = \frac{1000 \times 1.8 \times 2.53}{0.88 \times 90} = 58g\ mol^{-1}$$

3. Depression of Freezing point

➢ The temperature at which the vapour pressure of a substance in its liquid phase is equal to its vapour pressure in the solid phase is known as freezing point of that substance.

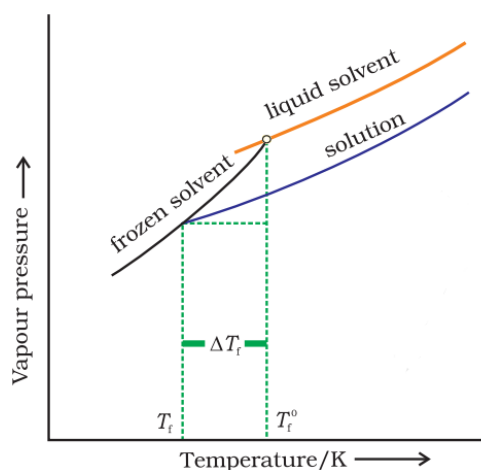➢ The lowering of vapour pressure of a solution causes a decrease in freezing point.

Figure: Depression of the freezing point of a solvent in a solution

➢ Let T°_f be the freezing point of pure solvent and T_f is freezing point when non-volatile solute is added to pure solvent.

The decrease in freezing point-

$$\Delta T_f = T°_f - T_f$$

➤ For dilute solution- $\Delta T_f \propto m$

$$\Delta T_f = K_f \, m$$

Here K_f = Freezing point depression constant or Molal depression constant or cryoscopic constant.

➤ It w_2 g of the solute having molar mass M_2 present-

$$m = \frac{w_2/M_2}{w_1/\,1000}$$

$$\Delta T_f = \frac{K_f \times w_2/M_2}{w_1/\,1000}$$

$$\Delta T_f = \frac{K_f \times 1000 \times w_2}{M_2 \times w_1}$$

$$M_2 = \frac{K_f \times 1000 \times w_2}{\Delta T_f \times w_1}$$

The value of K_f and K_b

$$K_f = \frac{R \times M_1 \times T^2_f}{1000 \times \Delta_{fus} H}$$

$$K_b = \frac{R \times M_1 \times T^2_b}{1000 \times \Delta_{vap} H}$$

Here R = gas constant

M_1 = molar mass of the solvent

T_f = Freezing point of pure solvent in kelvin

Tb = Boiling point of pure solvent in kelvin

$\Delta_{fus} H$ = Enthalpy for fusion

$\Delta_{vap} H$ = Enthalpy for vaporisation

4. Osmosis and osmotic pressure

➤ The process in which solvent molecules flow through semipermeable membrane towards solute, is known as osmosis.

➤ This process will continue till the equilibrium is attained.

➤ The flow of solvent towards solute through semipermeable membrane could be stopped if extra pressure is applied on solution. This extra

pressure that stops the flow of solvent is called osmotic pressure of the solution.
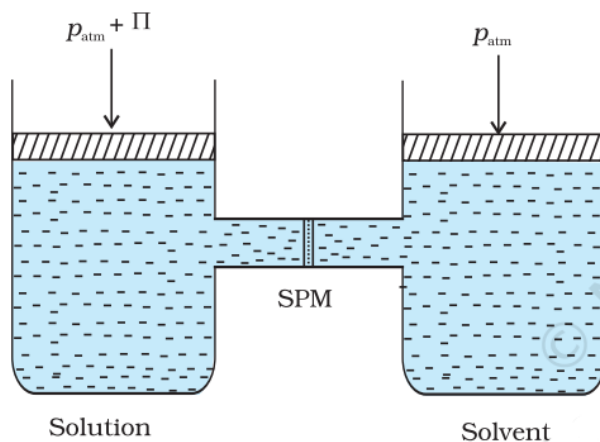

Figure: The excess pressure equal to the osmotic pressure must be applied on the solution side to prevent osmosis.

➢ Osmotic pressure is a colligative property as it depends on the number of solute molecules and not on their identity.
➢ Osmotic pressure is proportional to the molarity, C of the solution at a given temperature T.

$$\pi = C\,R\,T$$

Here π is the osmotic pressure
R = gas constant

$$\pi\,(n_2/\,V)\,R\,T$$

Here V = volume of solution in litres, containing n_2 moles of solute. If w_2 g of solute of molar mass M_2 is present in the solution then $n_2 = w_2/M_2$

$$\pi\;V = \frac{w_2\,R\,T}{M_2}$$
$$M_2 = \frac{w_2\,R\,T}{\pi\;V}$$

➢ Two solutions having same osmotic pressure at a given temperature are called isotonic solutions.
➢ If we put a cell which has semi permeable membrane in a solution of 0.9% (mass/volume) NaCl solution, water will flow out of the cell and they will shrink. Such solution is known as hypertonic.
➢ If the NaCl concentration is less than 0.9%, the solution is called hypotonic. The water will flow inside the cell and the cell will swell.

Reverse osmosis and water purification-

- The phenomenon, in which the pure solvent flows out of the solution through semi permeable membrane, is called reverse osmosis.
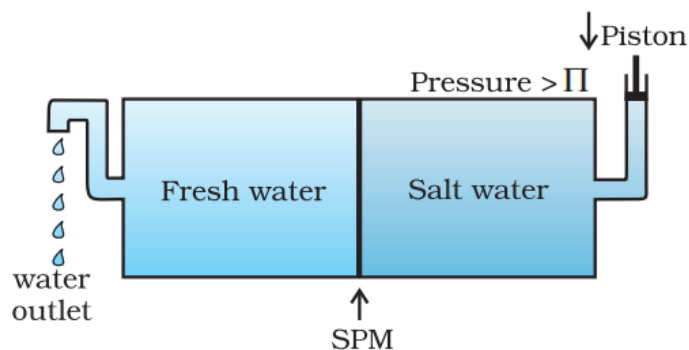

Figure: Reverse osmosis occurs when a pressure larger than the osmotic pressure is applied.

- This phenomenon is used in desalination of sea water. When a pressure more than osmotic pressure is applied, pure water is squeezed out of the sea water through membrane.
- Many countries are using desalination plants to meet their potable water requirements.

Abnormal molar mass

- Ionic compounds dissociate into cation and anion when we dissolve them in water.
- When there is dissociation of solute into ions the experimentally determined molar mass is always lower than the true value.
- Ethanoic acid dimerise in benzene due to H-bonding. Usually, dimerization takes place in solvent of low dielectric constant.

$$2 \ CH_3COOH \ \rightleftharpoons \ (CH_3COOH)_2$$
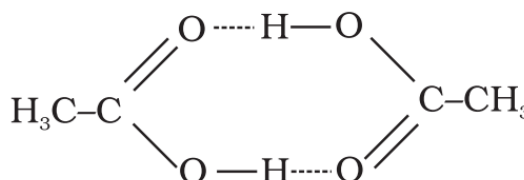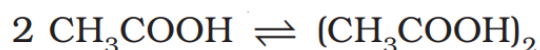
- Due to association of Ethanoic acid molecules in benzene the value of ΔT_b or ΔT_f becomes half of the normal value. Thus, molar mass will also decrease or increase from the normal value is called abnormal molar mass.

- In 1880 Van't Hoff introduced a factor i, known as the van't Hoff factor, indicates extent of dissociation or association. This factor i is defined as:

$$i = \frac{\text{Normal molar mass}}{\text{Abnormal molar mass}}$$

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

$$i = \frac{\text{Total number of moles of particles after association/dissociation}}{\text{Number of moles of particles before association/dissociation}}$$

In case of association, value of i is less than unity while for dissociation it is greater than unit. Example – The value of i for aqueous solution of KCl is 2. The value for Ethanoic acid is benzene is 0.5.

➢ Relative lowering of vapour pressure of solvent

$$\frac{p°_1 - p_1}{p°_1} = i. \frac{n_2}{n_1}$$

➢ 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 $\pi = i\, n_2\, R\, T/\, V$