6

SOLUTIONS

6.1 INTRODUCTION

We have seen that salt and sugar disolve in water. The resulting mixture is called a solution. Solutions play an important role in our daily life. In industry, solutions of various substances are used in carrying out a large number of chemical reactions.

Study of solutions of various substances is very interesting. There are certain properties of liquid solutions which depend only on the number of solute (you will learn about solute in this lesson) particles. These properties are independent of the nature of the solute.

In this lesson, let us learn about the various components of a solution and the ways in which concentration of solutions is expressed. We shall also learn about some properties of solutions which are only dependent on the number of solute particle

6.2 OBJECTIVES

After reading this lesson, you will be able to :

- identify the components of different types of solution
- express the concentration of solutions in different ways
- list different types of solutions
- state Henry's law
- define vapour pressure
- state and explain Raoult's law for solutions

- define ideal solutions
- give reasons for non-ideal behaviour of solutions
- state reasons for positive and negative deviations from ideal behaviour.
- explain the significance of colligative properties
- state reasons for the elevation of boiling point and depression in freezing point of solutions.
- solve numerical problems.

6.3 COMPONENTS OF A SOLUTION

When we put sugar into water, it dissolves to form a solution. We do not see any more sugar in it. Like sugar, a large number of other substances such as common salt, urea, potassium chloride etc dissolve in water forming solutions. In all such solutions, water is the *solvent* and the substances which dissolve are the *solutes*.

Thus, solute and solvent are the components of a solution. Whenever a solute homogeneously mixes in a solvent, a solution is formed.

Solute + Solvent ⟶ Solution

A solution is a homogenous mixture of two or more substances.

Solvent is that component of a solution that has the same physical state as the solution itself.

Solute is the substance that is dissolved in a solvent to form a solution.

6.4 THE CONCENTRATION OF A SOLUTION

Some of the properties of solutions, eg. the sweetness of a sugar solution or the colour of a dye solution, depend on the amount of solute compared to solvent. This is called the solution concentration. There are several methods for describing concentration of solutions. They include molarity, molality, normality, mole fraction and mass percentage.

Molarity : Molarity is expressed as the number of moles of solute per litre of **solution and** is usuelly denoted by M. It is expressed as :

$$M = \frac{n}{V}$$

Here n is the number of moles of solute and V is the volume of the solution in litres. A 2.0 molar solution of sulphuric aced would be labelled 2.0 M H_2SO_4. It is prepared by adding 2.0 mol of H_2SO_4 to enough water to make a litre of solution. Molarity of a solution changes with temperature because of expansion or contraction of the liquid.

Molality : It is expressed as the number of moles of solute per kilogram of solvent. It is designated by m. The label 2.0 m, H_2SO_4 is read "2.0 molal" and is prepared by adding 2.0 mol of H_2SO_4 to of solvent. Molality is expressed as.

$$m = \frac{1000\, n_B}{W_A}$$

Where n_B is the number of moles of the solute and W_A is the number of grams of solvent. The molality of a solution does not change with te ure.

Example 6.1 Find out the molarity of the solution whi ns 32.0 gm of methyl alcohol (CH_3OH) in 200 ml.

Solution : Molecular mass of CH_3OH = $12 + 1 \times 3 + 16 + 1 = 32$

Number of moles of $CH_3OH = \frac{32}{32} = 1$

Volume of the solution = 200 ml = 0.2 litre

$$\therefore \text{ Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}} = \frac{1}{0.2} = 5\text{ M}$$

Example 6.2 : What is the molality of a sulphuric acid solution of density 1.20 g/cm^3 containing 50% sulphuric acid by weight.

Solution : Weight of 1cm^3 of H_2SO_4 solution = 1.20 g

Mass of 1 litre (1000 cm^3) of H_2SO_4 solution = 1.2×1000

= 1200 g

Mass of H_2SO_4 in 100 g solution of H_2SO_4 = 50 g

Mass of H_2SO_4 in 1200 g solution of $H_2SO_4 = \frac{50}{100} \times 1200$

= 600 g

$\therefore$ Mass of water in the solution = 1200 – 600 = 600 g

Molecular mass of H_2SO_4 = 98

$$\text{No. of moles of } H_2SO_4 = \frac{\text{Mass in grams}}{\text{Molecular mass}} = \frac{600}{98}$$

$$\therefore \text{ Molarity} = \frac{\text{No. of moles of } H_2SO_4}{\text{Mass of water in grams}} \times 1000$$

$$= \frac{600}{98} \times \frac{1}{600} \times 1000 = 6.8\text{ m}$$

Normality : Normality is a concentration unit. It is defined as the number of gram equivalent weights of solute dissolved per litre of the solution.

The number of parts by weight of a substance (element or compound) that will combine with or displace, directly or indirectly 1.008 parts by weight of hydrogen, 8 parts by weight of oxygen and 35.5 parts by weight of chlorine is known as equivalent weight. Like atomic weight and molecular weight, equivalent weight is also a number and hence no units are used to express it. However, when equivalent weight is expressed in grams, it is known as gram equivalent weight or gram equivalent of the substance.

$$\textbf{Equivalent weight} = \frac{\textbf{Atomic or molecular weight}}{\textbf{valency}}$$

Normality is denoted by the symbol N.

$$\therefore \text{Normality (N)} = \frac{\text{No. of gram equivalent weights of the solute}}{\text{volume of the solution in litres}}$$

$$= \frac{\text{Amount of the solute in grams}}{\text{Equivalent weight of the solute}} \times \frac{1}{\text{volume of the solution in litres}}$$

$$= \frac{\text{Strength of the solution in grams/litre}}{\text{Equivalent weight of the solute}}$$

The label 0.5N $KMnO_4$ is read "0.5 normal" and represents a solution which contains 0.5 gm equuivalent of $KMnO_4$ per litre of solution.

Mole Fraction : The mole fraction is the ratio of the number of moles of one component to the total number of moles in the solution. If a solution contains 2 mol of alcohol and 3 mol of water, the mole fraction of alcohol is $\frac{2}{5}$, and that of water $\frac{3}{5}$. The sum of mole fractions of all the components of a solution is equal to one. The mole fraction (x_a) of a component A in solution with B is :

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

Where n_A and n_B are the number of the moles of A and B respectively.

Mass Percentage : Mass per cent is the amount of solute present in 100g of solution. Thus, 5% solution of $KMnO_4$ in water means that 5 g of $KMnO_4$ is present in 100 g of the aqueous solution of $KMnO_4$.

Example 6.3 : A solution contains 36.0 g water and 46.0 g ethyl alcohol (C_2H_5OH). Determine the mole fraction of each component in the solution.

Solution : Molecular weight of water = 18

Molecular weight of C_2H_5OH = 46

No. of moles of water = $\frac{36}{18} = 2.0$

No. of moles of C_2H_5OH = $\frac{46}{46} = 1.0$

Totqal number of moles in the solution = 2.0 + 1.0 = 3.0

$$\text{Mole fraction of water} = \frac{\text{No. of moles of water}}{\text{Total no. of moles in the solution}}$$

$$= \frac{2.0}{3.0} = 0.67$$

$$\text{Mole fraction of } C_2H_5OH = \frac{\text{No. of moles of } C_2H_5OH}{\text{Total no. of moles in the solution}}$$

$$= \frac{1.0}{3.0} = 0.33$$

Example 6.4 : Calculate the normality of a solution of NaOH if 0.4 gm of NaOH is dissolved in 100 ml of the solution.

Solution : Amount of NaOH present in 100 ml of the solution = 0.4 g

$\therefore$ Mass of NaOH present in 1000 ml of the solution $\frac{0.4}{100} \times 1000 = 4.0$ g

Mol. wt. of NaOH = 23 + 16 + 1 = 40 amu

$$\therefore \text{ Eq. Wt. of NaOH} = \frac{\text{Mol. Wt.}}{\text{Basicity}} = \frac{4.0}{1} = 40$$

$$\therefore \text{ Normality} = \frac{\text{Strength in g/litre}}{\text{Eq. wt.}} = \frac{4}{40} = \frac{1}{10} \text{ N}$$

Hence, the normality of the solution $= \frac{1}{10}$ N or 0.1 N

INTEXT QUESTIONS 6.1

1. What are the various methods of expressing the concentration of a solution ?

....

2. Define the following

   (i) Molarity (ii) Molality (iii) Normality

   ....

   ....

6.5 TYPES OF SOLUTIONS

Solutions can be solid, liquid or gaseous. Depending upon the physical state of the solute and the solvent, there are nine possible types of solutions consisting of two components (binary solution). Different types of solutions are given in Table 6.1.

Table 6.1 Different Types of Solutions

SOLUTE	SOLVENT	EXAMPLE
Gas	Gas	Air
Gas	liquid	Soda water
Gas	solid	Hydrogen in palladium
Liquid	Gas	Humidity in air
Liquid	Liquid	Alcohol in water
Liquid	Solid	Hydrated salt, e.g. $CuSO_4 . 5H_2O$
Solid	Gas	Camphor in air
Solid	Liquid	Sugar in water
Solid	Solid	Alloys such as brass (zinc in copper) and bronze (tin in copper)

Generally, we come across only the following three types of solutions:

(a) Liquids in Liquids : In the solution of liquids in liquids such as alcohol in water, the constituent present in smaller amounts is usually designated as solute. The constituent present in larger amounts is called the solvent, when two liquids are mixed, three different situations may arise :

(i) Both the Liquids are completely miscible, i.e, when two liquids are mixed. they dissolve in each other in all proportions, e.g., alcohol and water, benzene and toluene.

(ii) The liquids are partially miscible, i.e, they dissolve in each other only to a certain extent, e.g., water and ether, water and phenol.

(iii) The liquids are immiscible, i.e, they do not dissolve in each other, e.g. water and benzene, water and toluene, water and kerosene.

The solubility of liquids in liquids increases with rise in temperature.

(b) Gases in Liquids : Gases are generally soluble in liquids. Oxygen is sufficiently soluble in water, which allows the survival of aquatic life in ponds, rivers and oceans. Gases like CO_2 and NH_3 are highly soluble in water. The solubility of a gas in a liquid depends on the pressure, temperature and the nature of the gas and the solvent. These factors are discussed in detail below :

(i) Effect of Pressure : The variation of solubility of a gas in a liquid with pressure is governed by Henry's law. Henry's law states that

The mass or mole fraction, of a gas dissolved in a solvent is directly proportional to the partial pressure of the gas.

Henry's law is represented by

$$K = \frac{p}{x}$$

where K is a constant, p is the partial pressure of the gas and x is the mole fraction of the gas in the solution. Let us now see what are the conditions for the validity of Henry's law.

(i) **Conditions for validity of Henry's law :** It is found that gases obey Henrys law under the following conditions.

- (i) the pressure is not too high.
- (ii) the temperature is not too low.
- (iii) the gas does not dissociate, associate or enter into any chemical reaction with the solvent.

(ii) **Effect of temperature :** The solubility of a gas in a liquid at constant pressure decreases with rise in temperature. For example, the solubility of CO_2 in water at 20°C is 0.88 cm³ per cm³ of water. Where as it is 0.53 cm³ per cm³ of water at 40°C. This happens because on heating a solution containing a dissolved gas, some gas is usually expelled from the solution.

(iii) **Effect of nature of the gas and the solvent :** Gases like CO_2, HCl and NH_3 are highly solube in water where as H_2, O_2 and N_2 are sparingly soluble.

(c) **Solids in liquids :** When a solid is dissolved in a liquid, the solid is referred as the solute and the liquid as the solvent.For example, in a solution of sodium chloride in water, the solute is sodium chloride and water is the solvent. Different substances dissolve to different extent in the same solvent.

6.6 VAPOUR PRESSURE

If we keep a beaker over a small beaker containing a pure liquid, it is found that the molecules of the liquid start evaporating in the form of vapours and fill the empty space above the beaker containing the liquid. A time comes when the number of molecules evaporating per unit time is equal to the number of molecules condensing during that time (Fig 6.1). An equilibrium is thus establihed between the vapour and the liquid phases. The pressure exerted by the vapours of the liquid in such a case is called the vapour pressure of the liquid.

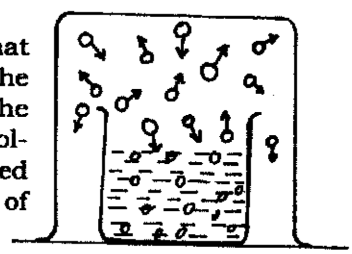

Fig 6.1 An equilibrium state

6.7 RAOULT'S LAW FOR SOLUTIONS

Did you ever think that if you mix two miscible volatile liquids A and B, what would be the vapour pressure of the resulting solution ? The vapour phase now consists of vapours of both the liquids A and B. The partial vapour pressure of each liquid will depend upon its mole fraction in the solution. Let the mole fractions of the liquids A and B be x_A and x_B respectively. Also, if p_A and p_B are the partial vapour pressures of A and B respectively, then

$$p_A \propto x_A \text{ or } p_A = p^0_A\, x_A$$

Similarly, $$p_B = p^0_B\, x_B$$

where p^0_A and p^0_B represent the vapour pressures of pure liquids A and B respectively

The relationship between vapour pressure of a liquid and its mole fraction is given by Raoult's law.

Raoult's law states that for a solution of volatile liquids, the partial vapour pressure of each liquid in the solution is directly proportional to its mole fraction.

Raoult's law is applicable only if the liquids are miscible.

If the values of p_A and p_B are plotted against the values of x_A and x_B for a solution, two straight lines are obtained as shown in Fig 6.2 the total vapour pressure p of the solution is given by the sum of partial vapour pressures p_A and p_B.

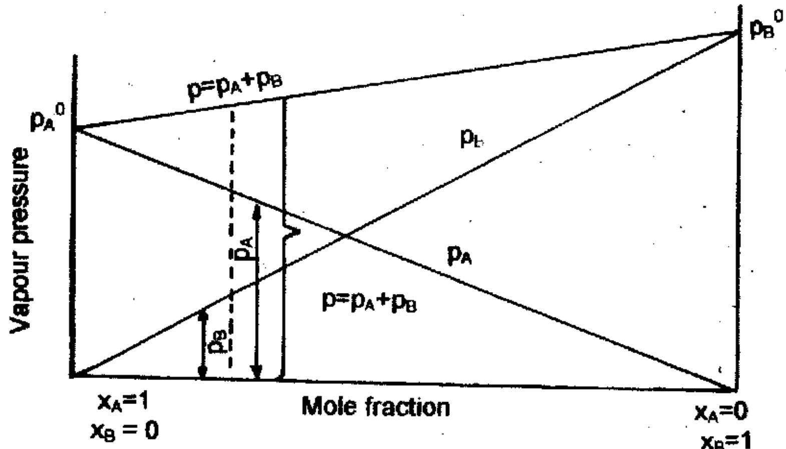

Fig. 6.2 Relationship between vapour pressure and mole fraction in a solution

Thus, $p = p_A + p_B$

or $p = p^0_A x_A + p^0_B x_B$

The total vapour pressure (p) of a solution is represented by the line joining p^0_A and p^0_B. The solutions which obey Raoult's law are known as **ideal solutions.**

A solution which obeys Raoult's law over the entire range of concentration at all temperatures is called an ideal solution.

INTEXT QUESTIONS 6.2

1. State Raoult's law.

   ....

2. State Henry's law and list the conditions necessary for the validity of Henry's law

   ....

6.8 RAOULT'S LAW FOR SOLUTIONS CONTAINING NON-VOLATILE SOLUTE

If we have an aqueous solution containing a non-volatile solute, such as sugar or salt, what do you think about the vapour pressure exerted by such a solution ? The vapour phase of such a solution consists of vapours of solvent (A) only because the solute is non-volatile. Since the mole fraction of the solvent in solution is less than one, therefore according to Raoult's law, the vapour pressure of the solution will be less than the vapour pressure of the pure solvent. If the total vapour pressure of the solution is p, then

$$p = p_A x_A \quad (1)$$

for a binary mixture

$$x_A + x_B = 1$$

therefore, $x_A = 1 - x_B$

Substituting the value of x_A in equation (1) we get

$$p_A = p^0_A(1 - x_B)$$

or $$\frac{p_A}{p^0_A} = 1 - x_B$$

therefore, $$\frac{p^0_A - p_A}{p^0_A} = x_B$$

In the above equation, $(p^0_A - p_A)$ represents the lowering of the vapour pressure and $\frac{p^0_A - p_A}{p^0_A}$ is called the relative lowering of the vapour pressure of the solution. An alternative statement of Raoult's law for solutions of non-volatile solute is :

The relative lowering of vapour pressure for a solution is equal to the mole fraction of the solute, when only the solvent is volatile.

6.9 IDEAL AND NON-IDEAL SOLUTIONS

Non-ideal solutions are those solutions which do not obey Raoult's law and whose formation is accompanied by changes of heat and volume.

Most of the real solutions are non-ideal. They show considerable deviation from the ideal behaviour. Generally deviations are of two types.

(i) Positive deviation : Positive diviations are shown by liquid pairs for which the A-B molecular interactions are lower than the A-A and /or B-B molecular interactions. The total vapour pressure for such solutions is greater than predicted by Raoults law. The total vapour pressure for such a slolution will be maximum for a particular inter mediate composition (Fig 6.3)

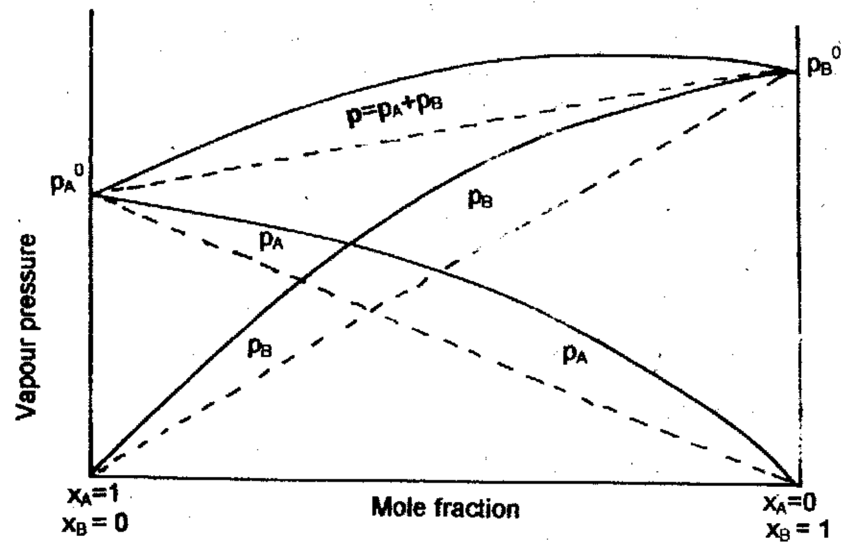

Fig. 6.3 Positive deviation for a liquid pair

Examples of non-ideal solutions showing positive deviation from the ideal behaviour are mixtures of liquids such as water-propanol, ethanol-chlorofrm, acetone- carbon disulphide, ethanol-cyclohexane etc.

(ii) Negative Deviation : Negative deviations are shown by liquid pairs for which the A-B molecular interactions are higher than A-A or B-B molecular interaction The total vapour pressure for such solutions is less than that predicted by Raoult's law. For a particular intermediate composition, the total vapour pressure of such a solution will be minimum. (Fig 6.4) Examples of such liquid pairs are Chloroform-acetone, water-sulphuric acid chloroform-benzene, water- HCl etc

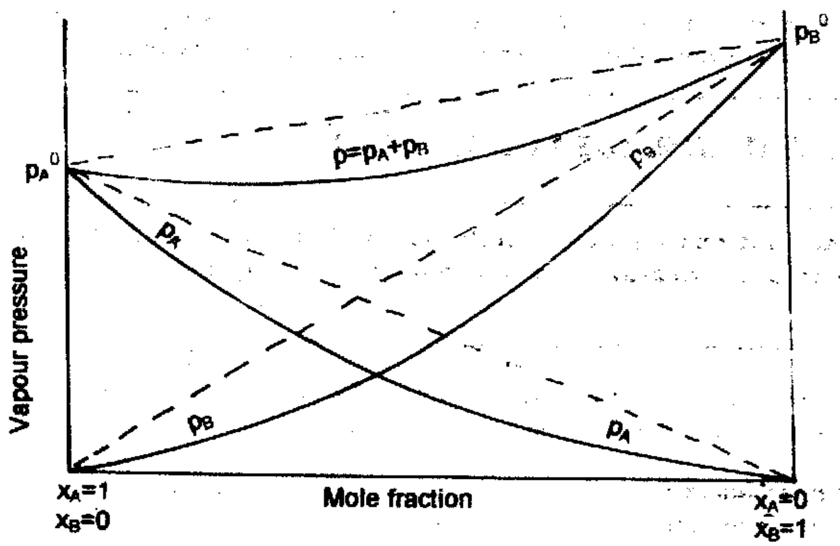

Fig. 6.4 Negative deviation for a liquid pair

6.10 COLLIGATIVE PROPERTIES

Do you know that there are certain properties of dilute solutions which depend only on the member of particles of solute and not on the nature of the solvent nd the solute ? Such properties are called **colligative properties**.

Some important colligative properties are :

relative lowering of vapour pressure, elevation in boiling point, depression in freezing point, osmotic pressure.

We shall discus the first three colligative properties i.e, relative lowering of vapour pressure, elevation in boiling point and depressio n freezing point in detail.

6.10.1 Relative Lowering of Vapour Pressure

According to Raoults law for solutions containing on-volatile solute

$$\frac{p_A^0 - p_A^0}{p_A^0} = x_B \text{ (see section 6.8)} \qquad (2)$$

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

Therefore equation (2) can be writen as

$$\frac{p_A^0 - p_A}{p_A^0} = x_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

In dilute solution $\frac{W_B}{M_B} << \frac{W_A}{M_A}$. Therefore the term $\frac{W_B}{M_B}$ can be neglected from the denominator.

Hence, $$\frac{p_A^0 - p_A}{p_A^0} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

The above expression can be used to determine the molecular mass of the solute B, provided the relative lowering of vapour pressure of a slolution of known concentration and molecular mass of the solvent are known. However, the determination of molecular mass by this method is often difficult because the accurate determination of lowering of vapour pressure is difficult.

Example 6.5 : The relative lowering of vapour pressure produced by dissolving 7.2 g of a substance in 100g water is 0.00715. What is the molecular mass of the substance ?

Solution : We know that

$$\frac{p_A^0 - p_A}{p_A^0} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

On substituting the values we get

$$0.00715 = \frac{7.2 \times 18}{M_B \times 100} \quad \text{or} \quad M_B = \frac{7.2 \times 18}{0.00715 \times 100} = 181.26$$

∴ Molecular mass of the substance = 181.26 amu

6.10.2 Elevation of Boiling Point

Boiling point of a liquid is the temperature at which the vapour pressure of the liquid becomes equal to the atmospheric pressure.

As you know, the vapour pressure of a pure solvent is always higher than that of a solution. So, the boiling point of the solution is always higher than that of the pure solvent. If you see the vapour pressure curves for the solvent and the solution (Fig 6.5), you will find that there is an elevation in the boiling point of the solution.

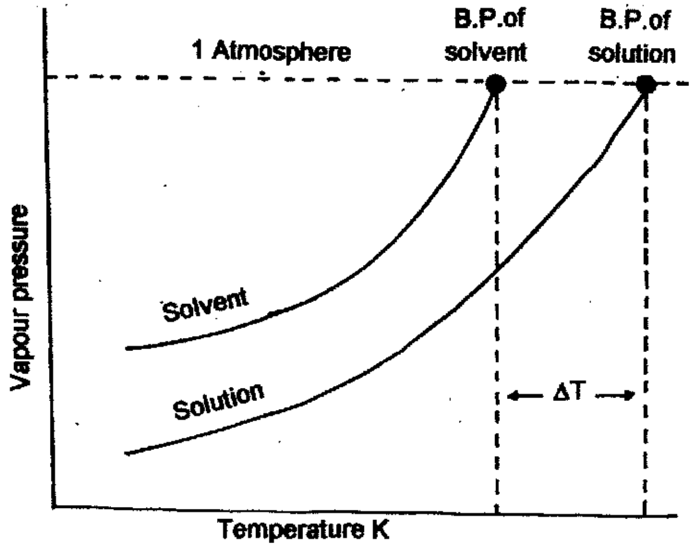

Fig. 6.5 Vapour pressure curves for solvent and solution

Now let ΔT_b be the elevation in boiling point and Δp be the lowering in vapour pressure. Then,

$$\Delta T_b \ \alpha \ \Delta p \ \alpha \ x_B \text{ or } \Delta T_b = K x_B \qquad (3)$$

(K is the proportionality constant

As you know $x_B = \dfrac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$

In a dilute solution, $\frac{W_B}{M_B} << \frac{W_A}{M_A}$ and thus the term $\frac{W_B}{M_B}$ is neglected in the denominiator.

Thus, $x_B = \dfrac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}} = \dfrac{W_B}{M_B} \times \dfrac{M_A}{W_A} = n_B \times \dfrac{M_A}{W_A}$

$$\left(\text{since } n_B = \frac{W_B}{M_B}\right)$$

On substituting the value of x_B in the equation (3) we get

$$\Delta T_b = K \times n_B \times \frac{M_A}{W_A}$$

If we take the mass of the solvent W_A in kilograms the term $\frac{n_B}{W_A}$ is molality m. Thus

$$\Delta T_b = K \cdot M_A \cdot m = K_B\ m$$

The constant K_b is called the **molal elevation constant** for the solvent. K_b may be defined as **the elevation in boiling point when one mole of a solute is dissolved in one kilogram of the solvent.** K_b is expressed in degree per molality.

6.10.3 Depression in Freezing Point

Freezing point is the temperature at which the solid and the liquid forms have the same vapour pressure.

The freezing point of the solution is always less than that of the pure solvent. Thus, there is a depression in the freezing point of the solution This is because the vapour pressure of the solution is always less than that of the pure solvent. This is evident from the Fig 6.6.

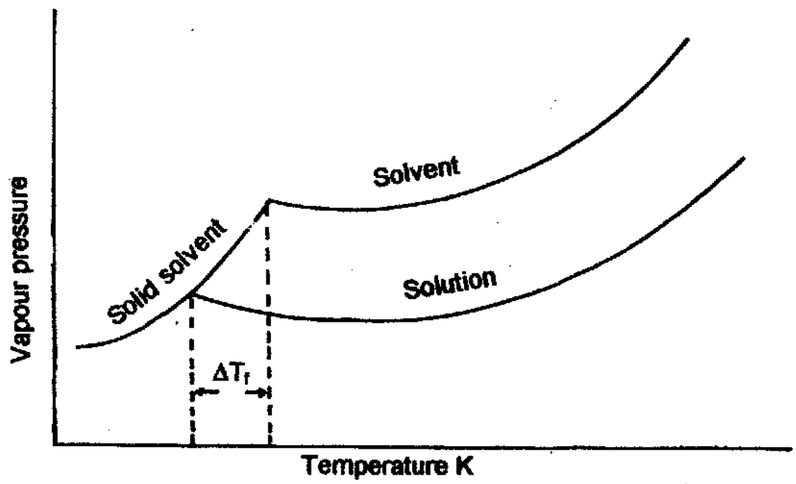

Fig 6.6 Vapour pressure curves for solid, solvent and solution

Let ΔT_f be the depression in freezing point. Then :

$$\Delta T_f \ \alpha\ x_B$$

$$\text{or}\quad \Delta T_f = K\, x_B \qquad (4)$$

Where K the proportionality constant

You know that

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

In dilute solutions $\frac{W_B}{M} << \frac{W_A}{M_A}$.

Therefore, the term $\frac{W_B}{M_B}$ can be neglected from the denominiator. Thus,

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

$$\left(\text{since } n_B = \frac{W_B}{M_B}\right)$$

On substituting the value of x_B in equation (4) usee get

$$\Delta T_f = K \times n_B \times \frac{M_A}{W} = \frac{K n_B M_A}{W}$$

If the mass of the solvent W_A is taken in kg, then the term $\frac{n_B}{W}$ becomes molality m.

Thus, $\Delta T_f = K\ M_A \cdot m = K_f \cdot m$

The constant (K_f) for a solution is known as **molal depression constant or molal cryoscopic constant** for the solvent. K_f may be **defined as the depression in freezing point of a solution when one mole of a solute is dissolved in 1 kilogram of the solvent.**

Example 6.6 : Find the (i) boiling point and (ii) freezing point of a solution containing 0.520 g glucose ($C_6H_{12}O_6$) dissolved in 80.2g of water. (K_f = 1.86 K/m)

Solution : (i) K_b = 0.52 K/m

$$\text{Molality of glucose} = \frac{\text{Wt. of glucose}}{\text{mol. Wt.}} \times \frac{1000}{\text{Wt. of solvent}}$$

$$= \frac{0.52}{180} \times \frac{1000}{80.2} = 0.036$$

$$\Delta T_b = K_f m = 0.52 \times 0.036 = 0.018$$

∴ Boiling point = 373 + 0.018 = 373.018 = 373.02 K

(i) K_f = 1.86 K/m

$$m = \frac{0.52}{180} \times \frac{1000}{80.2} = 0.036$$

∴ $\Delta T_f = 1.86 \times 0.036 = 0.66$

∴ Freezing point = 373 – 0.66 = 372.34 K

INTEXT QUESTIONS 6.3

1. Define colligative property. List two colligative properties.

....

2. What type of liquid pairs show (i) positive deviation (ii) negative deviation.

....

6.11 WHAT YOU HAVE LEARNT

- Solution is a homogeneous mixture of two or more substances.
- Solvent is that component of a solution that has the same physical state as the solution itself.
- Solute is the substance that is dissolved in a solvent to form a solution.
- Molarity is expressed as the number of moles of solute per litre of solution

- Molality is expressed as the number of moles of solute per kilogram of solvent.
- Normality is a concentration unit which tells the number of gram equivalent of solute per litre of solution.
- Mole fraction is the ratio of the number of moles of one component to the total number of moles in the solution.
- Solutions can be solid, liquid or gaseous.
- Henry's law states that mass or mole fraction of a gas dissolved in a solvent is directly proportional to the partial pressure of the gas.
- Raoult's law states that for a solution of volatile liquids, the partial pressure of each liquid in the solution is directly proportional to its mole fraction.
- A solution which obeys Raoult's law over the entire range of concentration at all temperatures is called an ideal solution.
- The relative lowering of vapour pressure for a solution is equal to the mole fraction of the solute, when only the solvent is volatile.
- Those properties of dilute solutions which depend only on the number of particles of solute and not on the nature of the solvent are known as colligative properties.
- Molal elevation constant is the elevation in boiling point when one mole of solute is dissolved in one kilogram of the solvent.
- Boiling point of a liquid is the temperature at which the vapour pressure of the liquid becomes equal to the atmospheric pressure.
- Freezing point is the temperatures at which the solid and the liquid forms have the same vapour pressure.

6.12 TERMINAL EXERCISE

1. What do you understand by ideal and non-ideal solutions ?

   ....

2. Define freezing point and boiling point.

   ....

   Derive the relationship $\Delta T_b = K_b\, m$

   ....

4. A solution containing 7 g of a non-volatile solute in 250 g of water boils at 100.26°C. Find the molecular mass of the solute.

   ....

5. 2 g of a substance dissolved in 40 g of water produced a depression of 1.5°C in the freezing point of water. Calculate the molecular mass of the substance. The molar depression constant for water is 1.85°C per mole.

   ....

6. Calculate the mole fraction of the solute in a solution obtained by dissolving 10 g of urea (mol wt 60) in 100 g of water.

   ....

CHECK YOUR ANSWERS

INTEXT QUESTIONS 6.1

1. Refer to Section 6.4
2. Refer to Section 6.4

INTEXT QUESTIONS 6.2

1. Refer to Section 6.7
2. Refer to Section 6.5

INTEXT QUESTIONS 6.3

1. Refer to Section 6.10
2. Refer to Section 6.9

TERMINAL EXERCISE

1. Refer to Section 6.9
2. Refer to Section 6.10.3 and 6.10.2
3. Refer to Section 6.10.2
4. We know that $\Delta T_b = K_b\, m$

   on substituting the values we get

   $$100.26 - 100 = \frac{1000 \times 7}{M_B \times 200} \times 0.52$$

   $$\text{Or} \quad M_B = \frac{1000 \times 7 \times 0.52}{0.26 \times 250} = 56$$

   Molecular mass of the solute = 56 amu

5. We know that $\Delta T_f = K_f\, m$

   on substituting the values, we get

   $$1.5 = \frac{100 \times 2}{M_B \times 40} \times 1.85 \quad \text{or} \quad M_B = \frac{1000 \times 2 \times 1.85}{15 \times 40} = 61.5$$

   ∴ Molecular mass of the substance = 61.5 amu

6. No of moles of urea $= \dfrac{10\ \text{g}}{60\ \text{g mol}^{-1}} = \dfrac{1}{6}$ mol = 0.166 mol

   No. of moles of water $= \dfrac{100\ \text{g}}{18\ \text{g mol}^{-1}} = \dfrac{50}{9}$ mol = 0.555 mol

   Mole fraction of urea $= \dfrac{0.166}{0.166 + 0.555} = \dfrac{0.166}{0.721} = 0.23$