Unit-10

Solutions and Colloids

After studying this unit, the students will be able to:

- List the characteristics of colloids and suspensions that distinguish them fr solutions. (Applying)
- Define hydrophilic and hydrophobic molecules. (Remembering)
- Explain the nature of solutions in liquid phase giving examples of complet miscible, partially miscible and immiscible liquid-liquid solutions. (Applying
- Explain the effect of temperature on solubility and interpret the solubil graph. (Analysing)
- Express solution concentration in terms of mass percent, molarity, molali parts per million, billion and trillion and mole fraction. (Remembering)
- Define the term colligative properties. (Remembering)
- Describe on a particle basis why a solution has a lower vapour pressure th the pure solvent. (Applying)
- Explain on a particle basis how the addition of a solute to a pure solve causes an elevation of the boiling point and depression of the freezing poi

Teaching

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Assessment

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Weightage %

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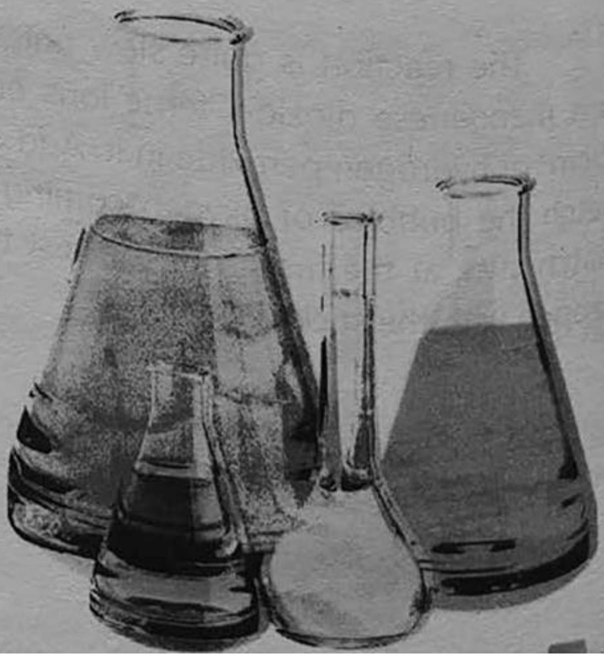

of the resultant solution. (Applying)

- Describe the role of solvation in the dissolving process. (Understanding)
- Define the term water of hydration. (Remembering)
- Explain concept of solubility and how it applies to solution saturation. (Applying)
- Distinguish between the solvation of ionic species and molecular substances. (Understanding)
- List three factors that accelerate the dissolution process. (Understanding)
- Define heat of solution and apply this concept to the hydration of ammonium nitrate crystals. (Applying)
- Explain how solute particles may alter the colligative properties. (Applying)
- Explain osmotic pressure, reverse osmosis and give their daily life applications. (Applying)
- Describe types of colloids and their properties. (Applying)
- List some colligative properties of liquids. (Understanding)

Introduction

In the grade IX, you have learnt the basic information about solution, like homogenous and heterogeneous mixture, types of solution, saturated and unsaturated solutions, concentration and its units, solubility and factors affecting solubility, dissolution process, and some introduction to colloids and suspension. On the basis of these concepts you will develop more knowledge about the solution in this unit.

10.1 General Properties of Solutions

A **solution** is a homogeneous mixture of two or more substances in which the ratio of the constituent substances remains the same throughout the solution. Generally a solution consists of two components, the solute and solvent. The substance that is present in lesser amount is the solute while the one with the greater amount is called solvent. Solute may be solid, liquid or gas or combination of these. The particles of solute may be at ionic or molecular level. The particle size ranges from 0.01 to 1 nm, which cannot be seen by naked eye. A solution is stable if it is not separated on standing. Solute and solvent cannot be separated by filtration or by centrifugation. It is composed of one phase. At a given temperature and pressure, the composition of solution can be varied to a maximum value up to the saturation value. A solution does not allow a beam of light to scatter when passed through it. The mixture of salt or sugar in water is an example of solution.

Colloids, on the other hand, are heterogeneous mixture of larger particles or aggregate of particles of size ranging from 1 to 1000nm dispersed in solvent or dispersion medium. They do not settle down on standing and cannot be separated by filtration. They scatter light when passed through them, a phenomena, called the Tyndal effect. Gum in water is an example of colloidal dispersion. Suspensions are also heterogeneous mixtures with size of the particles greater than 1000nm. Particles are settled down on standing and can be separated by filtration. They can scatter light or may be opaque, e.g., finely ground clay mixed with water is an example of suspension.

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When a bottle containing a carbonated beverage (soda water) is opened, bubbles of carbon dioxide rise to the liquid's surface. Sometimes violent release of carbon dioxide gas can be stopped if the bottle's cap is retightened. If the beverage is at room temperature, an energetic evolution of gas may occur. This observation can be explained as follows.

The solubility of gas depends on the pressure acting upon the system. If the pressure of the system is reduced, the dissolved gas rapidly leaves the solution phase as small bubbles. As these bubbles rise within the solvent, their size is increased because they are encountering less pressure. On reaching the liquid's surface, the bubbles burst.

When the container is closed, there is equilibrium between the gas above the liquid and the gas dissolved in the solvent. If the container is opened, the pressurized gas escapes. The reduced pressure on the liquid's surface allows additional gas molecules to leave the solution. Beverages are packaged in sealed bottles and cans to prevent the escape of carbon dioxide. The beverages can be kept indefinitely without losing their carbonation (CO_2) until they are opened by the consumer.

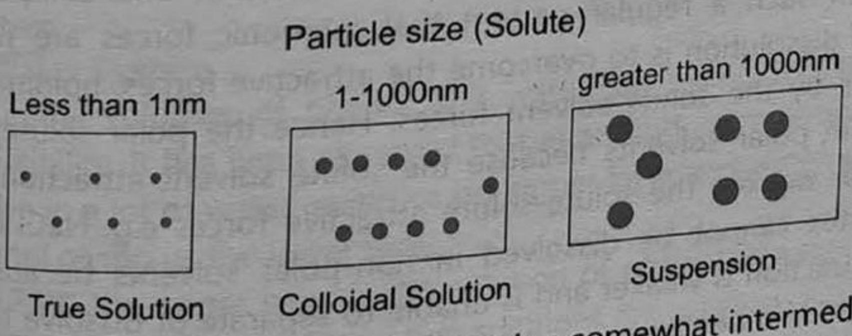
10.1.1 Solution, Suspension and Colloids

Colloids are mixtures of two substances which are intermediate between true solution and suspension. The particles in such a mixture are smaller than those in a suspension but larger than those in a true solution. A detailed discussion on colloids is given in section 10.5.

The name colloid was coined by Graham (Greek: kolla: glue, eidos, like) Colloids are substances consisting of a continuous, homogenous medium known as dispersion medium (or outer phase) and particles of a discontinuous medium termed as dispersed medium (or inner phase). The milky dispersion of sulphur, starch, gum in water and blood serum are all examples of colloids.

Tidbit

The general pattern of dissolution can be stated as "like dissolves like" that means inorganic solute such as sodium chloride, ammonia etc. are soluble in inorganic solvent e.g. water. Organic substances e.g. fats, oils etc are soluble in organic solvents such as benzene, alcohol etc.



Since colloidal dispersion of a substance is a somewhat intermediate stage between a true solution and a course suspension, it is impossible to draw a line of difference between true solution and colloidal dispersion at one end and between colloidal dispersion and suspension at the other end. There is a gradual change from one type of system to the other, however, colloidal system have certain properties which place them in a separate group. The properties and the behaviour of colloids depend chiefly on the size, shape and charge of the dispersed particles. Some distinctive properties of colloids are compiled and compared with true solutions and coarse dispersion in Table (10.3)

10.1.2 Hydrophilic and Hydrophobic Particles

The stability of colloids depends upon the charge of dispersed particles and particles of dispersion medium. When water is used as dispersion medium and there exist an attractive force between water molecules and dispersed particles, then colloidal particles are called hydrophilic particles which means "liking water". The hydrophilic particles containing colloids are much stable. When such colloids are once precipitated they can be directly return into the colloidal form and are, therefore, called reversible colloids.

Hydrophobic means "Fearing water". When water is used as dispersion medium and no attractive force exists between water molecules and dispersed particles, the particles to be dispersed, are called hydrophobic.

The colloids formed by hydrophobic particles are less stable. If, once precipitated they cannot be directly converted into the colloidal form and are thus termed as irreversible colloids.

10.1.3 The Nature of Solutions in Liquid Phase

When a solid comes in contact with a suitable liquid, it dissolves forming a solution. This process of dissolution can be explained in terms of attraction between particles of a solute and that of a solvent. In ionic compounds ions are arranged in such a regular pattern that interionic forces are maximum. The process of dissolution is to overcome the attractive forces, holding together the solute ions by the solute-solvent forces. Hence, the polar solutes can only be dissolved in polar solvents because the solute solvent attraction is enough to separate or weaken the solute-solute attractive forces e.g. NaCl in water. Polar solid solutes cannot be dissolved in non-polar solvents because the solute-solvent attraction is weaker and is unable to separate or dissolve the solid solute ions e.g. $CaCl_2$ in benzene.

Similarly, non-polar solid solutes cannot be dissolved in polar solvents because no attractive forces are established between solute and solvent molecules e.g. naphthalene in water. However, non-polar solutes can only be dissolved in non-polar solvents because the attractive forces found in molecules of solute are less than the attractive forces between solute-solvent molecules. e.g. naphthalene is soluble in benzene, both being non-polar.

Solutions of Liquids in Liquids

There are three types of liquids in liquids systems:

1. Completely miscible liquids.
2. Partially miscible liquids.
3. Completely immiscible liquids.

Completely Miscible Liquids

Miscible liquids mix in any proportion. Immiscible liquids do not mix completely; rather make two separate layers. Miscible liquids usually are similar in their chemical structures, e.g., alcohol and water when mixed they form homogeneous solutions. Their molecules are both polar and contain OH- groups that give rise to hydrogen bonding among them.

Partially Miscible Liquids

There are a number of liquids which are not completely miscible but they are partially miscible and mix with each other up to a limited extent.

Ether dissolves water upto the extent of about 1.2% and water dissolves ether upto the extent of about 6.5%. If you mix them in proportion greater than this you would find that they will separate into two layers. Each layer is a saturated solution of the other liquid. Such solutions are called conjugate solutions. The mutual solubility of these conjugate solutions is affected by temperature changes. Typical examples of such systems are phenol-water system, trimethylamine-water system, nicotine-water system.

Phenol-Water System

If equal volumes of phenol and water are mixed together, they show partial miscibility. It has been observed that at around room temperature, phenol will dissolve in a lot of water giving us the upper layer and water will dissolve in a lot of phenol giving us the lower layer.

At 25°C the upper layer is 5% solution of phenol in water and the lower layer is 30% water in phenol. These two solutions are *conjugate solutions* to each other. The lower layer has a greater density due to greater percentage of phenol. Water acts as a solute in the lower layer while phenol is a solute in the upper layer. When the temperature of water-phenol system is increased, the compositions of both layers change. Water starts travelling from upper to the lower layer and phenol travels from lower to the upper layer. When the temperature of this system approaches 65.9°C, a homogeneous mixture of two components is obtained. This homogeneous mixture contains 34% phenol and 66% water. *The temperature above which two conjugate solutions merge into* *one another, is called critical solution temperature.* Thus phenol-water system has critical solution temperature of 65.9°C. Other partially miscible pairs of liquids have their own critical solution temperature with definite compositions of solutions.

Completely Immiscible Liquids

Such liquids are incapable of being mixed or blended together. The reason is as mentioned earlier, that polar dissolve polar and non-polar dissolve non-polar (like dissolves like). Such liquids when shacked together will form two distinct layers. Examples of immiscible liquid pairs are as follow:

a. Water and benzene.
b. Carbon disulphide and water

10.1.4 The effect of Temperature and Pressure on Solubility

It is our common observation that some substances dissolve readily, while others do not. e.g glucose dissolve easily in water to form solution as compared to cane sugar. *The weight in grams of a solute necessary to satu [illegible] 100 grams of a solvent at constant temperature and pressure is called solubility.* Solubility of a solute depends on the following three factors.

1. Nature of solute and solvent.
2. Pressure.
3. Temperature.

1. Nature of solute and solvent

An ionic or polar solute dissolves in polar solvent while a non-polar solute dissolves in non-polar solvent. In all the cases the limit of solubility is the saturated solution. The concentrations of various solutes in a solvent necessary for saturation range over wide limits. Thus at 20 °C, 100 g of water dissolve 192 grams of NH_4NO_3, 6.5 grams of mercuric chloride ($HgCl_2$) and only $8.4x10^{-4}$ grams of Ag Br. In ethyl alcohol, on the other hand, the order of solubility is reversed that is 47.6 grams of $HgCl_2$ and only 3.8 grams of ammonium nitrate. Generally, most of the inorganic substances dissolve in water than in an organic solvent, while the reverse is true for organic substances because they dissolve in organic solvent like benzene.

2. Pressure

The effect of pressure on the solubility of solids in liquids is generally quite small. The solubility of gases is affected much by varying the pressure. At constant temperature the solubility of a gas in a liquid is directly proportional to the pressure of the gas above the liquid. CO_2 is filled in soft drinks at pressure of 3-5 atm. So when pressure is released on the bottle by opening it, CO_2 comes out with effervescence.

3. Temperature (Solubility curve)

The solubility of most of the solid and liquid solutes increase with a rise in temperature. There is, however, no general rule to correlate temperature and solubility. For example: the solubility of KNO_3 and $Al_2(SO_4)_3$ increases with the increase in temperature but the solubility of certain solids like $Ce_2(SO_4)_3$ and Li_2CO_3 decreases with the increase in temperature. The solubility of NaCl and KBr is almost not affected by increase or decrease in temperature but remains constant. The solubility of sodium sulphate increases upto 305.4 K and on raising the temperature further, it decreases. Sodium sulphate forms decahydrate ($Na_2SO_4.10H_2O$) in water. Below 305.4 K the decahydrate is present but above 305.4K it adopts the anhydrous form (Na_2SO_4). The maximum solubility is at 305.4K.

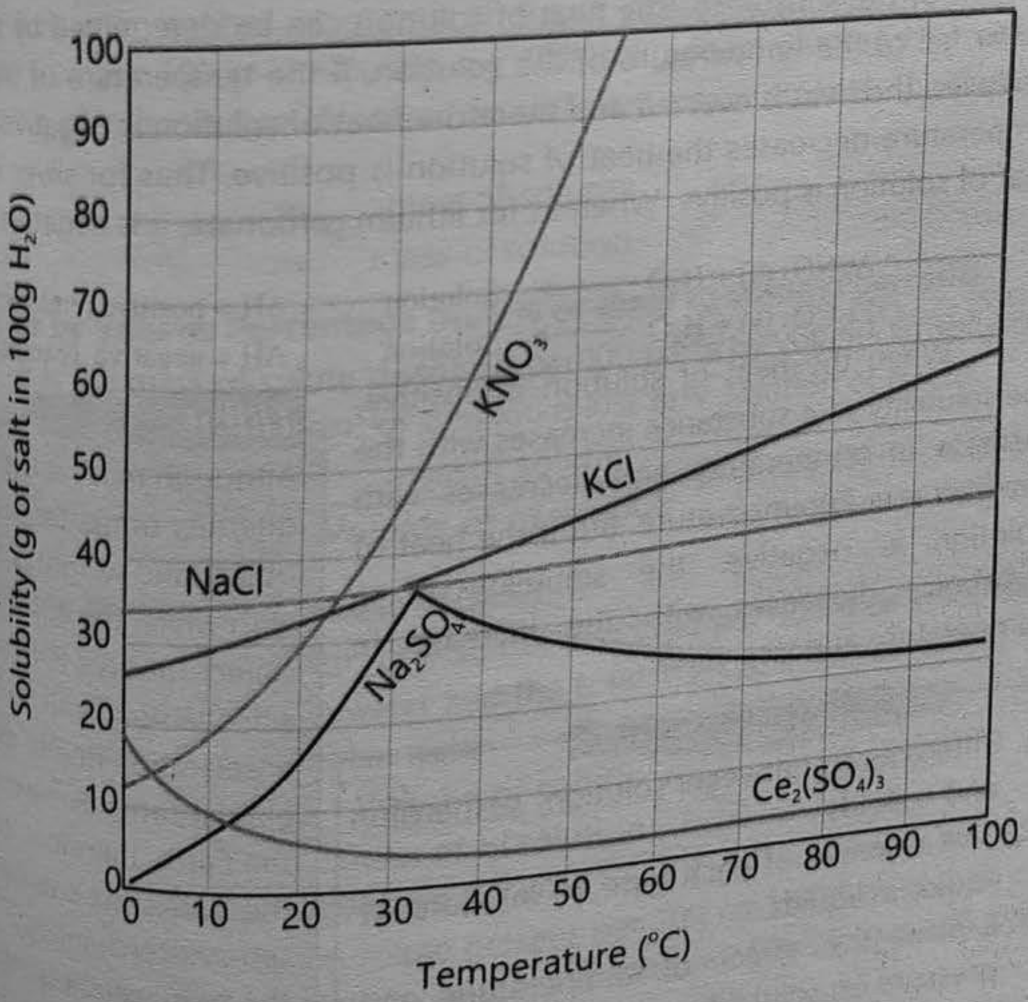


Figure 10.1 Solubility curves

Table : 10.1 Solubility of different Substances at different Temperatures

Substance	Solubility in grams per 100 grams of water					
	273K	283K	293K	303K	313K	323K
KNO_3	13.3	20.9	31.6	45.8	63.9	85.5
$Al_2(SO_4)_3$	31.2	33.5	36.4	40.4	46.1	52.2
$AgNO_3$	122	170	222	300	376	455
NaCl	35.7	35.8	36.00	36.3	36.6	37.0
$Ce_2(SO_4)_3$	20.75		10.08	6.79	-	4.67
Li_2CO_3	1.54	1.43	1.33	1.25	1.117	1.68
Na_2SO_4	5.0	9.0	19.4	40.8	48.8	44.7

10.1.5. Solubility and Heat of Solution

The change in solubility with temperature is closely related to the heat of solution of the substance. The heat of solution can be determined by noting the rise or fall of the temperature of the solution. If the temperature of the solution increases, the heat is evolved and therefore,heat of solution is negative and if the temperature decreases the heat of solution is positive. Thus, for silver nitrate the heat of solution is positive. Whereas for lithium carbonate, it is negative.

$$AgNO_3\,(s) + H_2O \longrightarrow \text{Solution} \qquad \Delta H = \text{positive (endothermic)}$$
$$Li_2CO_3\,(s) + H_2O \longrightarrow \text{Solution} \qquad \Delta H = \text{negative (exothermic)}$$

When the heat of solution is positive, the solubility of a substance increases with the increase in temperature and decreases with the decrease in temperature. But if the heat of solution is negative, the solubility of a substance decreases with the increase in temperature and vice versa.

Self-Assessment

1. Differentiate between solution, suspension and colloids.
2. Give example of each type of solutions of liquids in liquids.
3. Explain the effect of temperature and pressure on solubility.

Tidbit

Although mass is the quantity of matter in an object and weight is force with which the earth pull the object towards it, however, since W=mg, and acceleration due to gravity is almost same everywhere on the Earth, therefore, most of the books use these two terms interchangeably i.e. the term weight is mostly used for mass.

10.2 Concentration Units

The physical properties of solutions depend, to a large extent, upon the relative amounts of solute and solvent e.g. colour of a dye, the sweetness of a sugar solution or the saltish taste of common salt solution depends on the quantity of solute present. The amount of solute present in a given amount of solvent or solution is called concentration of a solution. Greater the amount of solute present in a solution the higher would be its concentration. The amount of solute, solvent and solution may be measured by volume, weight or number of moles. Accordingly, the concentration of a solution can be expressed in many ways.

10.2.1 Percentage Composition

The percentage of solution is expressed in four ways.

i) Percentage by Mass (m/m% or mass to mass percentage)

It is the number of grams of solute dissolved in sufficient amount of solvent to make 100 g of solution e.g 10% solution of glucose by mass means that 10 g of glucose are dissolved in sufficient water so that solution mass is 100g.

$$\frac{m}{m}\% = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

ii) Mass by Volume Percentage (m/V% or mass to volume percent)

It is the mass of solute dissolved per 100 parts by volume of solution. 10g of NaOH dissolved per 100 cm^3 of solution is 10% m/V solution of NaOH.

$$\frac{m}{V}\% = \frac{\text{mass of solute}}{\text{Volume of of solution}} \times 100$$

iii. Volume by mass Percentage (V/m%)

It is the number of cm^3 of a solute dissolved per 100 g of solution. If we dissolve 10 cm^3 of alcohol in water and the total mass of solution is 100g then it is 10% V/m solution of alcohol in water.

$$\frac{V}{m}\% = \frac{\text{Volume of solute}}{\text{mass of of solution}} \times 100$$

iv. Volume by volume Percentage (V/V%)

It is the volume of the solute present per 100 cm^3 of the solution e.g. 10 cm^3 of alcohol is dissolved in water so that the volume of solution is 100 cm^3. It is 10% V/V solution of alcohol in water.

$$\frac{V}{V}\% = \frac{\text{Volume of solute}}{\text{Volume of of solution}} \times 100$$

10.2.2. Molarity (M)

The molarity (M) of a solution is the number of moles of solute per dm^3 o solution.

$$\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{Volume of solution } (dm^3)}$$

$$1dm^3 = 1000cm^3 = 1 \text{ litre}$$

$$\text{Number of moles of solute} = \frac{\text{mass of solute in g}}{\text{Mol. mass of solute}}$$

Therefore,

$$\text{Molarity (M)} = \frac{\text{mass of solute in g}}{\text{Mol. mass of solute}} \times \frac{1}{\text{volume of solution } (dm^3)}$$

For example, one molar solution (1M) of sugar, $C_{12}H_{22}O_{11}$, contains 1 mole or 342 g of sugar per dm^3 of solution. It is prepared by adding 342 g of sugar in one dm^3 volumetric flask and then by adding water first to dissolve and then dilute the solution until the level of the liquid reaches the etched mark i.e when the volume of the solution becomes exactly one dm^3.

10.2.3 Molality (m)

Molality is another concentration unit of solution. Molality is defined as the number of moles of solute present in 1kg of solvent. In other words the molality of a solution is equal to the number of moles of solute dissolved in 1000g of solvent.

$$\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Number of kilogram of solvent}}$$

Or

$$= \frac{\text{mass of solute}}{\text{mol. mass of solute}} \times \frac{1}{\text{Kilogram of solvent}}$$

When 58.5grams (1 mole) of NaCl is dissolved in 1000 grams of water, the solution will be 1.0molal or 1.0m. Similarly, 1.0molal solution of glucose will contain 180g of glucose in 1000g of solvent. If 18g of glucose is dissolved in 1kg of water the molality (m) of glucose solution will be 0.1. Molality is independent of temperature.

Example 10.1

4.5g of glucose are dissolved in 100g of water. Calculate the molality of lution. Molecular mass of glucose is 180g mol^{-1}.

Solution

Mass of glucose = 4.5g.

Mass of water = 100g.

Molecular mass of glucose = 180g mol^{-1}

Calculations

$$\text{Moles of glucose} = \frac{4.5\ g}{180\ g.\ mol^{-1}} = 0.025\ mol$$

$$\text{mass of water in kilograms} = \frac{100\,g}{1000} = 0.1kg$$

$$\text{molality (m)} = \frac{0.025}{0.1kg} = 0.25\ mol\ kg^{-1}$$

Practice Problem 10.1

A sample has 3.50 dm^3 of solution that contains 90g of sodium chloride, NaCl. What is the molarity of that solution?

Practice Problem 10.2

What is the molality of a solution containing 18.2g HCl and 250 g of water?

10.2.4 Mole Fraction (X)

Mole fraction of any component of a solution is defined as the number of moles of that particular component divided by the total number of moles of all the components in the solution. For example, if n_1 is the number of moles of solvent and n_2 is the number moles of solute. X_1 and X_2 are mole fractions of solvent and solute respectively, then mathematically X_1 and X_2 are expressed as

$$\text{Mole fraction of solvent, } X_1 = \frac{n_1}{n_1 + n_2}$$

$$\text{Mole fraction of solute, } X_2 = \frac{n_2}{n_1 + n_2}$$

$$X_1 + X_2 = \frac{n_1}{n_1 + n_2} + \frac{n_2}{n_1 + n_2} = 1$$

Note that the sum of the mole fractions is equal to one. The importance of mole fraction is, like molality, this method of expressing concentration is also independent of temperature. Moreover, concentrations of more than one solutes can easily be expressed by this method.

Example 10.2

A solution contains 2.0 moles of alcohol and 3.0 moles of water. Calculate mole fractions of alcohol and water.

Solution

Number of moles of water: $(n_1) = 3.0$ moles

Number of moles of alcohol: $(n_2) = 2.0$ moles

Mole fraction of water and alcohol are represented by X_1 and X_2 respectively.

Formula

$$\text{Mole fraction (water)}\ X_1 = \frac{n_1}{n_1+n_2} = \frac{3.0}{3.0+2.0} = \frac{3}{5}$$

$$\text{Mole fraction (alcohol)}\ X_2 = \frac{n_2}{n_1+n_2} = \frac{2.0}{3.0+2.0} = \frac{2}{5}$$

Practice problem 10.3

Calculate the mole fraction of methyl alcohol in a solution composed of 1.46 mol of methyl alcohol and 2.19 mol of ethyl alcohol.

10.2.5 Parts per million, billion, and trillion

i) Parts per million (ppm)

It is defined as the number of parts (by mass or volume) of a solute per million parts (by mass or volume) of the solution. Parts per million is used for very low concentration of solution. The concentration of impurities of substances in water is expressed by this unit.

$$\text{ppm} = \frac{\text{mass or vol. of solute}}{\text{mass or vol. of solution}} \times 10^6$$

Example 10.3

Sea water contain 5.65×10^{-3}g of dissolved oxygen per kilogram of solution. Calculate the concentration of oxygen in parts per million in sea water.

Solution

Mass of solute (dissolved oxygen) = 5.65×10^{-3}g

Mass of solution =1kg

Concentration of solute in ppm=?

As

$$\text{Concentration (ppm)} = \frac{\text{mass of dissolved oxygen}}{\text{mass of solution}} \times 10^6$$

$$\text{ppm of oxygen} = \frac{5.65\times10^{-3}\ \text{g}}{1000\ \text{g}} \times 10^6$$

$$= \frac{5.65\times10^{-3}\ \text{g}}{10^3\ \text{g}} \times 10^6 = 5.65\ \text{ppm}$$

Parts per billion (ppb)

It is defined as the number of parts (by mass or volume) of a solute per lion parts (by mass or volume) of the solution. Parts per billion is used for very w concentration of solution.

$$\text{Concentration (ppb)} = \frac{\text{mass or vol. of solute}}{\text{mass or vol. of solution}} \times 10^9$$

ample 10.4

A 2.0dm^3 gas cylinder contains a mixture of various gases. If 5.0cm^3 of H_2 s is present in this mixture. Calculate the concentration of H_2 gas in parts per lion (ppb).

lution

Volume of H_2 gas =5.0cm^3.

Volume of cylinder= 2.0 dm^3 = (2000 cm^3).

Concentration of H_2 gas in ppb = ?

$$\text{Concentration of } H_2 \text{ gas (ppb)} = \frac{\text{Vol. of } H_2 \text{ in(cm}^3)}{\text{Vol. of solution in(cm}^3)} \times 10^9$$

$$\text{Concentration of } H_2 \text{ gas in ppb} = \frac{5.0\ \text{cm}^3}{2000\ \text{cm}^3} \times 10^9$$

$$= 2.5\times10^{-3}\times10^9$$

$$= 2.5\times10^6\ \text{ppb}$$

Parts per trillion (ppt)

It is defined as the number of parts (by mass or volume) of a solute per lion parts (by mass or volume) of the solution. Parts per trillion is used for wer concentrations.

$$\text{Parts per trillion(ppt)} = \frac{\text{mass or volume of solute}}{\text{mass or volume of solution}} \times 10^{12}$$

Self-Assessment

1. In how many ways, percentage of solution can be expressed.
2. Differentiate between molarity (M) and Molality (m).

10.3 Raoult's Law

When a solid solute dissolves in a solvent, it changes the vapour press of the solvent. Similar effect is observed for a liquid solute in a liquid solve However, there is a major difference; a solid solute is usually non-volatile, whi liquid solute is volatile. For example, if we analyse the vapour above a su solution in water, we would find only water vapours, whereas, the vapours abo a solution of ethanol and water always contain both water and ethanol.

This fact allows us to classify the solutions on the basis of solutes i.e. (i) t solution of volatile solute and solvent and (ii) the solution of non-volatile and no electrolyte solute in volatile solvent.

The quantitative relationship between the vapour pressure of solution a composition of components of solution was given by F. M. Raoult. *The Raou law states that the partial vapour pressure of any volatile component o solution is equal to the vapour pressure of the pure component multiplied by t mole fraction of that component in solution*. Those solutions which obey Raou law are called ideal solutions.

Mathematically, Raoult's law can be written as $P = P^\circ X$ where P and P° a the partial vapour pressure of the component and that of the pure compone respectively, and X is the mole fraction of the component in the solution.

10.3.1 The solution of Volatile Solute and Solvent (bo components are volatile)

Consider an ideal solution of two liquids A and B. According to Raou law partial vapour pressure of both components A and B is given by:

$$P_A = P_A^\circ X_A$$

$$P_B = P_B^\circ X_B$$

P_A° is the vapour pressure of pure component A and P_B° is vapour pres of pure component B. X_A and X_B are the mole fractions of components A a respectively. By applying Dalton's law of partial pressures, vapour pressu solution is

$$P_{Total} = P_A + P_B$$

$$P_t = P_A^\circ X_A + P_B^\circ X_B \qquad (10.1)$$

Since $X_A + X_B = 1$

Therefore $X_A = (1 - X_B)$

The eq.(10.1) becomes

$$P_t = P_A^o (1 - X_B) + P_B^o X_B$$

$$P_t = P_A^o - P_A^o X_B + P_B^o X_B$$

On rearranging the above equation.

$$P_t = (P_B^\circ - P_A^\circ) X_B + P_A^\circ$$

This equation shows that when $X_B = 0$, $X_A = 1$ it means pure component A is present and P_{Total} becomes equal to P°_A and if $X_A = 0$, then $X_B = 1$ then pure component B is present, then P_{Total} becomes equal to P^o_B.

The graph in Fig. 10.3 shows variation of total pressure (solid line) and partial vapour pressures of components (dotted lines) of solution with changes in concentration of both A and B components. At any concentration, the total pressure is the sum of the partial pressures of the two components.

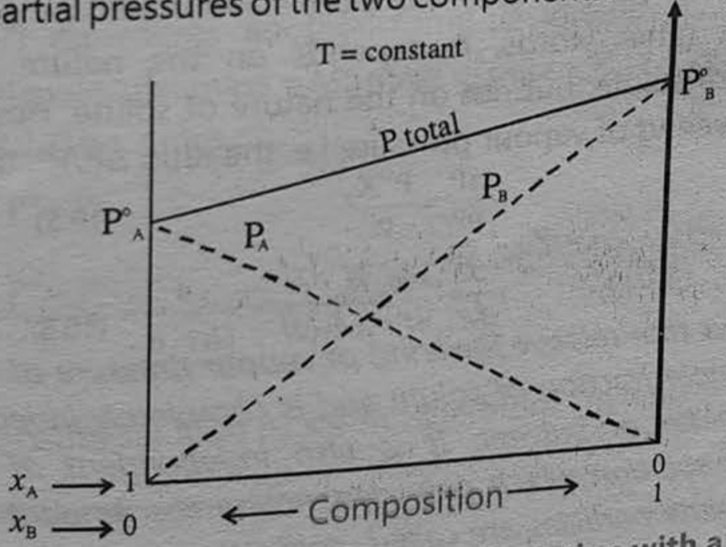


Figure 10.2: Variation of vapour pressure of solution with a change in concentration of A and B

10.3.2 Non-Volatile and Non-Electrolyte Solute in Volatile Solvent

In the case of solution of non-electrolyte and non-volatile solute the solute particles remain in the (un-dissociated) molecular form. Such solute lowers the vapour pressure of the solvent. The vapour pressure lowering, suffered by the solvent, can be readily understood in terms of Raoult's law.

Let us consider a solution of component A and B where A is solvent of mole fraction X_1 and B is solute of mole fraction X_2. P° is the vapour pressure of pure solvent and P is the vapour pressure of solution. According to Raoult's law, vapour pressure P is given by

$$P = P^{\circ}X_1 \qquad (10.2)$$

Since X_1 in any solution is less than unity, P must always be less than P°. Consequently, a solute in a solvent tends to lower the vapour pressure of the pure solvent. Furthermore, if the solute is non-volatile it does not contribute to the total vapour pressure, and hence Eq. (10.2) gives the total vapour pressure above the solution, which in this case is due to solvent only and is always less than P°. The extent of the vapour pressure lowering ΔP is

$$\Delta P = P^{\circ} - P \qquad (10.3)$$

$$\Delta P = P^{\circ} - P^{\circ}X_1$$

$$\Delta P = P^{\circ}(1 - X_1)$$

$$\Delta P = P^{\circ}X_2 \qquad (10.4)$$

According to Eq.(10.4), *the lowering of vapour pressure of solution depends both on vapour pressure of pure solvent and the mole fraction of solute in solution.* In other words, it depends on the nature of solvent and the concentration of solute, but not on the nature of solute. However, if we consider the relative lowering of vapour pressure i.e. the ratio $\Delta P/P^{\circ}$, then from Eq. (10.4).

$$\frac{\Delta P}{P^{o}} = \frac{P^{o}X_2}{P^{o}} \qquad (10.5)$$

$$\frac{\Delta P}{P^{o}} = X_2$$

(10.6)

It shows that relative lowering of vapour pressure of the solution depends only on the mole fraction of solute and is completely independent of either the nature of solute or solvent. It is also independent of temperature. Those properties of solution which only depend on the amount of the solute or the number of solute particles are called colligative properties. Thus the lowering in vapour pressure is a colligative property.

10.4 Colligative Properties of Dilute Solutions

Colligative (or collective) properties are those properties which depend only on the number of solute particles but not on the nature of the solute particles. The study of colligative properties is important especially because they provide methods for determining the molecular weight of dissolved substances. The colligative properties of a dilute solution containing a non-volatile non-electrolyte solute are:

a. Lowering of vapour pressure.
b. Elevation of boiling point.
c. Depression in freezing point.
d. Osmotic pressure.

10.4.1 Vapour Pressure lowering

The lowering of vapour pressure can be used to calculate the molecular mass 'M' of solute as follows.

Since $X_2 = \frac{n_2}{n_1 + n_2}$

So Equation (10.6) becomes. $X_2 = \frac{\Delta P}{P^\circ}$

$$\frac{P^\circ - P_s}{P^\circ} = \frac{\Delta P}{P^\circ} = \frac{n_2}{n_1 + n_2} \qquad (10.7)$$

Eq. (10.7) represents another form of Raoult's law. If mass of pure solvent = W_1 and molecular mass = M_1, mass of pure solute = W_2 and molecular mass = M_2 then Eq. (10.7) becomes

$P^\circ - P_s$

$$\frac{P^\circ - P_s}{P^\circ} = \frac{n_2}{n_1 + n_2} = \frac{W_2/M_2}{W_1/M_1 + W_2/M_2} \qquad (10.8)$$

For a very dilute solution $n_2 << n_1$: therefore, $(n_1 + n_2) = n_1$, Thus $\frac{W_2}{M_2}$ can be neglected from the denominator.

The Eq. (10.8) then can be written as

$$\frac{P^\circ - P_s}{P^\circ} = \frac{W_2 M_1}{W_1 M_2} \qquad (10.9)$$

Eq. (10.9) can be used to calculate molecular mass of the dissolved substance if relative lowering of vapour pressure and the definite mass of solute in a given amount of solvent are known.

10.4.2 Boiling Point Elevation and Freezing Point Depression

(i) Boiling Point Elevation

Boiling point of a solution is the temperature at which its vapour pressu becomes equal to external pressure or atmospheric pressure. Since the presenc of a non-volatile solute lowers the vapour pressure of the solution, therefore boiling point must increase. This fact has been qualitatively expressed by Raoult's law.

According to Raoult's law the vapour pressure of solution containing non-volatile solutes is always less than that of pure solvent. As a result of lowering of vapour pressure the solution will boil at a temperature higher than the boiling point of the pure solvent at the atmospheric pressure. *This difference in the boiling points of solution and its pure solvent is known as boiling point elevation of the solution.* It depends on the nature of solvent and the concentration of solute. The boiling point elevation in dilute solution is independent of the nature of solute if it does not ionize, associate, or react with the solvent. *For 1.0 mole of any solute if dissolved in 1 k of pure solvent, the elevation of boiling point is called molal boiling elevatio constant or ebullioscopic constant.* If 1 mole of a solute is dissolved in 1 kg o pure water the solution now boils at 100.52°C instead of 100°C. The differenc 0.52°C is the ebullioscopic constant, k_b, of water. Elevation of boiling point can b explained by plotting a graph, of vapour pressure verses temperature for pu water and an aqueous solution.

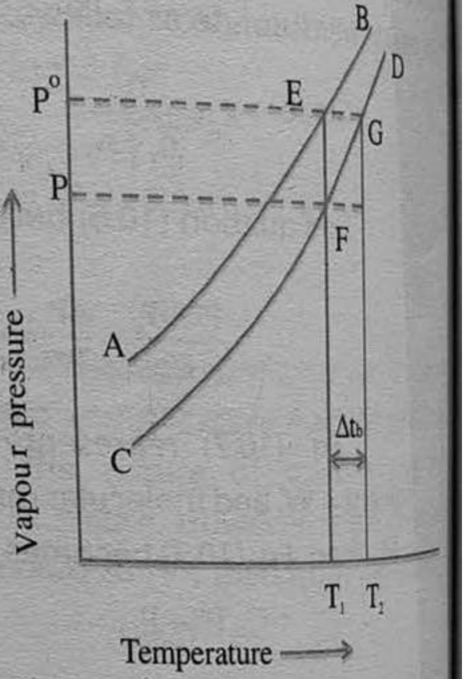


Figure 10.3 Boiling Point Elevatio

Consider Fig. 10.3 the curve AB and CD represent vapour pressure of th pure solvent and solution, respectively, as a function of temperatur Temperature T_1 on curve AB is boiling point of pure solvent. The solvent bo when its vapour pressure become equal to the external pressure represented b P^o. T_2 on curve CD is the boiling point of solution. CD curve is lower than th curve AB because vapour pressure of solution is less than that of pure solve Thus the solution will boil at higher temperature T_2 to equalize its vapo essure to P°. Hence elevation of boiling point ΔT_b is given as,

$$\Delta T_b = T_2 - T_1$$

Increase in concentration of solute particles will increase the boiling point f the solution. So $\Delta T_b \propto m$ where m is the molality of solution.

$$\Delta T_b = K_b m \qquad (10.10)$$

K_b is called molal boiling point elevation constant or ebullioscopic constant, when solution is 1 molal or m=1 then $\Delta T_b = K_b$.

It means that *if a solution is of unit molality its elevation of boiling point is lled ebullioscopic constant.*

The molality (m) of the solution containing W_2 gram of solute of molecular ass M_2 dissolved in W_1 gram of solvent of molecular mass M_2 is.

$$m = \frac{W_2}{M_2} \times \frac{1000}{W_1}$$

Substituting the value of "m" in Equation (10.10) we get.

$$\Delta T_b = K_b \left(\frac{1000 W_2}{M_2 W_1} \right) \qquad (10.11)$$

Or $$M_2 = \frac{K_b}{\Delta T_b} \times \frac{W_2}{W_1} \times 1000 \qquad (10.12)$$

Thus if masses of solute and solvent are known along with ebullioscopic nstant and elevation of boiling point then molecular mass of solute can be lculated.

ample 10.5

The boiling point of 50g of carbon tetrachloride was raised by 0.402 K hen 0.5126g of naphthalene (mol.mass =128g/mol) was dissolved in it. Calculate e ebullioscopic constant of carbon tetrachloride.

lution

Mass of carbon tetrachloride = W_1 = 50g
Mass of naphthalene = W_2 = 0.5126g
Molecular mass of naphthalene = M_2 = 128 g mol^{-1}
Boiling point elevation ΔT_b = 0.402K
Ebullioscopic constant" = K_b =?

As we have $$K_b = \frac{\Delta T_b \times W_1 \times M_2}{1000 W_2}$$

Putting values in the above equation.

$$= \frac{0.402\text{K}}{1000\text{gkg}^{-1}} \times \frac{50\text{g} \times 128\text{gmol}^{-1}}{0.5126\text{g}}$$

$$= 5.02\text{K kgmol}^{-1}$$

(ii) Freezing Point Depression

The freezing point of a liquid is the temperature at which the solid pha begins to separate out from the liquid phase. At this temperature solid and liqu phases are in equilibrium and have equal vapour pressure. A solution freezes lower temperature as compared to its pure solvent. This lowering of freezing point of solution as compared to its pure solvent is called depression of freezing point. This lowering is due to lowering in vapour pressure of solution as a result of addition of small amount of non-volatile solute. The depression of freezing point of a substance depends on:

i. The nature of solvent.

ii. Concentration of solute particles.

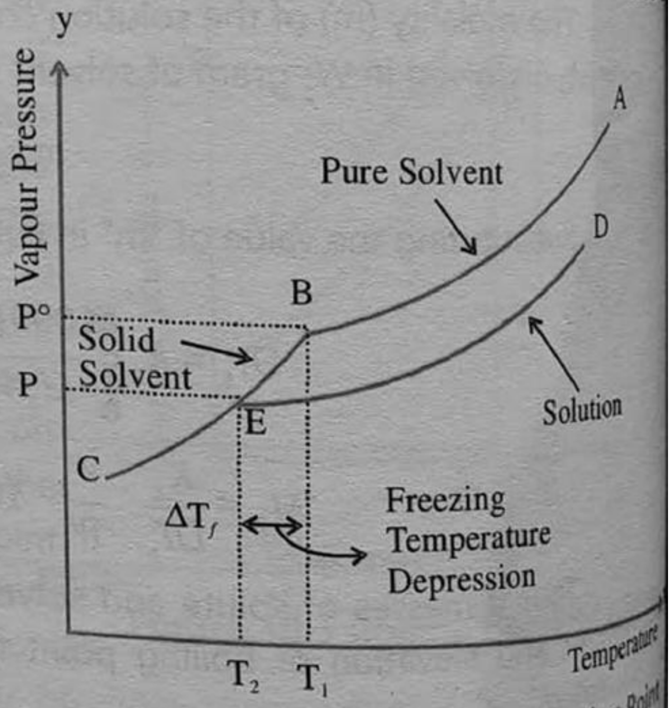


Figure 10.4 Depression in Freezing Point

In order to understand the relationship between lowering in freezin point, ΔT_f, and concentration in terms of molality 'm', consider Fig. 10.4 whi shows the vapour pressure of solution and pure solvent as a function temperature.

The curve AB is for the vapour pressure of the pure solvent and curve for that of the solution. The curve CB is the sublimation curve of the solid solv The solvent freezes at temperature T_1 corresponding to the point B where vapo pressure of freezing solvent is P°. The portion of the curve BC is for the s solvent. The curve DE for solution intersects curve BC at point E. This is freezing point of solution T_2. Hence

$$\Delta T_f = T_1 - T_2$$

ΔT_f, the depression in freezing point is related to the molality (m) of th

solution as

$$\Delta T_f \propto m$$

$$\Delta T_f = k_f m \qquad (10.13)$$

K_f is the molal freezing point elevation constant also called as the cryoscopi constant. We know that

$$\text{molality}\ (m) = \frac{1000W_2}{M_2.W_1}$$

Then from Equation 10.13

$$\Delta T_f = k_f \frac{1000W_2}{M_2.W_1} \qquad (10.14)$$

W_2 and M_2 are the mass and molar mass of solute, respectively, W_1 is the mass of solvent in gram and the factor of 1000 is for converting g into kg. In order to calculate M_2 molecular mass of solute, Eq. (10.14) is rearranged as

$$\boxed{M_2 = k_f \frac{1000\,W_2}{DT_f W_1} \qquad (10.15)}$$

Thus, if we know the mass of solute and solvent, molal freezing point constant (k_f) and depression in freezing point ΔT_f. We can calculate molar mass of solute. Moreover, this equation can also be used to ascertain the purity of solvent from the k_f value. From Eq. 10.13 ΔT_f is equal to k_f when molality is unity. Thus for solution of unit molality, if the calculated value of ΔT_f does not match with the literature value of k_f then this is an indication of impure liquid. Here it is instructive to define *the molal freezing point constant or cryoscopic constant 'kf' as the depression in freezing point of a solution of unit molality.* Different solvents have their own specific k_f values.

10.4.3 Molar Mass determination by Vapour Pressure lowering, Boiling Point Elevation, and Freezing Point depression

Molar mass of substances can be determined by vapour pressure lowering, boiling point elevation, and freezing point depression methods using equations 10.9, 10.12, and 10.15, respectively.

ample 10.6

If 3.60g of glucose is dissolved in 100g of water and the freezing point depression is found to be 0.372K. Calculate the molecular mass of glucose. k_f for water is 1.86 K kg mol^{-1}.

Solution

Weight of glucose = W_2 =3.60g

Weight of water = W_1 =100g

Freezing point depression = ΔT_f = 0.372K

Molal freezing point of water = k_f = 1.86K kg mol^{-1}

Molecular mass of glucose M_2 = ?

As we know $M_2 = \dfrac{K_f}{\Delta T_f} \dfrac{1000W_2}{W_1}$

$$M_2 = \frac{1.86\text{ ' }1000\text{ ' }360}{0.372\text{ ' }100}$$

$$M_2 = 180\text{g} / \text{mol}$$

Practice problem 10.4

A solution containing 1.66 g of nonionic solute in 171 g of water freezes a 0.602°C. Calculate the molar mass of the solute. K_f of water is 1.86 K kg mol^{-1}.

Osmosis

When a solution is separated from its pure solvent, by a semi-permeab membrane, it is observed that solvent tends to pass through the membran spontaneously into the solution and thereby dilute it. The phenomenon is calle osmosis. A membrane which allows the passage of solvent molecules only, no the solute, through it is called a semi-permeable membrane. Osmosis also take place when a dilute solution is separated from a concentrated solution th solvent molecules from dilute solution pass through the membrane towards th concentrated solution region. For low molecular mass solutes in water the be semi-permeable membrane is a film of copper ferrocyanide $Cu_2[Fe(CN)_6]$. Fo high molecular weight solutes in organic solvents the membranes used, mo frequently, are thin films of either cellulose or cellulose nitrate.

Osmosis can easily be understood with the help of an apparatus show Fig. 10.5. A semi-permeable membrane e.g. egg membrane, is tied over the e of a thistle funnel. It is, then, filled with a solution of sugar and dipped in beaker containing water (solvent). Due to osmosis, water will pass throug membrane and the level of sugar solution will rise until the hydrostatic pres of the liquid column equalizes the osmotic pressure of the solution.

Thus, the flow of solvent molecules from lower concentration regio higher concentration region is called osmosis or endosmosis.

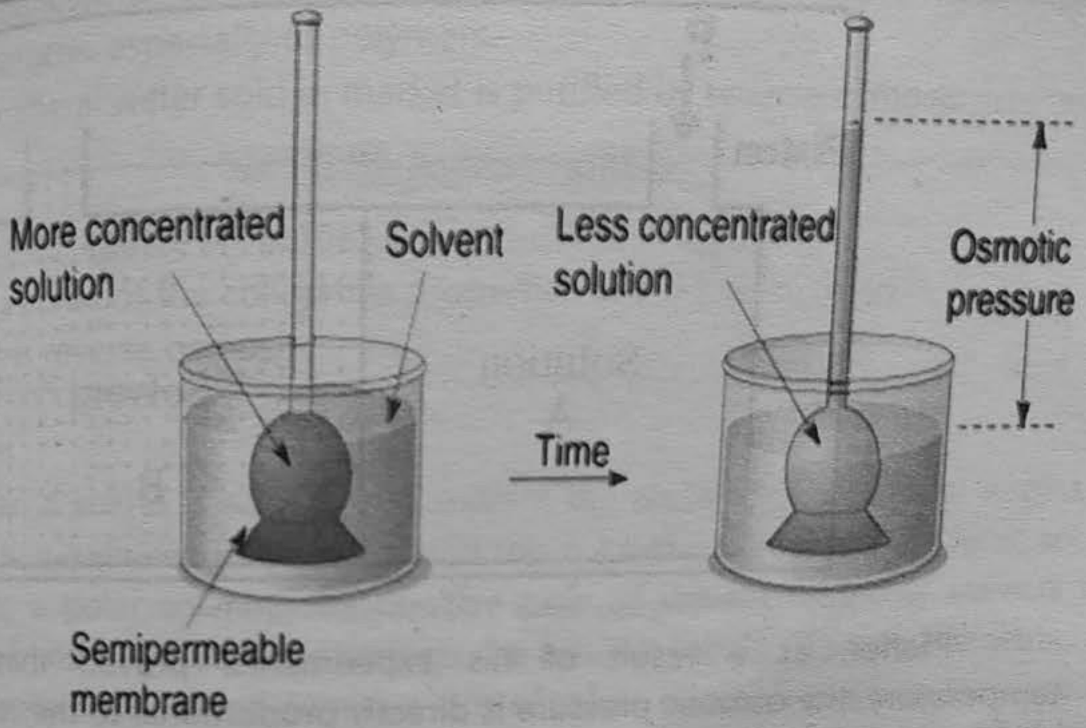


Figure 10.5 Osmosis

10.4.4 Osmotic Pressure and Reverse Osmosis

Another example may be given to illustrate the phenomenon of osmosis and osmotic pressure. Consider Fig.10.6, the chamber is divided by a semipermeable membrane into two compartments. The left compartment 'A' is filled with the solution and the opening of the compartment is provided with a moveable piston while the right compartment 'B' contains the pure solvent and its mouth is open to atmosphere. The solvent will pass through the semipermeable membrane into the solution compartment due to the osmosis and will push the piston upward. The upward movement of the piston can be prevented by applying pressure on the piston to keep it in the original position. *The pressure that must be applied on the solution in order to prevent the osmosis is called the osmotic pressure of the solution. It is denoted by 'π'*. If the solvent in the above experiment is replaced by a solution of different concentration, osmosis will still occur from dilute solution towards concentrated solution. If the solutions have same concentration on both sides, no osmosis will occur and the solutions are said to be isotonic (having same osmotic pressure). *If we apply external pressure on the concentrated solution side, more than the osmotic pressure, the solvent will move from concentrated region towards the dilute region. This process is called reverse osmosis or exosmosis.*

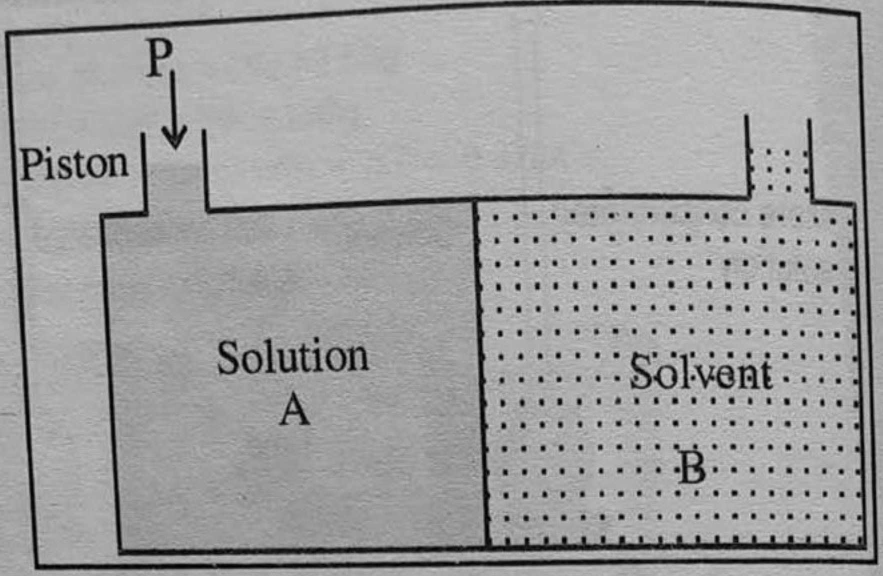


Figure 10.6 Osmotic Pressure

Pfeffer, as a result of his experiments, proved that at consta temperature, the osmotic pressure is directly proportional to the concentration solution as is shown in Table 10.2. This indicates that osmotic pressure is colligative property.

Table: 10.2: Variation of Osmotic Pressure with Concentration of Sucros

% Concentration (C)	Osmotic pressure (mm of Hg)
1	535
2	1016
2.47	1518
4	2082
6	3075

10.4.5 Application of Osmosis in Daily Life

1. The process is used in many industries.
2. The victims of ship wreckage purify seawater, which has very h concentration of salts, by reverse osmosis, to make it drinkable.
3. By osmotic pressure the plant cell become turgid and these turgid c provide support to weak parts of plant.
4. Ascent of sap in plants from roots to upper parts takes place by osm and osmotic pressure
5. Isotonic (same osmotic pressure) solution have been prepared to pre plasmolysis (shrinkage of cytoplasm) and hemolysis (deficiency of blood cells).

6. The concept of osmotic pressure can be used to calculate molar mass of substances, especially of polymers.
7. The mineral water sold in market is purified by reverse osmosis process.

Self-Assessment

1. Define colligative properties.
2. Name some of the colligative properties of a dilute solution.
3. What is reverse osmosis.

Solvation

When a solute is added to a solvent the dissolution of solute begins by he interaction between its particles and the solvent molecules. If an ionic solute added to a polar solvent, the positive pole of solvent molecule attracts the egative ion and negative pole attracts the positive ion. In this way, the attractive orces between positive and negative ions of solute diminish and now they are far part from one another. Both the positive and negative ions are surrounded by a phere of solvent molecules and keep them in the dissolved form. This process is alled solvation, on the other hand if water is used as solvent the solvation is called ydration.

The hydration of KCl, for example, is shown in Fig.10.7. It can be seen that ations, K^+, and anions, Cl^-, get separated from one another and are surrounded y oppositely charged poles of water molecules

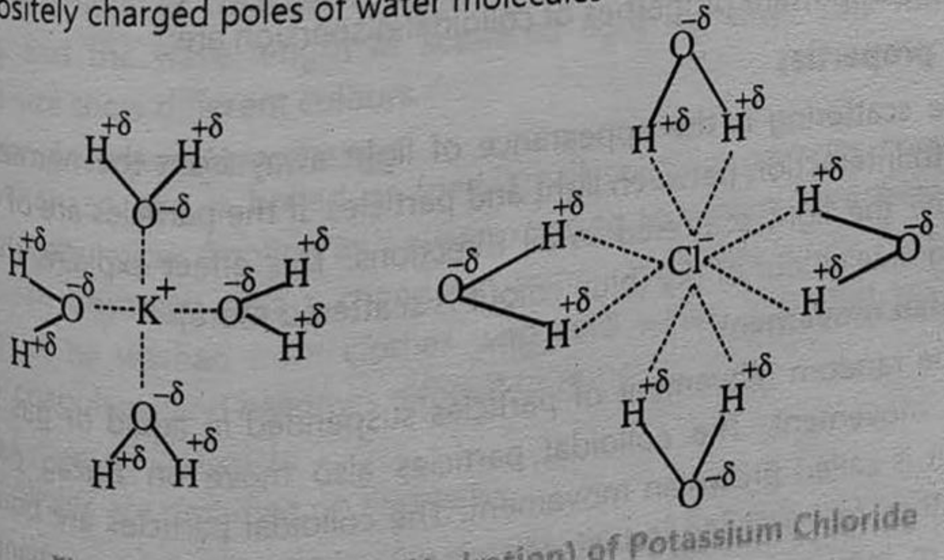


Figure 10.7 Solvation (Hydration) of Potassium Chloride

Solvation is a process of reorganizing solvent and solute molecule into solvent complex. Solvation involves bond formation, hydrogen bonding, and Van der Waals forces or dispersion forces. Ionic substances are solvated by polar solvents, like water. However, non-polar solutes are solvated by non-polar solvent. The forces involved in such case are predominantly Van der Waals forces.

Water of Crystallization

Whenever an acid-base neutralization reaction takes place, salt and water are formed i.e. we get a salt solution. In order to get salt, water is dried from it by evaporation. In some types of salts some of the water molecules remain inside the salt crystals, that is just a physical presence of water molecules and is called *water of hydration or water of crystallization*. Some salts absorb water molecules from the atmosphere if kept open and get hydrated. Examples of such salts are: Pentahydrate copper sulphate ($CuSO_4.5H_2O$), heptahydrate magnesium sulphate ($MgSO_4.7H_2O$) or decahydrate sodium carbonate ($Na_2CO_3.10H_2O$).

10.5 Colloids

A colloidal solution or colloidal dispersion is a heterogeneous mixture in which one substance is dispersed (called dispersed phase), as very fine particles, in another substance (called dispersion medium). The size of the dispersed particle is larger than that of true solution and smaller than that of suspension. The size ranges between 1 to 1000 nm. It is intermediate between suspension and true solution. The particles of collides remain suspended and do not settle down at the bottom of the mixture. They are visible under high resolving microscope.

10.5.1 Properties of colloids

Some important properties of colloidal dispersion are:

I. Optical properties

The scattering is the appearance of light away from the normal optical path due to interaction between light and particles. If the particles are of colloidal size we see the light scattered in all directions. This effect explains the tail of comet and blue colour of sky (blue colour is scattered more).

ii. Brownian movement

The random movement of particles suspended in liquid or gas is called Brownian movement. The colloidal particles also move on zigzag path. This movement is called Brownian movement. The colloidal particles are bombarded by the other fast moving molecules of the medium and thus change their direction of movement. The higher the temperature, the more vigorous is the movement and hitting by the molecules. A zigzag motion of colloidal particles can be observed by an ultramicroscope.

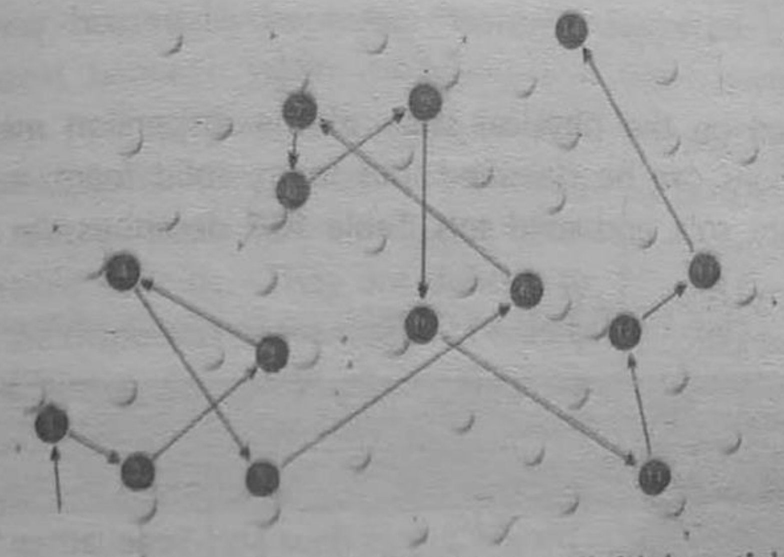


Figure 10.8 Brownian movement of a single colloidal particle

iii. Filterability

Colloidal particles cannot be separated by ordinary filtration but these particles can be removed generally by specially designed filters of extremely fine porosity called ultrafilters. This process of separation of colloidal particles by ultrafilters is called ultrafiltration.

iv. Diffusibility

The colloidal particles have very large size as compared to ordinary solute particles hence they have little power of diffusion.

v. Colour

The colour of colloidal solution depends on size and shape of colloidal particles and the wave length of scattered light e.g. silver sols with different particles size show different colours.

vi. Effect of temperature change

When an egg is boiled or fried, small colloidal particles convert into a lump by coagulation. Cooling and even exposing to sunlight may help jellies set. It has been observed that a change in temperature coagulate colloids. Blood stains must be washed from clothes with cold water. With hot water protein material coagulates and becomes difficult to remove.

10.5.2 Types of colloids

Colloids are classified on two basis.

Basis for classification of Colloids

Components (physical state of matter)

Behaviour

a. On the basis of physical states of dispersed phase and dispersion medium

Based on the physical state of the dispersion medium and dispersed phase, collides can be classified into foam, solid foam, aerosol, emulsion, gel, solid aerosol, sols, and solid sols. Table 10.3 describes the types along with the examples.

Table.10.3 Types of Colloids on the basis of Physical State

S#	Dispersed Phase	Dispersion Medium	Name of the colloidal system	Examples of collides
1	Gas	Liquid	Foam	Soap lather, Soda water, Froth, etc.
2	Gas	Solid	Solid foam	Cake, bread, lava.
3	Liquid	Gas	Aerosol	Mist, fog, clouds.
4	Liquid	Liquid	Emulsion	Milk, cream, butter, oil in water
5	Liquid	Solid	Gel	Curd, chees, jellies, butter, shoe polish
6	Solid	Gas	Solid aerosol	Smoke, dust
7	Solid	Liquid	Sol	Paint, ink
8	Solid	Solid	Solid sol	Alloys, coloured glass, gem stone

b. Classification of colloids on the basis of Interaction between Dispersed Phase and Dispersion medium

Classification of colloids on the basis of interaction between dispersed phase and dispersion medium can be made as *lyophilic and lyophobic colloids.*

Lyophilic colloids. Lyophilic means liquid loving or solvent attracting. If the dispersion medium is water then the term *hydrophilic colloids* is used. In these colloids dispersed particles generally have same chemical nature as that of solvent or dispersion medium, such as –OH group able to form hydrogen bond. When such colloids are once precipitated they can be directly returned into the colloidal form and are, therefore, called reversible colloids. A gel is a semisolid mass of a lyophilic sol, other examples are starch and rubber solutions.

Lyophobic colloids: The colloide is called *lyophobic* if the dispersed phase has little or no affinity for the dispersion medium. Lyophobic means 'liqui fearing' or 'solvent hating'. If water is used as a solvent then the term *hydrophobic colloids* is used. Hydrophobic means "Fearing water". In this case no attractive force exist between water molecules and dispersed particles. The colloids formed by hydrophobic particles are less stable. If, once, precipitated they cannot be directly reconverted into the colloidal form and are, thus, termed as irreversible colloids. They are difficult to prepare because the dispersed phase does not readily form colloide with the dispersion medium. They are unstable and require stabilizing agent for their preservation. Examples are sols of metals like silver and gold, sols of metal hydroxide such as $Al(OH)_3$ etc. Moreover, homogenized milk and lassi (casein suspended in water) are the examples of hydrophilic and hydrophobic sols, respectively.

ReadingCheck

- Differentiate between Lyophilic and Lyophobic colloids.
- What are the basis for the classification of colloids?

10.6 Heat of Solution and its application

The formation of a solution is accompanied with the change in temperature. If solid sodium hydroxide (NaOH) is dissolved in water, the temperature of water will rise. This is an exothermic process. On the other hand, when NH_4NO_3 is dissolved in water the temperature of the water decreases due to endothermic process. That is heat is either absorbed or evolved during the process of dissolution.

The dissolution process is comprised of three processes.

i) Breaking of solute-solute attraction (endothermic process, ΔH positive)
ii) Breaking of solvent-solvent attraction (endothermic process, ΔH positive)
iii) Resulting solute-solvent attraction (exothermic process, ΔH negative)

The value of heat of solution is the net effect of these three heats i.e. the algebraic sum of these three heats. If heat evolved is higher than the heat absorbed the dissolution process will be exothermic and vice versa *The heat of solution is defined as amount of heat evolved or absorbed when one mole of a substance is completely dissolved in large amount of solvent at constant temperature and pressure.*

In the gaseous state, the attractive forces between the gas molecules are very weak. Thus when gases are dissolved in solvent, the solute-solute interaction has little effect. So energy is always evolved when a gas is dissolved in solvent, as only the solute - solvent interactions are operative.

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Heat changes occur when a solute is dissolved in solvent. 'Hot and Cold Packs' use this property. These packs are used by patients to treat their muscle strain and sore joints and to reduce swellings in body muscles.

These packs contain solute and the solvent in separate bags inside the pack. At the time of use, the pack is squeezed and the solvent bag is broken. Solvent reacts with solute and the temperature effect is utilized by the patient.

Mostly, the hot packs use calcium chloride, which produce heat when dissolved in water as below

$$CaCl_{2(s)} \longrightarrow Ca^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \qquad \Delta H = -\ 82.8\ kJ\ mol^{-1}$$

The molar heat of solution of $CaCl_2$ is 82.8kJ mol^{-1} and it is an exothermic reaction.

The cold packs utilize ammonium nitrate (NH_4NO_3) which absorbs heat from surrounding when it is dissolved in water

$$NH_4NO_{3(s)} \longrightarrow NH_4^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)} \qquad \Delta H = +25.7\ kJ\ mol^{-1}$$

It is an endothermic reaction and reduce the pain by cooling effect.

KEY POINTS

- A homogenous mixture of two or more substance is called solution.
- When solute particles are greater in size and do not dissolve completely, then the mixture is called suspension.
- When a solution is in between true solution and suspension, it is called colloidal dispersion or colloidal system.
- Solution of liquids-in-liquids are of three types, completely miscible liquids, partially miscible liquids, completely immiscible liquids.
- Ideal solution is one which obeys Raoult's Law over all ranges of temperature and pressure.
- Molarity, molality, ppm, ppb, ppt, and mole fraction are also concentration units of solution.
- The properties which depends on number and not on nature of solute particles are called colligative properties.
- Lowering of vapour pressure, elevation of boiling point, depression of freezing point, and osmotic pressure are colligative properties.
- Colloids are mixtures whose particle size is between that of true solution and suspension.
- Colloids are classified on the basis of physical states and interaction between the particles of dispersed phase and those of dispersion medium i.e. lyophilic and lyophobic.

EXERCISE

Choose the correct answer

1. In atmospheric gaseous solution, nitrogen is a
   a) solvent b) solute
   c) mixture d) dispersed phase
2. In ice cream, sugar is
   a) solute b) mixture
   c) solvent d) none
3. Elevation of boiling point is a property
   a) additive b) constitutive
   c) colligative d) none
4. Change in enthalpy, ΔH_{sol}, is heat of solution which is
   a) positive b) negative
   c) zero d) may be positive or negative.
5. Freezing point of solution as compared to the solvent is
   a) higher b) lower c) variable d) remains the same
6. Phenol - water system is the example of
   a) completely miscible liquids b) completely immiscible liquids
   c) partially miscible liquids d) none
7. Fog is an example of
   a) solution b) colloid c) suspension d) coarse mixture
8. Which of the following solute would not form a solution if added to water?
   a) KCl b) CH_4 c) HCl d) NH_3
9. Vapour pressure of a given liquid will decrease if
   a) surface area of liquid is increased.
   b) total volume of liquid in the container is decreased.
   c) volume of the vapour phase is increased.
   d) the temperature is decreased.
10. For a dilute solution Raoult's law states that
    a) The lowering of vapour pressure is equal to the mole fraction of solute.
    b) The relative lowering of vapour pressure is equal to the mole fraction of solute.
    c) The relative lowering of vapour pressure is proportional to the amount of solute.

d) The vapour pressure of the solution is equal to the mole fraction solvent.

11. The boiling points of C_6H_6, CH_3OH, $C_6H_5NH_2$, and $C_6H_5NO_2$ are 80° 65°C, 184°C and 112°C respectively. Which one will show the highes vapour pressure at room temperature?
    a) C_6H_6. b) CH_3OH
    c) $C_6H_5NH_2$ d) $C_6H_5NO_2$
12. The process in which the water molecules surround and interact with the solute ions or molecules is called
    a) neutralization b) hydration
    c) hydrolysis d) hydrogenation
13. The molarity of pure water in itself is
    a) 18 b)50 c) 59.5 d)55.6

Short Questions

1. Define the following concentration units:
   i) ppm ii) ppb iii) ppt iv) mole fraction
2. Which concentration unit, from molarity and molality, depends on temperature.
3. Prove mathematically that sum of the mole fractions of all the components of a solution is equal to one.
4. Define and name four colligative properties.
5. Define osmosis and give it applications in daily life.

Numerical

1. Calculate Molarity (M) of the following solutions:
   i) 2.0 g of H_2SO_4 /2dm^3 of aqueous solution Ans. (0.01M)
   ii) 0.4 g of NaOH/l00 cm^3 of aqueous solution Ans. (0.1M)
   iii) 0.5 g of Na_2CO_3/250 cm^3 of aqueous solution Ans. (0.02M)
2. Calculate mole fraction of each component in the following solutions.
   i) 2.5 moles of water (H_2O)+1.5 moles of NaCl. Ans.(0.625, 0.375)
   ii) 500 g of water (H_2O)+600 g of glucose ($C_6H_{12}O_6$). Ans. (0.89,0.11)
   iii) 800 g of water (H_2O)+120 g of acetone $(CH_3)_2$ CO. Ans. (0.955, 0.044)
3. A 3.0 cm^3 cylinder contains a mixture of oxygen, nitrogen, and chlorine gases. If their amounts are 4.64×10^{-5}g, 3.5×10^{-4} g and 2.3×10^{-5}g respectively. Calculate the concentration of each in
   i) Parts per million (ppm)

ii) Parts per billion (PPb)

Answers

i) ppm O_2 15.46 N_2 116.6 Cl_2 7.7

ii) ppb O_2 1.546×10^4 N_2 1.16×10^5 Cl_2 7.7×10^3

4. In an experiment 6.4g of an organic compound was dissolved in 25.6g o benzene and 1.0 °C depression of freezing point was observed. Calculate the molecular weight of the substance. (k_f of benzene=5.12)

Ans. 1280 gmol

5. Calculate molecular weight of Iodine (I_2) when a solution containing 1.19 g of I_2 in 35.0 grams of ether shows a raise in boiling point by 0.296 °C (k_b for ether: 2.22)

(Ans. M.Wt of I_2 = 254g mol

Long Questions

1. How a given mixture can be differentiated into a true solution or coarse suspension.
2. Define solubility, what are the factors that affect the solubility of substances?
3. Classify colloids on the basis of their behaviour towards medium and the physical states of matter.
4. Give the statement of Raoult's law. Explain the lowering of vapour pressure of a solution based on this law.
5. Explain the phenomenon of osmosis.
6. Explain the phenomenon of depression of freezing point. How molecular mass of solute can be calculated on the basis of depression in freezing point?
7. Explain the elevation of boiling point, how molecular mass of solute can be calculated on the basis of elevation of boiling point?

Project

Take a potato of four to five centimetres length. Peel off it and cut rod shaped finger chips of equal length. Measure the length of these chips. Place three chips in distilled water and three in concentrated sodium chloride solution. After one hour take out the chips and measure the length of the chips. Record the difference in length of the chips before and after dipping. You will observe that the chips in distilled water will be longer than before and those in the solution of NaCl will be shorter than before. Guess what might be the reason for this observation?