10

SOLUTIONS AND COLLOIDS

This is a 12 days unit

After completing this lesson, you will be able to:

- Define hydrophilic and hydrophobic molecules
- Express solution concentration in terms of mass percent, molality, parts per million, billion and trillion and mole fraction.
- Define the term colligative.
- Define the term water of hydration.
- Describe the role of solvation in the dissolving process.
- Distinguish between the solvation of ionic species and molecular substances.
- List the characteristic of colloids and suspensions that distinguish them from solutions.
- Explain the nature of solutions in liquid phase giving examples of completely miscible, partially miscible and immiscible liquid- liquid solution.
- Describe on a particle basis why a solution has a lower vapour pressure than the pure solvent.
- Explain on a particle basis how the addition of a solute to a pure solvent causes an elevation of the boiling point and depression of the freezing point of the resultant solution.
- Explain concept of solubility and how it applies to solution saturation.
- Explain the effect of temperature on solubility and interpret the solubility graph.

Reading

INTRODUCTION

Most of the substances we encounter more often in daily life are mixtures. For example milk, wood, air, gasoline etc. are mixtures of substances. Such substances have properties different from those of the pure material they contain. Recall from grade IX-X knowledge that any mixture has two defining characteristics, its composition is variable and it retains some properties of its components. In this chapter we focus on two extremely common types of mixtures, solutions and colloids. We will discuss role of intermolecular forces in their formation. We will examine equilibrium nature of solubility and will discuss how temperature and pressure affect it. We have already learned various types of solutions in the grade IX-X. Here we will discuss ways of expressing concentration. We will also discuss some important properties of solutions which do not depend on the nature of solute and daily life application of such properties.

10.1. GENERAL PROPERTIES OF SOLUTIONS

A sample of matter having a fixed composition and uniform properties throughout is called a phase. For example pure sample of water under standard conditions exists as

single liquid phase. The properties of water e.g. density, vapour pressure etc are uniform throughout this liquid phase. When a small quantity of sugar or common salt dissolves in it, the entire sample remains a single phase. The composition and properties of this new phase are different from those of pure water. This is a mixture since it contains two different substances. Such a mixture is said to be homogeneous, because its properties are uniform throughout the liquid.

A homogeneous mixture of two or more pure substances, which has uniform composition throughout is called solution. Do you think, air is a solution?

10.1.1. Hydrophilic and Hydorphobic Molecules

Dissolving solid solutes in water is very common. We dissolve salt in water used to cook vegetables, meat, pulses etc. Particles of solute (atoms, ions, or, molecules) interact with the solvent molecules and become part of the structure of the solvent during the dissolving process. For example, when glucose is dissolved in water, the crystalline structure of glucose is broken down by the water molecules. When totally dissolved, glucose molecules are evenly distributed throughout the water. Oceans contain many dissolved solid substances such as sodium chloride, sodium bromide, magnesium chloride etc.

There are some substances that do not dissolve in water. For instance substances like benzene, cyclohexane, oils etc. do not dissolve in water. This is because these substances have non-polar molecules which do not have any interaction with water molecules. Therefore, molecules can be hydrophilic or hydrophobic in nature.

Hydro mean water, philic means loving. So hydrophilic literally means water loving. **Molecules that are miscible with water are known as hydrophilic molecules**. Such molecules can form hydrogen bond with water molecules. For example, molecules of methanol, acetone, acetic acid etc. are called hydrophilic molecules. Since phobic means disliking, therefore hydrophobic literally mean water disliking. **Molecules that do not dissolve in water are known as hydrophobic molecules**. For example, molecules of organic fats and oils are called hydrophobic molecules.

10.1.2 Nature of Solution in Liquid phase (Like dissolve like) (bond)

Solutions of liquids in liquids may be divided into three classes.

(a) Completely miscible (b) Partially miscible (c) Immiscible.

Completely Miscible Liquids

Liquids which are miscible in all proportions are called completely miscible. For instance water and methanol, acetone and water, benzene and cyclohexane are completely miscible liquid pairs. Substances that dissolve in each other possess similar types of inter-molecular interactions. For example water and methanol are miscible because molecules in the pure liquids and their mixtures form hydrogen bonds. The degree of hydrogen bonding in the solution is same as that in the pure liquids. Benzene and cyclohexane are completely miscible because the

The student should be guided to recall information.

molecules in the pure liquids and their mixtures interact through London dispersion forces. These dispersion forces between benzene and cyclohexane are about the same as those in pure liquids.

Partially Miscible Liquids

Liquids that dissolve in each other to a very small extent are called partially miscible liquids. On mixing such liquids two layers are formed. Each layer is a saturated solution of the other liquid. For example ether and water are partially miscible liquids. The mutual solubilities of these solutions change by temperature changes.

Typical examples of such systems are:-

(i) Phenol-water
(ii) Aniline-water
(iii) Aniline-n-hexane
(iv) Nicotine-water

Phenol-water system

When equal volumes of phenol and water are mixed into each other, two liquid layers are formed. The lower layer consists of a small amount of water dissolved in phenol, while the upper layer consists of small amount of phenol dissolved in water. It is observed that at 25^0C upper layer is 5% solution of phenol in water and the lower layer is 30% water in phenol. This means water is solute in lower layer and phenol is solute in the upper layer. Such solutions are called conjugate solutions. As the temperature increases, the mutual solubility of two liquids increases. Water starts moving from upper to the lower layer and phenol from lower layer to the upper layer. Thus composition of both the layers changes. When the temperature finally reaches 65.9^0C, the composition of both the layers become identical. Each layer contains 34% phenol and 66% water. Above this temperature two solutions merge into one another and two liquids become completely miscible in all proportions. The temperature at which two conjugate solutions merge into one another is called critical solution temperature or upper consolute temperature. For example:

Critical solution temperature for water-aniline system is 167°C with 15% water and that for aniline-hexane is 59.6 °C with 52% aniline. And that for water-phenol is 65.9^0C with 34% phenol.

Completely Immiscible Liquids

What happens when you mix oil in water?

Liquids which do not dissolve into each other in any proportion at any temperature are called completely immiscible. For instance water and benzene, carbon disulphide and water, cyclohexane and water. Benzene, carbon disulphide or cyclohexane are non-polar in nature. The only forces of attraction between their molecules are dispersion forces. In water hydrogen bonds are much stronger than dispersion forces. Since hydrogen bonds cannot be disrupted by the molecules of benzene, carbon disulphide or cyclohexane, these liquids are immiscible with water.

The student should be guided to distinguish between different parts of topics.

10.1.3 Dissolution Process

When a solute is added in a suitable solvent, it dissolves forming a solution. In the formation of a solution three types of interactions are involved. These are solute - solute, solvent-solvent and solute-solvent interactions. A solution forms only when the interactions between solute-solvent molecules are equal to or greater than the interactions between solute-solute and solvent-solvent molecules. This means, the process of dissolution can occur only when the interactions that hold together solute particles are weekend appreciably by the solute-solvent interactions. The speed with which a solute dissolves in a solvent is called dissolving rate. Four factors influence the rate at which substances dissolve. (1) Particle size (2) Temperature (3) Solution concentration (4) Stirring.

When a sugar cube is placed into water, it dissolves slowly than does an equal amount of finely granulated sugar. A sugar cube exposes less surface area to the water molecules than do the tiny particles of granulated sugar. Inner regions of cube dissolves only after the outer layers are in solution. There are fewer unexposed particles when sugar is granulated, so it dissolves faster. Thus, to increase the rate of dissolving, large piece of solid must be grinded into small particles.

Temperature is another factor which changes the rate at which solutes dissolve in a solvent. At higher temperature, increased molecular motion increases the interaction of solute and solvent particles which increases the dissolving rate.

As the concentration of solute in solution increases, the time needed for more solute to dissolve increases. Initially the dissolving rate is maximum. When the first solute is added to the solvent. With each addition of solute, the dissolving rate decreases until no more solute is observed to dissolve. At this stage solution is said to be saturated. Any addition of solute to a saturated solution remain un-dissolved and settle to the bottom of the container.

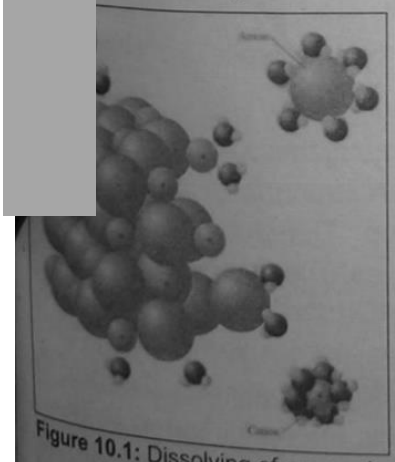


Figure 10.1: Dissolving of ammonium nitrate crystal in water

Stirring a solution increases the rate at which a solid dissolves decreasing the concentration of solute in the immediate region surrounding the solid solute. Stirring also increases the amount of exposed solute surface to the solvent.

10.1.4 Solvation of Ionic and Molecular Substances

Ionic solids are soluble only in solvents having polar molecules. When an ionic compound is dissolved in a polar solvent, it splits up into its ions. For instance, when NH_4NO_3 dissolves in water the resulting solution contains NH_4^+ and NO_3^- ions floating around independently. Solvent molecules surround these ions by directing their negative poles towards

The student should be guided to explain the concepts.

positive ions and their positive poles towards negative ions. This interaction is called as ion-dipole interaction (see figure 10.1). Circles with positive sign represent NH_4^+ ions and circles with negative sign represent NO_3^- ions. Process of solvation of ionic and molecular substances is similar in the sense that in both the cases the process of solvation involves the disruption of solute particles from the crystal lattice. Then these particles are surrounded by solvent molecules. However, the heat of solvation in both the cases is much different since ion-dipole interactions are much stronger than dipole-dipole or dispersion forces, heat of solvation for ionic species is generally larger than those for molecular solvents.

Molecular solids are held together by dispersion forces, dipole-dipole forces and sometimes hydrogen bonds. Such solids dissolve readily in solvents with similar types of intermolecular forces. For example when a polar molecular substance is mixed with a polar solvent. Polar ends of solvent molecules interact with the opposite polar ends of molecules of solute and break attractive forces between them. Finally solvent molecules surround these molecules. This interaction is called dipole-dipole interaction.

The process in which solvent molecules interact and surround solute ions or molecules is known as **solvation**. When water is the solvent, this process is known as **hydration**.

10.1.5 Role of Solvation in the Dissolving Process

Solution making process is accompanied by either the absorption or evolution of heat. Energy change in the formation of a solution depends upon three types of interactions between solute-solute, solvent-solvent and solute-solvent particles. In the formation of solution interactions between solute particles are broken. At the same time solvent molecules also move apart to accommodate the solute particles. Since energy is needed to break interactions. Therefore, both these processes are endothermic. Simultaneously interactions between the particles of solute and solvent are established i.e., solvation occurs. Solvent molecules surround solute particles from all sides. In these interactions energy is released, so it is exothermic. Thus the strength of these two types of interactions determines whether the process of dissolution is endothermic or exothermic.

Thus when a solute dissolves in water, attractive forces among these solute particles must be overcome to break solute particles from the solid lattice. A strong force must bind these separated solute particles to water molecules in the solution. Therefore, heat of solution measures the net energy flow that occurs as a substance dissolves. The energy needed to break solute particles is equal to lattice energy of solute ($\Delta H_{lattice}$). On the other hand the energy released when solute particles bind to the water molecules in the solution is called heat of hydration (ΔH_{hyd}). Heat of hydration increases with increasing charge on the ions and decreases with increasing size of ions. For other solvents this energy is known as heat of solvation. Heat of solution includes both the energy needed to move solvents molecules

The student should be guided to evaluate the information.

apart and energy release when they surround solute particles. Thus, heat of solution $\Delta H_{solutoin}$ is the difference in $\Delta H_{lattice}$ and ΔH_{hyd}.

$$\Delta H_{solutoin} = \Delta H_{lattice} - \Delta H_{hyd}$$

Because of the different attractive forces in both solute and solution, heats of solution vary from significantly exothermic to significantly endothermic. If the crystal binds solute particles more tightly then the solution, solute must absorb energy as it dissolves and we have endothermic solution process. When this $\Delta H_{solution}$ is very large, the solute is unlikely to be soluble. But if this $\Delta H_{solution}$ moderately positive, then solute dissolves. For example, when ammonium nitrate dissolves in water 25.7 KJmole^{-1}energy is absorbed and the flask becomes cold. This is because crystal binds NH_4^+ and NO_4^- ions more tightly than the solution.

$$NH_4NO_3(s) + \times H_2O_{(l)} \longrightarrow NH^+_{4(aq)} + NO^-_{3(aq)} \qquad \Delta H_{solution} = 25.7 \text{ kJmole}^{-1}$$

On the other hand, if solution binds solute particles more tightly than the crystal, energy is released as the solute dissolves and we have an exothermic solution process.

For example, when NaOH dissolves in water 44.5 kJmole^{-1}energy is released and the flask becomes hot

$$NaOH_{(s)} + xH_2O_{(l)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)} \qquad \Delta H_{soln} = -44.5 \text{ kJmole}^{-1}$$

Thus solution process can be endothermic or exothermic depending upon the difference in the lattice energy for solute and heat of solvation for solute particles. Other examples are given below

$$CaCl_{2(s)} + H_2O_{(l)} \longrightarrow Ca^{+2}_{(aq)} + 2Cl^-_{(aq)} \qquad \Delta H_{soln} = -83.0 \text{ kJmole}^{-1}$$

$$MgSO_{4(s)} + \times H_2O_{(l)} \longrightarrow Mg^{+2}_{(aq)} + SO^{-2}_{4(aq)} \qquad \Delta H_{soln} = -91.2 \text{ kJmole}^{-1}$$

The values of standard enthalpies of solution of some ionic solids in water are given in table (10.2)

Table 10.2: Enthalpies of Solutions of some ionic solids.

Substance	Enthalpy of Solution ($kJ\ mole^{-1}$)
LiCl	-37.0
NaCl	+2.98
KCl	+17.2
KI	+20.3
NH_4NO_3	+25.7
$AlCl_3$	-321.0

Daily Life Applications of Heat of Solution

Instant hot and cold packs are in common use today. Cold packs are used for the treatment of injuries and reduction of swelling. Hot packs are used for instant warmth for hikers and skiers and treatment of pulled muscles. These packs are excellent examples of basic science producing a technologically useful product. These packs are based on heat of solutions. These packs contain two separate compartments. One contains water and the other contains a salt,

NH_4NO_3 for cold packs and $CaCl_2$ or $MgSO_4$ for hot packs. When required these packs are kneaded, the wall between the compartments breaks, allowing the salts to mix with water. Heat is absorbed in cold packs and released in hot packs. Gradually these packs attain room temperature. For the amount of heat absorbed or released in these packs, see section 10.1.5.

10.1.6 Water of Hydration

A crystalline substance which has associated with each formula unit a definite number of water molecules is called a hydrate. Such water molecules are called water of crystallization or water of hydration. **The water molecules that combine with compounds as they are crystallized from aqueous solution are called water of crystallization or water of hydration**. For example,

(i) $CuSO_4.5H_2O$ (ii) $BaCl_2.2H_2O$

(iii) $MgSO_4.7H_2O$ (iv) $(COOH)_2.2H_2O$

(v) $Na_2CO_3.10H_2O$

How many moles of water are present in hydrated copper sulphate.

The percentage of water in a hydrate can be determined. For this purpose a known mass of hydrate is heated to expel water completely. Mass of anhydrous solid is determined. Difference in the two masses gives the mass of water present. From these masses, percentage of water in the hydrate is determined. An ion having higher charge density has greater ability to attract water molecules. In $CuSO_4.5H_2O$, the charge density of Cu^{++} ion in greater than that of SO_4 ion. Thus Cu^{++} ion has greater ability to attract water molecules than SO_4^{-2} ion. That is why, out of five water of hydration four are associated with Cu^{++} ion and only one with SO_4^{-2} ion. In hydrates, although new bonds between ions and water molecules are formed but no hydrogen oxygen bond of water is broken.

$$\text{Percent water in a hydrate} = \frac{\text{mass of water in the hydrate}}{\text{mass of hydrate}} \times 100$$

Example 10.1

250g of $CuSO_{4.} \times H_2O$ on heating produced 159.82g $CuSO_4$. Calculate the percent of water in the $CuSO_{4.} \times H_2O$. Also determine the value of x.

Problem solving strategy

1. Find the mass of hydrate by subtracting mass of anhydrite from that of hydrate.
2. Find % water in hydrate using formula.
3. Find the value of x using formula.

Solution

Masss of $CuSO_{4.} \times H_2O$ = 250g

Mass of $CuSO_4$ = 159.82g

Mass of water in the hydrate = 250-159.82
= 90.18g

Percent of water in the hydrate = $\frac{90.18\,g}{250\,g} \times 100$
= 36.07%

X = No. of moles of water = $\frac{\text{Mass of water in hydrate}}{\text{molar mass of water}}$

= $\frac{90.18\,g}{18\,g/mole}$

X = 5.01
X = 5

Do You Know

One of the water's most valuable properties is its ability to dissolve many substances. Without this property, many things on earth would be different. For example the occeans would not be salty. Plants and animals, which depend on moving dissolved substances into and out of their cells, would not exist.

10.1.7 Concept of Solubility

How can you sweeten drinks such as tea?

When a small amount of solute is mixed in excess of water, it dissolves to form a solution. If a little more solute is added it will also dissolve. Such a solution which can dissolve more solute under existing conditions is called **unsaturated solution**. As more and more solute is added, a stage is reached where further dissolving seems to cease leaving solid solute at the bottom of the container. Such a solution which contains maximum amount of dissolved solute under existing conditions is called **saturated solution**. The maximum amount of solute that dissolves in the given quantity of a solvent under given conditions is called **solubility**. Solubility is commonly expressed in number of grams of solute per cm^3 of solution or in terms of number of moles of solute per dm^3 of solution.

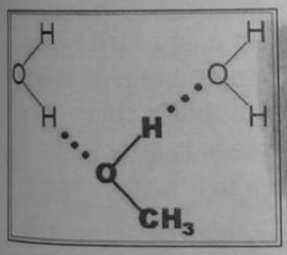


Fig 10.2: H-bonding between water and methanol

On molecular level a saturated solution is in equilibrium with excess of solute. At equilibrium the rate of solute going into the solution equals the rate of solute returning from the solution.

$$\text{Undissolved solute}_{(s)} \rightleftharpoons \text{Dissolved solute}_{(aq)}$$

In a saturated solution dissolving of solute infact continues. At the same time many dissolved particles moving freely in solution strike the solid solute. Such particles are recaptured by the solid solute present at the bottom of the container. Such particles are thus re-crystallized. A stage is reached when a dynamic equilibrium is established betwseen dissolved and undissolved solute.

A general rule useful for predicting solubilities is "like dissolves like." Substances which have similar structures and intermolecular forces tend to be soluble. While substances which have dissimilar structures and intermolecular forces are insoluble.

The student must be guided to argue about the information.

For instance, water and methanol have similar structures and have hydrogen-bonding between their molecules. They can form hydrogen bonds with each other, when they are mixed. See Fig 10.2. Thus water and methanol are miscible.

Molecules of benzene, C_6H_6 and carbon tetrachloride CCl_4 are non-polar and have London dispersion forces between their molecules. When these two liquids are mixed, their molecules attract each other with London dispersion forces, thus these are miscible liquids. Think what will happen when CCl_4 or Benzene is mixed with water?

Ionic solids have a crystal lattice structure composed of oppositely charged ions. When an ionic solid e.g. NaCl is placed in water which is a polar solvent. These ions are attracted by polar molecules. Water molecules break the crystal lattice of Na^+Cl^- and then surround the resulting Na^+ and Cl^- ions. These ions are called hydrated ions.

But when NaCl crystals are placed in CCl_4 or C_6H_6. Non-polar molecules of these liquids are unable to attract ions in Na^+Cl^- and cannot break apart the crystal lattice. Thus NaCl is insoluble in these solvents. Think, what will happen when solid NH_4NO_3, an ionic solid is placed in water?

Self Check Exercise 10.1

Which solvent, liquid ammonia, NH_3 or benzene, C_6H_6 is more likely to dissolve each of the following solute. Give reasons?

(a) AgCl (b) Wax (c) H_2O (d) NH_4OH

Effect of Temperature on Solubility

Generally, an increase in temperature increases solubility of a solid in a liquid. At higher temperature greater masses of solutes dissolve in a fixed mass of water than at lower temperature. A curve drawn between solubility and temperature is called **solubility curve** (see figure 10.3). The solubilites of solid solutes generally increase with the increase in temperature. However this is not always the case. The solubilties of some solids e.g. Na_2SO_4 and $Ce_2(SO_4)_3$ decrease with increasing temperature. Some substances like NaCl have relatively constant solubility with increasing temperature.

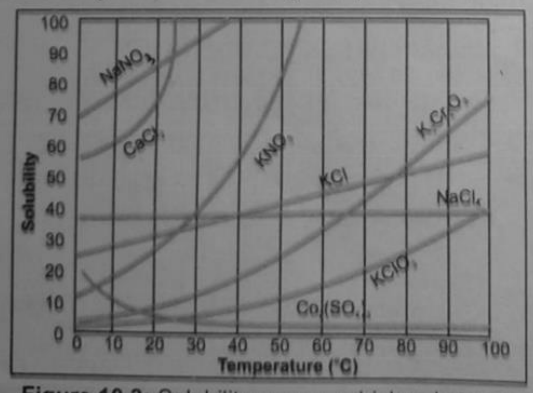


Figure 10.3: Solubility curves, which substance has higher solubility at room temperature?

The variation of solubility with temperature can be predicted from the enthalpy of solution. If ΔH°_{Soln} in negative, the solubility of solute decreases as the temperature increases. But if ΔH°_{Soln} is positive, the solubility increases with temperature.

Solubility of KNO_3 rapidly increases with increase in temperature, this is because it has large positive value of ΔH°_{Soln} (+36 kJ/mole) the solubility of NaCl does not vary much with temperature because of relatively small magnitude of ΔH°_{Soln} (+3 kJ/mole).

It is generally observed that the solubility of a gas decreases with increasing temperature The dissolving of gases in liquids can be understood in terms of two processes (a) th condensation of the gas which is exothermic and (b) the creation of holes in the liquid t accommodate the condensed gas molecules, it is endothermic. Because of the open structur of water, little work is required to accommodate gas molecules, this means dissolving proces is exothermic. Thus solubility of gases in water decrease with temperature.

Self Check Exercise 10.2

Consult solubility curve shown in figure 10.3 and explain the following

ΔHsolution of $Ce_2(SO_4)_3$ and that of $KClO_3$ is positive or negative.

ΔHsolution of $CaCl_2$ is greater or lesser than that of $NaNO_3$ at 20°C.

Indicate the temperature at which solubilities of NaCl and KCl are same.

Which of the following salts has rapid increase in solubility with the increase in temperature:

(a) $CaCl_2$ (b) $K_2Cr_2O_7$

Effect of Pressure on Solubility

Pressure has negligible effect on the solubility of solids or liquids. On the other hand it does significantly increase the solubility of a gas. At a given temperature, the solubility of gases which do not react with the solvent are directly proportional to the partial pressures of the gases above the solution. This relationship is known as **Henry's Law**. It is obeyed most accurately by dilute solutions of gases: Henry's law can be understood from Fig. 10.4.

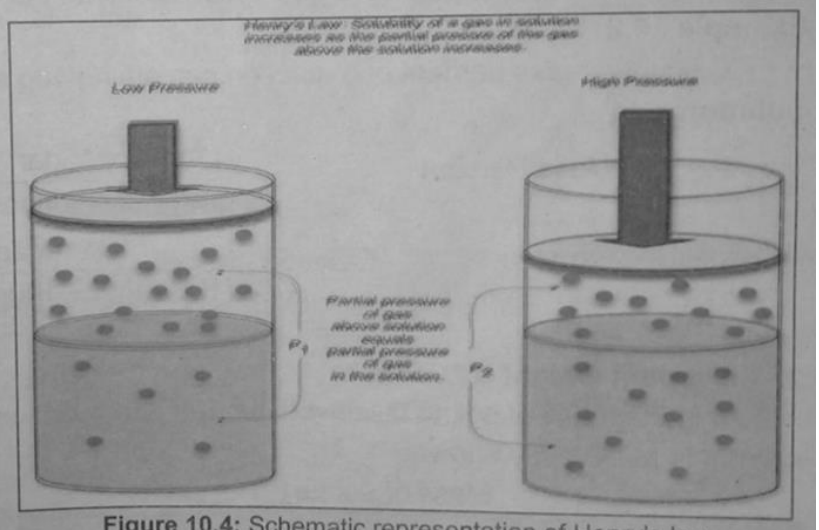


Figure 10.4: Schematic representation of Henry's law

Equilibrium between the gas above a liquid and the dissolved gas within the liquid is reached when the rates of evaporation and condensation of gas molecules become equal. (Fig.10.4). When the pressure is suddenly increased by pushing the piston, the number of molecules per unit volume increase in the gaseous state, this causes an increase in the rate at which the gas enter the solution, so the concentration of the dissolved gas increases. The greater gas concentration in the solution causes an increase in the rate of evaporation, until a new equilibrium is reached.

Carbonated drinks are bottled at high pressure of carbon dioxide. When the cap is removed, the effervescences results from the fact that the partial pressure of carbon dioxide in the atmosphere is much less than used in the bottling process. As a result, the equilibrium quickly shifts to one of the lower gas solubility. This causes effervescence.

10.2. CONCENTRATION UNITS

We sometimes speak of dilute or concentrated solutions, but these terms have no precise meaning. A dilute solution is one which contains relatively small amount of solute per unit volume of solution. Whereas a concentrated solution is one which contains relatively greater amount of solute per unit volume of solution. We need to define solution composition more precisely in order to perform calculations. The amount of solute, solvent and solutions may be expressed by volume, mass or number of moles. Therefore, there are various types of concentration units. Some of these are given below.

1. Mass Percent

It is defined as the mass of solute present in 100g of solution. It is also referred as percent weight / weight.

$$\text{Mass Percent} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100$$

15% NaCl solution means 15g of NaCl dissolved per 100 g of solution or 85g of water.

Example 10.2

Calculate mass percent of a solution containing 10g sugar dissolved in 100g of water.

Solution

$$\text{Mass Percent} = \frac{\text{grams of sugar}}{\text{grams of solution}} \times 100$$

$$= \frac{10\text{g sugar}}{(10\text{g sugar} + 100\text{g water})} \times 100$$

$$= 9.09\%$$

2. Percent weight by volume

It is the mass of solute dissolved per 100 part by volume of solution. In this case volume of solvent is not exactly known.

$$\text{Percent w/v} = \frac{\text{Mass of solute (g)}}{\text{vol. of solution (cm}^3\text{)}} \times 100$$

For example, a 10% (w/v) NaCl solution contains 10g NaCl dissolved in 100 cm^3 of solution.

3. Percent volume by volume

It is the volume of solute dissolved per 100 parts by volume of solution. In such solutions volumes of solute and solvent may not be necessarily equal to the volume of solution

$$\text{Percent v/v} = \frac{\text{cm}^3 \text{ of solute}}{\text{cm}^3 \text{ of solution}} \times 100$$

For example a 5% (v/v) ethanol solution means 5cm^3 ethanol dissolved per 100 cm^3 of solution.

4. Percent volume by weight

The volume of a solute dissolved per 100g of solution is called percent volume by weight. In this case total volume of solution is not known.

The student should be guided to recall information

For example. a 5% (v/w) Ethanol contains 5 cm^3 ethanol per 100g of solution.

$$\% \text{ v/w} = \frac{\text{volume of solute } (cm^3)}{\text{Mass of solution (g)}} \times 100$$

5. Molarity (M)

It is defined as the number of moles of solute dissolved per dm^3 of solution. Note that the reference volume is that of solution and not the pure solvent. If 0.5 mole of NaOH (20g) is dissolved in enough water to make one dm^3 of solution, 0.5 molar or 0.5 M NaOH solution is obtained.

depend upon temperature because when temperature increases, Volume also increases.

$$M = \frac{\text{moles of solute}}{dm^3 \text{ of solution}}$$

$$M = \frac{\text{grams of solute}}{\text{molar mass of solute} \times dm^3 \text{ of solution}}$$

Example 10.3

What is the molarity of a solution of 0.25g of $NaHCO_3$ in 100 cm^3 of solution.

Mass of $NaHCO_3$ = 0.25g

Molar mass of $NaHCO_3$ = 23+1 + 12 + 16 × 3
= 84g / mole

Volume of solution = 100 cm^3
= 100/1000 = 0.1 dm^3

$$\text{Molarity} = \frac{\text{grams of } NaHCO_3}{\text{molar mass of } NaHCO_3 \times dm^3 \text{ of solution}}$$

$$= \frac{0.25g}{84\,g/mole \times 0.1\,dm^3}$$

$$= 0.0298 \text{ M}$$

6. Molality (m)

It is defined as the number of moles of solute dissolved per kilogram of solvent. For example when 58.5g NaCl (1 mole) is dissolved in one kilogram of water, the resulting, solution would be one molal or 1m NaCl solution.

$$m = \frac{\text{moles of solute}}{\text{kg. of solvent}}$$

Example 10.4

Ethanol is an excellent organic solvent. It is used to prepare tinctures and in the extraction of medicinal compounds from plants. For this purpose either pure ethanol or its aqueous solutions are used. A solution is prepared by mixing 1.00g of ethanol (C_2H_5OH) with 100g of water. Calculate molality of this solution.

Solution

Mass of ethanol = 1.00g

Molar mass of ethanol = 12 × 2 + 1 × 6 + 16
= 46g/mole

Moles of ethanol = $\frac{1.00g}{46g/mole}$

= 2.17×10^{-2} mole

Mass of water = 100g

= 100/1000 = 0.1 kg

Molality of solution = moles of C_2H_5OH / kg of water

= 2.17×10^{-2} mole / 0.1 kg

= 0.217 m.

Notice that molality is independent of temperature whereas molarity depends on temperature. This is because molarity is based upon volume of solution. As the temperature increases, the amount of solute remains constant but the volume of solution increases slightly, thus molarity decreases slightly.

Self Check Exercise 10.3

Any fluids infused intravenously into an individual must be isotonic with the blood cells and the blood plasma. Such infusions are either 5% dextrose (glucose) or 0.9% normal saline. The first solution is composed of 5.0 gram of glucose per 100 cm^3 of solution and the other of 0.9 gram NaCl per 100 cm^3 of solution. Calculate the molarity of these intravenous solutions.

(**Ans**: 0.28M, 0.15M)

7. Mole Fraction (X)

It is sometimes necessary to use a concentration unit in which the amount of all the solution components are described on a number basis. This can be done through mole fraction concentration unit.

It is defined as the ratio of the number of moles of a given component to the total number of moles of solution. Suppose a solution contains n_A and n_B moles oaf two components A and B. Mole fraction of each component is given by

Mole fraction of component A = $X_A = \frac{n_A}{n_A + n_B}$

Mole fraction of component B = $X_B = \frac{n_B}{n_A + n_B}$

Example 10.5

An aqueous solution containing 100g ethanol per dm^3 of solution has a density of 0.984 cm^{-3}. Calculate mole fraction of each component of solution.

Solution

Volume of solution = $1dm^3$ = $1000\ cm^3$

Mass of ethanol = 100g

Mass of solution = Density × Vol of solution of solution

= 0.984 g. $cm^{-3}\times$ 1000 cm^3

= 984g

Mass of water (solvent) = 984g – 100g

$= 884g$

No of moles of $H_2O = \frac{884g}{18g/mole}$

$= 49.1$ mole

No. of moles of ethanol $= \frac{100g}{46\,g/mole}$

$= 2.17$ mole

Total No. of moles in solution $= 49.1 + 2.17$

$= 51.27$ mole

$X_{H_2O} = \frac{49.1 mole}{51.27 mole}$

$= 0.958$

$X_{ethanol} = \frac{2.17 mole}{51.27 mole}$

$= 0.042$

8. Parts per million (ppm)

It is defined as the number of parts by weight (or volume) of a solute per million parts by weight (or volume) of the solution.

$$ppm = \frac{\text{Mass or volume of solute}}{\text{Mass or volume of solution}} \times 10^6$$

9. Parts per billion (ppb)

It is defined as the number of parts by weight (or volume) of a solute per billion parts by weight (or volume) of the solution.

$$ppb = \frac{\text{Mass or volume of solute}}{\text{Mass or volume of solution}} \times 10^9$$

10. Parts per trillion (ppt)

It is defined as the number of parts by weight (or volume) of a solute per trillion parts by weight (or volume) of the solution.

$$ppt = \frac{\text{Mass or volume of solute}}{\text{Mass or volume of solution}} \times 10^{12}$$

Rare components of a solution are expressed by ppm, ppb and ppt. A pollutant concentration of 1 ppm in the atmosphere means that out of one million molecules of air there is one molecule of the pollutant, whereas 1 ppb concentration of it means that out of one billion molecules of air there is one molecule of that pollutant.

Example 10.6

An atmospheric chemist reports that one dm^3 of air in an urban area contained 3.5×10^{-4} cm^3 of CO. What was the concentration of CO in ppm?

$$\text{ppm} = \frac{\text{Volume of CO } (cm^3)}{\text{Volume of air } (cm^3)} \times 10^6$$

$$= \frac{3.5 \times 10^{-4}\ cm^3}{10^3\ cm^3} \times 10^6$$

$$= 0.35 \text{ ppm.}$$

Example 10.7

If the concentration of ozone in atmosphere reaches 0.5 ppb. What mass of ozone would be present per kg of air.

Solution

$$\text{ppb} = \frac{\text{Mass of ozone (g)}}{\text{Mass of air (g)}} \times 10^9$$

$$0.5 = \frac{\text{Mass of ozone}}{1000\ g} \times 10^9$$

$$\text{Mass of ozone} = \frac{0.5 \times 1000}{10^9}$$

$$= 0.5 \times 10^3 \times 10^{-9}$$

$$= 5 \times 10^{-7}\ g$$

Self Check Exercise 10.4:

1. Calculate the ppm by mass of calcium in a 2.5g tablet that contains 500mg calcium. (**Ans:** 0.2×10^6 ppm)
2. Waste water from a cement factory contains 22mg Ca^{+2} ions and 0.006g Mg^{+2} ions per kg of solution. Calculate the concentration of both these ions in ppm. (Ans: 22ppm, 6ppm)
3. Gold occurs in sea water at an average concentration of 1.09×10^{-2} ppb. How many kg of water must be processed to recover 1gm of Gold. (**Ans**: 9.17×10^7 kg)

10.2.1 Inter Conversion of Various Concentration Units of Solution.

Sometimes we get prepared solutions whose molarity is given, but we need its w/w percentage or molality. For this purpose, we need to convert one unit of concentration into other. How it can be done? It is explained by the following examples.

Example 10.8

Calculate the molality of 15% (w/w) glucose ($C_6H_{12}O_6$) solution.

Solution

Since solution of glucose is 15% w/w

Mass of glucose = 15g

Mass of solution = 100g

The student should be guided to interpret information.

Mass of water = 100g – 15g
= 85g.
= 85/1000 = 0.085 kg.

Molar mass of glucose = $12 \times 6 + 1 \times 12 + 16 \times 6$
= 180 g / mole

$$\text{Molality} = \frac{\text{Mass of solute}_{(g)}}{\text{Molar mass of solute g / mole} \times \text{Mass of solvent(kg)}}$$

$$\text{Molality} = \frac{15\ g}{180\ g/mole \times 0.085 kg}$$

Molality = 0.98 moles per kg.

Molality = 0.98 m

Example 10.9

Sulphuric acid is known as king of chemicals it is used in the manufacture of many chemicals, drugs, dies, plastics, paints, disinfectants, explosives, synthetic fibres etc. It is prepared commercially by contact process and is normally 98% by weight. If its density is 1.84g cm^{-3}, what is its molarity.

Solution

Molar Mass of H_2SO_4 = $1\times2+32\times1+16\times4$
= 98g/mole

Percentage of H_2SO_4 solution = 98%

Density of H_2SO_4 solution = 1.84g cm^{-3}

1 cm^3 of H_2SO_4 solution contains 1.84g H_2SO_4 solution

1000 cm^3g H_2SO_4 solution would have = 1.84×1000
= 1840g H_2SO_4 solution

Of this total mass 98% is H_2SO_4

Mass of H_2SO_4 in solution = $1840 \times 98 / 100$
= 1803.2g

Moles of H_2SO_4 in solution = 1803.2g / 98 g/mole
= 18.4 moles

Since solution has 18.4 moles of H_2SO_4 per 1000 cm^3 or 1dm^3 of solution, the molarity of this solution is 18.4 M.

Example 10.10

Commercial HCl is 12 molar (density = 1.17 g cm^{-3}). Calculate the mass percent of HCl in this solution.

The student should be guided to operate information in practical uses.

Solution

Molarity of HCl = 12 M

As d = m / v

m = d × v

Thus mass of HCl solution present in its 1000 cm^3 (1dm^3) would be

1.17 g / cm^3 x 1000 cm^3 = 1170g

Molar mass of HCl = 1 + 35.5

= 36.5g/mole

As HCl solution is 12M,

Mass of HCl in solution = 12 × 36.5

= 438g

Mass percent of HCl = $\frac{438g \times 100}{1170g}$

= 37.44%

Often more concentrated solutions are diluted to obtain less concentrated solutions. Such dilutions are done after calculating the amount of water that is mixed with the more concentrated solution to give the desired lower concentration.

Example 10.11

Sodium hydroxide solutions are used to neutralize acids, to treat cellulose in the preparation of rayon and to remove potato peals in commercial use. 250cm^3 of 2M NaOH is mixed with 250cm^3 of water. Calculate the molarity of resulting solution.

Solution

250 cm^3 of 2M NaOH contain = $\frac{250\ cm^3 \times 2}{1000 cm^3}$

= 0.5 moles of NaOH

Total volume of solution = 250 cm^3 + 250 cm^3

= 500 cm^3

= 0.5 dm^3

Molarity of resulting solution = 0.5 moles / 0.5 dm^3

= 1M

Self Check Exercise 10.5

1. Urea (NH_2CONH_2) is a white solid produced commercially as a fertilizer and starting material for plastics. It is a non-volatile and non-ionizable compound. What is the mole fraction of Urea in an aqueous solution that is 0.25m Urea. (**Ans**: 4.49×10^{-3})
2. Potassium hydroxide (KOH) is used in the manufacture of liquid soaps, in paints and varnish removers. An experimenter needs 25cm^3 of 0.015M KOH. What mass of KOH he will dissolve to make this solution. (**Ans**: 0.021 g)

10.3 RAOULT'S LAW

A French chemist Raoult discovered a quantitative relationship between vapour pressure of solution and concentration of solution. This relationship is known as Raoult's law.

It states that when a nonvolatile solute is dissolved in a solvent, the vapour pressure of solution (P) is directly proportional to the mole fraction of solvent (X_1)

$P \propto X_1$

$P = P^o X_1$ ---------- (1) where P^o is constant of proportionality and is the vapour pressure of the pure solvent. A more common form of Raoult's law is obtained by simple substitution. In a binary solution $X_1 + X_2 = 1$

And $X_1 = 1 - X_2$ Where X_2 is mole fraction of solute.

Substituting value of X_1 in equation (1), we get

$P = P^o (1 - X_2)$

$P = P^o - P^o X_2)$

$P - P^o = - P^o X_2$

$P^o - P = P^o X_2$ ---------- (2)

$\Delta P = P^o X_2$

Where $P^o - P$ = lowering in vapour of solvent (ΔP).

Therefore Raoult's law can also be stated as the lowering in vapour pressure is directly proportional to the mole fraction of solute. Rearranging equation (2) we get another form of Raoult's law.

$$\frac{P^o - P}{P^o} = x_2$$

$$\frac{\Delta P}{P^o} = x_2$$

$\frac{P^o - P}{P^o}$ or $\frac{\Delta P}{P^o}$ is relative lowering in vapour pressure of the solvent. Raoult's law can also be stated as the relative lowering in vapour pressure is equal to the mole fraction of solute. Lowering in vapour pressure depends on temperature whereas relative lowering in vapour pressure is independent of temperature.

$P_t = P_A + P_B$

Roult's law for ideal solution of two volatile components (liquids) $P_t = P_A^o x_A + P_B^o x_B$

Raoult's Law also governs the vapour pressure of mixtures of two or more volatile components. The vapour pressure of each component is determined by its mole fraction and the vapour pressure of the pure component. The total vapour pressure always lies between the vapour pressures of the pure components, and is determined by the mixture composition

Example 10.12

What are the partial pressures of benzene and toluene in a solution in which the mole fraction of benzene is 0.6? What is the total vapour pressure? The vapor pressure of pure benzene is 95.1 mm Hg and the vapour pressure of pure toluene is 28.4 mm Hg at 25°C.

The student should be guided to choose useful information to solve various problems.

Problem Solving Strategy

1. Remember that the sun of mole fractions of all the components of solution is unity.
2. Mole fraction of benzene is given find the mole fraction of toluene on the besis of point 1.
3. Find the partial pressure of each component of solution using formula.
4. Find total pressure by adding the two partial pressures.

Solution

If $x_{benzene}$ = 0.6, then $x_{toluene}$ = 0.4 because 1 - 0.6 = 0.4.

Now that we know the mole fractions and vapor pressures.

$P_{benzene} = x_{benzene}P^o_{benzene}$ = (0.6) (95.1 mm Hg) = 57.1 mm Hg $P_{toluene} = x_{toluene}P^o_{toluene}$ = (0.4)(28.4 mm Hg) = 11.4 mm Hg

The total vapor pressure is simply the sum of the partial pressures:

$P_{total} = P_{benzene} + P_{tuolene}$ = 57.1 mm Hg + 11.4 mm Hg = 68.5 mm Hg

10.3.1 Causes of Lowering in Vapour Pressure

Every liquid has a definite vapour pressure at a particular temperature (see Chap 5). In a pure liquid, all the surface particles are that of the liquid. But in a solution containing a nonvolatile solute, both solute and solvent particles occur on the surface. The solute particles decrease the number of solvent surface particles, (see figure 10.5). Which beaker contains less numbers of solvent particles on the surface? In which beeker solvent will have higher vapour pressure?

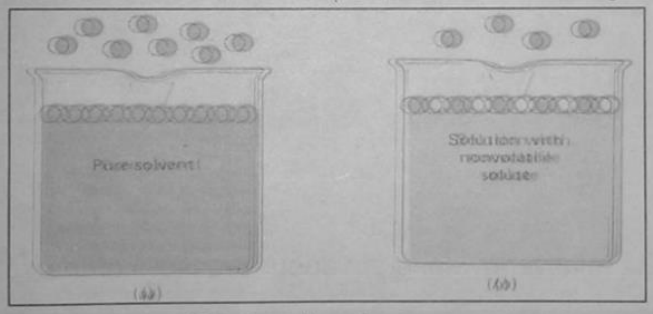


Figure 10.5: lowering of vapour pressure

Decreasing the number of surface solvent particles decreases the rate of evaporation of the solvent, which decreases the vapour pressure. Thus the vapour pressure of a solution containing a nonvolatile solute is always less than the vapour pressure of the pure solvent.

10.4 COLLIGATIVE PROPERTIES OF SOLUTIONS CONTAINING NON-ELECTROLYTE SOLUTES

Have you ever seen salting roads in Muree during snowing? Have you ever added salt to water when cooking pasta? If so, you have already been exposed to a group of four physical phenomena called colligative properties.

The student should be guided to differentiate between parts of topic.

In the preceding sections we saw how the presence of solute always lowers the vapour pressures of the solutions. Since a lower vapour pressure implies a higher boiling point, we can also conclude that the presence of nonvolatile solute always raises the boiling point of solvent. It also decreases the freezing point of the solvent. Thus certain properties of solutions depend only on number of solute particles and not on their nature. These properties are called as **colligative properties**. Colligative properties are properties of solutions that depend upon the ratio of the number of solute particles to the number of solvent molecules in a solution, and not on the type of chemical species present. The word colligative means "to collect". Four such related properties are:

(i) Vapour pressure lowering
(ii) Boiling point elevation
(iii) Freezing point depression
(iv) Osmotic pressure.

The study of colligative properties has provided methods of molecular weight determination.

10.4.1 Lowering of Vapour Pressure

When a non-volatile solute is dissolved in a solvent, the escaping tendency of solvent molecules from the surface of the solution decreases. Thus its vapour pressure is lowered (see section 10.3.1).

According to the Raoult's law relative lowering of vapour pressure is equal to mole fraction of solute.

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

If n_1 and n_2 are the number of moles of the solvent and solute respectively, then

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

So $$\frac{\Delta P}{P^o} = \frac{n_2}{n_1 + n_2}$$

Since for dilute solutions $n_2 << n_1$, n_2 can be ignored in the denominator

$$\therefore \frac{\Delta P}{P^o} = \frac{n_2}{n_1}$$

If W_1 and W_2 are masses of solvent and solute while M_1 and M_2 are their molecular masses respectively, then

$$n_1 = \frac{W_1}{M_1} \quad \text{and} \quad n_2 = \frac{W_2}{M_2}$$

So $$\frac{\Delta P}{P^o} = \frac{W_2/M_2}{W_1/M_1}$$

$$\frac{\Delta P}{P^o} = \frac{W_2 \times M_1}{M_2 \times W_1}$$

$$\text{Or} \quad M_2 = \frac{W_2 \times M_1 \times P^o}{W_1 \times \Delta P}$$

Thus molecular mass of a non-volatile solute can be calculated from this equation.

Example 10.12

When 106.3g of an organic compound M dissolved in 863.5g of benzene, the vapour pressure of benzene is lowered from 98.6 to 86.7torr. Calculate the molecular mass of M.

Solution

$P^o = 98.6$ torr

$P = 86.7$ torr

$\Delta P = P^o - P$

$= 98.6 - 86.7$

$= 11.9$ torr

$W_2 = 106.3$g

$W_1 = 863.5$g

$M_1 = 78$g/mole

$$M_2 = \frac{W_2 \times M_1 \times P^o}{W_1 \times DP}$$

$$= \frac{106.3\text{g} \times 78\text{g/mole} \times 98.6\text{torr}}{863.5\text{g} \times 11.9\text{torr}}$$

$= 79.56$ g/mole

10.4.1 Causes of Boiling Point Elevation and Freezing Point Depression

A liquid boils at a temperature where its vapour pressure becomes equal to the atmospheric pressure. On the other hand, the freezing point of a substance is the temperature at which liquid and solid phases of a substance co-exist in equilibrium. At this temperature liquid and solid phases of the substance have the same vapour pressure. Solute particles effect both the boiling point and freezing point of a pure solvent. A pure liquid has only one type of particle. But when a non-volatile and non-electrolyte solute is added to a solvent, its vapour pressure is decreased (see section 10.3.1). This is because in solution both solute and solvent particles occur on the surface. Thus solute particles decrease the number of solvent surface particles. This decreases the rate of evaporation of solvent, which decreases the vapour pressure. Therefore, a solution must be heated to a higher temperature than the boiling point of pure solvent to equalize its

The student should be guided to appraise the significance of the information.

vapour pressure to the atmospheric pressure. Thus addition of solute to a pure solvent causes an elevation of the boiling point of solution.

Similarly, decrease in vapour pressure of a pure solvent on the addition of a solute also affects the freezing point of the solution. The solution will freeze at a temperature at which vapour pressure of both solution and solid solvent are the same. This means solution will freeze at a lower temperature than that of the pure solvent. Thus addition of a non-volatile solute also causes a decrease or depression in freezing point of solution.

10.4.2 Quantiative Aspects of Boiling Point Elevation

Experiments have shown that when one mole of non-electrolyte and nonvolatile solute is dissolved in 1 kg of water boiling point is raised 0.52°C. This value is known as the molal boiling point elevation constant (ebullioscopic constant) for water. Figure 10.6 shows how the vapour pressure influences the boiling point. Curve AB represents variation in vapour pressure of pure solvent with temperature. The solvent boils at temperature T_1 when its vapour pressure becomes equal to the external pressure P^o. The curve CD represents variation in vapour pressure of solution with temperature. This curve must lie below that of pure solvent. This is because the vapour pressure of solution at all temperatures is lower than that of the pure solvent. The solution will boil at higher temperature T_2 to equalize its vapour pressure to external pressure P^o.

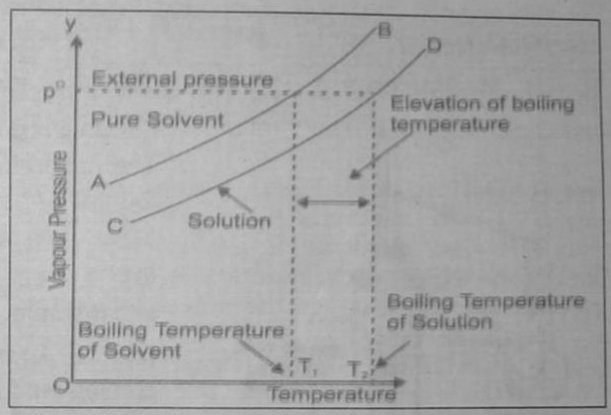


Figure 10.6: Elevation of boiling point – temperature curve

The difference of two boiling points gives the elevation of boiling point ($\triangle T_b$)

$$\triangle T_b = T_1 - T_2$$

The magnitude of the boiling point elevation is directly proportional to the molality of solution.

$$\triangle T_b \times m$$

$$\triangle T_b = \triangle K_b m \ ) \ldots\ldots\ldots (1)$$

The constant K_b is called molal boiling point elevation constant or ebullioscopic constant. For a 1 molal solution.

$$\triangle T_b \times K_b$$

Thus elevation of boiling point when 1 mole of non-volatile, non-electrolyte solute is dissolved in 1 kg of solvent is called molal boiling point elevation constant. The value of K_b depends only on the nature of the solvent and is independent of nature of solute used. The nature of solute does not effect boiling point of elevation as long as the solute does not ionize.

The student should be guided to use information to solve different data

The molality m of the solution containing W_{2g} of solute of molecular mass M_2 dissolved in W_{1g} of the solvent is given by

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

Substituting the value of m in eq(1) we have

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

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

This equation is used to determine the molecular mass of solute.

Self Check Exercise 10.6

Glycerol ($C_3H_8O_3$) is a syrupy, sweat tasting liquid used in cosmetics and candy. It is a non-volatile and non-electrolyte compound. What is the freezing point of an aqueous solution that is 0.25m glycerol? **(Ans:** 0.465°C)

10.4.3 Measurement of Boiling Point Elevation:

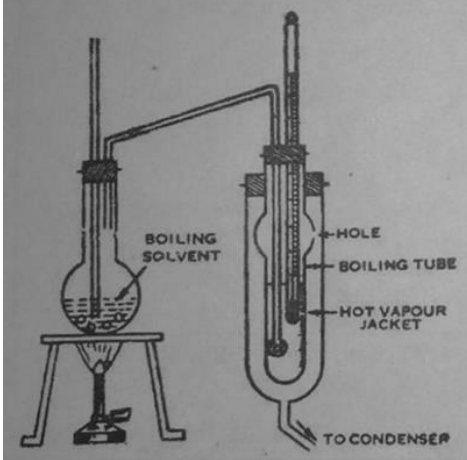


Figure10.7: Landsbeger's apparatus for measurement of elevation of boiling points.

Lands Berger's Method

The apparatus used is shown in Fig 10.7. It consists of four major parts.

(a) A graduated inner tube with a hole in its side.

(b) A boiling flask which sends the solvent vapours into the inner tube though a rosehead.

(c) An outer tube, which receives solvent vapours coming out from the side hole of the inner tube.

(d) A Beckmann thermometer which can read upto 0.01K

Lands Berger's method can be understood by the following activity.

ACTIVIY:

Cholesterol is an important compound in our body. Its excess has been implicated as a cause of heart disease. Take 100 cm^3 of pure benzene (solvent) in the inner tube. Boil benzene in the boiling flask and pass its vapours through the benzene in the inner flask. These vapours will boil benzene in the inner tube by its latent heat of condensation. Record the temperature at which benzene is boiled in the inner tube. You will observe that benzene will boil at 5.5°C. Now stop the supply of vapours temporarily. Drop 4.5g of cholesterol in the inner tube. Pass vapours of the benzene from boiling flask again to boil the solution. Record boiling point of the solution. Thermometer will show 6.07 °C. Now stop the supply of benzene vapours. Remove thermometer and rose head from the inner tube and note the volume of solution. It will be 121.5 cm^3. Determine density of this solution. Its density will be 0.897g cm^{-3}. From the volume and density

determine the mass of solution. From the mass of solution determine the mass of benzene. From this data calculate the molecular mass of Cholesterol.

Solution:

Mass of Cholesterol = $W_2 = 4.5$ g
Mass of Benzene = $W_1 = ?$
Boiling Point of Benzene = $T_1 = 5.5^oC$
Boiling Point of Solution = $T_2 = 6.07^oC$
K_b for Benzene = 5.12
Volume of solution = $V = 121.5\ cm^3$
Density of solution = $d = 0.897\ g\ cm^{-3}$
Molecular mass of Cholesterol = $M_2 = ?$

As,

$$d = \frac{m}{V}$$

$$m = dxV$$

$\therefore$ Mass of Solution = 0.897 x 121.5
= 108.99g

Mass of Benzene (W_1) = Mass of solution – Mass of Cholesterol

W_1 = 108.99- 4.5
= 104.486g

As,

$$\triangle T_b = T_2 - T_1$$

$\therefore$ $\triangle T_b = 6.07 - 5.5$
= 0.57°C

Now

$$M_2 = \frac{K_b \times W_2 \times 1000}{\Delta Tb \times W_1}$$

$$= \frac{4.5 \times 5.12 \times 1000}{0.57 \times 104.486}$$

$= 386.86g\ mole^{-1}$

Self Check Exercise 10.7

Calculate the boiling point of a solution containing 12.5g of benzoic acid, $C_7H_6O_2$ in 110g of benzene. Boiling point and K_b of benzene are 80.1°C and 2.53 respectively.

(**Ans:** 82.637°C)

10.4.4 Quantitative Aspects of Freezing Point Depression

The difference in the freezing points of pure solvent and solution is called the depression of freezing point.

Figure 10.8 shows how the vapour pressure influences the freezing point. Curve ABC is for the pure solvent. The solvent freezes at temperature T_1 corresponding to the point B when its vapour pressure is P^o. The portion of the cure BC is for the solid solvent. Greater slope of curve BC indicates a rapid change of vapour pressure with temperature.

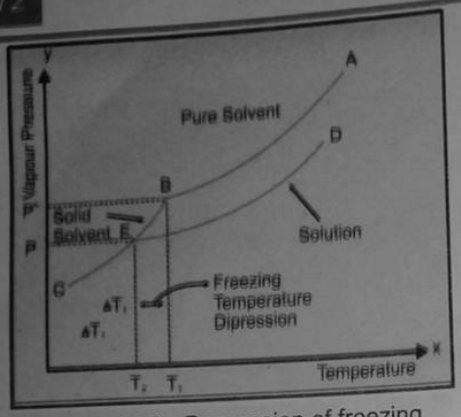


Figure 10.8: Depression of freezing point – temperature curve

The curve DEC is for the solution. It meets curve BC at point E which is freezing point of solution T_2 when its vapour pressure is P which is lower than P^o. This is because vapour pressure of solution is always less than that of the pure solvent.

The Difference between the two freezing points gives the depression of the freezing point ΔT_f.

$$\Delta T_f = T_1 - T_2$$

Experimentally $\Delta T_f \propto m$

$$\Delta T_f = K_f m \quad \text{---------- (1)}$$

Where K_f is constant called the molal freezing point constant or the cryoscopic constant. For I molal solution.

$$\Delta T_f = K_f$$

Thus the depression of freezing point when I mole of non-volatile, non-electrolyte solute is dissolved in 1 kg of solvent is called molal freezing point depression constant.

The value of K_f depends upon the nature of the solvents and is independent of the solute used. The nature of solute does not affect the freezing point depression as long as the solute doesn't ionize.

The molality m of the solution containing W_{2g} of solute of molecular mass M_2 dissolved in W_{1g} of the solvent is given by

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

Substituting the value of m in aq (1)
We have

$$\Delta T_f = \frac{K_f \times W_2 \times 1000}{M_2 \times W_1}$$

Rearranging

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

This equation is used to determine the molecular mass of solute.

10.4.5 Measurement of Freezing Point Depression:

Beckmann's Method

The apparatus used is shown in Fig 10.9. It consists of four parts.

(a) An inner freezing tube with a side arm. It is fitted with a stirrer.
(b) An outer tube in which freezing tube is adjusted. This tube serves as air jacket and helps to achieve a slower and more uniform rate of cooling.
(c) A large vessel containing a freezing mixture.
(d) A Beckmann thermometer which can read up to 0.01K.

Beckmann's method can be understood by the following activity.

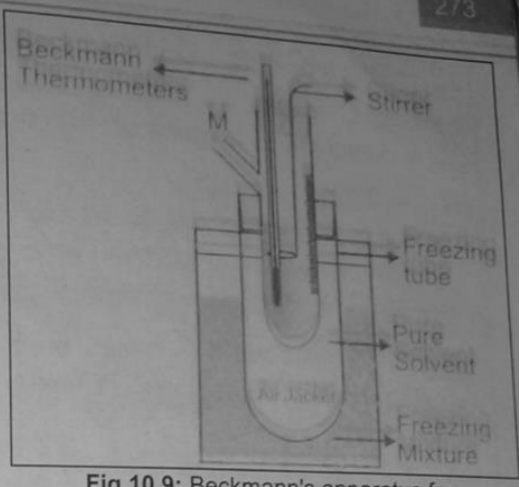


Fig 10.9: Beckmann's apparatus for measurement of depression of freezing point.

ACTIVITY:

In large quantities, the nicotine is a deadly poison. Take 13.4 g of pure water (solvent) in the freezing tube. Now fix thermometer in it, such that the bulb of the thermometer immerse in the solvent. Place freezing tube in the air jacket and air jacket in the large vessel containing a freezing mixture. Record accurate freezing point of solvent. After this remove freezing tube from the air jacket and re-melt the solvent. Add 3.62g of nicotine in the solvent through the side arm. Place freezing tube again in the air jacket and record freezing point of solution while stirring the solution constantly. Thermometer will show –0.563°C. Find the difference of the two freezing points. This will give the value of ΔT_f. From this data determine molecular mass of nicotine.

Solution

Mass of Nicotine = W_2 = 3.62 g

Mass of water = W_1 = 73.4g

K_f for water = 1.86

Freezing point of water = T_1 = 0°C

Freezing point of solution = T_2 = -0.563°C

Molecular mass of nicotine = M_2 = ?

$$\Delta T_f = T_2 - T_2$$
$$= 0 - (-0.563)$$
$$= 0.563^oC$$

Now

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{1.86 \times 3.62 \times 1000}{73.4 \times 0.563}$$

$$M_2 = 162.93 g\ mole^{-1}$$

Thus molecular mass of nicotine is 162.93g $mole^{-1}$

Self Check Exercise 10.8

1. Vitamin K_1 is a substance found in green leafy vegetables. It is needed by the body to produce a blood-clotting factor. A solution of 55.8 mg of vitamin K_1 in 1.048g of Benzene (freezing point of benzene is 5.455°C) has a freezing point of 4.850 °C. K_f for benzene is 5.12. What is the molecular weight of Vitamin K_1? (Ans: 450.596 $g.mole^{-1}$)

2. Ethylene glycol $C_2H_6O_2$ is added to automobile radiator to prevent cooling water from freezing. Calculate the freezing point of ethylene glycol solution that contains 2kg ethylene glycol dissolved in 5 kg water. (K_f = 1.86). Also calculate the boiling point of this solution (K_b = 0.52). **(Ans:** F.P = -11.99°C, B.P = 103.35°C)

Do You Know

De-icing of aeroplane is based on freezing point depression. Think how?

Ice on frozen roads and sidewalks melts when sprinkled with salts such as NaCl and $CaCl_2$. This is because it depresses the freezing point of the water.

Daily Life Applications of Depression of Freezing Point and Elevation of Boiling Point

Ethylene glycol is used as an antifreeze. It is nonvolatile in character and completely miscible with water. When mixed with water it not only lowers the freezing point but also raises the boiling point. In winter it protects a car by preventing the liquid from freezing whereas in hot summer it protects the radiator from over heating.

The principle of freezing point depression is also used to prepare a freezing mixture for use in an ice cream machine. For this purpose NaCl or $NaNO_3$ is used to lower melting point of ice.

10.4.6 Osmotic Pressure

You should not drink seawater, if you are stranded in an ocean. Sea water has higher osmotic pressure than most of the fluids in you body. If you drink it, will pull water out of your body cells. Your cells will die of thirst and you will also die.

Osmotic pressure is the pressure which needs to be applied to a solution to prevent the inward flow of water across a semipermeable membrane. It is also defined as the minimum pressure needed to nullify osmosis. Movement of water molecules from higher concentration to the lower concentration through semi-permeable membrane.

The osmotic pressure of an ideal solution with low concentration can be approximated using the Morse equation (named after Harmon Northrop Morse).

$$\pi = MRT$$

Where

π *is osmotic pressure*

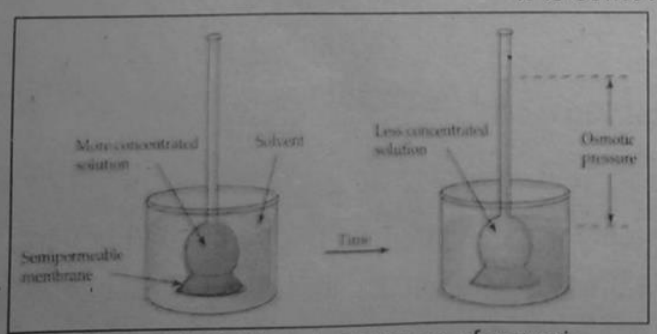


Figure 10.10: The phenomenon of osmosis

M is the molarity

R=.08205746 dm^3 atm K^{-1} mol^{-1} is the gas constant

T is the thermodynamic (absolute) temperature

Certain animal membranes such as that of bladder or outer covering of intestines are semipermeable. They allow the passage of

The student should be guided to interpret information.

ater but not any solute dissolved in the water. A man made material such as cellophane can e used as semipermeable membrane.

A solution inside the bulb is separated from pure solvent in the beaker by a semipermeable s shown in figure 10.10.

As time passes the volume of solution increases and that of solvent decreases. The rocess continues until the hydrostatic pressure due to the extra height of the solution prevents urther osmosis. Therefore the osmotic pressure of a solution is defined as the pressure, which ust be applied above the solution to prevent passage of solvent through a semipermeable embrane into the solution.

Most important applications of osmosis are found in living organisms. If red blood cells are laced in pure water, the cells expand and ultimately rupture as a result of water entering the ells through osmosis. The osmotic pressure of the fluid inside the cell is equivalent to that of .95% NaCl solution. Thus if cells are placed in 0.95% solution of NaCl, there is no net flow of ater through the cell walls and the cell remain stable, this solution is said to be isotonic. olutions having same osmotic pressures are called isotonic solution. If concentration of NaCl olution is greater than 0.95%, water flows out of the cell walls and the cell shrinks. This NaCl olution is hypertonic. If concentration of NaCl solution is less than 0.95%, water flows into the ell and the solution is called hypotonic.

aily Life Applications of Osmosis

The osmosis phenomenon manifests itself in many interesting daily life applications.

1. Biochemists use a technique called haemolysis to study the contents of red blood cells. These cells are protected from the external environment by a semi permeable membrane. Red blood cells are placed in a hypotonic solution. Since hypotonic solution is less concentrated than the interior of the cell, water moves into the cells. The cells swell and eventually bursts releasing haemoglobin and other molecules.
2. Osmosis is the major mechanism for transporting water upward in plants. Because leaves constantly loose water to the air by a process called transpiration. As a result, leaf fluids become more concentrated than the ground water. Thus water enters membranes in the roots and rises into the tree. It can create an osmotic pressure that can exceed 20 atm in the tallest trees.
3. A large quantity of sugar is essential to preserve jam and jelly. This is because the sugar helps to kill bacteria that may cause botulism. When a bacterial cell is in a hypertonic solution. The intracellular water tends to move out of the bacterial cell to the more concentrated sugar solution by osmosis. This causes the bacterial cell to shrink and eventually to cease functioning.
4. Food is preserved by coating with salt, which produces hypertonic solution. Thus food coated with salt causes microbes on the surface to shrivel and die from loss of cell water.

0.4.7 Reverse Osmosis

If a solution in contact with pure solvent across a semipermeable membrane is subjected o an external pressure equal to osmotic pressure, it stops osmosis. If external pressure is

greater than solution's osmotic pressure, it will force solvent to flow from solution to the solvent. This process is called reverse osmosis.

Daily Life Application of Reverse Osmosis

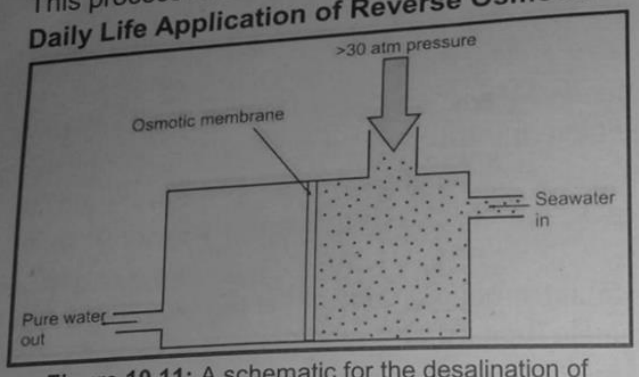


Figure 10.11: A schematic for the desalination of seawater by reverse osmosis

Sea water is highly hypertonic to body fluids and thus is not drinkable. By reverse osmosis it is subjected to desalination (removal of dissolved salts) to make it drinkable. Desalination plant (as shown in figure 10.11), remove large amounts of dissolve salts from sea water. The sea water is pumped under high pressure (20 atm) through the semipermiable membrane, which allow water molecules to pass and stop ions.

Self Check Exercise 10.9

Rank the following aqueous solutions:

i) 0.02m urea ii) 0.02m glucose in order of decreasing

(a) boiling point (b) freezing point

(c) vapour pressure (d) osmotic pressure

Colloidal Particles
Dispersed phase = Solute
Dispersion medium = Solvent

10.5 COLLOIDS

Colloids play a very significant role in nature and our life. For instance medicines in colloidal form are easily adsorbed by body tissues and hence are more effective. Do you think milk is a homogeneous mixture?

A heterogeneous mixture of tiny particles of a substance dispersed through a medium is called colloidal dispersion or a colloid. The particles are called colloidal particles or the dispersed phase and the medium is called the dispersion medium or continuous phase (solvent). In a colloid dispersed phase is like solute in a true solution and dispersion medium corresponds to the solvent in the true solution. The dispersed or suspended particles are single large molecules or aggregates of molecules or ions ranging in size from 1 to 10^3 nm. This means these particles are bigger in size than those of true solutions. These particles are visible in electron microscope but invisible in ordinary microscope. For example, Milk, mayonnaise, smoke, rubber, foam, soap, fog, mist, gem stones like black diamond, opal etc.

When a protein crystal is dropped into water. The polymer molecules dissolve to produce a colloidal dispersion. Such colloidal dispersions in which dispersed phase shows an affinity or attraction for the dispersion medium are called lyophillic (means solvent loving). The molecules of dispersed phase are called **lyophillic molecules**. Examples of lyophillic colloids are proteins, gum, gelatin, starch etc in water.

When solid AgCl is brought in contact with water, it does not spontaneously disperse to form a colloid. Such a colloid, which cannot be made by spontaneous dispersion is called

lyophobic (means solvent hating). The molecules of dispersed phase of such colloids have very little or no attraction for the dispersion medium, are called **lyophobic molecules or particles**. Examples of lyophobic colloids are sulphur, gold, Iron (III) hydroxide in water.

10.5.1 Suspensions

Mix one tea spoon full of sand or mud in a glass of water. What happens? Leave it undisturbed for some time. What happens?

A dispersions in which particles of the dispersed substance are visible to the naked eye are called suspensions. Sizes of these particles are greater than 10^3 nm. This means these particles are bigger than these of colloids. These particles cannot pass through ordinary filter paper. They settle down under the influence of gravity. Suspensions scatter and reflect light. For example, Mud or sand in water are suspensions.

Science Titbits

Blue colour of sky is due to the scattering of light by colloidal dust particles in air. Similarly sea water looks blue due to scattering effect of light by the colloidal impurities in sea water.

10.5.2 Properties of Colloids

General properties of colloids are as follows:

1. Most colloids are cloudy or opaque, but some are transparent to the naked eye
2. When light passes through a colloid, it is scattered by the dispersed particles because their sizes are similar to the wavelength of visible light. When viewed from the side, the scattered beam is visible and broader than that one passing through a solution. This phenomenon is known as Tyndall effect (see Fig 10.12).
3. Colloidal particles exhibit Brownian motion. This can be seen under low magnification. This is because colloidal particles are pushed this way and that way by molecules of dispersing medium.
4. Colloidal particles do not coagulate into larger particles and settle out. This is due to the fact that particles have charged surfaces. These surfaces interact with the molecules of dispersing medium. For example, aqueous protein has charge amino acid groups facing the water and uncharged groups buried within the molecule.

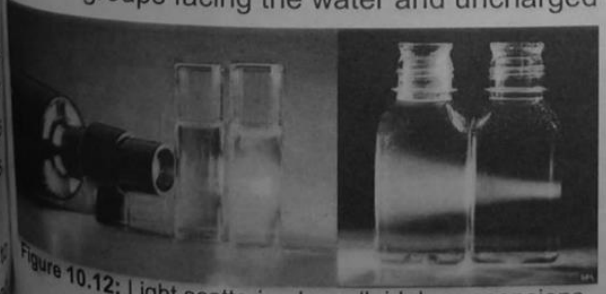


Figure 10.12: Light scattering by colloidals suspensions

5. Temperature changes affect colloids. Increase in temperature makes the colloidal particles move faster and colloid more often and convert little particles into a lump. This means increase in temperature causes coagulation. For example, heating milk causes coagulation of casein in milk.

6. Colloidal particles have little power of diffusion. This is because the colloidal particles have very large size as compare to ordinary solute particles (see table 10.4).
7. Colloidal particles can pass through ordinary filter paper but cannot pass through ultra filter papers.
8. Colloidal particles have high ratio of surface area to volume as compared to the particles of true solutions.
9. Colloidal particles do not settle under the influence of gravity.

10.5.3 Types of Colloids

Colloids are classified according to whether the dispersed and dispersing substances are gases, liquids or solids. Table 10.3 lists some common types of colloids and their examples.

10.3 Types of colloids

Colloid Type (Common Name)	Nature and Composition	Example
Sol	Solid particles dispersed through a liquid	Milk of magnesia, starch dispersed in water, paints, coloured glasses, gems
Gel	Continuous network of solid through out the liquid medium	Halwa, jellies, gelatin
Aerosol	Either a solid or a liquid dispersed in a gas	Smoke (solid expressed in air), fog (liquid dispersed in air)
Emulsion	Liquid dispersed in an another liquid	Milk, mayonnaise
Foam	Bubbles of a gas suspended in a liquid or a solid	Canned shaving cream, soap leather, whipped cream

10.5.4 Comparison of Colloids, Suspensions and True Solutions

The properties of a colloid depend mainly on the size, shape and charge of dispersed particles. Table 10.4 shows, some of the distinctive properties of colloids, suspensions and true solutions.

Table 10.4: Some properties of colloids, suspensions and true solutions

Sr. No.	Properties	Colloids	Suspensions	True solutions
1.	Size of particles	1-10^3 nm	> 10^3 nm	0.1 – 1nm
2.	Phase	Heterogeneous	Heterogeneous	Homogenous
3.	Aggregates	Particles are composed of 10^3 to 109 atoms	Particles are composed of more than 10^9 atoms	Particles are composed of 1 to 10^3 atoms
4.	Charge on the particles	Positive or negative	Positive or negative or may be neutral	Both positive and negative or may be neutral
5.	Visibility of particles	Invisible by naked eye and ordinary micro-scope but visible in electron microscope	Visible by the naked eye and in ordinary micro-scope	Indivisible by the naked eye, ordinary micro-scope as well as electron micro-scope
6.	Filterability	Particles can pass through ordinary filter paper, but cannot pass through ultra filter paper	Particles cannot pass through ordinary as well as ultra filter paper.	Particles can pass through ordinary as well as ultra filter paper.
7.	Dispersion of light	Scatter light	Scatter light	Cannot scatter light
8.	Effect of gravity	Particles do not settle under the influence of gravity	Particles settle under the influence of gravity	Particles do not settle under the influence of gravity

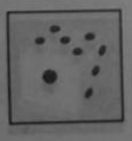
Summary of the Key Terms

1. A solution is a homogenous mixture of two or more substances, which may be solids liquids or gases. The solubility of a solute in a given solvent depends on the interactions between the solute and the solvent particles. Substances that have similarly intermolecular forces tend to dissolve in each other, leading to the generalization "like dissolves like".
2. A rise in temperature usually increases the solubility of solids and liquids but decrease the solubility of gases.

3. Raoult's Law states that when a non-volatile solute is dissolved in a solvent, the vapour pressure of the solution (P) is directly promotional to the mole fraction of the solvent. (X_1) $P = P^0 X_1$
4. The concentration of a solution can be expressed as percent by mass, percent by volume, molarity, molality, mole fraction, ppm, ppb and ppt. The unit of concentration most commonly used by the chemist is molarity, which is defined as the number of moles of solute dissolved per dm^3 of solution.
5. Sometimes it is useful to know the ratio of the amount of solute to the amount of solvent. For this purpose units of molality and mole fraction are used. These units are particularly important in the study of colligative properties that depend only on the number of solute particles that are present and not on their nature.
6. Vapour pressure lowering, boiling point elevation, freezing point depression and osmotic pressure are colligative properties of solution. Adding a solute to a solvent lowers the vapour pressure of the solvent. This causes an elevation in boiling point and depression in freezing point of solution.
7. Osmotic pressure causes the solvent to flow across a semi-permeable membrane when the concentrations of solute on either side of the membrane are unequal.
8. Colligative properties can be used to determine the molecular weight of the solute. This is because these properties depend only on the number of solute particles in a solution and not on their identity.
9. Sea water can be made drinkable by the process of reverse osmosis.
10. A colloid is a heterogeneous mixture of tiny particles of a substance dispersed through a medium. These particles are only visible in electron microscope. When light passes through a colloid, it is scattered by the dispersed particles. This phenomenon is known as Tyndall effect.
11. A dispersion in which particles of the dispersed substance are visible to the naked eye is called suspensions.

References for additional information:

- Silberberg, Chemistry the Molecular Nature of Matter and Energy
- Olmsted and Williams, Chemistry, the Molecular Science
- Zumdahl, Introductory Chemistry
- Raymond Chang, Essential Chemistry
- Caret, Demniston, Topping, Principles and Applications of Inorganic, Organic and Biological Chemistry



Exercise

1: Choose the correct answer

(i) 18g glucose is dissolved in 180g of water. The relative lowering of vapour pressure is
(a) 1 (b) 1.8 (c) 0.01 (d) 0.001

(ii) 100g of a 10% (W/W) NaOH contains 10g of NaOH in

(a) 100g of H_2O (b) 110g of H_2O (c) 10g of H_2O (d) 100g solution.

(iii) Which of the following W/W solutions has the lowest freezing point?
(a) 18% glucose (b) 6% urea (c) 34.2% sucorose (d) 9% glucose

(iv) All solutions containing 1g of non-volatile solutes will have
(a) Same vapour pressure (b) Same boiling point
(c) Same freezing point (d) all of a, b, c

(v) Molarity of pure water is
(a) 1 (b) 1.8 (c) 55.5 (d) 18

(vi) A solution of urea (Mol.Wt=60) is 10% (W/V). the volume in which 1 mole of it is dissolved will be
(a) 6 dm^3 (b) 0.6 dm^3 (c) 60 cm^3 (d) 0.54 dm^3

(vii) The solubility of a substance decreases with increase in temperature if heat of solution is
(a) positive (b) negative
(c) zero (d) cannot be predicted

(viii) Molarity of glucose solution when 9g of it is dissolved in 250 cm^3 of solution is
(a) 0.25M (b) 0.2m (c) 0.5M (d) None of these

(ix) Sea water has about 6 ppm dissolved oxygen. What mass of dissolved oxygen is present in one kg of sea water.
(a) 6g (b) 6×10^{-3}kg (c) 6×10^{-3}g (d) 6×10^{-3}mg

(x) Which of the following is volatile
(a) glucose (b) urea (c) sucrose (d) water

(xi) Which of the following is not true of a colloid.
(a) it is heterogeneous (b) scatter light
(c) can pass through ultra filter paper
(d) Its particles cannot be seen under ordinary microscope.

(xii) Aerosol are colloids which contain
(a) A solid dispersed in a liquid (b) A solid dispersed in a gas
(b) A liquid dispersed in a gas (c) A liquid in another liquid

1. a 2. a, b 3. b, c 4. c, d

2. Determine the molality of the following solutions
i) 3.5% (W/W) glucose (aq) (**Ans:** 0.2 m)
ii) 9.9g $NaNO_3$ dissolved in 940g of water (**Ans:** 0.124 m)

3. What are the freezing and boiling points of a solution prepared by dissolving 13.3g ethylene glycol in 100g of water. (**Ans:** F.P =-3.99°C, B.P = 101.115°C)

4. When cooking, what effect does adding salt to water have on the time required to boil foo

5: Concentrated sulphuric acid is 98% (W/W) H_2SO_4 (aq). Its density is 1.84g.cm^{-3}. Calculate

i) The molarity of the solution (**Ans:**18.4 M)

ii) The molality of the solution (**Ans:**10molal)

iii) What quantity of conc.H_2SO_4 is required to prepare 500cm^3 of 0.1M H_2SO_4.

(**Ans:** 2.717 cm^3)

6. What would you expect potassium chloride, KCl, an ionic solid to be more soluble in H_2O or CCl_4? Explain your prediction.
7. List three factors that accelerate the dissolution process.
8. Explain the nature of solutions in liquid phase giving examples of completely miscible, partially miscible and immiscible liquid-liquid solutions.
9. Express solution concentration in terms of mass percent, molality, molarity, parts per million, parts per billion and parts per trillion and mole fraction.
10. Distinguish between the solvation of ionic species and molecular substances.
11. Describe the role of solvation in the dissolving process.
12. Define the term water of hydration.
13. Define the term colligative.
14. List the characteristic of colloids and suspensions.
15. Define hydrophilic and hydrophobic molecules.
16. Describe on a particle basis why a solution has a lower vapour pressure than the pure solvent.
17. Explain osmotic pressure, reverse osmosis and give their daily life applications.
18. Define heat of solution and apply this concept to the hydration of ammonium nitrate crystals.
19. Interpret the phenomenon of freezing in a mixture of ice and salt.
20. Interpret how solute particles may alter the colligative properties.
21. Interpret the effect of temperature on solubility.
22. Compare molal and molar solutions.
23. Explain the following

(a) Molality is independent of temperature but malarity depends on it.

(b) One molal solution of glucose in water is dilute as compared to one molar solution of glucose, but the number of particles of solute is the same.

(c) The total volume of solution by mixing 50cm^3 of ethanol and 50cm^3 of water may not be equal to 100cm^3, why?

(d) NaCl and $NaNO_3$ are used to lower the melting point of ice.

(e) In summer the antifreeze solutions protect the radiator from boiling over.

(f) One molal and two molal solutions of urea boil at different temperatures.

(g) Relative lowering of vapour pressure is independent of the temperature.

(h) The sum of mole fractions of all the components is always equal to one.

24. 1.89g of an organic compound, A was dissolved per $85cm^3$ of water ($d=0.998g\ cm^{-3}$). The boiling point under one atmospheric pressure of this solution is increased to 100.106°C. What is the molecular weight of A. (**Ans:** 109.2978.4)

25. What freezing point do you expect for water in which 17.9g sucrose $C_{12}H_{22}O_{11}$ is dissolved per 47.6g of H_2O. (**Ans:** -2°C)

26. The vapour pressure of pure water is 23.756 torr at 25°C. How much glucose would be added to 100g of water to bring the vapour pressure down to 23.00 torr? (**Ans:** 31.82g)

27. You are provided 98%(W/W) H_2SO_4 having density 1.84 gcm^{-3}. How much volume of this H_2SO_4 is required to obtain one dm^3 of 30% (W/W) H_2SO_4 which has density 1.25g cm^{-3}. This solution is used in lead storage batteries as electrolyte.

28. Caffeine is about 10 times more soluble in hot water than in cold water. Is ΔH°_{soln} of caffeine in water is positive or negative?

29. PCl_5 can react with itself in a reversible reaction to form a salt that contains the PCl_4^+ and PCl_6^- ions.

$$2PCl_5 \rightleftharpoons PCl_4^{1+} + PCl_6^{1-}$$

The extent to which this reaction occurs depends on the solvent in which it is run. Predict whether a non polar solvent such as CCl_4 favours the product or the reactants of this reaction. Predict what would happen to this reaction if we use a polar solvent. (**Ans:** -8.5°C)

30. An aqueous solution of a compound boil at 102.4 °C. At what temperature will this solution freeze? (**Ans:** 8.5 85 °C)

31. Water and carbon tetrachloride are not missible. When mixed they form two layers. If an aqueous solution of iodine I_2 is shaken with CCl_4, the iodine is extracted into the CCl_4 layer. Interpret this behavior on the basis of your knowledge of intermolecular forces.

32. A Cucumber placed in concentrated brine (concentrated NaCl solution in water) shrivels into a pickle. Give argument to support this observation?

33. Two beakers are placed in a sealed bell jar. One beaker contains water and the other contained concentrated glucose solution. With time volume of water decreases and solution volume increases. Suggest reason?

34. Consider two aqueous solutions one is sucrose and the other of glucose. Both of these solutions boil at 101.52 °C. List some common properties of these solutions.

35. Design on experiment to determine water of crystallizeation in a compound.