Chapter # 10

Solutions and Colloids

Major Concepts

10.1 General Properties of Solutions
10.2 Concentration Units
10.3 Raoult's Law
10.4 Colligative Properties of Non-Electrolyte in Solutions
10.5 Colloids

Learning Outcomes

The students will be able to:

- List the characteristics of colloids and suspensions that distinguish them from solutions. (Applying)
- Define hydrophilic and hydrophobic molecules. (Remembering)
- Explain the nature of solutions in liquid phase giving examples of completely miscible, partially miscible and immiscible liquid-liquid solutions. (Applying)
- Explain the effect of temperature on solubility and interpret the solubility graph. (Analyzing)
- Express solution concentration in terms of mass percent, molality, molarity, parts per million, billion and trillion and mole fraction. (Remembering)
- Define the terms colligative. (Remembering)
- Describe on a particle basis why a solution has a lower vapor pressure than the pure solvent. (Applying)
- Explain on a particle basis how the addition of a solute to a pure solvent causes a elevation of the boiling point and depression of the freezing point of the resulta 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 substan (Understanding)

Introduction

Anything that has mass and occupies space is called matter. Matter is of two types: (I) substance (pure substance), and (ii) mixture (impure substance). Mixtures are classified as either homogeneous (uniform composition throughout) or heterogeneous (different composition in different regions). Solution is a homogeneous mixture of two or more than two different substances consisting of ions or molecules. A solution has at least two components. The component of solution which is in lesser amount is called solute and the component which is in larger amount is called solvent. For example, NaCl is solute and water is solvent in salt-water solution. In the everyday routine, we make or use solutions very often. In the start of the day, we drink a cup of tea which is the solution of water (solvent), sugar (solute) and soluble extracts of tea (solute). Other common examples of solutions are natural gas (homogeneous mixture of gaseous lower hydrocarbons), gasoline (homogeneous mixture of liquid hydrocarbons), soda water (solution of CO_2 in water), milk, sea water, steel, shampoo, and the air we breathe. The amount of solute in the given amount of solvent or solution is called concentration of solution. The solution which has relatively lesser amount of solute is called dilute solution and the solution which has relatively a larger amount of solute is called concentrated solution. For example, 5% aqueous solution of sugar is dilute than 10% aqueous solution of sugar. Likewise, the 10% aqueous solution of sugar is concentrated than 5% aqueous solution of sugar.

The solution of a substance in water is called aqueous solution. Aqueous solutions are the most common solutions and play an important role in chemistry. Common examples of aqueous solutions are sugar-water and salt-water solutions. Solutions may be gases, liquids or solids. There are nine types of solutions on the bases of physical states (i.e. solid, liquid and gaseous nature) of solute and solvent.

Table 10.1: Types of Solutions on the basis of physical states

State of Solute	**State of Solvent**	**State of Resulting Solution**	**Examples**
Gas	Gas	Gas	Air, natural gas
Gas	Liquid	Liquid	CO_2 in water (Soda water)
Gas	Solid	Solid	Hydrogen in palladium, hydrogen in platinum
Liquid	Gas	Gas	Mist, fog, clouds, sprays
Liquid	Liquid	Liquid	Alcohol in water, milk, gasoline, gas oil, antifreeze

State of Solute	State of Solvent	State of Resulting Solution	Examples
Liquid	Solid	Solid	Mercury in Silver (dental amalgam), Cheese, Butter
Gas	Gas	Gas	Dust in smoke
Solid	Liquid	Liquid	Sugar in H_2O, Jellies, Paints, vinegar, sea water
Solid	Solid	Solid	Metal alloys, steel, pearls, opals, ruby, garnet, Brass (Cu/Zn), Solder (Sn/Pb)

Keep in Mind

In the atmosphere there are at least three kinds of solutions. These are (i) gas in gas-oxygen dissolved in nitrogen, (ii) liquid in gas-water vapour dissolved in air, and (iii) solid in gas-tiny particles of smoke dissolved in the air.

In general, polar solutes are dissolved in polar solvents and non-polar solutes are dissolved in non-polar solvents. The common polar solvents that are used in the laboratory are water, acetone, methyl alcohol, and ethyl alcohol. The common non-polar solvents are ether, chloroform, carbon tetrachloride, hexane, toluene benzene.

10.1 General Properties of Solutions

Some of the general properties of solutions are:

i) A solution has at least two components: a solvent and a solute. It may have more than two components: a solvent and two or more solutes.

ii) A solution does not have fixed composition; that is, the ratio of solute can be different from solvent.

iii) The solutes particles consist of ions or molecules.

iv) Solutions may either be colourless or coloured and are usually transparent.

v) The properties of a solution do not remain same when the ratio of solute to solvent is changed.

vi) The solute particles are uniformly distributed in the whole solution and the solution is homogeneous throughout. The solute particles will not "settle out" after a time. The composition and properties of the every part of the solution is exactly the same as every other part.

vii) The solute can generally be separated from the solvent by physical techniques such as evaporation and distillation.

10.1.1 Solution, Suspension and Colloids

Mixtures are classified into homogeneous mixtures and heterogeneous mixtures. Homogeneous mixtures are further classified into solutions and colloids.

A solution is a homogenous mixture of two or more components. The solute particles of solutions are so small that they cannot be seen by the naked eye and microscope. Due to small size, the solute particles pass through filter paper and are not separated from the solution by filtration. The size of solute particles ranges from about 0.01nm to 1nm. Examples of some common solutions are sugar in water, soft drinks, gasoline, and air.

A suspension, on the other hand, is a heterogeneous mixture and the particles are so large that can be seen with microscope and can often be seen with the naked eye. Due to bigger size, the particles cannot pass through filter paper and can be separated from heterogeneous mixtures. The particles of suspension are heavy enough and will settle out under the influence of gravity after a time. The size of particles is more than 1000nm. A common example of suspension is muddy water. Other examples of suspensions are soot, chalk powder in water, sand in water, flour in water and milk of magnesia. Many of the medicinal solutions are marketed as a suspension that are labeled as "suspension" and instructed "shake well before using". You may be familiar with antibiotics that are aqueous suspensions and are well shaken before using. The main purpose of shaking the antibiotics is to mix the substance uniformly in a solvent.

A colloid (or colloidal dispersion) is also a homogeneous mixture of two or more components like solution. The word colloidal means "glue-like" and was originally applied to sticky substances such as starch, gelatin and glue. We cannot use the terms solute and solvent for the components of colloids because the solute particles are not dissolved in the solvent and is present in the form of tiny particles that are suspended in the medium. We, therefore, use the terms dispersed phase for solute and dispersing medium (or continuous phase) for solvent. The dispersed particles are present in lesser amounts than dispersing medium in colloids. Examples of colloids are milk, hair sprays, fog, smoke, clouds, butter, face cream, toothpaste, starch and blood serum. The dispersed particles are intermediate in size between those of a true solution and those of suspension. The dispersed particles of colloids are so small that are not usually seen by the naked eye and ordinary microscope. However, they can be detected by electron microscope. Due to relatively small size, the dispersed particles pass through filter paper of large pores and will not settle out under the influence of gravity over time. The size of dispersed particles is about 2 nm to 1000 nm.

10.1.2 Hydrophilic and Hydrophobic Molecules

Hydrophilic (adj.) (Greek = 'water loving') molecules are polar and having affinity for water. They have ability to form hydrogen bonds or ion dipole forces with water and are readily soluble in water. Hydrophilic substances are also known as polar substances. Examples of hydrophilic molecules are sugars, salts (NaCl), alcohols (ethanol).

Hydrophobic (adj.) (Greek = 'water hating') molecules are non-polar and are repelled by water. The word hydrophobic does not actually mean that they hate water but it means that they are unable to form hydrogen bonds with water and are not dissolved in water. They are dissolved in non-polar solvents such as ether, carbon tetrachloride, benzene. Hydrophobic substances are also known as non-polar substances. Examples of hydrophobic molecules are oils, fats, alkanes (hexane), and greasy substances.

What happens when we add a lump of sugar in water and stir it? The sugar forms hydrogen bonds with water and is dissolved. What happens when we add half test tube of oil in another test tube containing water? The oil does not mix with water. Sugar is hydrophilic and oil is hydrophobic.

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The use of Soap and Detergents

The soap is generally used for bathing and detergents are used for washing. Soaps are mixtures of sodium or potassium salts of fatty acids. A soap molecule is amphipathic and has two ends, the head and the tail. The head of the soap is hydrophilic and the tail of the soap is hydrophobic. When we put soap in water. The hydrophilic end is attracted to water and the hydrophobic end is repelled by water. Hydrophobic end that consists of long chains hydrocarbon and structurally similar to oil is out of water and is attracted by other hydrophobic substances such as oil droplets and grease that are present at the surface of skin or cloth. The soap in this way forms an emulsion (colloid) with water which is used to wash skin, clothes or greasy dishes to remove stain or dirt from their surfaces. Like soaps, detergents also have molecules with a polar end and a non-polar end.

10.1.3 The Nature of Solutions in Liquid Phase

Solutions may be gaseous, liquid, or solid as we know. But here we will focus on liquid solutions because they are in the interests of chemists as lots of reactions of great significance occur in liquid solutions and are undoubtedly the most common and important. Liquid solutions are formed by dissolving a gas liquid, or solid in a liquid solvent. The process of dissolving of solute in a solvent is called dissolution. Examples of liquid solutions are soda water, iodine tincture, tap water, sea water, house hold ammonia, brine, sodium-potassium alloy. Liquid solutions are the most common types that you found in laboratory of chemistry, super market and at your home. In our daily life we dissolve sugar in water to make tea, salt in water to cook food, CO_2 in water to make soda water. Soda water is an example of gas-liquid solution where carbon dioxide gas is dissolved in water. Antifreeze is an example of liquid-liquid solution where ethylene glycol (a viscous hygroscopic liquid) is dissolved in water. Brine is an example of solid-liquid solution where NaCl (a solid salt) is dissolved in water.

The solutions of liquids in liquids are divided into three classes as follows:

(a) Completely Miscible liquids
(b) Partially Miscible liquids
(c) Practically Immiscible liquids

(a) Completely Miscible Liquids

Liquids which dissolve in each other in all proportions are called completely miscible liquids. For example:

i) Alcohol and water
ii) Alcohol and ether
iii) Benzene and toluene
iv) Phenol and H_2O above 67°C.

The properties of such solutions are not additive. Generally, the volume decreases on mixing but sometimes it increases. During the formation of solution, heat may either be evolved or absorbed. Such solutions can usually be separated by fractional distillation.

(b) Partially Miscible Liquids

Liquids which are slightly soluble in each other are called partially miscible liquids. For example:

i) Ether and water
ii) Bromine and water
ii) Phenol and water below 67°C
iv) Aniline and water

When equal volumes of ether and water are shaken in separating funnel, then two layers are formed. Each liquid layer is a saturated solution of the other liquid. Such solutions are called conjugate solutions. Solubility of these conjugate solutions is affected by temperature changes.

(c) Practically Immiscible Liquids

Liquids which are completely insoluble in one another are called immiscible liquids. For example:

i) Benzene and water
ii) Carbon tetrachloride and water
iii) Chloroform ($CHCl_3$) and water
iv) Carbon disulphide (CS_2) and water

10.1.4 Solubility

The concentration of the solute in the solution when it is in equilibrium with the solid substance at a definite temperature is called solubility. For example, the solubility of NaCl in 100g of water at 0°C is 35.7g.

It can also be defined as:

The amount of solute dissolved in 100g of solvent to make a saturated solution at a definite temperature is called solubility.

Units of solubility are:

i) No. of grams of solute in 100g of solvent
ii) Mole dm^{-3}

Various methods are used for the determination of solubility. One of them is given here:

The saturated solution of a solid mass is prepared at a constant temperature. Then this solution is filtered. A known volume of this solution is evaporated in a china dish and from the mass of residue, the solubility is determined.

The extent of solubility of solute in a particular solvent depends on the nature of solute and solvent particles whether they are polar or non-polar, the temperature at which the solution is formed, and the pressure of the gaseous solute.

The Nature of Solute and Solvent

The intermolecular attractive forces that are present among liquids and solids play a key role in the formation of solutions. The intermolecular forces will enable us to understand, why water dissolves table salt (or sugar) but not dissolve cooking oil (or motor oil). A general rule of thumb that, "like dissolves like" is applied here. This means that polar solvents dissolve polar solutes and non-polar solvents dissolve non-polar solutes. For example, polar solutes such as NaCl and sugar are dissolved in water (polar solvent) and non-polar solutes such as cooking oil and motor oil are not dissolved in water (polar solvent).

When NaCl an ionic compound is dissolved in water, the inter-ionic attractive forces among NaCl particles are overcome and they are split up into positive and negative ions. Positive ions are attracted by negative pole of water molecules and negative ions are attracted by positive pole of water molecules. In this way all negative and positive ions are surrounded by water molecules. The process in which an ion or molecule is surrounded by water molecules is called hydration. In a broader way, the process in which an ion or molecule is surrounded by solvent molecules is called solvation.

Sugar (sucrose) and water are soluble in each other because the major force between molecules is hydrogen bonding involving -OH groups. When they both are mixed, the sugar molecules that have -OH groups form hydrogen bonds with water molecules and are dissolved. In this case each molecule of sugar (sucrose) stays together as a single unit and the ions in the solution do not form. The non-polar solutes have London dispersion forces while water has hydrogen bonding that is why neither cooking oil nor motor oil is dissolved in water and float on the surface of water because they do not attract each other significantly.

The non-polar solutes such as cooking oil and motor oil are dissolved in non-polar solvents such as gasoline and benzene. The solutes and solvents both, in this case, has the same attractive forces (London dispersion forces), that is why; they are soluble in one another.

Keep in Mind

The solution is formed when solute-solvent attractions are either comparable to or stronger than solute-solute and solvent-solvent attractions.

We have main focus on solid-liquid solutions and liquid-liquid solutions because the solutions of gases in water are unusual.

10.1.4.1 The Effect of Temperature on Solubility

We have observed that the solubility of tea leaves, coffee beans or sugar increases with rise in temperature and decreases with fall in temperature; that is, they are more soluble in hot water and less soluble in cold water. But the solubilities of some substances such as cerium sulphate and cesium sulphate decrease with rise in temperature.

Keep in Mind

The temperature is not only affecting the rate of dissolving, it also affects the solubility. The rate of dissolving of solute is directly proportional to temperature; that is, the rate of dissolving of solute increases by increase in temperature and decreases by decrease in temperature. But the amount of dissolved solute in the solution may decrease or increase by raising temperature of the solution.

Solubility of Solids in Liquids

The solids have strong attractive forces among their particles. The energy is required to overcome these attractive forces among particles of solids when a solid is dissolved in liquid. The rise in temperature usually increases the solubility of solids in liquids. For example, the solubility of sugar (sucrose) in 100g of water at 0°C is 179g, at 40°C is 238g and at 100°C is 487g. The solubility of some solids like NaCl is little affected by rise in temperature. The solubility of NaCl in 100g of water at 0°C is 35.7g, at 40°C is 36.4g and at 100°C is 39.2g. The solubility of sodium chloride increases only about 3.5g. On the other hand, the solubility of some solids decreases by rise in temperature. For instance, the solubility of cerium sulphate at 0°C is 20.8g, at 40°C is about 6.7g and at 100°C is about 3.87g. The solubility of cerium sulphate decreases about 17g. The solubility curves are shown in the figure. The solubility curve is a graphical representation between temperature and solubility of solute in a particular solvent.

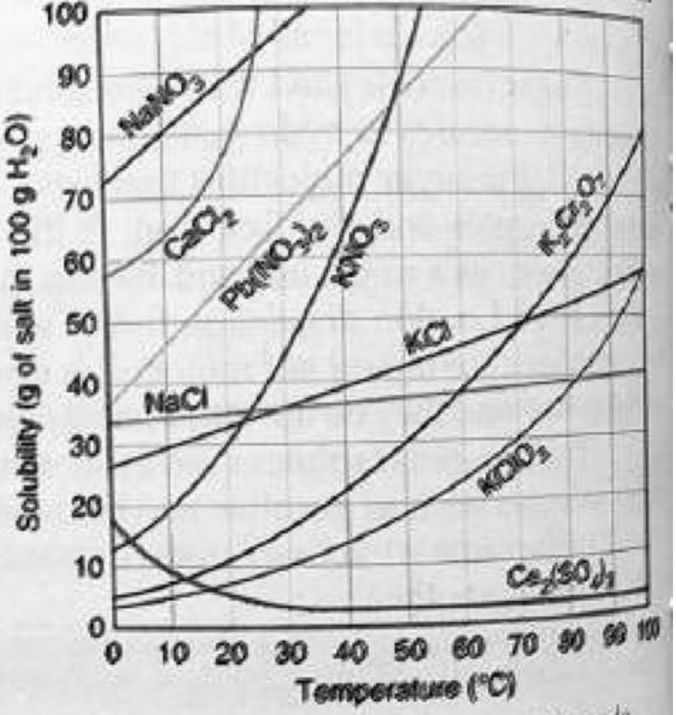

Fig. 10.1: Solublity of most compounds increase by increasing temperature

Table 10.2: Solubility in Gram of Solutes in 100g of Water

Substance	Temperature					
	0°C	20°C	40°C	60°C	80°C	100°C
NaCl	35.7	35.9	36.4	37.1	38.0	39.2
$C_{12}H_{22}O_{11}$	179	204	238	287	362	487
KNO_3	13.9	31.6	61.3	106	167	245
$AgNO_3$	122	216	311	440	585	733
$LiCO_3$	1.54	1.33	1.17	1.01	0.85	0.72
$Ce_2(SO_4)_3$	20.8	10.1	–	3.87	–	–
CO_2 at SP	0.335	0.169	0.0973	0.058	–	–
O_2 at SP	0.00694	0.00537	0.00308	0.00227	0.00138	0.00

Solubility of Liquids in Liquids

The attractive forces among the particles of liquids are weaker than solids. Th additional energy is not required to overcome these attractive forces amo particles of liquids when it is dissolved in another liquid. We can say that the rise in temperature usually does not affect the solubility of most of the liquids in liquids.

Solubility of Gases in Liquids

There are weak attractive forces between gas molecules, so the additional energy is not required to overcome these forces among the molecules of gases when they are dissolved in liquid such as water. The energy as heat is released when a gas is dissolved in water due to forces of attractions between gases molecules and solvent molecules. It means that this is an exothermic process and is favoured by fall in temperature. We may say that the solubilities of gases in water are decreased by rise in temperature (except helium). This is because as the temperature increases the kinetic energy of the gas molecules increase and as a result of this the forces of attractions between gases and water molecules decrease.

Heat of Solution

The energy changes during the formation of a solution are called energetics of solutions or heat of solution or enthalpy of solution.

The solution formation may either be endothermic or exothermic.

The heat of solution is negative for an exothermic process and is positive for an endothermic process. Heat of solution measure the intermolecular attraction between solute and solvent molecules. When KNO_3 is added to water, the temperature of the solution decreases which shows that it is an endothermic process.

$$KNO_{3(s)} + H_2O_{(l)} \rightleftharpoons K^+_{(aq)} + NO^-_{3(aq)} \qquad \Delta H = +35.7\ \text{kJ mol}^{-1}$$

The solubility of potassium nitrate in water is increased by rise in temperature. The solubility of solutes that has positive heats of solution is usually raised by rise in temperature.

When LiCl is added to water, the temperature of the solution increases which shows that it is an exothermic process.

$$LiCl_{(s)} + H_2O_{(l)} \rightleftharpoons Li^+_{(aq)} + Cl^-_{(aq)} \qquad \Delta H = -37\ \text{kJ mol}^{-1}$$

The solubility of lithium chloride in water is increased by fall in temperature. The solubility of solutes that has negative heats of solution is usually raised by fall in temperature. By increasing temperature, the solubility of solutes usually falls in this case.

Table 10.3: Heat of Solution of Some Substances at Room Temperature

Substance	Heat of Solution (kcal/mol)	Substance	Heat of Solution (kcal/mol)
$NaCl_{(s)}$	+3.88	$CH_3COOH_{(l)}$	−1.51
$NaOH_{(s)}$	−44.51	$H_2SO_{4(l)}$	−74.32
$KNO_{3(s)}$	+34.89	$HCl_{(g)}$	−74.84
$NH_4NO_{3(s)}$	+25.69	$NH_{3(g)}$	−30.50
$KOH_{(s)}$	−57.61	$NH_4Cl_{(s)}$	+14.78
$KClO_{3(s)}$	+41.38	$AgNO_{3(s)}$	+22.59
$CuSO_4.5H_2O_{(s)}$	+11.51	$NaNO_{3(s)}$	+20.50
$MgSO_4.7H_2O_{(s)}$	+15.90	$Na_2SO_4.10H_2O_{(s)}$	+78.54

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Use of Cold and Hot Pouches

Heat is either given off or absorbed when ionic substances are dissolved in water. The endothermic solution process is used in the "instant cold pouches" and the exothermic solution process is used in the "instant hot pouches". These instant cold and hot pouches are used in hospitals and at other places for the treatment of minor injuries, sprains, and bruises to reduce swelling. In instant cold pouches, the NH_4NO_3 (or NH_4Cl) is packed in a stronger bag that is placed inside a plastic bag of water. When the inner bag that contains ammonium nitrate is broken by squeezing, the ammonium nitrate dissolves in the water in the outer bag. In this case, the heat is absorbed and the pouch becomes cold. On the other hand, the hot pouches generally contains $CaCl_2$ (or $MgSO_4$), which on dissolving in water, release heat and the pouch becomes hot.

10.1.4.2 The Effect of Pressure on Solubility

The solubilities of solids or liquids are affected very little or not affected by changing pressure that is why the effect of pressure on the solubility of solids or liquids in liquid solvent is neglected. However, pressure has a great effect on the solubility of gases. The English chemist William Henry studied the effect of gases on the solubility of gases. According to Henry's law, the solubility of a gas in a liquid is directly proportional to the pressure of the gas above the liquid at a given temperature. -

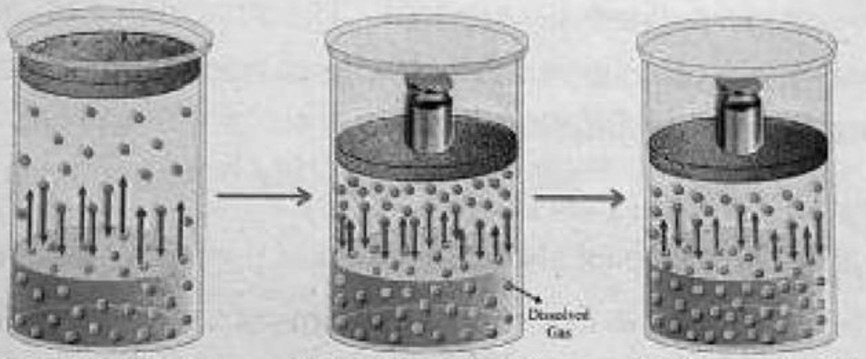

William Henry (1774-1836)

$$S \propto P \quad \text{or} \quad S = K.P$$

Where, K is proportionality constant, S is the solubility of the gas in the solvent, and P is the partial pressure of the gas.

The solubility of the gas is doubled by doubling the partial pressure of the gas above the solution and is tripled by tripling the partial pressure of the gas above the solution, and so on. The solubility of CO_2 gas at one atmosphere in 100g of water is 0.169g at 20°C. The solubility of CO_2 gas will be doubled (0.338g) by doubling the pressure of the gas to two atmospheres at same temperature.

Suppose we have a gas in equilibrium with a solution at which the number of molecules entering the solution is equal to the number of molecules leaving the solution. When the pressure of the gas above the solution is increased, the more molecules are forced into the solution. Thus, the concentration of the gas in the solution increases until the equilibrium is restored again; that is, the solubility increases until the rate at which the number of gas molecules entering the solution is equal to the number of gas molecules leaving from the solution. Thus, we can say that the solubility of gas in a liquid increases by increasing pressure above the solution.



(a) Dynamic equilibrium (b) Gas pressure increased (c) Dynamic equilibrium

Figure: 10.2: This diagram shows the concentration of dissolved gas is proportional to the pressure of the gas above the solution

Interesting Information

Carbonated beverages are the saturated solutions of carbon dioxide in water under pressure. Carbonated beverages are bottled at a pressure of 5-10 atmospheres. When a bottle of carbonated beverage is uncapped, the pressure on the surface of the beverage is reduced to one atmosphere, and the bubbles of carbon dioxide come out suddenly of the solution. This rapid escape of dissolved gas from a liquid is known as effervescence.

0.2 Concentration Units of Solutions

oncentration of solution means the amount of solute present in a given amount of olution. We commonly deal with dilute (a solution which has relatively small amount of solute) and concentrated solution (a solution which has relatively large amount of solute) in our daily life. Sometimes we use unsaturated (a solution which can dissolve further solute on addition), saturated (a solution which is fully loaded up with solute and cannot dissolve more solute at a given temperature) and supersaturated (a solution which has more of the solute than the usual maximum amount and are unstable). They are prepared at an elevated temperature or pressure to describe the concentration of solution. But in chemistry lab we have to know the exact concentration of a solution; that is, the exact amount of solute in the given amount of solution. Chemists use several different concentration units, each of which has its own benefits and drawbacks. The amount of solutes or solvents in a solution is generally measured in terms of volume, mass or moles and these units are usually called concentration units. The concentration of solution can be expressed by percentage composition, molarity, molality, mole fraction, and parts per million, billion and trillion.

10.2.1 Percentage Composition

The term percent means "parts per hundred". The number of parts of solute present in 100 parts of solvent or solution is called %age composition.

It can be expressed by four different ways:

i) Weight by Weight Percent (W/W %)

The weight by weight percent also called weight percent or percent by weight or percent by mass is defined as the number of grams of solute present in 100 grams of solution. For example, a 10% W/W sugar solution means contains 10 g of sugar in 100g of solution. The solution has 10 g of sugar (solute) and 90 g of water (solvent).

$$\%\text{ age by weight} = \frac{\text{Mass of solute}}{\text{Mass of solution (mass of solute + mass of solvent)}} \times 100$$

The solute and solution masses must be measured in the same unit; that is, in grams.

Example 10.1

What is weight percent of solution containing 100g of water and 2.5g of glucose?

Solution:

Mass of glucose = 2.5g
Mass of water = 100g
Mass of solution = 100 + 2.5 = 102.5g

Weight percent of glucose solution = ?

Weight percent of glucose solution $= \dfrac{\text{Mass of glucose}}{\text{Mass of solution}} \times 100$

$= \dfrac{2.5\text{g}}{102.5\text{g}} \times 100$

$= 2.45\%$

Practice Exercise 1:
Calculate the weight percent of sugar (sucrose) when 15g of sugar is added to 85g of water?

i) Weight by Volume Percent (W/V %)

The weight by volume percent is defined as the number of grams of solute present in 100mL of solution. For example, a 10% W/V sugar solution means that 10g of sugar in each 100mL of solution.

$$\%\text{age by weight} - \text{volume} = \frac{\text{Mass of solute (g)}}{\text{Volume of solution (mL)}} \times 100$$

The units do not cancel because the concentration unit of solute is always given in grams and that of solution is milliliters. This concentration unit is normally used when the solute is a solid and the solvent or resulting solutions are liquids.

Example 10.2

Calculate the mass by volume percent of 70mL of solution that has 20g of NaOH.

Solution:

Mass of NaOH = 20g
Volume of solution = 70mL
Mass by volume percent of solution = ?

Mass by volume percent of solution $= \dfrac{\text{Mass of NaOH in g}}{\text{Volume of solution in mL}} \times 100$

$= \dfrac{20\text{g}}{70\text{mL}} \times 100$

$= 28.57\%$

Practice Exercise 2:
A 50mL solution is prepared by dissolving 0.75g of NaCl in enough water. What is the W/V% of the solution?

iii) Volume by Weight Percent V/W %

The Volume by Weight Percent is defined as the number of milliliters of solute in 100 g of solution. For example, a 10% V/W solution means that 10mL of alcohol in water to make the total mass of the solution 100g.

$$\%\text{age by volume} - \text{weight} = \frac{\text{Volume of solute}}{\text{Mass of solution}} \times 100$$

In this type of solution, we know the total mass of solution but we do not know the total volume of solution.

iv) Volume by Volume Percent (V/V %)

The volume by volume percent (also called volume percent or percent by volume) is defined as the volume in milliliter of solute in 100 milliliter of solution. For example, 10% V/V alcohol-water solution means that 10mL of alcohol dissolved in enough water to make the total volume of solution 100mL. Keep in mind that 10% V/V alcohol-water solution cannot be prepared by dissolving 10mL of alcohol in 90mL of water because the total volume of the solution may not be necessarily equal to the sum of volumes of solute and solvent.

$$\%\text{age by volume} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

The solute and solution masses must be measured in the same unit; that is, in milliliters.

Example 10.3

An alcoholic beverage has 25mL of ethyl alcohol. What is V/V % of 75mL of the beverage?

Solution:

Volume of ethyl alcohol = 25mL
Volume of beverage = 75mL
Volume percent = ?

$$\text{Volume percent} = \frac{\text{Volume of ethyl alcohol}}{\text{Volume of beverage}} \times 100$$

$$= \frac{25\text{mL}}{75\text{mL}} \times 100$$

$$= 33.33\%$$

Practice Exercise 3:

Calculate the V/V % of 0.250L of vinegar solution containing 12.5mL of acetic acid.

10.2.2 Molarity

Molarity is the most useful unit of concentration in chemistry. The number of moles of solute dissolved in one liter of solution is called molarity. It is denoted by 'M'. The molarity has the units of mole per liter (mol/L). The mathematical equation for the calculation of molarity is as follows:

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute (n)}}{\text{Volume of solution in liter (V)}} \qquad \text{or}$$

$$\text{Molarity (M)} = \frac{\text{Mass of solute in gram}}{\text{Molar mass of solute}} \times \frac{1}{\text{Volume of solution in liter}}$$

In order to calculate the molarity of solution, the number of moles of solute and the volume of the solution in liters are required. Keep in mind that a solution of a given molarity is prepared by dissolving the solute in enough solvent to make the total volume of the solution one liter. We are not allowed to dissolve the solute in one liter of solvent, because the total volume of solution in this case might be a bit larger than one liter due to additional volume of solute. If the volume of the solution is larger than one liter, then the molarity will be less than 1M for the solution.

Keep In mind

To prepare one molar (1M) aqueous solution of NaOH, we are required to dissolve 1 mol (40g) of NaOH in enough water to make the total volume of the solution one liter. If we want to prepare 2M NaOH solution, then we are required to dissolve 2 moles (80g) of NaOH in enough water to make the total volume of the solution one liter. The solution whose molarity is one is called one molar solution.

Molarity is very useful to us because the number of moles of solute dissolved can be calculated by multiplying the molarity of solution with the volume of solution.

$$\text{No. of moles of solute (n)} = \text{Molarity (M)} \times \text{Volume of solution in liter (V)}$$

The number of moles of solutes enables us to calculate the number of particles and to solve those problems that involve quantities in chemical reactions.

Table 10.4: Concentration of Solution in Molarity

Solute	of Solute	Moles of Solute	Solvent	Volume of Solution (mL)	Volume of Solution (L)	Molarity (M)
Iodine	2.54	0.01	CCl_4	500	0.50	0.02
KNO_3	750	7.42	H_2O	15×10^2	1.50	4.95
Sucrose	250	0.73	H_2O	750	0.75	0.97
NH_3	8.5	0.50	H_2O	180	0.18	2.78

Example 10.4

What is the molarity of 2.5L of solution containing 2.0mol of ethyl alcohol?

Solution:

Molarity of solution = ?
Volume of solution = 2.5L
Moles of ethylalcohol = 2.0mol

$$\text{Molarity (M) of solution} = \frac{\text{Moles of ethyl alcohol}}{\text{Volume of solution in dm}^3}$$

$$= \frac{2.0\text{mol}}{2.5\ \text{L}}$$

$$= 0.8\text{M}$$

Practice Exercise 4:

Calculate the molarity of the sulphuric acid solution when 49g of sulphuric acid is dissolved in 2.5×10^2mL of solution.

Example 10.5

Prepare 1 M and 0.50M solutions of glucose.

Solution:

Molar mass of glucose $(C_6H_{12}O_6) = 72 + 12 + 96 = 180\text{g mol}^{-1}$

- **<u>1 M glucose solution</u>** is prepared by dissolving 180g (1 mol) of glucose one liter of solution.
- **<u>0.50M glucose solution</u>** is prepared by dissolving 0.50mol (0.50×180=90 of glucose in one liter of solution.

Practice Exercise 5:

How can you prepare 0.25M and 2.50M NaOH solution?

10.2.3 Molality

The number of moles of solute dissolved in one kilogram of solvent is called molality. It is denoted by lower case 'm'. The molality has the units of moles per kilogram (mol/kg). The mathematical equation for the calculation of molality is as follows:

$$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kilogram}} \quad \text{or}$$

$$\text{Molality (m)} = \frac{\text{Mass of solute in gram}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in kilogram}}$$

The one molal solution of a solute is prepared by dissolving one mole of solute in one kilogram of solvent.

For example, 1 molal HCl solution is prepared when one mole of HCl (36.5g) is dissolved in 1kg of solvent. If we dissolve one mole (60g) of urea, $(NH_2)_2CO$ in 500 g (0.50kg) of water, the solution in hand will be 2 molal.

Note that the molarity of solution is temperature dependent but molality is temperature independent. It means that molarity of a solution changes slightly when temperature of the solution changes, but molality does not. Hence, molality is more convenient in those experiments that involve temperature changes.

The mass of one liter of water at room temperature is nearly equal to one kilogram that is why, in dilute solutions the number of moles of solute in one liter is about equal to the number of moles of solute in one kilogram of water. Hence, molarity is about equal to molality in dilute aqueous solutions.

Table 10.5: Concentration of Solution in Molality

Solute	Amount of Solute (g)	Moles of Solute	Solvent	Amount of Solvent (g)	Amount of Solvent (kg)	Molality (m)
CH_3OH	16.0	0.50	C_2H_5OH	1000	1.00	0.50
$CuSO_4$	159.6	1.00	H_2O	10^4	10.00	0.10
H_2O_2	68	2.00	H_2O	2×10^3	2.00	1.00
HCl	3.65	0.10	H_2O	1×10^2	0.10	1.00

Example 10.6

What is the molality of solution made by dissolving 290g of acetone (CH_3COCH_3) in 500 g of water?

Solution:

Molality of acetone solution = ?

Mass of acetone = 290g

Mass of water = 500g = 0.50kg

$$\text{Molality of acetone solution} = \frac{\text{Given mass of acetone}}{\text{Molar mass of acetone}} \times \frac{1}{\text{Mass of solvent in kg}}$$

$$\text{Molality of acetone solution} = \frac{290\text{ g}}{58\text{ g}} \times \frac{1}{0.50\text{kg}} = 10\text{m}$$

Practice Exercise 6:

What is the molality of solution that is made by dissolving 96g of SO_2 gas in 750g of water?

10.2.4 Mole Fraction

It can be defined as:

The number of moles of a particular component divided by total number of moles of all the components present in the solution is called mole fraction. It is denoted by 'X'. Mole fraction has no unit.

$$\textbf{Mole fraction (X)} = \frac{\textbf{Number of moles of one component}}{\textbf{Total number of moles of all the component}}$$

The mole fraction of a particular component can be calculated as:

$$X_A = \frac{n_A}{n_A + n_B + n_C} = \frac{n_A}{n_t}$$

i) The mole fraction of a particular component is always less than 1.

ii) The sum of mole fraction of the entire components is always unity.

$$X_A + X_B + X_C = 1$$

If a solution has 2 moles of NH_3 in 7 moles of water, the mole fraction of NH is shown as X_{NH_3} and the mole fraction of water is shown as X_{H_2O}. The total numbe of moles of ammonia and water in the solution are (2 mol + 7 mol) = 9.00 moles. Th mole fractions of ammonia and water in the solution are calculated as:

$$\text{The mole fraction of ammonia} = X_{NH_3} = \frac{\text{No. of moles of ammonia}}{\text{Total number of moles}} = \frac{2\text{mol}}{9\text{mol}} = 0.2$$

$$\text{The mole fraction of water} = X_{H_2O} = \frac{\text{No. of moles of water}}{\text{Total number of moles}} = \frac{7\text{mol}}{9\text{mol}} = 0.7$$

The sum of the mole fractions of both the components of solution is equal to one.

Mole fractions are specially used for the calculations of concentrations of gas mixtures but not for liquid solutions because other units of liquid solutions are generally more appropriate. Mole fractions do not depend on temperature.

Example 10.7

Calculate the mole fraction of each component of solution prepared by dissolving 1.22g of benzoic acid (C_6H_5COOH), in 11.95g of chloroform ($CHCl_3$).

Solution:

Mole fraction of benzoic acid = $X_{C_6H_5COOH} = ?$

Mole fraction of chloroform = $X_{CHCl_3} = ?$

Mass of benzoic acid = 1.22g

Mass of chloroform = 11.95g

$$\text{Moles of benzoic acid} = \frac{\text{Mass of benzoic acid}}{\text{Molar mass of benzoic acid}} = \frac{1.22\text{ g}}{122\text{gmol}^{-1}} = 0.01\text{ mol}$$

$$\text{Moles of chloroform} = \frac{\text{Mass of chloroform}}{\text{Molar mass of chloroform}} = \frac{11.95\text{ g}}{119.5\text{gmol}^{-1}} = 0.1\text{ mol}$$

$$\text{Total number of moles} = 0.01\text{mol} + 0.1\text{mol} = 0.11\text{moles}$$

$$\text{Mole fraction of benzoic acid} = \frac{\text{Moles of benzoic acid}}{\text{Total number of moles}} = \frac{0.01\text{mol}}{0.11\text{mol}} = 0.09$$

$$\text{Mole fraction of chloroform} = \frac{\text{Moles of chloroform}}{\text{Total number of moles}} = \frac{0.1\text{ mol}}{0.11\text{mol}} = 0.91$$

Practice Exercise 7:

A can of 2kg antifreeze contains 0.8 kg of ethylene glycol ($C_2H_6O_2$), and 1.2kg of water. What are the mole fractions of both the components of antifreeze?

10.2.5 Parts Per Million, Billion and Trillion

These units are used for very dilute solutions in which solutes have very low mass concentrations. These units can be used to express the amounts of pollutant gases in the air, the impurities present in the water, and the drugs in the human body.

Parts per million

The number of parts by mass (or volume) of solute present in one million (10 lac) parts by mass (or volume) of the solution is called parts per million. It is abbreviated as ppm. **One ppm is equal to one milligram of solute in one kilogram of solution (1 ppm = 1 mg/kg).**

$$\text{Parts per million (ppm)} = \frac{\text{Mass of solute in grams}}{\text{Mass of solution in grams}} \times 10^6 \quad \text{or}$$

$$\text{Parts per million (ppm)} = \frac{\text{Volume of solute in mL}}{\text{Volume of solution in mL}} \times 10^6 \quad \text{or}$$

$$\text{Parts per million (ppm)} = \frac{\text{Mass of solute in grams}}{\text{Volume of solution in mL}} \times 10^6$$

It is used for very low concentrations of the solution. The impurities of the substances are usually expressed in ppm. For example, the maximum allowable amount of arsenic in water is 0.05ppm. A concentration of 0.05ppm means that if you drink 10 lac grams of water, then 0.05g of those is arsenic and other 999999.95g are consist of water and other minerals and contaminants.

Table 10.6: The Maximum Allowable Amounts of Minerals in Drinking Water

Mineral	Amount in ppm	Mineral	Amount in ppm
Sodium	160.00	Lead	0.05
Barium	1.00	Cadmium	0.01
Arsenic	0.05	Selenium	0.01
Chromium	0.05	Mercury	0.002
Silver	0.05		

Parts per Billion

The number of parts by mass (or volume) of solute present in one billion (1000million) parts by mass (or volume) of the solution is called parts per billion. It is abbreviated as ppb. **One ppb is equal to one microgram of solute in one kilogram of solution (1 ppb = 1μg/kg).**

$$\text{Parts per billion (ppb)} = \frac{\text{Mass of solute in grams}}{\text{Mass of solution in grams}} \times 10^9 \quad \text{or}$$

$$\text{Parts per million (ppb)} = \frac{\text{Volume of solute in mL}}{\text{Volume of solution in mL}} \times 10^9 \quad \text{or}$$

$$\text{Parts per million (ppb)} = \frac{\text{Mass of solute in grams}}{\text{Volume of solution in mL}} \times 10^9$$

It is used for very low concentrations of the solution.

Parts per Trillion

The number of parts by mass (or volume) of solute present in one trillion (1000 billion) parts by mass (or volume) of the solution is called parts per billion. It is abbreviated as ppt. One ppt is equal to one nanogram of solute in one kilogram of solution (1 ppt = 1ng/kg).

$$\text{Parts per trillion (ppt)} = \frac{\text{Mass of solute in grams}}{\text{Mass of solution in grams}} \times 10^{12} \quad \text{or}$$

$$\text{Parts per million (ppt)} = \frac{\text{Volume of solute in mL}}{\text{Volume of solution in mL}} \times 10^{12} \quad \text{or}$$

$$\text{Parts per million (ppt)} = \frac{\text{Mass of solute in grams}}{\text{Volume of solution in mL}} \times 10^{12}$$

It is also used for very low concentrations of the solution.

Example 10.8

The concentration of gold in one liter of sea water is 0.001mg. What is the concentration of this solution is ppm, ppb, and ppt?

Solution

Mass of gold = 0.001mg = 0.000001g

Volume of sea water = 1L = 1000mL

$$\text{Concentration of gold in sea water in ppm} = \frac{0.000001\text{ g}}{1000\text{ mL}} \times 10^{6} = 0.001\text{ ppm}$$

$$\text{Concentration of gold in sea water in ppb} = \frac{0.000001\text{ g}}{1000\text{ mL}} \times 10^{9} = 1\text{ ppb}$$

$$\text{Concentration of gold in sea water in ppt} = \frac{0.000001\text{ g}}{1000\text{ mL}} \times 10^{12} = 1000\text{ ppt}$$

Practice Exercise 8:

A college student added one tea spoon of table sugar in the swimming pool. The concentration of table sugar is about 4×10^{-6}g per liter of water. Determine concentration of table sugar in water of swimming pool in ppm, ppb, and ppt.

0.3 Raoult's Law

1886, the French chemist François Marie Raoult (1830-1901) discovered a lation between the vapour pressure of pure solvent and that of a solution. He served that the vapour pressure of solvent above a solution of a non-volatile, non-ctrolyte solute depends on the mole fraction of solvent in the solution.

10.3.1 Non-volatile, Non-electrolyte Solute in Volatile Solvents

Raoult's law can be defined in three ways:

i) The vapour pressure of solvent above a solution is equal to the product of the vapour pressure of pure solvent and its mole fraction in the solution.

$$P = P^{o}X_1$$

Where,

P is vapour pressure of solvent above solution, P° is vapour pressure of pure solvent, and X_1 is mole fraction of the solvent in the solution.

ii) The lowering of vapour pressure of a solvent is directly proportional to the mole fraction of solute.

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

Where, Lowering of vapour pressure $(\Delta P) = P^{\circ} - P$

The lowering of vapour pressure depends on the number of non-volatile, non-electrolyte solute particles (atoms, molecules, or ions) in the solution but does not depend on the nature of the solute particles; that is, no matter what types of particles (atoms, molecules, or ions) are present in the solution.

iii) The relative lowering of vapour pressure $(\Delta P / P^{\circ})$ is equal to the mole fraction of solute. By increasing mole fraction of solute, vapour pressure of solvent or solution decreases.

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

The solutions that obey Raoult's law are said to be ideal solutions.

The relative lowering of vapour pressure $(\Delta P / P^{\circ})$ has three properties.

i) It is independent of temperature.
ii) It is proportional to the concentration of solute.
iii) It is constant when equimolar proportions of different solutes are dissolved in the same solvent.

The composition of vapours formed above a solution is different from the composition of liquid mixture at a given temperature. In order to obey the Raoult's law for the ideal solutions, the solute should be non-electrolyte and solution should be dilute.

10.3.2 When Both of the Components are Volatile

According to Dalton's law of partial pressure, the total pressure of a mixture of two or more volatile liquids such as petrol is equal to the sum of the individual vapour pressures of all the components (liquids) present in the mixture (liquid solutions).

$$P_{total} = P_A + P_B \quad \text{..................(i)}$$

The individual vapour pressures P_A and P_B are determined by Raoult's law. In case of components A and B, it makes no difference which we call the solute and which the solvent. According to Raoult's law, the individual vapour pressure of a particular component above a solution is equal to the product of the vapour pressure and its mole fraction in the solution.

$$P_A = P_A^\circ . X_A$$

$$P_B = P_B^\circ . X_B$$

By putting the values of P_A and P_B in equation (i) we get

$$P_t = P_A^\circ X_A + P_B^\circ X_B$$

Since,

$$X_A + X_B = 1$$

or $$X_A = 1 - X_B$$

Therefore,

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

or $$P_t = P_A^\circ - P_A^\circ X_B + P_B^\circ X_B$$

or $$P_t = P_B^\circ X_B - P_A^\circ X_B + P_A^\circ$$

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

Graphical Representation of Raoult's Law

Consider the following six solutions with their mole fraction and vapour pressure values in the table.

Table 10.7: Six solutions with their Mole Fraction and Vapour Pressure Values

Mole Fraction of Component (X_A)	Vapour Pressure of Component A (P_A°)	Mole Fraction of Component B (X_B)	Vapour Pressure of Component B (P_B°)	Vapour Pressure of Solution ($P_t = P_A^\circ + P_B^\circ$)
0.0	0	1.0	50	50
0.2	5	0.8	40	45
0.4	10	0.6	30	40
0.6	15	0.4	20	35
0.8	20	0.2	10	30
1.0	25	0.0	0	25

All of the above solutions have their own vapour pressures. We know that the total vapour pressure is equal to the sum of the vapour pressures of all the components present in the solution.

When a graph is plotted between mole fractions of A and B; that is, X_A and X_B on x-axis and vapour pressures of solutions; that is, P_A° and P_B° on y-axis, then a straight line is obtained.

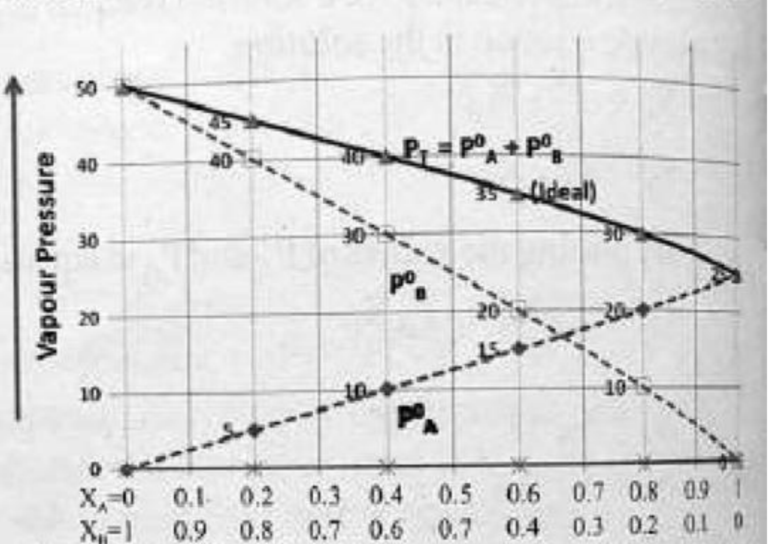


Figure 10.3: Graphical Representation of Raoult's Law

The graph shows that component B is more volatile than component A. Hence, it has high vapour pressure and low boiling point as compared to component A. The two dotted lines show the partial pressures of the individual components of the solution. The vapour pressure of the components increases with increase in its mole fraction in the solution.

The total pressure is equal to the sum of the vapour pressures of the individual components ($P_t = P_A^\circ + P_B^\circ$) as shown in the graph.

Deviations of Solutions

Only the ideal solutions obey Raoult's law. The ideal behaviour is shown by those solutions which have nearly the same forces of attractions between the particles of solute-solute, solute-solvent, solvent-solvent. Hence, the enthalpy change during the formation of solution is zero. For real solutions, the forces of attractions between the particles of solute-solvent are much stronger or weaker than solute-solute and solvent-solvent. Consequently, the enthalpy change may either be negative or positive. Therefore, a non-ideal solution shows two types of deviations from Raoult's law.

Positive Deviation

A non-ideal solution in which solute-solvent attraction is weaker than solute-solute and solvent-solvent attraction shows positive deviation from Raoult's law

In this type of deviation,

i) Vapour pressure of solution is higher than its individual components.

ii) Boiling point of solution is lower than its individual components. For example, the boiling point of the solution of ethanol + water is 78.1°C which is lower than the boiling point of both ethanol (78.5°C) and water (100°C).

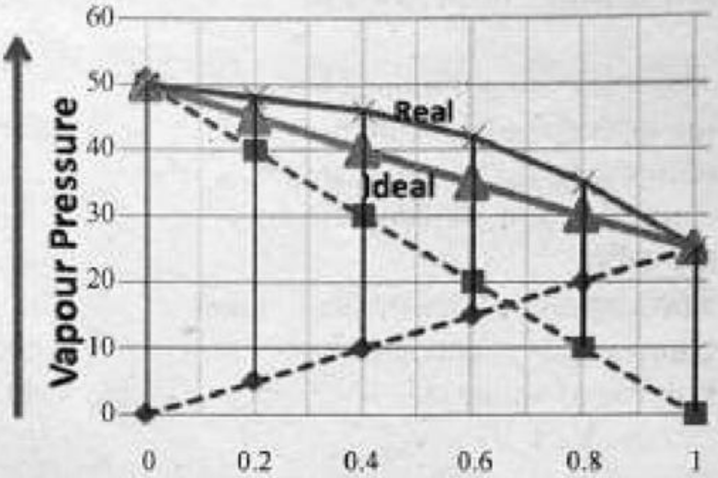


Figure 10.4: Positive Deviation from Raoult's Law

Negative Deviation

A solution in which solute-solvent attraction is stronger than solute-solute and solvent-solvent attraction shows negative deviation from Raoult's law.

In this type of deviation:

i) The vapour pressure of solution is lower than that of its individual components.

ii) The boiling point of solution is higher than that of its individual components. For example, the boiling point of the solution of Water + HCl is 108.5°C which are higher than the boiling point of both water and HCl.

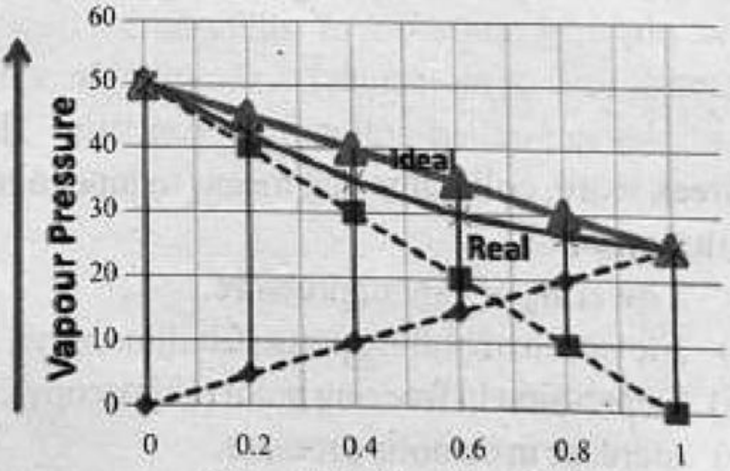


Figure 10.5: Negative Deviation from Raoult's Law

Table 10.8: Difference Between Ideal and Non-ideal Solutions

Ideal Solution	Non-ideal Solution
A solution which obeys Raoult's law is called ideal solution.	A solution which does not obey Raoult's law is called non-ideal solution.
The forces of attractions between the molecules of different components (i.e. solute-solute, solute-solvent, solvent-solvent) are same.	The forces of attractions between the molecules of different components are not same.
The sum of volumes of individual components is equal to the volume of solution. $V_1 + V_2 = V_{solution}$	The sum of volumes of individual components is not equal to the volume of solution. $V_1 + V_2 \neq V_{solution}$
There is no enthalpy change during the formation of solution. $\Delta H = 0$	There is an enthalpy change during the formation of solution. $\Delta H \neq 0$
Examples are: Benzene-Toluene, Benzene-ether.	Examples are: Acetone-water, ethanol-Hexane.

10.4 Colligative Properties of Dilute Solutions

The physical properties of solutions that depend on the number of solute particles (atoms, ions or molecules) in the solution and do not depend on the kind of solute particles are called colligative properties. The word colligative is derived from Greek word, colligatus that means to bind together. Colligative properties of dilute solution are:

i) Lowering of vapour pressure.
ii) Elevation of boiling point (Ebullioscopy).
iii) Depression in freezing point (Cryoscopy).
iv) Increase in osmotic pressure.

It is important to keep in mind the conditions for colligative properties given below:

i) Solution should be dilute; that is, the concentration of the solution should be less than or equal to 0.2 M.
ii) Solute should be non-volatile.
iii) Solute should be non-electrolyte.

10.4.1 Lowering of Vapour Pressure

Vapour pressure is the pressure exerted by the vapour in equilibrium on liquid in a closed container. A substance which has a measureable vapour pressure is volatile and a substance which does not have a measureable vapour pressure is non-volatile. If a non-volatile and non-electrolyte solute is added to a solvent, then evaporating tendency of solvent molecules from the surface of solution decreases. Thus vapour pressure of solution decreases. This is known as lowering of vapour pressure. The vapour pressure of a volatile solvent that has a non-volatile solute is inversely proportional to number of solute particles in the solution.

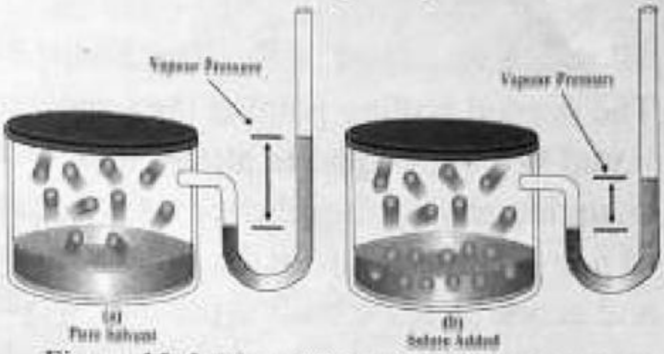


Figure 10.6: The addition of solute to the liquid solvent decreases the evaporating tendency of solvent

$$\text{Vapour pressure} \propto \frac{1}{\text{No. of non-volatile solute particles}}$$

It means the vapour pressure of solvent decrease by increase in the number of solute particles and vice versa.

According to Raoult's law,

The relative lowering of vapour pressure is equal to the mole fraction of solute.

$$\Delta P / P^\circ = X_2$$

We know that, $X_2 = \dfrac{n_2}{n_1 + n_2}$

Therefore,

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

So, for a dilute solution, n_2 can be ignored in denominator, hence

$$\frac{\Delta P}{P^\circ} = \frac{n_2}{n_1}$$

As we know, $n_1 = W_1 / M_1$ and $n_2 = W_2 / M_2$

Hence, $$\frac{\Delta P}{P^\circ} = \frac{W_2 / M_2}{W_1 / M_1}$$

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

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

The molecular mass of a non-volatile solute (M_2) can be calculated by the above equation.

10.4.2 Elevation of Boiling Point of the Solvent

The normal boiling point is the temperature at which the vapour pressure of the liquid becomes equal to atmospheric pressure. Boiling point can be changed by changing either atmospheric pressure or vapour pressure of the liquid. The presence of non-volatile solute in the solution decreases the vapour pressure of the solvent and hence increase boiling point of solvent. This is known as elevation of boiling point of the solvent. The elevation of boiling point (ΔT_b) is the difference between the normal boiling point of pure solvent and elevated boiling point of a solution.

$$\Delta T_b = T_{b(solution)} - T_{b(solvent)}$$

Where, ΔT_b is the elevation in the boiling point of solution, $T_{b(solution)}$ is the boiling point of solution, and $T_{b(solvent)}$ is the boiling point of the pure solvent.

The elevation of boiling point (ΔT_b) is directly proportional to the concentration of the solute in the solution (molality).

$$\Delta T_b \propto m \qquad \text{or}$$

$$\Delta T_b = K_b m \quad \dots\dots\dots\dots\dots\dots\dots\dots \quad (i)$$

Where,

K_b is called molal boiling point constant or ebullioscopic constant.

We know,

$$\text{Molality (m)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in Kg}}$$

or $$m = \frac{W_2}{M_2} \times \frac{1}{W_1}$$

or $$m = \frac{W_2}{M_2 W_1}$$

By putting the value of 'm' in equation (i), we get

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

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

This equation can be used to determine the molar mass of a non-volatile non-electrolyte solute in a volatile solvent.

Example 10.9

Calculate the boiling point of the solution when 18.5g of phenol (C_6H_5OH) is dissolved in 500g of water. The boiling point of pure water is 100°C and the boiling point elevation constant for water is 0.51°C kg/mol.

Solution:

Elevation in the boiling point of the water $= \Delta T_b = ?$
Mass of phenol $= W_2 = 18.5g$
Molar mass of phenol (C_6H_5OH) $= M_2 = 72 + 5 + 16 + 1 = 94 g mol^{-1}$
Boiling point elevation constant for water $= K_b = 0.51°C\ kg\ mol^{-1}$
Mass of water $= W_1 = 500g = 0.5kg$

The equation for the calculation of ΔT_b is:

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

By putting the values, we obtain

$$\Delta T_b = 0.51°C\ kg\ mol^{-1} \times \frac{18.5g}{(94g\ mol^{-1})(0.5kg)}$$

$$\Delta T_b = \frac{9.435°C}{47}$$

$$\Delta T_b = 0.20°C$$

The boiling point of the solution is equal to the sum of the original boiling point and elevation in the boiling point of the solvent.

The boiling point of the solution $= 100.00°C + 0.20°C = 100.20°C$

Practice Exercise 9:

A solution is made in which 100g of ethylene glycol ($C_2H_6O_2$) is dissolved in 500g of water. The boiling point elevation constant for water is 0.51°C kg/mol. What is the boiling point elevation of the water? What is the boiling point of aqueous solution?

Keep in Mind

We use molality instead of molarity in boiling point elevation calculations because the molality of a solution does not change by changing temperature, but the molarity of solution changes.

Table 10.9: Molal Boiling Point Elevation Constants (K_b) for Some Common Substances

Substance	**Formula**	**Normal Boiling Point (°C)**	**Molal Boiling Point Constant, K_b (°C Kg/mol)**
Acetic acid	CH_3COOH	118.5	3.07
Acetone	CH_3COCH_3	56.0	1.71
Aniline	$C_6H_5NH_2$	184.4	3.22
Benzene	C_6H_6	80.15	2.53
Camphor	$C_{10}H_{16}O$	207.4	5.61
Carbon disulphide	CS_2	46.3	2.34
Carbon tetrachloride	CCl_4	76.5	5.02
Chloroform	$CHCl_3$	61.2	3.63
Diethyl ether	$CH_3CH_2OCH_2CH_3$	34.42	2.02
Ethyl alcohol	CH_3CH_2OH	78.26	1.22
Naphthalene	$C_{10}H_8$	218.0	5.65
Nitrobenzene	$C_6H_5NO_2$	210.9	5.24
Phenol	C_6H_5OH	181.8	3.56
Water	H_2O	100.0	0.51

Example 10.10

A solution consists of 270g of solute dissolved in 250g of water. The boiling point elevation and the boiling point elevation constant values for water are 3.06°C and 0.51°C kg/mol respectively.

Solution

Elevation in the boiling point of the water $= \Delta T_b = 3.06°C$

Mass of glucose $= W_2 = 270g$

Molar mass of solute $= M_2 = ?$

Boiling point elevation constant for water $= K_b = 0.51°C\,kg\,mol^{-1}$

Mass of water $= W_1 = 250g = 0.25kg$

The equation for the calculation of molar mass of solute is:

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

By putting the value, we have

$$M_2 = \frac{0.51^\circ\text{C kg mol}^{-1}}{3.06^\circ\text{C}} \times \frac{270\text{g}}{0.25\text{kg}}$$

$$M_2 = 180\text{g mol}^{-1}$$

Practice Exercise 10:

The boiling point of chloroform is increased by 2.42°C when 15.2g of camphor is dissolved in 150g chloroform. What is the molar mass of camphor? The value of K_b for chloroform is 3.63 °C kg/mol.

10.4.3 Depression in Freezing Point

The temperature at which the solid and liquid states of a substance exist in equilibrium with each other is called freezing point. It may also be defined as: the temperature at which its solid and liquid phases have the same vapour pressure is called freezing point.

When a non-volatile solute is added to a solvent, its vapour pressure is decreased. As a result of this freezing point of solvent decreases, this is known as depression of freezing point. At the freezing point, there are two things in the vessel; that is, liquid solution and the solid solvent. The solution will freeze at that temperature at which the vapour pressures of both liquid solution and solid solvent are same. It means that a solution should freeze at lower temperature than pure solvent.

The depression of a freezing point is directly proportional to the concentration (molality) of the solution.

$$\Delta T_f \propto m$$

or

$$\Delta T_f = K_f m \qquad \text{.........................(i)}$$

Where,

K_f is called molal freezing point constant or cryoscopic constant.

We know,

$$\text{Molality (m)} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}} \times \frac{1}{\text{Mass of solvent in kilogram}}$$

or

$$m = \frac{W_2}{M_2} \times \frac{1}{W_1}$$

or

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

By putting the value of 'm' in eq. (i), we have

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

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

The molar mass of solute can be calculated by this equation.

Table 10.10: Molal Freezing Point Depression Constants (K_f) for Some Common Substances

Substance	Formula	Normal Freezing Point (°C)	Molal Freezing Point Constant, K_f (°C Kg/mol)
Acetic acid	CH_3COOH	16.6	3.90
Acetone	CH_3COCH_3	–94.8	-
Aniline	$C_6H_5NH_2$	–6.1	5.87
Benzene	C_6H_6	5.5	5.12
Camphor	$C_{10}H_{16}O$	178.8	–39.70
Carbon disulphide	CS_2	–111.5	3.80
Carbon tetrachloride	CCl_4	–2.96	-
Chloroform	$CHCl_3$	–63.5	4.70
Diethyl ether	$CH_3CH_2OCH_2CH_3$	–116.3	1.79
Ethyl alcohol	CH_3CH_2OH	–114.6	1.99
Naphthalene	$C_{10}H_8$	80.2	6.90
Nitrobenzene	$C_6H_5NO_2$	5.7	5.12
Phenol	C_6H_5OH	40.9	7.27
Water	H_2O	0.0	1.86

Applications of Freezing Point Depression and Boiling Point Elevation

i) Ethylene glycol is used as antifreeze in automobile radiators because it lowers the freezing point of the coolant. In cold climates, it prevents the solution in the radiators, coolant lines and engine blocks from freezing. Besides lowering the freezing point, it elevates the boiling point of the solution when added to water and prevents the water in the radiator from boiling off in hot summer days and the engine is not overheated.

ii) A salt such as NaCl or $CaCl_2$ is sprinkled on icy roadways and walkways in the winter to lower the melting point of ice and snow so that the water on the roads will not freeze at normal freezing point of water (0°C).

iii) A freezing mixture (Ice + NaCl or KNO_3) is used in making ice cream and kulfa. A freezing mixture lowers the freezing point of the solution well below the freezing point of pure water and freezes the ice cream rapidly.

iv) Sea water has freezing point below 0°C because it is the solution and contains higher concentration of various salts.

Example 10.11

A salt solution is prepared by adding 58.5g of NaCl in 0.75kg of water. Determine the freezing point of the solution. The freezing point constant of water is 1.86°C kg/mol and the normal freezing point of water is 0.00°C.

Solution:

Depression in the freezing point of the water $= \Delta T_f = ?$
Freezing point of solution $= ?$
Mass of sodium chloride $= W_2 = 117g$
Molar mass of sodium chloride (NaCl) $= M_2 = 23 + 35.5 = 58.5 g\,mol^{-1}$
Freezing point depression constant for water $= K_f = 1.86°C\,kg\,mol^{-1}$
Mass of water $= W_1 = 0.75kg$
The equation for the calculation of ΔT_f is:

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

By putting the values, we get

$$\Delta T_f = 1.86°C\ Kg\ mol^{-1} \times \frac{117g}{(58.5gmol^{-1})(0.75kg)}$$

$$\Delta T_b = \frac{217.62°C}{43.87}$$

$$\Delta T_b = 4.96°C$$

The freezing point of the solution is equal to the difference of the original freezing point and depression in the freezing point of the solvent.
The freezing point of the solution $= 0.00°C - 4.96°C = -4.96°C$

Example 10.12

A solution is prepared in which 51.2g of sulphur is placed in 0.22kg of carbon disulphide. i.e. ΔT_f value for carbon disulphide is 3.45°C and K_f value for carbon disulphide is 3.80°C kg/mol. What is the molar mass of sulphur?

Solution:

Depression in the freezing point of the carbon disulphide $= \Delta T_f = 3.45°C$

Mass of sulphur $= W_2 = 51.2g$

Molar mass of sulphur $= M_2 = ?$

Freezing point depression constant for carbon disulphide $= K_f = 3.80°C\,kgmol^{-1}$

Mass of carbon disulphide $= W_1 = 0.22kg$

The equation for the calculation of molar mass of sulphur is:

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

By putting the value, we have

$$M_2 = \frac{3.80°C\,kg\,mol^{-1}}{3.45°C} \times \frac{51.2g}{0.22kg}$$

$$M_2 = 256g\ mol^{-1}$$

Practice Exercise 11:

An aqueous solution of glycerin contains 64.4g of glycerin in 0.644kg of water. What is the molar mass of the glycerin? The ΔT_f value for water is 2.02°C and K_f value for water is 1.86°C kg/mol.

Practice Exercise 12:

Antifreeze contains 300g of ethylene glycol in 700g of water. What is the freezing point of antifreeze? The molar mass of ethylene glycol is 250g. The freezing point constant of water is 1.86°C kg/mol and the normal freezing point of water is 0.00°C.

10.4.4 Osmotic Pressure and Reverse Osmosis

Osmosis is the spontaneous process whereby the solvent molecules pass through a semipermeable membrane from a solution of lower concentration of solute (higher concentration solvent) into a solution of higher concentration of solute (lower concentration of solvent). Osmosis takes place, when two solutions of different concentrations (a pure solvent and a solution) are separated by semipermeable membrane. A semipermeable membrane is a thin layer of material that allows solvent particles to pass through it in the both directions and prevent solute particles to pass through. The rate of passage of solvent particles from the pure solvent side toward the solution side is high enough, because of this, the quantity of liquid on the pure solvent side decreases and the quantity of liquid on the solution side increases; hence the concentration of solution decreases. Animal bladder, cell membranes in the cells of living things (plants and animals) and cellophane (a polymer derived from cellulose) are the examples of semipermeable membranes. Consider an experiment in which a 10% concentrated sugar (glucose) solution is taken in a thistle funnel whose mouth is sealed with a semipermeable membrane. The funnel containing sugar solution is then place in a beaker of pure water. In the start the rate of flow of solvent (water) from beaker into the funnel (sugar solution) is greater than the rate of flow of solvent from funnel to beaker. The solvent starts flowing from the beaker through the semipermeable membrane into the sugar solution side (funnel).

Overtime, the volume of the sugar solution increases in the stem of the funnel while the water level in the beaker decreases. The process remains in operation until the pressure created by the weight of the water on the semipermeable membrane is enough to stop the upward flow of water. At this stage the rate of flow of water from beaker to funnel becomes equal to the rate of flow of water from funnel to beaker and the solution level stops rising and the equilibrium is established. Now the solvent flows at an equal rate in both directions across the semipermeable membrane and the volume of the liquid on each side of the membrane does not change. The amount of pressure required to attain equilibrium is called osmatic pressure. Osmotic pressure is defined as the pressure that must be applied on the high solute concentration side to stop the flow of solvent through a semipermeable membrane. It is denoted by capital pi (π). The osmotic pressure of the solution is given by:

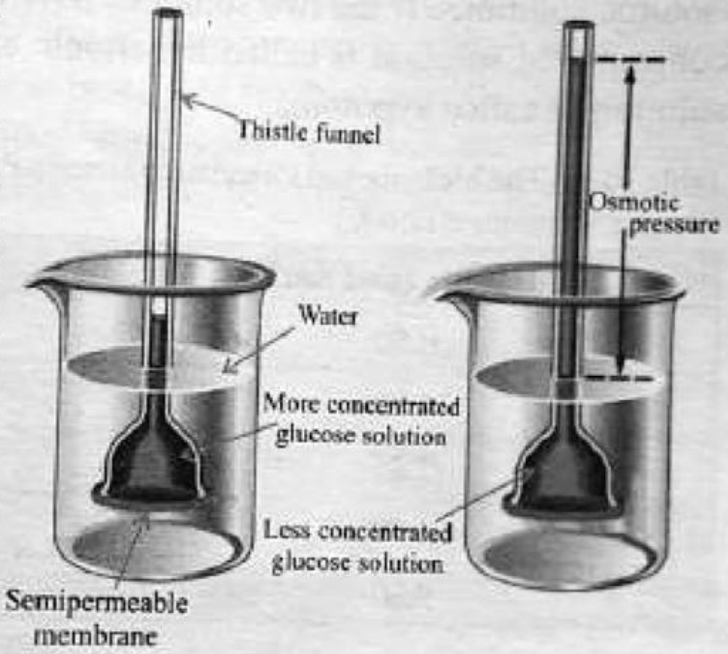


Figure: 10.7: Demonstration of Osmotic Pressure

$\pi = (n/V)RT$ or

$\pi = MRT$

Where, M is the molarity of solution, R is the general gas constant (0.08206atm dm^3/mol K) and T is the absolute temperature. The osmotic pressure is expressed in atmospheres. The osmotic pressure increases by rise in the concentration of solute particles in the solution at a specific temperature. It also increases by rise in temperature and decrease by fall in temperature. Osmotic pressure depends on the number of solute particles in the solution but not on the type of solute particles. The two different solutions of equal concentration have the same osmotic pressure. The solutions that have the same osmotic pressures are called isotonic solutions. If the two solutions have different osmotic pressures, the more concentrated solution is called hypertonic and less concentrated solution (dilute solution) is called hypotonic.

Table 10.11: The Molarity and Calculated Osmotic Pressure Values of Variable Concentrations of Glucose Solutions at (298K)

Molarity (mol/dm^3)	Osmotic Pressure (atm)
0.50	12.23
1.00	24.45
2.25	55.02
4.50	110.04
9.00	220.08

Some of the Interesting Examples of Osmosis

i) The leaves of plants give off water vapours that lead to the rise in solute concentration in the leaf fluids. Water rises up from soil into plant leaves through roots, trunk, branches and stem creating an osmotic pressure.

ii) If plants are watered with salt water whose concentration is greater than the solution in the root, the roots will lose water rather than absorb it. Hence, the plants become dehydrated and sooner or later the plants die.

iii) The meat can be preserved by treating its surface with salt. The meat becomes dehydrated and does not go bad and remain in a good condition for a long time and there is no need of refrigeration for meat. The salt present on the surface of meat produces a concentrated solution which is hypertonic to bacteria cells. The bacteria on the surface of salted meat then tend to shrivel up and die.

iv) You are advised never drink excessive amount of sea water, even if your boat is badly damaged and you are in the middle of the sea. When sea water (concentrated salt solution) is taken into stomach and intestine, the sea water extracts the w out of the stomach and intestine and the tissues become dehydrated. As a consequence of dehydration thirst increases but not decreases because the body will lose water instead absorb it.

Reverse Osmosis

The reverse of osmosis is said to be reverse osmosis. In normal osmosis, the solvent flows through a semipermeable membrane from a lower solute concentration to one of higher solute concentration but in reverse osmosis, the pressure greater than osmotic pressure is applied to flow the solvent through a semipermeable membrane from a higher solute concentration to one of lower solute concentration. The process of reverse osmosis is used to purify water. This process is particularly used to remove contaminates from sea water to make it fit for drinking. The world largest desalination plants are present in Saudi Arabia that treat salt water by reverse osmosis to produce fresh water and fulfill at about 50% requirements of the country.

10.5 Colloids

10.5.1 Types of Colloids

There are eight types of colloids on the basis of phase (solid, liquid, or gas) of the dispersed substance and of the dispersing medium.

i) Liquid aerosols may be formed when liquid droplets are dispersed in a gas. Examples of liquid aerosols are fog, mist, and aerosol sprays.

ii) Solid aerosols may be formed when solid particles are dispersed in a gas. Examples of solid aerosols are dust, and smoke in air.

Keep in mind

A colloid (or colloidal dispersion) is a homogeneous mixture of two or more components like solution.

iii) Liquid foams may be formed when a gas is dispersed in a liquid. Examples of liquid foams are beaten egg white, and shaving lather.

iv) Solid foams may be formed when solid particles are dispersed in a liquid. Examples of solid foams are Styrofoam, and marshmallow.

v) Emulsions may be formed when liquid particles are dispersed in another liquid. Examples of emulsions are mayonnaise, and face cream.

vi) Gels may be formed when liquid particles are dispersed in a solid. Examples of gels are butter, and jellies.

vi) Liquid sols may be formed when solid particles are dispersed in a liquid. Examples of liquid sols are pudding, and ink.

vii) Solid sols may be formed when solid particles are dispersed in a solid. Examples of solid sols are steel, and pearls.

Table 10.12: Types of Colloids and their examples

State of Dispersed Phase	State of Dispersing Medium	State of Resulting Colloid	Name	Example
Gas	Liquid	Liquid	Foam	Whipped cream, shaving lather, soap lather, beaten egg white
Gas	Solid	Solid	Solid foam	Styrofoam, marshmallow, sponge, rubber
Liquid	Gas	Gas	Aerosol	Mist, fog, clouds, sprays
Liquid	Liquid	Liquid	Emulsion	Mayonnaise, milk, face cream
Liquid	Solid	Solid	Gel	Jelly, Cheese, Butter, gelatin
Solid	Gas	Gas	Aerosol	Smoke, airborne viruses
Solid	Liquid	Liquid	Sol	Paint, ink, puddings
Solid	Solid	Solid	Solid sol	Pearls, opals, ruby, porcelain, coloured glass, some metal alloys such as steel

Colloids in which the dispersing phase is water may be classified as hydrophilic or hydrophobic. A colloid in which there is strong attractive force between dispersed phase and the dispersing medium (water) is known as hydrophilic colloid. The examples of hydrophilic colloids are gelatin in water, and jellies. A colloid in which there is a very weak attractive force between dispersed phase and the dispersing medium (water) is known as hydrophobic colloid. The examples of hydrophobic colloids are milk and mayonnaise.

10.5.2 Properties of Colloids

Some of the important properties of colloids are as follows:

I) Colloids are also homogeneous like that of solutions but borderline; that is, the particles of dispersed phase may be uniformly distributed among the particles of dispersing medium. A colloid is also known as colloidal dispersion because the particles of dispersed phase remain dispersed in the dispersing medium but not dissolved.

ii) The colloidal suspensions are often translucent or opaque but may be transparent.

iii) Colloidal particles are intermediate in size; that is, the size of colloidal particles is greater than the size of the particles of the true solutions and smaller than particles of suspensions that can be seen by naked eye. Colloidal particles have dimensions ranging from approximately 2nm to 1000nm. Particles with diameters 0.01nm to 1nm are true solution particles and those larger than 1000nm are suspension particles.

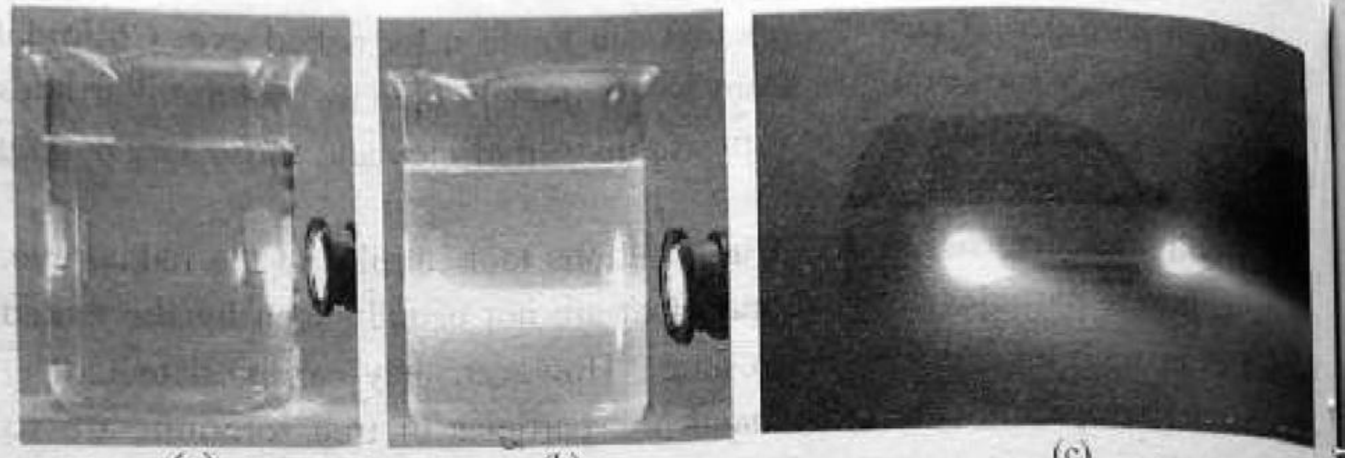
iv) The colloidal suspensions and true solutions look similar to the naked eye. But the colloidal particles are too small that are not usually seen by the naked eye and even with an ordinary microscope. However, they can be detected by electron microscope. On the other hand, the particles of true solutions are so small that they cannot be seen by the naked eye and microscope but the particles of suspensions are so large that can be seen with microscope and can often be seen with the naked eye.

v) The particles of colloidal suspensions are not settling out under the influence of gravity because their particles are in constant random motion, called Brownian motion, which keeps them from settling. This is due to small sized-suspended particles of colloids. The motion of colloidal particles increases by rise in temperature and decreases by increase the viscosity of the colloidal suspension. The motion of colloidal particles can be detected by ultra-microscope.

vi) The diffusion of colloidal suspensions is very slow due to large sized-suspended particles of dispersed phase than that of solution which is fast due to small size of solute particles. The suspensions do not diffuse. Diffusion is due to Brownian motion of particles which is more rapid for small sized-particles.

vii) Colloidal suspensions pass through filter paper but do not pass through semipermeable membrane.

viii) One of the most important properties that distinguish colloids from solutions is its ability to scatter light. When a beam of light is passed through a true solution, we cannot see the path of the beam of visible light because the light is not scattered due to small size of solute molecules. However, when a beam of light is passed through a colloidal suspension, we can see the path of the beam of visible light because the light is scattered by the dispersed phase. This phenomenon was first observed by the Irish physicist John Tyndall (1820-1893), and is termed as the Tyndall effect. Consider the examples of Tyndall effects. The visibility of headlights of cars is not better in the fogging weathers because the colloidal-sized particles in the fog scatter the light. The particles of dust can easily be seen in a beam of sunlight in the air of a room because the colloidal-sized dust particles



(a) (b) (c)

Figure 10.8: (a) Light passing through solution, (b) Light Scattering by Colloidal Suspensions, (c) Light Scattering by Colloidal Suspensions

scatter the light. The sky looks blue because the colloidal-sized particles in the air scatter the light.

Some of the properties of solutions, colloids and suspensions are compared in the table 10.13.

Table 10.13: properties of solutions, colloids and suspensions

Property	Solution	Colloid	Suspension
Type of mixture	Homogeneous	Homogeneous	Heterogeneous
Size of particles	0.01 nm to 1 nm	2 nm to 1000 nm	Greater than 1000 nm
Effect of light	Cannot scatter light	Scatter light	Scatter light
Transparency	Transparent	Cloudy (translucent)	Opaque
Settling properties of particles	Do not settle down on standing	Do not settle down on standing	Settle down on standing
Filterability	Non-filterable	Non-filterable	Filterable
Separation	The solute can be separated from the solvent by evaporation and distillation.	The dispersed phase can be separated from dispersing medium by semipermeable membrane; that is, cellophane and cell walls	Can easily be separated by filtration
Visibility of particles	Invisible by naked eye and microscope	Invisible by naked eye and with an ordinary microscope	Visible by naked eye

Summary of Facts and Concepts

- Solution is a homogeneous mixture of two or more substances that may be gases, liquids or solids. The component of solution that is in larger amount is the solvent. The component of solution that is in smaller amount as compared to solvent is called solute. The solution of substances in water is called aqueous solution (aqua means water). Liquids solutions may be coloured or colourless but clear and transparent. The solute particles of solutions are very small and cannot be seen by the naked eye and microscope.
- A suspension is a heterogeneous mixture and the particles are so large that can be seen with microscope and can often be seen with the naked eye. The particles of suspensions are large enough that they settle out under the influence of gravity.
- A colloid is a homogeneous mixture of two or more components like solution. The dispersed particles are intermediate in size between those of a true solution and suspension and do not settle out under the influence of gravity. The dispersed particles of colloids are so small that are not usually seen by the naked eye and ordinary microscope. But can be detected by electron microscope.
- The solubility of solute is the maximum amount of solute dissolved in given amount of solvent to make a saturated solution at a given temperature. The solubility of solids and liquids in a solvent is usually increases by increase in temperature. In contrast, the solubility of gases decreases by increasing temperature. The solubility of gases increases by increasing the pressure over the solution of the gas.
- Concentration is the amount of solute dissolved in the given amount of solvent or solution at a given temperature. The concentration of the solution can be expressed by percentage composition, molarity, molality, mole fraction, parts per million, parts per billion, and parts per trillion.
- A substance which has measureable vapour pressure is said to be volatile while a substance which has no measureable vapour pressure is said to be non-volatile.
- Osmosis is the flow of solvent molecules through a semipermeable membrane from a low concentrated solution to a high concentrated solution.
- Osmotic pressure is the pressure that is applied on the high solute concentration side to prevent osmosis from happening.
- In *reverse osmosis*, the pressure on the more concentrated side is applied to pass the solvent molecules but not the solute particles through a semipermeable membrane. This process is used for desalination of sea water.

Questions and Problems

Q.1. Four answers are given for each question. Select the correct one:

i) One molar solution has one mole of solute in the solution of:

(a) 1000 ml (b) 1L

(c) both a and b (d) 1000g

ii) Which one of the following is more concentrated:

(a) 0.1M (b) 0.25M (c) 0.5M (d) 0.75M

iii) Which one of the following is the example of colloid:

(a) Ink (b) milk of magnesia

(c) glucose solution (d) soda water

iv) The mole fraction of ethyl alcohol in the solution consisting 23g ethyl alcohol and 3.2g methyl alcohol is:

(a) 0.17 (b) 0.37 (c) 0.63 (d) 0.83

v) Which one of the following is **NOT** soluble in water:

(a) Benzene (b) phenol

(c) ethyl alcohol (d) table salt

vi) Which one of the following is a colligative property:

(a) Melting point (b) density

(c) boiling point (d) elevation in boiling point

vii) Elevation of boiling point is directly proportional to:

(a) molal concentration of solution

(b) molal concentration of solvent

(c) molar concentration of solvent

(d) molar concentration of solution

viii) The solubility of which one of the following substances increases by rise in temperature:

(a) CO_2 (b) $Ce_2(SO_4)_3$ (c) KNO_3 (d) O_2

ix) The solution cools when water is added:

(a) KNO_3 (b) $CuSO_4$ (c) H_2SO_4 (d) LiCl

x) Which of the following can **NOT** scatter light:

(a) sugar in water (b) milk of magnesia

(c) starch solution (d) fog

Q.2. Fill in the blanks with suitable words given in the brackets:

i) The oxygen gas is __________ soluble in water. (completely / slightly)

ii) The solubility of the gas is __________ by doubling the partial pressure of the gas above the solution. (doubled/quadrupled)

iii) Sponge is an example of __________. (suspension/colloid)

iv) The component of solution which is in __________ amount is called solute. (larger/smaller)

v) The concentration of solution of 6g of urea per liter is __________. (0.1 M/1 M)

vi) The relative lowering of vapour pressure is __________ proportional to the mole fraction of solute. (directly/inversely)

vii) 15% W/W sugar solution contains 15g of sugar in 100g of __________. (solvent/solution)

viii) __________ is a process in which an ion or molecule is surrounded by solvent molecules. (hydration/solvation)

ix) The freezing mixture, which is used to freeze the ice cream, is prepared when __________ is mixed with ice. (KNO_3/$AgNO_3$)

x) A semipermeable membrane is a thin layer of material that allows __________ particles to pass through it. (solute/solvent)

Q.3. Label the following sentences as True or False:

i) AgCl is completely soluble in water.

ii) Glycerin is used as antifreeze in the radiators of automobiles.

iii) The amount of NH_4Cl needed to make 500mL of 0.25M solution is 6.69g.

iv) We can see the path of visible light when a beam of light is passed through a colloidal suspension.

v) Beaten egg white is the example of liquid foam.

vi) The impurities of the substances are usually expressed in parts per hundred.

vii) 20% V/V alcohol-water solution means that 20mL of alcohol dissolved in enough water to make the total volume of solution 100mL.

viii) The heat is absorbed when sulphuric acid is dissolved in water.

ix) A solution which can dissolve further solute on addition is known as saturated solution.

x) The completely miscible solutions can usually be separated by fractional distillation.

Q.4: Define solution. What are the general properties of solutions?

Q.5: Distinguish between hydrophilic and hydrophobic molecules. What is the function of soaps and detergents?

Q.6: What are liquid solutions? Explain in your own words with examples.

Q.7: What are the different types of liquid solutions? Explain in detail.

Q.8: What is the meaning of "like dissolves like"? Explain with examples.

Q.9: Explain why table sugar is dissolved in water but not in benzene?

Q.10: What is the effect of temperature on the solubility of liquids in liquids and solids in liquids?

Q.11: What is the effect of the pressure on the solubility of gases in liquids?

Q.12: What is the effect of the temperature on the solubility of gases in liquids?

Q.13: What happens:

(a) When the pressure of CO_2 gas above the solution of carbonated beverage is increased and

(b) When the temperature of the carbonated beverage is increased?

Q.14: Why we drunk carbonated beverages only cold but not hot?

Q.15: Why does a bottle of carbonated beverage effervescence when you uncapped the bottle?

Q.16: The carbonated beverage is kept in closed vessel, why?

Q.17: Discuss and explain the heat of solution. The heat of solution is either negative or positive for solution, explain.

Q.18: What do you think, the rise in temperature will increase or decrease the solubility of NH_4Cl and NaOH in water?

Q.19: What do you know about concentration units of solution? Describe the compositions of solutions in terms of weight by weight, weight by volume, volume by weight, and volume by volume percent.

Q.20: What do you mean by molarity and molality of solution? Give their mathematical equations to calculate molarity and molality of solutions.

Q.21: The molarity of solution is affected by temperature change and molality of solution is not affected by temperature change..Explain.

Q.22: Why the sum of the volumes of individual components is not necessarily equal to the total volume of solution?

Q.23: What is mole fraction? Why the sum of all the components present in the solution is equal to one?

Q.24: Describe and explain the concentration units, parts per million, billion, and trillion.

Q.25: Explain Raoult's law with the help of three statements when solute is non-volatile and solvent is volatile.

Q.26: Explain Raoult's law when both of the components of solution; that is, solute and solvent are volatile.

Q.27: A non-ideal solution shows two types of deviations from Raoult's law; that is positive and negative. Explain with the help of graphs.

Q.28: What is the difference between ideal and non-ideal solutions?

Q.29: What is the meaning of the term colligative? Define the term colligative property and name the four colligative properties of solution. What are the conditions for colligative properties?

Q.30: The vapour pressure of a solution containing a non-volatile solute is always less than that of pure solvent, justify.

Q.31: Why is the vapour pressure of sea water less than that of pure water?

Q.32: How can you calculate the molar mass of non-volatile and non-electrolyte solute in a volatile solvent?

Q.33: Explain freezing point depression and boiling point elevation in your own words. Give applications of freezing point depression and boiling point elevation.

.34: Explain osmosis, osmotic pressure, and reverse osmosis. Give their daily life applications.

.35: What is colloid? Describe the types and properties of colloids.

.36: What is the difference between solution, colloid and suspension?

.37: Explain why a solution and a colloid are considered as homogeneous mixtures while suspension is considered as heterogeneous mixture?

38: What is the Tyndall effect? Give two examples of this effect.

39: Define the terms solute and solvent. Identify the solute and solvent in the solutions given below:

(a) 10g of sugar and 100g of water
(b) 5g of oxygen and 20g of nitrogen
(c) 5mL of water and 30mL of ethanol
(d) Air
(e) Soda water

40: Calculate the weight percent of gold ring containing 2g of platinum and 7g of gold.

41: Calculate the W/V % of 0.5 liter of soft drink that has 55g of dissolved carbon dioxide gas.

Q.42: A cup of 200mL of hot milk tea contains 50mL of milk. What is the V/V% of the tea?

Q.43: A solution contains 12.25g of $KClO_3$, dissolved in 5×10^3 g of water. What is the molality of the solution?

Q.44: How many grams of calcium chloride are required to prepare 0.01M and 0.025M solutions?

Q.45: If 15g of urea (N_2H_4CO) is dissolved in 5×10^2mL of solution, calculate the molarity of the urea solution.

Q.46: Calculate the molality of each of the following solutions containing:

(a) 0.1 mol of solute in 0.25kg of solvent

(b) 0.25 mol of solute in 0.1kg of solvent

Q.47: Calculate the molarity of each of the following solutions containing:

(a) 1.50 mol of NaCl in 0.50L of water

(b) 2.25 mol of rock candy in 4.50L of water.

Q.48: A solution is prepared by dissolving 9.2g of toluene (C_7H_8) in 39g of benzene (C_6H_6). What is the mole fraction of toluene in this solution?

Q.49: The bottle of one liter mineral water contains 78mg calcium, 24mg magnesium, 5mg sodium, and 1mg potassium. Calculate the concentration of each mineral in the mineral water in ppm, ppb, and ppt.

Q.50: What is the boiling point of the solution that contains 23.25g of aniline ($C_6H_5NH_2$) in 0.75kg of ethyl alcohol? The boiling point of pure ethyl alcohol is 78.26°C. The boiling point elevation constant for ethyl alcohol is 1.22°C kg/mol.

Q.51: A solution prepared by adding 11.30g of an organic compound in 0.1kg of ether has ΔT_b and K_b values for ether 3.06°C and 2.02°C kg/mol respectively. Find out the molar mass of the organic compound.

Q.52: Calculate the freezing point of sugar solution containing 855g of sucrose ($C_{12}H_{22}O_{11}$) and 2.50kg of water. The K_f value for water is 1.86°C kg/mol.