

Lesson 1

1.1 Homogeneous And Heterogeneous Mixtures

At the end of this section, you should be able to:

- ☑ define the terms mixture, homogeneous and heterogeneous mixture, solute, solvent, solution;
- ☑ distinguish between homogeneous and heterogeneous mixtures;
- ☑ describe suspensions and colloids.

A **mixture** is a substance made up of two or more pure substances (elements or compounds) physically mixed together in any proportion. It has the following characteristics

- ☞ Its composition is variable
- ☞ Individual properties of the components are observable. Each substance retains its own chemical identity and hence its properties
- ☞ Its components can be separated by a simple physical means
- ☞ Its properties would be the average properties of its components

There are two broad classes of mixtures: homogeneous and heterogeneous

1. Heterogeneous Mixture

A heterogeneous mixture is a type of mixture whose constituents (individual components) can be recognized or seen by our naked eyes or a magnifying glass

- ☞ Has more than one phase. It has one or more visible physical boundaries between components with different properties
- ☞ Its composition is not uniform throughout.

Examples include: Milk, Blood, Rock, Soil, Petroleum, Coal, Ocean, Ink, Oil in water, etc.

2. Homogeneous Mixture (Solution)

A **homogeneous mixture (also known as a solution)** is a type of mixture whose constituents (individual components) cannot be seen by our naked eyes or a magnifying glass

- ☞ It has a uniform composition and properties throughout.
- ☞ It has a single phase-no physical or visible boundaries among its components

Examples include: Air, Syrup, Soft drink, Alloy, Salt in water, Alcohol in water, etc.

- ☞ A typical solution consists of at least two substances: Solute and solvent

2.1 Solute - is a substance dissolved in another substance called a solvent

- ☞ is usually a solid but also could be a liquid or a gas
- ☞ there could be more than one solute component in a solution

2.2 Solvent - is a substance which does the dissolving or in which the solute(s) dissolve

- ☞ is usually a liquid (commonly water) but also a gas or solid
- ☞ only one of the components of the solution is usually taken as a solvent
- ☞ if all the substances in a solution are in the same physical state, the substance present in larger amount is usually called the solvent

Example: In a solution containing 70% ethanol and 30 % water which both are liquids, the solvent is an alcohol (ethanol) and the solute is water

- ☞ the solvent is often but not always present in excess or large amount

Example: In a solution containing 40 g of sugar(solid) and 30 g of water(liquid), the solvent is water and the solute sugar

Note: *Solutions may be colorless or colored but transparent; the dissolved molecules or ions are invisible, and they do not settle out on standing.*

Note: *In this unit we focus on 3-common types of mixture*

- ☞ *Solution (A homogeneous mixture)*
- ☞ *Suspension and colloids (heterogeneous mixtures)*

1.1.1 Suspension And Colloids

Suspension - *is a heterogeneous mixture from which some of the particles (Components) will settle down (sediment) slowly (gradually) to the bottom of the beaker (container) upon standing the mixture to stand for sometime*

- ♣ Is a dispersion of solid particles in a liquid or gas, removable by filtration. Suspensions Are cloudy and are usually translucent
- ♣ Examples include: mud in water, flour and water, chalk in water, clay and water, paint and water, fine soot or dust in air, droplets of oil air, milk of magnesia (anti acid, $\text{Mg}(\text{OH})_2$ shaken in water) and sand in water

What is the difference between suspensions and solutions?

Colloids: *are heterogeneous mixtures that exhibit properties intermediate between those of suspensions and solution.*

- ♠ A Colloidal system may be solid, liquid or gaseous. Particles in colloid can not be retained by filter paper and do not settle out with time
- ♠ Most colloids are cloudy (milky) or opaque, but some are transparent to the naked eye.
- ♠ Examples include: blood, milk, paint, glue, aerosol sprays, smoke, cloud, fog, mist, shaving cream, rubber, cheese, butter, alloys (coins), and pigmented plastics

(Dear student you are kindly requested to do Exercise 1.1 on page 4 in your text book)

<i>Table 1.2: Properties Of Solutions, Colloids And Suspensions</i>				
<i>No</i>	<i>Property</i>	<i>System</i>		
		<i>Solution</i>	<i>Colloid</i>	<i>Suspension</i>
<i>1.</i>	<i>Particle Type</i>	<i>Ions, atoms, small molecules</i>	<i>Large molecules or particles</i>	<i>Large particles or aggregates</i>
<i>2.</i>	<i>Particle size (approximate)</i>	<i>0.1-1 nm</i>	<i>1-100 nm</i>	<i>100 nm and larger</i>
<i>3.</i>	<i>Effect of light</i>	<i>No scattering</i>	<i>Exhibit Tyndall effect</i>	<i>Exhibit Tyndall effect</i>
<i>4.</i>	<i>Effect Of Gravity</i>	<i>Stable, does not separate</i>	<i>Stable, does not separate</i>	<i>Unstable, sediment forms</i>
<i>5.</i>	<i>Filtration</i>	<i>Particles not retained on filter</i>	<i>Particles not retained on filter</i>	<i>Particles retained on filter</i>
<i>6.</i>	<i>Uniformity</i>	<i>Homogeneous</i>	<i>Borderline</i>	<i>Hetrogeneous</i>

1.2 Type of solutions

(N:B Before studying the following note please try to do Activity 1.1)

Solutions are homogeneous mixtures that can be grouped according to their “physical states”. They exist in any of the three states of matter (solid, liquid and gaseous). The possible physical states of solute and solvents, determines the class of these solutions

Note: In most cases (or usually) the physical state of the solvent determines the type (class) of the solution

1. Gaseous Solutions: Non-reactive gases can mix in all proportions to give a gaseous mixture. Gaseous solutions results when a gaseous salute dissolves in a gaseous solvent.

2. Liquids solutions: Most liquid solutions are obtained by dissolving a gas, liquid or solid solute in a given liquid solvent. They have gaseous, liquid and solid solutes but the solvent is always a liquid.

♣ **Mineral water** – contains CO₂ (gas solute) and water as a solvent

3. Solid solutions: have no restriction on the state of a solute (could be: gas, liquid or solid) but the solvent has to be the solid

♣ **Bronze** – an alloy of 25% tin solute and 75% copper solvent is a solid solution

TABLE 1.3: Examples of Solutions

State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold

Lesson 2

1.3 The Solution Process And Solubility

At the end of this section, you should be able to: •

- ☑ explain how the “like dissolves like” rule depends on inter-particle forces of interactions and predict relative solubility;
- ☑ define rate of dissolution;
- ☑ define heat of solution, solvation energy and hydration energy;
- ☑ apply the concept of heat of solution to a solution of ammonium nitrate and sodium hydroxide crystals;
- ☑ explain how heat of solution is influenced by the inter-particle interaction forces.

(N:B Before studying the following note please try to do Activity 1.2 on page 8 in your text book)

What are the major factors that affect the solubility of substances in liquid solvents?

1.3.1 Solubility

The solubility of a substance is the maximum amount of a substance (or a solute) that dissolves in a given quantity of a solvent at a given temperature to produce a saturated solution.

Solubility is usually expressed in grams of solute per 100 grams of solvent. For example, the statement “solubility of sugar is 204 g at 40 °C” means that at 40° C, 204 g of sugar dissolves in 100g of water. Common units are: g/lit, g/ml, mol/lit

Note: *A substance is said to be soluble in a given water solvent if 1g or more solute dissolves in 100ml of H₂O*

Note: *The solubility of one substance in another is determined by three major factors*

- ♣ *The strength of intermolecular forces between species (molecules or ions)*
- ♥ *Energy (heat of solution)*

♠ **Entropy (disorder)**

1.3.2 Intermolecular forces: are attractive forces between 2 or more molecules, ions, ions and molecules (same or different).

- ☑ Are attractive forces that holds 2 or more spears together. They could be : ion-dipole, dipole-dipole, ion-induced dipole, hydrogen bond, dipole-induced dipole and London dispersion force)
- ☑ Are not the actual chemical bonds (ionic, covalent, metallic) that holds atoms in a single compound
- ☑ Influence the physical properties of substances

1.3.3 The Solubility Rule (“like dissolves like”) - is a macroscopic (or an empirical observation stating that substances having similar kinds of intermolecular forces dissolve in each other. It often provides a sound qualitative way to predict solubility of a substance

- ♣ Polar solutes dissolve in polar solvents
- ♠ Non polar solutes dissolve in non polar solvents

Note:: From the knowledge of Intermolecular forces and Solubility rule we can predict:

- ♣ Which solute will dissolve in which solvent (Solubility) and
- ♠ The solution process (Process of solution formation)

Note: *Major factor that determine whether a solution forms or not (soln process) and the tendency (ability) to mix (dissolve) is the relative strength of the intermolecular forces within and between solutes or solvents. This is also affected by the relative strengths of 3-types of interactions*

- ❖ Solvent - Solvent interactions
 - ❖ Solute - Solute interactions
 - ❖ solvent - Solute interactions
- } *These are attractive interactions involved in solution process or solution formation*

Note the following Conclusions:

- 1. Strong forces of attractions between Solute particles (solute-solute interaction) or between solvent particles (solvent-solvent interaction)**
 - ♠ Tends to keep like particles close together and
 - ♣ Reduces the solubility of the solute in the solvent
- 2. On the other hand, strong attraction forces between solute and solvent particles**
 - ♠ makes the dissolving easier and
 - ♣ helps to keep particles in a solution
- 3. During the process of solution formation or when a solute is dissolved in a solvent, the:**

$$\left(\begin{array}{c} \text{Solute - Solvent} \\ \text{Interaction} \end{array} \right) > \left(\begin{array}{c} \text{Solute - Solute Interaction} \\ \text{And} \\ \text{Solvent - Solvent Interaction} \end{array} \right) \Rightarrow \text{The solution is formed}$$

$$\text{If the } \left(\begin{array}{c} \text{Solute - Solvent} \\ \text{Interaction} \end{array} \right) = \left(\begin{array}{c} \text{Solute - Solute Interaction} \\ \text{And} \\ \text{Solvent - Solvent Interaction} \end{array} \right) \Rightarrow \left(\begin{array}{c} \text{An Ideal solution Is formed} \\ \text{Where, the } \Delta H_{sol} = 0 \end{array} \right)$$

1.3.4 Solutions Of Liquids In Liquids

The “like dissolves like” principle is helpful in predicting the solubility of a substance in a given solvent. What this expression means is that two substances with intermolecular forces of similar type and magnitude are likely to be soluble in each other

1. Non Polar Solutes Dissolve In Non Polar Solvents (Nonpolar - Nonpolar)

Each pairs of molecules given below are completely miscible to each other, because both are non polar and experience only relatively weak London forces.

- ♣ Benzene (C₆H₆) and carbon tetrachloride (CCl₄)
- ♣ Benzene (C₆H₆) and toluene (C₇H₈)
- ♣ Carbon tetrachloride (CCl₄) and Octane (C₈H₁₈)

2. Polar solutes dissolve in polar solvents (Polar - Polar)

Each pairs of molecules given below are completely miscible to each other, because both are polar and both forms hydrogen bonding in each other

- ♣ Water (H₂O) and methanol (CH₃OH)
- ♣ Water (H₂O) and ethanol (C₂H₅OH)

3. Polar – Nonpolar : Each pairs of molecules given below are nearly immiscible to each other

- ♣ Water(H₂O) and octane (C₈H₁₈)
- ♣ Ethanol(C₂H₅OH) and carbon tetrachloride (CCl₄)

This is because water and ethanol are polar, whereas octane and carbon tetrachloride are non-polar. Hence they exhibit different intermolecular attractive forces

1.3.5 Solutions Of Solids In Liquids

In order for the solute to enter into a solution, the solute-solvent forces of attractions must be sufficient enough to overcome the strong attractive forces that hold the solid together.

Lesson three

1.3.6 Rate of Dissolution

What are the major factors that affect the rate of dissolution?

Rate of dissolution is the speed with which a solute goes in to a solution. It largely depends up on the inter-particle forces (the extent to which solvent molecules interact with particles of the solute)

and to a lesser extent on conditions such as surface area of the solid, temperature and pressure of the system

When the, $\left(\begin{array}{c} \text{Solute - Solvent} \\ \text{Interaction} \end{array} \right) > \left(\begin{array}{c} \text{Solute - Solute Interaction} \\ \text{And} \\ \text{Solvent - Solvent Interaction} \end{array} \right) \Rightarrow \left(\begin{array}{c} \text{the dissolution procees} \\ \text{becomes easier} \end{array} \right)$

Increasing surface area will increase the rate of dissolution because it increases the number of solute particles in contact with the solvent

1.3.7 Energy Changes In The Solution Process

The formation of a solution is often accompanied by either by absorption or a release of energy. It is accompanied by a temperature change

1. Heat of solution (Enthalpy Of Solution) = $\Delta H_{\text{Solution}}$

For simplicity we can imagine that “the solution process or the solubility of a solute in a solvent” is taking place in 3 distinct steps

Step - 1: Solute Particles Separate From Each Other.

♠ The step involves overcoming intermolecular attractions, so it is “Endothermic”

Solute (aggregated) + heat $\rightarrow$ Solute (separated), $\Delta H_{\text{solute}} > 0$, + Ve

Step - 2: Solvent Particles Separate From Each Other

♣ This step also involves overcoming attractions, so it is “Endothermic” too

Solvent (aggregated) + Heat $\rightarrow$ Solvent (separated), $\Delta H_{\text{solvent}} > 0$, + Ve

Step - 3: Solute And Solvent Particles Mix (Dissolve)

♣ the particles attract each other, so this step is “Exothermic “

Solute (separated) + Solvent (separated) $\rightarrow$ Solution + Heat, $\Delta H_{\text{mix}} < 0$, - Ve

The amount of energy that is absorbed or release when one mole of a substance (solute) enters into a solution is called heat of solution and is given a symbol, $\Delta H_{\text{solution}}$. It is obtained:

$$\Delta H_{\text{Solution}} = \Delta H_{\text{Solute}} + \Delta H_{\text{Solvent}} + \Delta H_{\text{Mix}}$$

Where: ΔH_{solute} is energy required (absorbed) to separate solute particles

$\Delta H_{\text{solvent}}$ is energy required (absorbed) to separate solvent molecule

ΔH_{mix} is energy required (released) to bind a solute and a solvent particles. It is also called salvation energy

Note: If the sum of the endothermic terms ($\Delta H_{\text{Solute}} + \Delta H_{\text{Solvent}}$)

a) is smaller than the exothermic term (ΔH_{mix}), then

$\Delta H_{solution}$ is negative, i.e., the process is exothermic

$$\Delta H_{Solute} + \Delta H_{Solvent} < \Delta H_{mix} = \Delta H_{Solution} = -Ve (< 0) = \text{exothermic (releases)}$$

b) is larger than the exothermic term (ΔH_{mix}), then

$\Delta H_{solution}$ is positive i.e., the process is endothermic

$$\Delta H_{Solute} + \Delta H_{Solvent} > \Delta H_{mix} = \Delta H_{Solution} = +Ve (> 0) = \text{endothermic (absorbes)}$$

Note: As the value of $\Delta H_{solution}$ becomes more and more positive, the solubility declines- the solute may not dissolve to any significant extent in that solvent

♣ Most soluble compounds have an exothermic (negative) $\Delta H_{solution}$

c) is equal to the exothermic term (ΔH_{mix}), then

$\Delta H_{solution}$ is zero i. e (the solution is ideal)

2. Hydration of Ionic Solids in Water And Heat Of Hydration (ΔH_{hydr}) (Dissolving Ionic Solids In water)

Solvation: is the binding of a solute (ion or molecule) to one or more solvent molecules, especially in a solvent other than water.

♣ If water is the solvent, solvation is called more specially hydration.

Hydration: is the attraction of solute ions for water molecule. Solvation in water is called hydration)

The Hydration of ions favors the dissolution of an ionic solid in water .The ions in an ionic solid (crystal) very strongly attracted to one another. Therefore, the solubility of an ionic solid depends on the following two factors

i) Energy of the hydration (Heat of hydration = (ΔH_{hydr}) of ions and

ii) Lattice energy (energy holding the ions together in the crystal lattice = ($\Delta H_{Lattice}$)

Lattice energies depend on the charge on the ions and also the distance between the centres of the neighboring positive and negative ions. As the magnitude of the charge on the ions increases the lattice energy also increases. For this reason, you can expect substances with single charged ions to be more comparatively soluble, and those with multiple charged ions to be less soluble in water

(Dear student you are kindly requested to do Activities 1.4, 1.5 and Exercise 1.3)

Lesson four

1.4 Solubility As An Equilibrium Process.

- ☑ *At the end of this section, you should be able to:*
- ☑ describe the distinctions among unsaturated, saturated and supersaturated solutions;
- ☑ prepare a supersaturated solution of sodium phosphate;
- ☑ explain the equilibrium nature of a saturated solutions;
- ☑ define solubility and describe the factors that affect the solubility of substances;

(N:B Before studying the following note please try to do Activity 1.6 on page 16 in your text book)

Imagine what take place on the molecular level when an ionic solid (solute) dissolves in a given solvent (liquid). Two opposing processes occur in a solution

- i) *Some dissolved ions leave the solid and become dispersed in the solvent*
- ii) *Some dissolved ions collide occasionally with un dissolved solute and re-crystallize*



Types of solution

1. Saturated Solution is a solution:

- ♣ in which the maximum amount of solute has been dissolved in a given solvent at a given temperature
- ♣ in which the dissolved and undissolved solute are in dynamic equilibrium.
- ♣ that cannot dissolve any more solute at a given temp.

Note: It is, of course possible to dissolve less solute than that needed to form a saturated soln. Such solutions are called unsaturated.

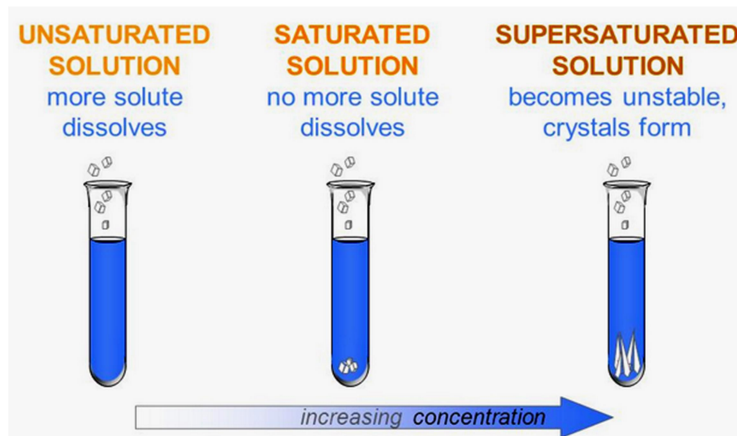
2. Unsaturated Solution

Is a solution before equilibrium is reached and contain (dissolve) less solute than saturated solution at a given temperature

Note: Under suitable conditions it is possible to form solns that contain a greater amount of solute than that needed to form a saturated soln. Such solns are said to be supersaturated.

3. **Supersaturated Solution** – is a solution containing more solute than saturated solutions. It is a state of a solution that contains more of the dissolved material than could be dissolved by the solvent under normal circumstances.

(Dear student you are kindly requested to do Activities 1.7)



Lesson five

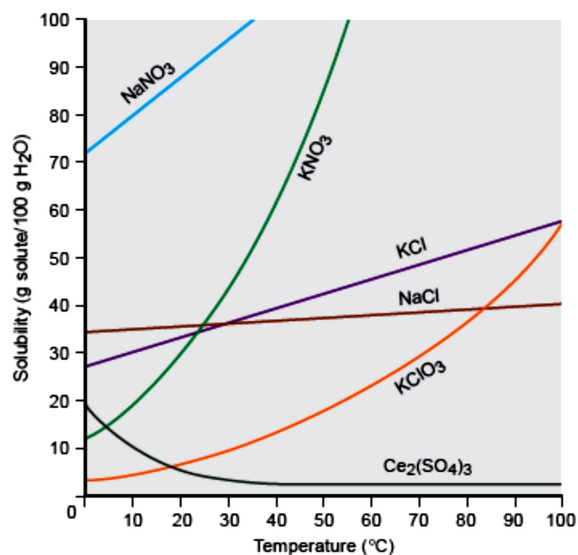
1.4.1 Effect Of Temperature On Solubility Of a Solute

(N:B Before studying the following lesson please try to do Activity 1.8 on page 20)

Temperature affects the solubility of most substances. You may have noticed, for example, that not only does sugar dissolve more quickly in hot tea than in iced tea, but *more* sugar dissolves; in other words, the solubility of sugar in tea is greater at higher temperatures. Let's examine the effects of temperature on the solubility of solids and of gases.

Temperature And The Solubility Of Solids (Solids In Liquids)

- ♣ *The solubility of most solids in a given solvent (like water) increases with increasing (rising) temperature. But the increasing solubility of these solids with temperature varies as shown in the figure below*



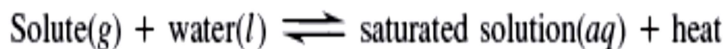
The Figure shows the solubility of several ionic compounds in water as a function of temperature.

KNO₃ shows a dramatic increase where as NaCl shows a slight increase

There are few solids that become less soluble as the temperature increases. Ce₂(SO₄)₃ is an example

Temperature and Gas Solubility in Water (Gas In Liquids)

The effect of temperature on gas solubility is much more predictable. When a solid dissolves in a liquid, the solute particles must separate, so energy must be added; thus, for a solid, $\Delta H_{\text{solute}} > 0$. In contrast, gas particles are already separated, so $\Delta H_{\text{solute}} \approx 0$. Because the hydration step is exothermic ($\Delta H_{\text{hydr}} < 0$), the sum of these two terms must be negative. Thus, for all gases in water, $\Delta H_{\text{soln}} < 0$:



1.4.2 The Effect Of Pressure On Solubility

Because liquids and solids are almost incompressible, pressure has little effect on their solubility; but has a major effect on gas solubility

When pressure increases, gas particles collide with the liquid surface more often. Thus, more gas particle enter than leave solution per unit time. *The solubility of gases increases with increasing pressure.*

- ♣ Henry's Law expresses the quantitative relationship between gas pressure and solubility: It states that: "The solubility of a gas (S_{gas}) is directly proportional to the partial pressure of the gas (P_{gas}) above the soln.

$$S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$$

where k_H is the Henry's law constant and is specific for a given gas-solvent combination at a given temperature. With S_{gas} in mol/L and P_{gas} in atm, the units of k_H are mol/L·atm.

PROBLEM The partial pressure of carbon dioxide gas inside a bottle of cola is 4 atm at 25°C. What is the solubility of CO_2 ? The Henry's law constant for CO_2 dissolved in water is 3.3×10^{-2} mol/L·atm at 25°C.

Given: Partial pressure of carbon dioxide gas, $P_{\text{CO}_2} = 4$ atm

Henry's Law constant for CO_2 dissolved in water (K_H) = 3.3×10^{-2} mol/L·atm

Required: Solubility (S_{CO_2}) = ?

Solution: Henry's Law equation can be applied in a straight forward fashion

$$S_{\text{CO}_2} = k_H \times P_{\text{CO}_2} = (3.3 \times 10^{-2} \text{ mol/L}\cdot\text{atm})(4 \text{ atm}) = 0.12 \text{ mol/L} = 0.12 \text{ M}$$

(Dear student you are kindly requested to do Exercise 1.4 page 23)

Lesson Six

1.5 Ways Of Expressing Concentrations Of Solution

At the end of this section, you should be able to:

- ☒ define terms like solution, concentration, mass percentages, ppm and ppb of a solute in a solution;
- ☒ calculate the mass percentage, ppm and ppb of a solute in a solution from given set of information;
- ☒ define mole fraction, molarity, equivalent mass, number of equivalents, normality and molality;
- ☒ calculate mole fraction, molarity, normality and molality of a solution from given set of information; and • *interconvert various concentration expressions.*

What does concentration mean? How do you determine it?

Concentration - is defined as the amount (quantity) of a solute present (or dissolved) in a given quantity of solvent or solution. It refers to the relative amount of solute and solvent present in a solution. The concentration of solution can be expressed either qualitatively or quantitatively.

1.5.1 Percent Concentrations

Chemists frequently express concentrations of solutions in terms of percentage (parts per hundred) of a solute. It is the percentage of solute relative to the final solution on either mass by mass (% m/m) or % w/w), volume by volume (%v/v) or mass by volume (% m/v). Out of these three, one of them is treated in this grade

1. Mass percentage (percent by mass, percent by weight or weight percent).

The percent by mass, also called percent by weight or weight percent, is the ratio of the mass of a solute to the mass of the solution, multiplied by 100.

$$\left(\begin{array}{c} \text{Percentage by mass} \\ \text{of a Solute} \end{array} \right) = \left(\frac{\text{mass of a solute}}{\text{mass of a solution}} \right) (100\%) = \% (W/W) = \% (m/m) \quad \text{Or}$$

$$\left(\begin{array}{c} \text{Percentage by mass} \\ \text{of a Solute} \end{array} \right) = \left(\frac{\text{mass of a solute}}{\text{mass of a solute} + \text{mass of a solvent}} \right) (100\%)$$

1.1 Parts per million (PP_m) and parts per billion (PP_b)

When the mass (concentration) of the solute in the solution is very small (like concentrations of pollutants (solutes) in air and water) a widely used expression is parts per million (PP_m) or parts per billion (PP_b).

$$\left(\begin{array}{c} \text{PPm} \\ \text{of a Solute} \end{array} \right) = \left(\frac{\text{mass of a solute}}{\text{mass of a solution}} \right) (10^6) \quad \text{and} \quad \left(\begin{array}{c} \text{PPb} \\ \text{Of a solute} \end{array} \right) = \left(\frac{\text{mass of a solute}}{\text{mass of a solution}} \right) (10^9)$$

☞ It is the numbers of parts by mass of solute per million (or per billion) parts by mass of the solution.

(Dear student you are kindly requested to do Exercise 1.5 page 25)

Lesson seven

1.5.2 Mole Fraction (x)

Mole fraction is defined as the ratio of moles of one components to the total no. of moles of all the components present in solution.

It is a unit less and dimension less quantity.

Mole Fraction formula :-

If a substance A of n_A moles is dissolves in solution of B having n_B , then mole fraction is expressed as denoted by greek letter χ (chi)

For two components

$$\chi^A = \frac{n_A}{n_A + n_B}$$
$$\chi^B = \frac{n_B}{n_A + n_B}$$
$$\chi^A + \chi^B = 1$$

Mole Fraction

where, n = moles

moles = Weight/MolarMass

Example Questions based on Mole Fraction Formula

Question 1

Find the mole fraction of Methanol CH_3OH and water in a solution prepared by dissolving 4.5 g of alcohol in 40 g of H_2O .

Molar Mass of H_2O is 18gm/mole and Molar mass of CH_3OH is 32gm/mole

Solution

Moles of $\text{CH}_3\text{OH} = 4.5 / 32 = 0.14$ mole

Moles of $\text{H}_2\text{O} = 40 / 18 = 2.2$ moles

Therefore, according to the equation

Mole fraction of $\text{CH}_3\text{OH} = 0.14 / 2.2 + 0.14 = 0.061$

Mole fraction of $\text{H}_2\text{O} = 2.2 / 2.2 + 0.14 = 0.940$

Question 2

Calculate the mole fraction of HCl and H_2O in a solution of hydrochloric acid in water, containing 30% HCl by weight.

Solution

The solution contains 30 grams of Hydrochloric acid and 70 grams of water. Also,

Molar mass of HCl is 36.5 grams/mole.

Molar mass of water is 18 grams/mole.

Moles of $\text{HCl} = 30/36.5 = 0.821$ moles of HCl

Moles of water = $70/18 = 3.88$ moles of H_2O .

Mole fraction of $\text{HCl} = 0.821/(0.821 + 3.88) = 0.174$

Mole fraction of $\text{H}_2\text{O} = 3.88/(0.821 + 3.88) = 0.825$

Question 3

In a solution, we have 1 mole of A, 1 mole of B and 2 mole of C. Find the mole fraction of each of them

Solution

This is an extension of two component solution which we discussed above Mole fraction will be defined as

$$\chi_A = n_A / (n_A + n_B + n_C) = 1/(1+1+2) = 0.25$$

$$\chi_B = n_B / (n_A + n_B + n_C) = 1/(1+1+2) = 0.25$$

$$\chi_C = n_C / (n_A + n_B + n_C) = 2/(1+1+2) = 0.5$$

Also

$$\chi_A + \chi_B + \chi_C = 1$$

(Dear student you are kindly requested to do Exercise 1.6 page 27)

Lesson Eight

1.5.3 Molarity (M) of a solution

The molarity of a solution, represented by “M”, is defined as the number of moles of a solute per liter (or in one liter) of solution. Molarity is simply a measure of the "strength" of a solution. A solution that we would call "strong" would have a higher molarity than one that we would call "weak"

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

$$\text{But, Number of moles of solute} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}}$$

Therefore,

$$\text{Molarity} = \frac{\text{Mass of solute}}{\text{Molar mass of solute} \times \text{Volume (L) of solution}}$$

Note: The common unit for molarity is mol/Liter or Molar (i.e., mol/L or M)

Solved Example of Molarity Formula

Question 1

Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution.

Solution

As per Molarity Formula

$$M = \frac{W}{M \times 1000/V \text{ ml}}$$

$$= (4 \text{ g} / 40 \text{ g}) \times (1000/250) = 0.1 \text{ mol}/0.250 \text{ L}$$

$$= 0.4 \text{ mol/L}$$

$$= \mathbf{0.4 \text{ M}}$$

Question 2

If 2 gm NaOH is dissolved in water to make solution up to 250 cc. Give molarity of the solution:

Solution

$$\text{NaOH} = W = 2 \text{ g}$$

$$\text{Molar Mass of NaOH} = 23 + 16 + 1 = 40$$

$$M = \frac{W}{M \times 1000/(V \text{ ml})} = \frac{2}{40} \times \frac{1000}{250} = \mathbf{1/5 \text{ M}}$$

Question 3

Give the mass of NaOH dissolute to make solution of 500 cm³, 1/2 M molar.

Solution

$$M = W/M \times 1000/(V \text{ ml})$$

$$\text{Molar Mass of NaOH} = 23 + 16 + 1 = 40$$

$$M=1/2 \text{ and } V=500 \text{ ml}$$

Therefore,

$$1/2 = W/40 \times 1000/500$$

$$W = 20/2 = 10 \text{ gm}$$

Question 4

20% (w/W) NaOH solution (density of solution = 1.2 g/cc). Calculate the Molarity

Solution

20% (w/M) means 20 gm in 100gm of solution

Now Density of solution = 1.2g/cc

So Volume of Solution = 100/1.2 cc

Now

$$M = W/M \times 1000/(V \text{ ml})$$

$$M = 20/40 \times 1000/100 \times 1.2 = 6 \text{ M}$$

Question 5

Give the volume of the solution if 10 gm NaOH, 1/20 M NaOH sol. is present.

Solution

$$M = W/M \times 1000/(V \text{ ml})$$

$$1/20 = 10/40 \times 1000/(V \text{ ml})$$

$$4/20 = 1000/(V \text{ ml})$$

$$V = 5000 \text{ ml} = 5 \text{ L}$$

Question 6

What is the amount of solute present in M/40, 100 ml Na OH sol.

Solution

$$M = W/M \times 1000/(V \text{ ml})$$

$$1/40 = W/40 \times 1000/100$$

$$4/40 = W$$

$$0.1 \text{ gm} = W$$

Question 7

Calculate the morality of 2% (m/v) glucose sol.)

Solution

2 gm Glucose in 100 ml sol.

Now Molar mass of glucose($C_6H_{12}O_6$) = $12 \times 6 + 1 \times 12 + 6 \times 16 = 72 + 12 + 96 = 180$ gm

$$M = \frac{W}{V} \times \frac{1000}{M} \quad (V \text{ in ml})$$

$$M = \frac{2}{180} \times \frac{1000}{100}$$

$$M = \frac{1}{9}$$

(Dear student you are kindly requested to do Exercise 1.7 page 28)

Lesson Nine

1.5.3 Normality (N) Of a solution

We have studied about of the measure of concentration like Molarity, Mole Fraction, Molality. We are going to see another measure of concentration called Normality. Normality is defined as the number of gram equivalent present in per litre solution. It is used for acid or base solution

$$N = \frac{\text{Gram eq. of Solute}}{\text{Volume of sol. in litre}} \\ = \frac{\text{Weight}}{\text{Equivalent weight}} \times \frac{1000}{V \text{ ml}}$$

$$\text{Equivalent Weight} = \frac{\text{Molar Mass}}{n}$$

where n = number of H^+ in Acid, OH^- in base and for salt, charge present in ionic forms. Equivalent Weight is new thing in the formula. Let's check how it is defined

Equivalent Weight Definition and Gram Equivalent Weight

Equivalent Weight is defined as the ratio of [Molar Mass](#) of substance to the valence of the substance. Valence is also denoted as equivalence factor. The valence is the number of H^+ in Acid, OH^- in base and for salt, charge present in ionic forms. For reduction/oxidation (redox) reactions it is number of electrons that an oxidizing or reducing agent can accept or donate are counted as valence or equivalence factor. When the Equivalent Weight is expressed in gram using molar mass in gram, it is called Gram equivalent weight.

Few more important things about Normality

- Normality unit is N or eq/L. It depends on the chemical reaction being studied and depends on temperature too. So, it is meaning with respect to the chemical reaction
- It is the reactive capacity of the solution and it is called equivalent concentration
- This is used in acid-base reactions and redox reactions

Difference Between Normality and Molarity

Molarity	Normality
number of moles of solute per liter of solution.	number of gram equivalent of solute per liter of solution.

Molarity is a measurement of the Moles in the total volume of the solution,	Normality is a measurement of the gram equivalent in relationship to the total volume of the solution.
Unit is Moles L ⁻¹	Unit is eq L ⁻¹

Relation between Normality and Molarity

- Here is Normality formula in terms of molarity

$$\text{Normality} = n \times \text{Molarity}$$

where n = number of H⁺ in Acid, OH⁻ in base and for salt, charge present in ionic forms

How to calculate Normality of the Solution

- Find the equivalent weight of the solute based on the chemical reaction it is going to be used
- Calculate No. of Gram Equivalent of Solute
- Calculate the volume in litre
- Calculate the Normality using the formula given above

Solved Examples

Question 1

Calculate the equivalent weight of following:-

- H₂SO₄
- NaCl
- NaOH
- Na₂CO₃

Solution

- Molar Mass of H₂SO₄=98

n = 2 as it has 2 H⁺ ion

$$\text{Equivalent Weight} = \frac{\text{Molar Mass}}{n} = \frac{98}{2} = 49$$

- Molar Mass of NaCl = 23 + 35.5 = 58.5 n = 1

$$\text{Equivalent Weight} = \frac{\text{Molar Mass}}{n} = \frac{58.5}{1} = 58.5$$

- Molar Mass of NaOH=40

n = 1 as it has 1 OH⁻ ion

$$\text{Equivalent Weight} = \frac{\text{Molar Mass}}{n} = \frac{40}{1} = 40$$

- Molar Mass of Na₂CO₃=106

The salt Na₂CO₃ ionizes to form 2Na⁺ and CO₃⁻². So, the charge present on both is 2

$$n = 2$$

$$\text{Equivalent Weight} = \text{Molar Mass} / n = 106 / 2 = 53$$

Question 2

Calculate the normality of NaOH solution Formed by dissolving 0.2 gm NaOH to make 250 ml solution.

Solution

$$\text{Normality}(N) = (\text{no. Gram Equivalent of solute}) / (\text{Volume of Solution in litre})$$

$$\text{No. of Gram Eq. of Solute} = \text{weight} / \text{Equivalent weight}$$

$$\text{Now, Equivalent weight} = \text{Molar Mass} / n = 23 + 16 + 1 = 40$$

$$\text{So, } N = (\text{No. gram eq. mass}) / (\text{Vol (l)})$$

$$= \text{Weight} / (\text{Equivalent weight}) \times 1000 / (V \text{ ml})$$

$$= (2/40) \times (1000/250) = 0.2 \text{ N}$$

Question 3

Calculate the number of moles & molarity of N/2 500 ml solution of H_2SO_4

Solution

$$N = n \times M$$

$$1/2 = 2 \times M, 1/4 = M$$

Now Molarity of solution

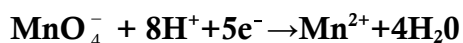
$$= \text{No of Moles} / \text{Volume in litre}$$

$$1/4 = n / .5$$

$$\text{or } n = 1/8$$

Question 4

Find the equivalent Weight of Potassium permanganate(KMnO_4) in redox reaction with Oxalic Acid($\text{H}_2\text{C}_2\text{O}_4 \cdot \text{H}_2\text{O}$) in presence of dilute H_2SO_4



Solution

KMnO_4 acts as oxidizing agent in the acidic medium and gain electron. Here it gain 5 electron. So valence or equivalence factor is 5

$$\text{Now Molar Mass of } \text{KMnO}_4 = 158$$

$$\text{Equivalent Weight} = \text{Molar Mass} / n = 158 / 5 = 31.6$$

Note: Remember the following mathematical equations while doing problems related with normality of a solution

Worked Example - 4

1. Calculate the normality of a solution made by dissolving 294 grams of calcium hydroxide, $\text{Ca}(\text{OH})_2$ in enough water to make 600 ml of solution? Atomic mass: Ca = 40, O = 16, H = 1

Given: Mass of $\text{Ca(OH)}_2 = 294.4 \text{ g}$

Molar mass of $\text{Ca(OH)}_2 = (1 \times 40) + (2 \times 16) + (2 \times 1) = 74 \text{ g/mol}$

Volume of solution = $600 \text{ ml} = 0.6 \text{ L}$

Number of replaceable hydroxide ions (OH^-) of a base, $\text{Ca(OH)}_2 = 2$

Equivalent mass of $\text{Ca(OH)}_2 = \text{Molar mass/number of "OH"}^- \text{ ions} = 74/2 = 37 \text{ g/equiv}$

Number of equivalents of $\text{Ca(OH)}_2 = \text{given mass/equivalent mass} = 294/37 = 7.95 \text{ equiv.}$

Required: Normality of a solution = ?

Solution:

$$(\text{Normality})_{\text{solution}} = \left(\frac{\text{number of equivalents of a solute}}{\text{volume of solution in liters}} \right) = \frac{7.95 \text{ equiv}}{0.6 \text{ L}} = 13.25 \text{ equiv/L} = 13.25 \text{ N}$$

(Dear student you are kindly requested to do Exercise 1.8 page 33)

Lesson Ten

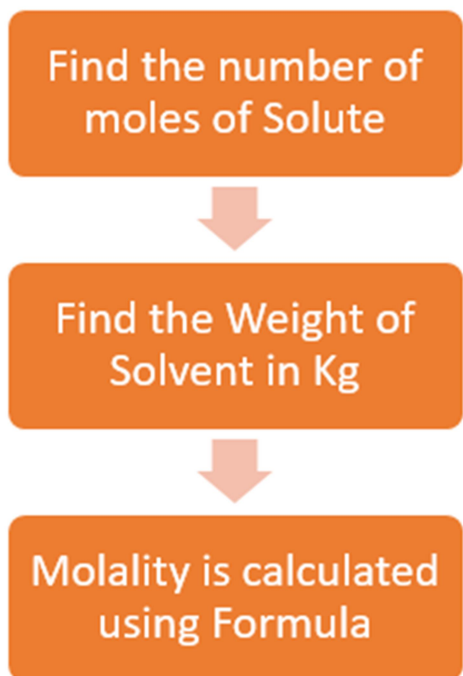
1.5.5 Molality (m) Of A solution

Molality Definition & Formula

- ☑ Molality is also known as Molal Concentration and It is one measure for Solute concentration
- ☑ Molality is defined as the number of Moles of solute dissolved per 1000 g (1 kg) of solvent.
- ☑ Molality is expressed as 'm' unit = Moles kg^{-1}
- ☑ Molality Formula is given as
$$m = \frac{\text{no of moles of solute}}{\text{weight of solvent (kg)}}$$
- ☑ Molality Formula can also be expressed as
- ☑ Molality Formula

$$m = \frac{\left(\frac{W}{\text{Molar Mass}} \right) \text{ solute}}{\frac{W}{1000} \text{ gm solvent}}$$
$$m = \left(\frac{W}{\text{Molar Mass}} \right) \times \frac{1000}{W \text{ gm solvent}}$$

The molality of a solution, designated by "m" is a concentration term expressed as the number of moles of solute dissolved in 1000 g (or 1 Kilogram) of a solvent



- ☑ Find the Number of moles of Solute. if Mass is given ,then find the Molar mass, then calculate the Moles
- ☑ Find the mass of the solvent in Kg. If volume and density is given, then mass can be calculated using Density=mass/volume
- ☑ Now that both moles and Solvent mass in kg is known, we can calculate the Molality using Molality Formula

Difference Between Molarity and Molality

Molarity	Molality
number of moles of solute per liter of solution.	number of moles of solute per kilogram of solvent.
Molarity is a measurement of the moles in the total volume of the solution,	molality is a measurement of the moles in relationship to the mass of the solvent.
Unit is Moles L ⁻¹	Unit is Moles kg ⁻¹

Solved Examples on Molality Formula

Question 1

5% (w/w) glucose solution is given .Find the molarity and molality of solution (density = 1.5gm/cc)

Solution

Molar mass of glucose (C₆ H₁₂ O₆) = 6 × 12 + 12 × 1 + 6 × 16 = 180 g mol⁻¹

5% (w/w) glucose solution means 5 gm Glucose in 100 gm solution.

Therefore, Solvent is 95 gm

Now Density = mass /volume

So volume = Mass/Density = $100/1.5 = 1000/15$ ml

For molarity:-

$$M = W/M \times 1000/(V \text{ ml})$$

$$M = 5/180 \times (1000/1000) \times 15 = 5/12$$

$$\text{Molality (m)} = W/MM \times 1000/(W \text{ gm in solvent})$$

$$= 5/180 \times 1000/95$$

$$= 100/(18 \times 19) = 50/171$$

Question 2

A solution of glucose in water is labelled as 10% w/w, Find out

(a) molality

(b) mole fraction of each component in the solution

(c) Molarity

Given the density of solution is 1.2 g mL⁻¹

Solution

10% w/w solution of glucose in water means that 10 g of glucose is present in 100 g of the solution i.e., 10 g of glucose is present in (100 - 10) g = 90 g of water.

Molar mass of glucose (C₆ H₁₂ O₆) = $6 \times 12 + 12 \times 1 + 6 \times 16 = 180$ g/mol

Then, number of moles of glucose = $10 / 180$ mol

$$= 0.056 \text{ mol}$$

Therefore, Molality of solution using Molality Formula = $0.056 \text{ mol} / 0.09 \text{ kg} = 0.62 \text{ m}$

Number of moles of water = $90 \text{ g} / 18 \text{ g mol}^{-1} = 5 \text{ mol}$

Mole fraction of glucose (m_g) = $0.056 / (0.056 + 5) = 0.011$

And, mole fraction of water m_w = $1 - m_g$

$$= 1 - 0.011 = 0.989$$

Now Density = mass /volume

So, volume = Mass/Density

$$= 100 \text{ g} / 1.2 \text{ g mL}^{-1} = 83.33 \text{ mL} = 83.33 \times 10^{-3} \text{ L}$$

Therefore, Molarity of the solution = $0.056 \text{ mol} / 83.33 \times 10^{-3} \text{ L} = 0.67 \text{ M}$

(Dear student you are kindly requested to do Exercise 1.9 page 34)

Lesson eleven

1.5.6 Inter Conversion of Concentration Units

All the terms we just discussed represent different ways of expressing concentration, so they are interconvertible. Keep these points in mind:

- ♣ To convert a term based on amount (mol) to one based on mass, you need the molar mass. These conversions are similar to the mass-mole conversions you've done earlier
- ♠ To convert a term based on mass to one based on volume, you need the solution *density*. Given the mass of a solution, the density (mass/volume) gives you the volume, or vice versa.
- ♦ Molality involves quantity of *solvent*, whereas the other concentration terms involve quantity of solution.

☞ Worked Example - 6

An aqueous solution of hydrogen peroxide, H_2O_2 is 30.0 % by mass and has a density of 1.11 g/ml. Calculate its: Molar mass of H_2O_2 = 34.02 g/mol; H_2O = 18.02 g/mol

- (a) Molality
- (b) Mole fraction of H_2O_2
- (c) Molarity

SOLUTION (a) From mass % to molality. Finding mass of solvent (assuming 100.0 g of solution):

$$\text{Mass (g) of } \text{H}_2\text{O} = 100.0 \text{ g solution} - 30.0 \text{ g } \text{H}_2\text{O}_2 = 70.0 \text{ g } \text{H}_2\text{O}$$

Converting from grams of H_2O_2 to moles:

$$\text{Moles of } \text{H}_2\text{O}_2 = 30.0 \text{ g } \text{H}_2\text{O}_2 \times \frac{1 \text{ mol } \text{H}_2\text{O}_2}{34.02 \text{ g } \text{H}_2\text{O}_2} = 0.882 \text{ mol } \text{H}_2\text{O}_2$$

Calculating molality:

$$\text{Molality of } \text{H}_2\text{O}_2 = \frac{0.882 \text{ mol } \text{H}_2\text{O}_2}{70.0 \text{ g} \times \frac{1 \text{ kg}}{10^3 \text{ g}}} = 12.6 \text{ m } \text{H}_2\text{O}_2$$

(b) From mass % to mole fraction:

$$\text{Moles of } \text{H}_2\text{O}_2 = 0.882 \text{ mol } \text{H}_2\text{O}_2 \text{ [from part (a)]}$$

$$\text{Moles of } \text{H}_2\text{O} = 70.0 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} = 3.88 \text{ mol } \text{H}_2\text{O}$$

$$X_{\text{H}_2\text{O}_2} = \frac{0.882 \text{ mol}}{0.882 \text{ mol} + 3.88 \text{ mol}} = 0.185$$

(c) From mass % and density to molarity. Converting from solution mass to volume:

$$\text{Volume (mL) of solution} = 100.0 \text{ g} \times \frac{1 \text{ mL}}{1.11 \text{ g}} = 90.1 \text{ mL}$$

Calculating molarity:

$$\text{Molarity} = \frac{\text{mol } \text{H}_2\text{O}_2}{\text{L soln}} = \frac{0.882 \text{ mol } \text{H}_2\text{O}_2}{90.1 \text{ mL} \times \frac{1 \text{ L soln}}{10^3 \text{ mL}}} = 9.79 \text{ M } \text{H}_2\text{O}_2$$