

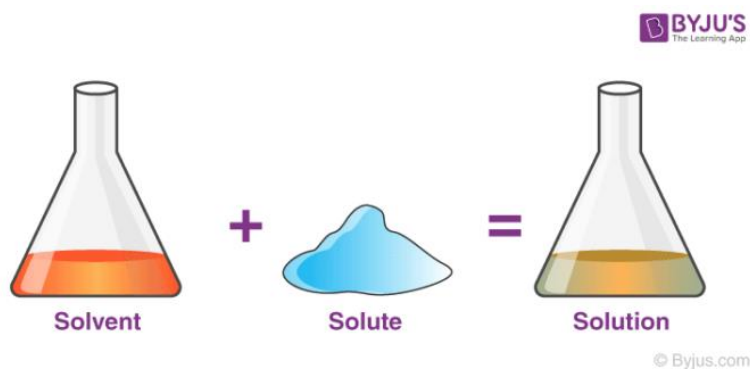
## Chapter I: Solutions and Solvents

### Introduction

Most liquids are not pure substances consisting of a single chemical species but are mixtures. The chemical species in excess is called the solvent, and all the other species present are solutes.

### Solution

A **solution** is a homogeneous mixture formed by dissolving one or more solutes (minority components) in a solvent (majority component). During dissolution, the solute disperses until it becomes invisible, creating a single phase. A solution is considered **saturated** when it contains the maximum amount of solute that the solvent can dissolve. The effectiveness of dissolution depends on the compatibility between solute and solvent, and while dissolution can occur naturally, it may be accelerated by heating or stirring.



### Example:

- Sugar (a solid solute) and water (solvent) form a good solute/solvent pair because their mixture forms a single phase.
- Oil (a liquid solute) or sand (a solid solute) do not form good solute/solvent pairs because, in each case, the phases do not mix.

### Aqueous Solutions:

An aqueous solution refers to a solution in which one or more solutes (in solid, liquid, or gaseous form) are dissolved in water (the solvent). This is the case in all living systems.

### Example:

- Saltwater is an aqueous solution where the solvent is water, and the solute is salt.

- Mineral water is an aqueous solution where the solvent is water, and the solutes are mineral salts, such as calcium, magnesium, and potassium ions.

### Solute

A solute is the substance that is dissolved in the solvent; it can be solid, liquid, or gas, and can be ionic or molecular. When the solvent used is water, this solution is called an aqueous solution.

### Solvents

A solvent is a liquid compound in which other substances can be dissolved or dispersed at the molecular level. In general, solvents are considered very important in chemistry due to their:

-Influence on the position of chemical equilibria.

-Importance on the absorption spectra of organic compounds. They have the ability to vary the speed and order of chemical reactions. This provides considerable possibilities in analysis.

**General Properties of a Solvent:** A solvent can act in two ways:

- **Through its chemical properties:**

If the solvent is polar, depending on its structure, it will act as a donor or an acceptor and thus transform the bonds between the atoms of the dissolved substance from covalent to ionized forms.

- **Through dielectric constant:**

The attraction between two ions depends on the dielectric constant of the medium. The force that binds two charged particles,  $q_1$  and  $q_2$ , separated by a distance "r" is given by Coulomb's law:

$$F = \frac{q_1 \cdot q_2}{r^2}$$

F is the magnitude of the electrostatic force (in Newtons, N).

$q_1$  et  $q_2$  are the magnitudes of the two charges (in Coulombs, C).

When th r is the distance between the charges in meters (m).

When ions are in a solvent with a dielectric constant or permittivity " $\xi$ ", the attractive force is divided by " $\xi$ ":

$$F = \frac{1}{\xi} \cdot \frac{q_1 \cdot q_2}{r^2}$$

- In solvents with " $\xi$ " > 40, the bonds are completely broken and the ion pairs are fully dissociated.
- In solvents with " $\xi$ " < 20, the ions mostly remain associated.

## Classification of Solvents

### A- Classification by chemical nature:

A solvent dissolves a compound that contains identical or analogous chemical groups; in other words, similar dissolves similar. The categories include:

- **Hydrocarbons:**

- Aliphatic hydrocarbons like hexane (C<sub>6</sub>H<sub>14</sub>) and heptane (C<sub>7</sub>H<sub>16</sub>).
- Alicyclic hydrocarbons like cyclohexane (C<sub>6</sub>H<sub>12</sub>).
- Aromatic hydrocarbons like benzene (C<sub>6</sub>H<sub>6</sub>).

They also include their derivatives:

- Halogenated compounds like chloroform.
- Nitro compounds like nitrobenzene.
- Sulfur compounds like carbon disulfide.

- **Hydroxylated solvents:**

Contain OH groups; examples include alcohols like methanol and ethanol, and carboxylic acids like formic acid and acetic acid.

- **Oxygenated solvents:**

- Ethers and oxides like methyl oxide.
- Esters like ethyl acetate.
- Ketones: the most well-known is acetone.
- Sulfoxides: dimethyl sulfoxide (DMSO).

- **Nitrogenated solvents:**

- Amides: formamide, dimethylformamide.
- Amines: pyridine.
- Nitriles: acetonitrile.

### B- Classification by polarity

- Solvents that have a permanent dipole moment are defined as polar solvents.

**Examples:**

- DMSO:  $\mu = 3.9$  D
- Acetone:  $\mu = 2.7$  D
- Water:  $\mu = 1.8$  D
- Solvents that do not have a permanent dipole moment are said to be nonpolar:  
**Examples:** Carbon disulfide, carbon tetrachloride, benzene, and cyclohexane.

Between these two classes, there are solvents of medium polarity such as ketones and esters.

### C- Classification by dissociating character:

- If  $\xi > 40$ , the solvent is dissociating: The dielectric constant is high, and the ionic bonds of the solutes are almost completely broken.

#### Examples:

- N-methylformamide:  $\xi = 182.4$
- Water:  $\xi = 78.5$
- If  $15 < \xi < 40$ , the solvent is intermediate: In this case, the concentration of ions in solution is negligible compared to the concentration of reactants that are present in molecular form.  
**Example:** Ethanol:  $\xi = 24.3$
- If  $\xi < 15$ , the solvent is very weakly dissociating:  
**Example:** Acetic acid:  $\xi = 6.2$

### Mechanism of Solution:

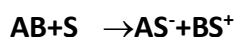
When a solid comes into contact with a solvent, the solvent can exert a threefold action.

- **Solvation**

The molecules of the solvent tend to form bonds with the particles constituting the solid. These bonds are of the Van der Waals type, involving charge transfer.

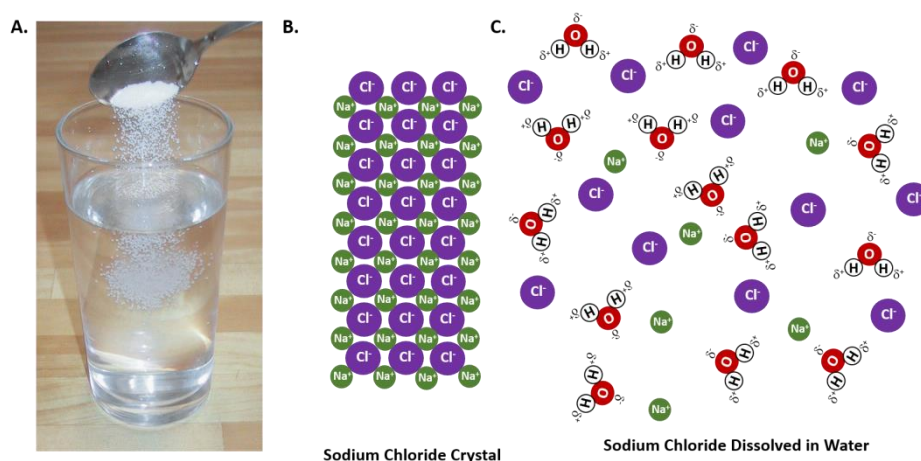
- **Ionization**

There is a breaking of the A-B bond and a profound modification of the structure of the solute (AB: solute molecule). If  $\xi$  is high:



- **Solvolysis**

New bonds are formed with the molecules of solvent S. When the solvent is water, this process is referred to as hydrolysis.



**Figure 1.** illustrates the process of dissolving, where water dissociates sodium chloride into ions, transforming its solid crystalline lattice into a dissolved ionic state.

### Factors Influencing Solubility

- **Influence of Structure:**

Solubility is favored when the solute and the solvent have the same polarity.

- Polar solvents dissolve polar solutes.
- Nonpolar solvents dissolve nonpolar solutes.

- **Influence of Pressure:**

There is little influence on the solubility of liquids or solids but notable influence on gases (solubility increases with pressure).

- **Influence of Temperature:**

- **In the case of gases:**

The solubility of a gas in water decreases as the temperature increases.

- **In the case of solids:**

The dissolution of a solid occurs more rapidly when the temperature increases.

### 6. Ideal and Non-Ideal Solutions:

In thermodynamics and chemistry, the distinction between **ideal** and **non-ideal (real)** solutions is important for understanding how different mixtures behave. This concept is based on how the components of a solution interact with each other.

## Ideal Solutions

An **ideal solution** follows **Raoult's law** perfectly and is characterized by several key features:

1. **No change in enthalpy:** When the components of an ideal solution are mixed, there is no heat absorbed or released, meaning the enthalpy of mixing  $\Delta H_{mix} = 0$
2. **No volume change on mixing:** The total volume of the solution is simply the sum of the volumes of the individual components before mixing, meaning the volume change  $\Delta V_{mix} = 0$
3. **Similar intermolecular interactions:**

The forces between like molecules (A-A or B-B) are similar to those between unlike molecules (A-B). For example, if the solution consists of two components A and B, the interaction energies between A-A, B-B, and A-B molecules are nearly identical.

4. **Obeys Raoult's law:**

In an ideal solution, the partial vapor pressure of each component is directly proportional to its mole fraction in the solution.

$$P_A = P_{A^*} \cdot x_A$$

$$P_B = P_{B^*} \cdot x_B$$

Where:

$P_A$  et  $P_B$  are the partial vapor pressures of components A and B.

$P_A^*$  et  $P_B^*$  are the vapor pressures of pure components A and B.

$x_A$  and  $x_B$  are the mole fractions of components A and B.

## Examples of Ideal Solutions

- Benzene and toluene
- Hexane and heptane
- Ethanol and methanol (in small concentrations)

## Non-Ideal Solutions

A **non-ideal solution** does not obey Raoult's law and exhibits deviations due to differences in molecular interactions between the components.

### Types of Deviations:

#### 1. Positive deviation from Raoult's law:

- Occurs when the interactions between unlike molecules (A-B) are weaker than those between like molecules (A-A and B-B).
- This results in an increase in vapor pressure compared to an ideal solution because the molecules escape more easily into the vapor phase.
- The enthalpy of mixing  $\Delta H_{mix} > 0$ , meaning heat is absorbed during mixing.

**Examples:** Ethanol and cyclohexane, Acetone and carbon disulfide.

#### 2. Negative deviation from Raoult's law:

- Occurs when the interactions between unlike molecules (A-B) are stronger than those between like molecules (A-A and B-B).
- This results in a decrease in vapor pressure because the stronger intermolecular forces make it harder for the molecules to escape into the vapor phase.
- The enthalpy of mixing  $\Delta H_{mix} < 0$ , meaning heat is released during mixing.

**Examples:** Acetone and water, Nitric acid and water.

### Key Features of Non-Ideal Solutions:

- **Changes in enthalpy:** Mixing components either releases or absorbs heat.
- **Volume change:** The volume of the solution may increase or decrease upon mixing.
- **Deviations from Raoult's law:** The vapor pressure of the components does not follow Raoult's law.

### Difference between Ideal and Non-ideal Solution

The table below shows the main differences between Ideal and Non-ideal Solutions:

Ideal Solution	Non-ideal Solution
<b>Raoult's Law</b>	
They obey Raoult's law	They do not obey Raoult's law
<b>Molecular Interactions</b>	
Intermolecular interaction between solute and solvent is the same as that of pure components	Intermolecular interaction between solute and solvent is weaker or stronger than that of between pure components
<b>Total Vapour Pressure</b>	
The total vapour pressure is the same as predicted by Raoult's law	The total vapour pressure increases or decreases from the predicted value according to Raoult's law
<b>Enthalpy of mixing</b>	
No heat is released or absorbed, so the enthalpy of mixing is zero, $\Delta H_{\text{mix}} = 0$	Heat is either absorbed or released, so the enthalpy of mixing is either positive or negative, $\Delta H_{\text{mix}} \neq 0$
<b>Volume of mixing</b>	
The total volume is equal to the sum of the volume of components (solute and solvent), so the volume of mixing is zero, $\Delta V_{\text{mix}} = 0$	The volume of mixing is not zero, $\Delta V_{\text{mix}} \neq 0$ . There is either expansion or contraction.
<b>Examples</b>	
Benzene and toluene, hexane and heptane, etc. All the dilute solutions nearly behave as an ideal solution	Ethanol and acetone, carbon disulphide and acetone, phenol and aniline, chloroform and acetone, etc.

### Expression of Concentrations

Concentration is a measure of how much solute is present in a given quantity of solvent or solution. It can be expressed in various ways depending on the context, and each form of expressing concentration is useful for specific types of chemical calculations.

#### 1. Molarity (M)

Molarity is one of the most commonly used units of concentration in chemistry, especially for solutions. It is defined as the number of moles of solute per liter of solution.

$$M = \frac{\text{mole of solute}}{\text{volume of solution in litre}} = \frac{n}{V}$$

$n$  = is the number of moles of solute,

$V$  = is the volume of the solution in liters.



**Example:**

If you dissolve 1 mole of sodium chloride (NaCl) in 1 liter of water, the concentration is 1 M.

**2. Molality (m)**

Molality is the number of moles of solute per kilogram of solvent. Unlike molarity, molality does not depend on temperature because it is based on mass, not volume.

$$m = \frac{\text{moles of solute}}{\text{mass of the solvent in kilograms.}} = \frac{n}{m(\text{solvent})}$$

n is the number of moles of solute,

$m_{\text{solvent}}$  is the mass of the solvent in kilograms.

**Example:** If you dissolve 1 mole of NaCl in 1 kilogram of water, the molality is 1 m.

**3. Mass Percent (w/w % or m/m %)**

Mass percent is the mass of solute per mass of solution, expressed as a percentage. It is widely used in industry and food science.

$$\% \text{ w/w} = \frac{\text{mass of solute}}{\text{mass of solution}} \cdot 100$$

Where:

- Mass of solute is typically in grams,
- Mass of solution is the total mass of solute + solvent in grams.

**Example:**

If you have 5 grams of NaCl dissolved in 95 grams of water, the mass percent is 5%.

$$\% \text{ w/w} = \frac{5}{5 + 95} \cdot 100 = 5\%$$

**4. Volume Percent (v/v %)**

Volume percent is the volume of solute per volume of solution, expressed as a percentage.

This is commonly used for liquid-liquid solutions, such as in the preparation of alcohol solutions.

$$\% \text{ v/v} = \frac{\text{volume of solute}}{\text{volume of solution}} \cdot 100$$

**Example:** If you mix 40 mL of ethanol in enough water to make 100 mL of solution, the volume percent is 40%.

### 5. Mole Fraction (X)

Mole fraction is the ratio of the number of moles of one component to the total number of moles in the solution. It is commonly used in thermodynamic calculations and gas mixtures.

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

Where:

- $X_A$  = mole fraction of component A
- $n_A$  = number of moles of component A (solute)
- $n_B$  = number of moles of component B (solvent)

**Example:** If you have a solution with 1 mole of solute and 4 moles of solvent, the mole fraction of the solute is 0.2.

### 6. Normality (N)

Normality is the number of equivalents of solute per liter of solution. It is particularly useful in acid-base titrations and redox reactions.

$$N = \frac{\text{equivalents of solute}}{\text{volume of solution in liters}}$$

Where:

- N = normality of the solution
- Equivalents of solute = the number of equivalents based on the specific reaction in which the solute participates
- Volume of solution = the total volume of the solution in liters

**Example:** For a 1 M sulfuric acid solution ( $\text{H}_2\text{SO}_4$ ), which has 2 equivalents of  $\text{H}^+$  per mole of  $\text{H}_2\text{SO}_4$ , the normality would be 2 N.

### 7. Parts per Million (ppm) and Parts per Billion (ppb)

Parts per million (ppm) and parts per billion (ppb) are used for very dilute concentrations, typically in environmental chemistry and analytical chemistry.

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

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$$

**Example:**

If a water sample contains 2 mg of lead in 1 liter of water, the concentration of lead is 2 ppm.

### 8. Osmolarity

Osmolarity refers to the concentration of osmotically active particles in a solution. It is commonly used in biology and medicine.

$$\text{Osmolarity} = \frac{\text{moles of particles}}{\text{volume of solution in liters}}$$

Where:

- **Moles of particles** refers to the total number of solute particles that contribute to osmotic pressure.
- For example, one mole of NaCl dissociates into two particles ( $\text{Na}^+$  and  $\text{Cl}^-$ ), resulting in a total of two moles of particles.
- **Volume of solution** is the total volume of the solution in liters.

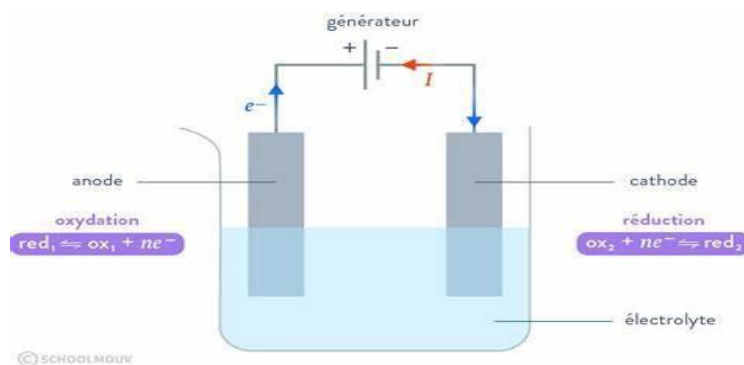
#### Summary Table

Concentration Expression	Definition
Molarity (M)	Moles of solute per liter of solution
Molality (m)	Moles of solute per kilogram of solvent
Mass Percent (w/w)	Mass of solute per mass of solution, expressed as a percentage
Volume Percent (v/v)	Volume of solute per volume of solution, expressed as a percentage
Mole Fraction (X)	Ratio of moles of a component to total moles in the solution
Normality (N)	Equivalents of solute per liter of solution
ppm/ppb	Parts per million (ppm) or billion (ppb) parts of solution
Osmolarity	Moles of particles per liter of solution

#### Ionic and Molecular Aqueous Solutions

An aqueous solution is classified as **ionic** if it contains dissolved ions, either directly from the solute or produced by its reaction with water, while it is considered **molecular** if it consists of dissolved molecules. Both types of solutions are electrically neutral.

**Electrolysis** is a chemical process that decomposes substances by adding a solvent to a solution, enabling ion breakdown. The resulting mixture is placed in a bath with an anode and cathode, where an electric current is passed through to facilitate the reaction. current to pass through the bath.



**Figure 2.** Schematic of Electrolysis

a. **Electrolyte:**

An electrolyte is a compound that, when dissolved in water, produces ions (anions and cations).

b. **Strong Electrolytes:**

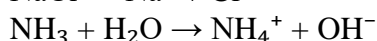
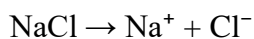
Strong electrolytes are solutes for which ionization or dissociation by the solvent is complete. Examples include NaCl, NaOH, KOH, HCl, etc.

c. **Weak Electrolytes:**

Weak electrolytes are solutes whose ionization or dissociation is only partial, due to a reverse reaction. Examples include ethanoic acid (CH<sub>3</sub>COOH), ammonia (NH<sub>3</sub>), etc. The solution, therefore, contains ions "produced by the ionization of the solute," molecules of the solute, and those of the solvent.

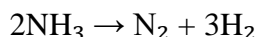
→ **Ionization:**

This term is used when the products of the reaction are ions, such as:



→ **Dissociation:**

This term is used when the products of the reaction are neutral, such as:



### 5. Dissociation Rate (or Degree of Dissociation/Ionization):

Weak electrolytes have a bond with a strong covalent character, meaning they dissociate very little in water, and their dissociation and ionization are rarely complete. Such a solution, therefore, contains both neutral molecules and ions.

The dissociation rate  $\alpha$ , or "Dissociation Coefficient," of a weak electrolyte is defined as:

$$\alpha = \frac{\text{Nombredemolesdissociés}}{\text{Nombredemolestotaldanslasolution}}$$

$\alpha$ : dimensionless (no unit).

**Note:**

- ✓  $\alpha = 0$ : Dissociation is zero (neutral).
- ✓  $0 < \alpha < 1$ : Dissociation is partial (weak electrolyte).
- ✓  $\alpha = 1$ : Dissociation is complete (strong electrolyte).

### Activity and Activity Coefficient

In the study of solutions, especially in electrolyte solutions and non-ideal systems, the concept of activity and activity coefficient is used to describe the effective concentration of a substance, accounting for non-ideal behavior. This helps to correct deviations from ideal behavior observed in real solutions.

**Activity:**

- Activity (a) is a corrected form of the concentration or partial pressure of a species in a solution. It accounts for interactions between molecules or ions, which alter how the substance behaves compared to its ideal state.

- For a substance (i) in solution, its activity  $a_i$  related to its mole fraction  $x_i$  or concentration ([i]) by the activity coefficient  $\gamma_i$  :

$$a_i = \gamma_i \cdot x_i \quad \text{or} \quad a_i = \gamma_i \cdot [i]$$

In ideal solutions, the activity is equal to the concentration ( $a_i = x_i$  or  $[a_i] = [i]$ ) because the activity coefficient  $\gamma_i = 1$ . In non – ideal solution  $\gamma_i \neq 1$

### Activity Coefficient:

- The activity coefficient ( $\gamma$ ) is a factor that quantifies the deviation of a solution from ideal behavior. It corrects the concentration or mole fraction to give the **effective concentration** (activity).

- If ( $\gamma > 1$ ), the solution exhibits positive deviation from ideality, meaning the species interact less favorably with the solvent (they are less “active”).

- If ( $\gamma < 1$ ), the solution exhibits negative deviation, meaning the species interact more favorably with the solvent (they are more “active”).

### - Example:

In a non-ideal solution, if the concentration of an ion is 1 mol/L but due to interactions between ions, the effective concentration behaves as if it were 0.8 mol/L, the activity coefficient would be ( $\gamma = 0.8$ ).

### Debye-Hückel Theory and Debye-Hückel Law

The Debye-Hückel theory provides a quantitative explanation for the behavior of ions in dilute electrolyte solutions, particularly the deviation from ideality due to electrostatic interactions between ions. The theory introduces the Debye-Hückel law, which allows the calculation of the activity coefficient for ions in dilute solutions.

### Debye-Hückel Limiting Law (for dilute solutions):

For a dilute solution of electrolytes, the Debye-Hückel limiting law gives the activity coefficient of an ion ( $\gamma_i$ ) as:

$$\log \gamma_i = - \frac{Az_i^2 \sqrt{I}}{1 + B \cdot a_i \sqrt{I}}$$

where:

$Z_i$  is the charge of the ion i.

A and B are constants that depend on the solvent (e.g., for water at 25°C,  $A \approx 0.509$  and  $B \approx 0.328$ ).

$a_i$  is the effective ion size (in Å).

I is the **ionic strength** of the solution, which is a measure of the total concentration of ions in the solution, given by:

$$I = \frac{1}{2} \sum_{i=1}^n C_i Z_i^2$$

where  $C_i$  is the concentration of ion  $i$  and  $Z_i$  is its charge.

$$\log \gamma_i = -Az_i^2 \sqrt{I}$$

This simplified form can be used for solutions with ionic strengths less than 0.01 mol/L, where the term  $Ba_i \sqrt{I}$  becomes negligible.

### Key Insights from Debye-Hückel Theory:

#### 1. Ionic Atmosphere:

In electrolyte solutions, ions do not behave independently. Each ion is surrounded by a cloud of oppositely charged ions, called the ionic atmosphere, which reduces the electrostatic potential experienced by each ion. This leads to deviations from ideal behavior.

#### 2. Ionic Strength:

The activity coefficient of an ion depends on the ionic strength of the solution. As the ionic strength increases, interactions between ions become more significant, leading to a greater deviation from ideal behavior (i.e.,  $\gamma_i$  deviates further from 1).

#### 3. Dilute Solutions:

The Debye-Hückel theory applies best to dilute solutions where ionic interactions are relatively simple. For highly concentrated solutions, more complex models are needed to account for short-range interactions and ion pairing.

#### 4. Charge Dependence:

The activity coefficient decreases more rapidly for ions with higher charges ( $z_i^2$ ) because stronger electrostatic interactions occur between ions with larger charges.

### **Applications:**

#### **- Electrochemistry :**

The Debye-Hückel theory is widely used in electrochemical systems, such as understanding electrode potentials, ion transport, and conductivity.

#### **- Biological Systems:**

It helps in calculating the activities of ions in biological systems like cell membranes and enzyme functions.

#### **- Industrial Chemistr :**

It's useful in designing processes involving electrolytes, such as water treatment and battery technology.

### **Summary of the Debye-Hückel Law:**

- Debye-Hückel Law quantifies the deviation from ideal behavior for ions in a solution based on electrostatic interactions.

- It provides a way to calculate the activity coefficient of ions in dilute solutions using the ionic strength and charge of the ions.

- The theory becomes less accurate for concentrated solutions, where more complex interactions are present.

The combination of activity, activity coefficient, and Debye-Hückel theory provides a robust framework for understanding and calculating the behavior of ions in non-ideal solutions, particularly in electrolyte chemistry.



## Solutions and Solvents in Pharmacy

Solutions and solvents play a crucial role in pharmacy for several reasons:

1. **Drug Formulation:** Solutions are essential for formulating various dosage forms, including liquids (e.g., syrups, solutions, and injections), enabling the effective delivery of active pharmaceutical ingredients (APIs).
2. **Solubility Enhancement:** Many drugs have low solubility, which can limit their bioavailability. Solvents help dissolve these drugs, improving their absorption and therapeutic effectiveness.
3. **Stability:** Proper solvent selection can enhance the stability of drugs, preventing degradation and extending shelf life. This is particularly important for sensitive compounds.
4. **Dosing Precision:** Solutions allow for accurate dosing and uniform distribution of the drug throughout the formulation, ensuring consistent therapeutic effects.
5. **Controlled Release:** Solvents can influence the release profile of drugs, allowing for controlled or sustained release formulations that improve patient adherence and outcomes.
6. **Compatibility:** The choice of solvent affects the compatibility of different drug components, reducing the risk of adverse reactions and ensuring the safety of the final product.
7. **Pharmaceutical Applications:** Solvents are used in various pharmaceutical processes, including extraction, purification, and crystallization, making them vital for the manufacturing of drugs.
8. **Patient Comfort:** Solutions can enhance the palatability of medications, making them easier to administer, especially in pediatric and geriatric populations.

In summary, solutions and solvents are fundamental to pharmaceutical science, affecting drug formulation, effectiveness, stability, and patient compliance.