

SOLUTION

There are several types of solutions based on the states of the solute and solvent:

1. Gaseous Solutions:

Examples include air, which is a solution of gases like nitrogen, oxygen, carbon dioxide, and others in the gaseous state. Another example is hydrogen gas (H_2) dissolved in palladium, forming a gaseous solution.

2. Liquid Solutions:

Common liquid solutions include sugar (solute) dissolved in water (solvent) to form a sugar solution. Alcoholic beverages, where ethanol (alcohol) is dissolved in water, are also examples of liquid solutions.

A solution is a homogeneous mixture composed of two or more substances, where one substance (**the solute**) is uniformly distributed in another substance (**the solvent**). In a solution, the solute particles are evenly dispersed at the molecular or ionic level throughout the solvent.

TRY TO SPEAK OUT

What is solute & what is solvent?

3. Solid Solutions:.....

Solid solutions involve the uniform distribution of solute particles in a solid solvent. An alloy, such as brass (a mixture of copper and zinc), is a solid solution. Steel, which is a mixture of iron and carbon, is another example of a solid solution.

4. Aqueous Solutions:.....

Aqueous solutions involve water as the solvent. Many chemical reactions and biological processes occur in aqueous environments. Well known examples include salt ($NaCl$) dissolved in water and hydrochloric acid (HCl) dissolved in water.

5. Supersaturated Solutions:

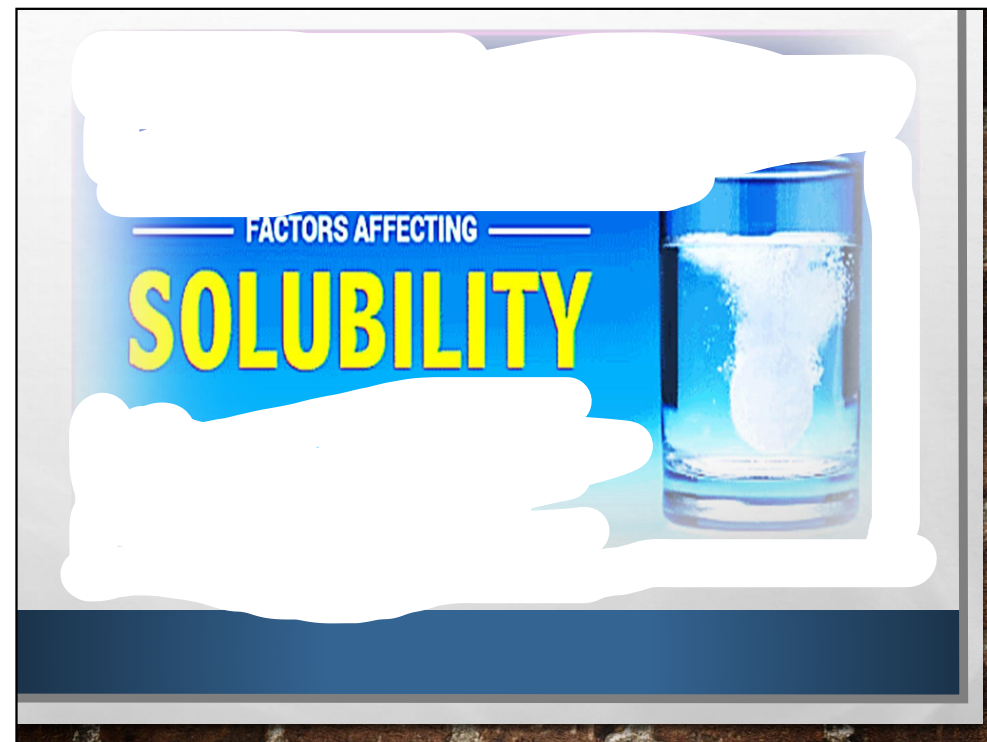
These solutions contain more solute than would normally dissolve at a particular temperature. They are often created by dissolving a solute at an elevated temperature and then cooling the solution.

6. Unsaturated Solutions:

These solutions have not reached their maximum solute concentration at a given temperature, meaning more solute could dissolve.

7. Saturated Solutions:

Saturated solutions have reached the maximum amount of solute that can dissolve at a particular temperature.



Rest of them

Polar Solvents with Polar Solute: Examples include water with sugar.

Nonpolar Solvents with Nonpolar Solute: Examples include hexane with oil.

Polar Solvents with Nonpolar Solute: Examples include acetone with oil.

1. Nature of Solute and Solvent:

The chemical nature of both the solute and solvent plays a crucial role. Like dissolves like, meaning polar solvents tend to dissolve polar solutes, and nonpolar solvents tend to dissolve nonpolar solutes.

2. Temperature:

In general, solubility increases with an increase in temperature for most solid solutes in liquid solvents. However, this is not universally true for all substances. For gases in liquids, solubility often decreases with an increase in temperature.

3. Pressure (for Gases):

For gases dissolved in liquids, solubility is influenced by pressure. **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of that gas above the liquid.

4.Surface Area:

Increasing the surface area of a solid solute can enhance its solubility. This is particularly relevant for finely powdered or crushed solutes, as it provides more contact points for interaction with the solvent.

5.Stirring or Agitation:

Stirring or agitating a solution helps in speeding up the dissolution process by increasing the contact between the solute and solvent. This is especially important for solid solutes.

6.Chemical Interactions:

Chemical interactions between the solute and solvent molecules can affect solubility. For example, the formation of hydrogen bonds or other attractive forces can enhance solubility.

Way of dissolution

7.Common Ion Effect:

The presence of common ions can influence the solubility of a substance. If a solution already contains an ion that is part of the solute, it can reduce the solubility of that solute.

8.pH of the Solution:

For substances that can undergo ionization or dissociation, the pH of the solution can affect solubility. Some substances are more soluble under acidic or basic conditions.

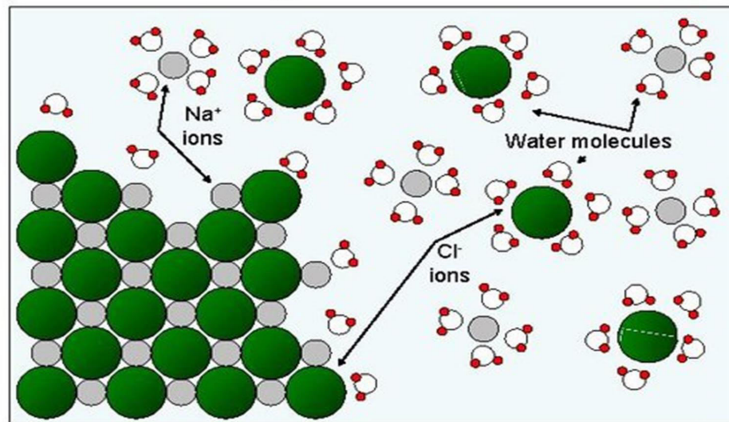
9.Complex Formation:

Some solutes can form complexes with solvent molecules or other solutes, affecting their overall solubility.

The process of dissolution involves the breakdown of solute particles and their integration into the solvent to form a homogeneous mixture, often referred to as a solution.

The mechanism of dissolution depends on the nature of the solute and solvent and is influenced by factors like **temperature, pressure, and molecular interactions.**

Dissolving of Salt in Water



Henry's Law provides a mathematical relationship describing the solubility of a gas in a liquid at a constant temperature. The equation is as follows:

$$C = k \cdot P_x$$

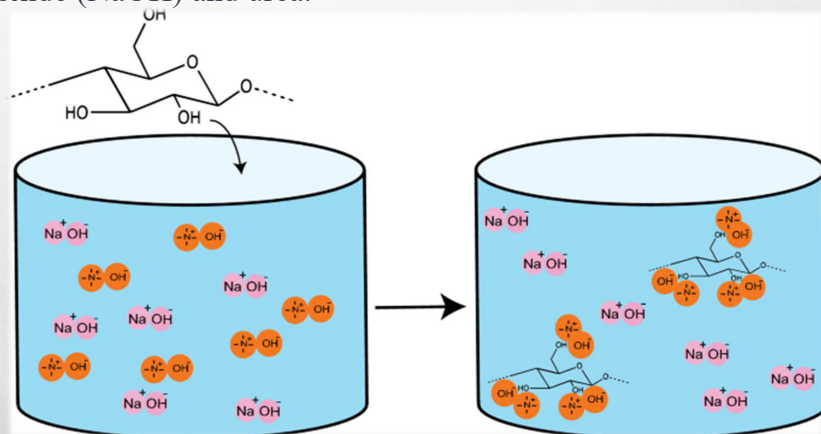
where:

C is the concentration of the dissolved gas in the liquid (usually in mol/L or another concentration unit),

k is Henry's Law constant, which is specific for a particular gas-solvent pair at a given temperature,

P_x is the partial pressure of the gas above the liquid.

While cellulose is insoluble in water and most common solvents, there are certain treatments that can modify its structure to make it more soluble. One such treatment involves using a solution of sodium hydroxide (NaOH) and urea.



Some common examples of solutions involving gases in liquids:

1. Carbonated Beverages: Carbon dioxide (CO₂) dissolves in water to form carbonic acid, leading to the fizzy or effervescent nature of carbonated beverages. The higher pressure in a sealed container, such as a soda bottle, increases the solubility of CO₂ in the liquid.

2. Aeration of Water: Oxygen (O₂) from the air can dissolve in water, especially in bodies of water such as rivers, lakes, and oceans. This dissolved oxygen is crucial for aquatic life and supports various biochemical processes.

3. Oxygen Dissolved in Blood Plasma: Oxygen is transported in the blood by dissolving in the liquid component known as plasma. The solubility of oxygen in blood is influenced by factors such as partial pressure and temperature.

4.Nitrogen Dissolved in Body Fluids: Nitrogen, a major component of air, can dissolve in body fluids. Scuba divers are familiar with the concept of nitrogen dissolving in the bloodstream at higher pressures, which can lead to decompression sickness if not managed properly during ascent.

5.Ammonia Dissolved in Water: Ammonia (NH₃) can dissolve in water, forming ammonium hydroxide. This is important in various industrial processes and environmental chemistry.

6.Chlorine Dissolved in Water: Chlorine gas (Cl₂) can dissolve in water to form a solution of hydrochloric acid (HCl) and hypochlorous acid (HOCl). This is utilized in water treatment for disinfection.

7.Sulfur Dioxide Dissolved in Water: Sulfur dioxide (SO₂) dissolves in water to form sulfurous acid (H₂SO₃). This process is relevant in environmental chemistry and the formation of acid rain.

1.Mass/Mass Percentage (% w/w): This is the mass of the solute divided by the total mass of the solution, multiplied by 100.

$$\text{Mass \% (w/w)} = (\text{Mass of Solute} / \text{Total Mass of Solution}) \times 100$$

2.Volume/Volume Percentage (% v/v): This is the volume of the solute divided by the total volume of the solution, multiplied by 100.

$$\text{Volume \% (v/v)} = (\text{Volume of Solute} / \text{Total Volume of Solution}) \times 100$$

3.Mass/Volume Percentage (% w/v): This is the mass of the solute divided by the volume of the solution, multiplied by 100.

$$\text{Mass \% (w/v)} = (\text{Mass of Solute} / \text{Volume of Solution}) \times 100$$

Different units of concentration

4.Molarity (M): This is the number of moles of solute per liter of solution.

$$\text{Molarity (M)} = \text{Moles of Solute} / \text{Liters of Solution}$$

5.Molality (m): This is the number of moles of solute per kilogram of solvent.

$$\text{Molality (m)} = \text{Moles of Solute} / \text{Mass of Solvent (in kg)}$$

6.Normality (N): This is the number of equivalents of solute per liter of solution.

$$\text{Normality (N)} = \text{Equivalents of Solute} / \text{Liters of Solution}$$

7.Parts Per Million (ppm): This is the number of parts of solute per million parts of solution.

$$\text{ppm} = \text{mg of solute} / \text{litters of solution}$$

$$\text{Or, ppm} = \mu\text{g of solute} / \text{mLs of solution}$$

- I. 5gNaCl dissolved in 145mL of water. What is the molarity, normality and ppm of the solution?
- II. If 105g of ferric chloride dissolved in 587mL of solvent, what will be the molarity of the solution?
- III. 98g of H₂SO₄ was dissolved in 1L water, what will be the molarity, normality of this solution?
- IV. Convert 0.001M NaOH to ppm....
- V. In a river water sample there are 5ppm O₂ was found. What is the molarity of the oxygen in this sample?

The Distribution Law, often referred to as the Law of Distribution or the **Partition Law**, describes the distribution of a solute between two immiscible solvents at equilibrium.

This law is commonly applied in the field of chemistry, particularly in the study of liquid-liquid extraction or separation techniques.

The law is based on the assumption that the distribution of a solute between two solvents is proportional to their respective concentrations in each solvent.

Distribution Law

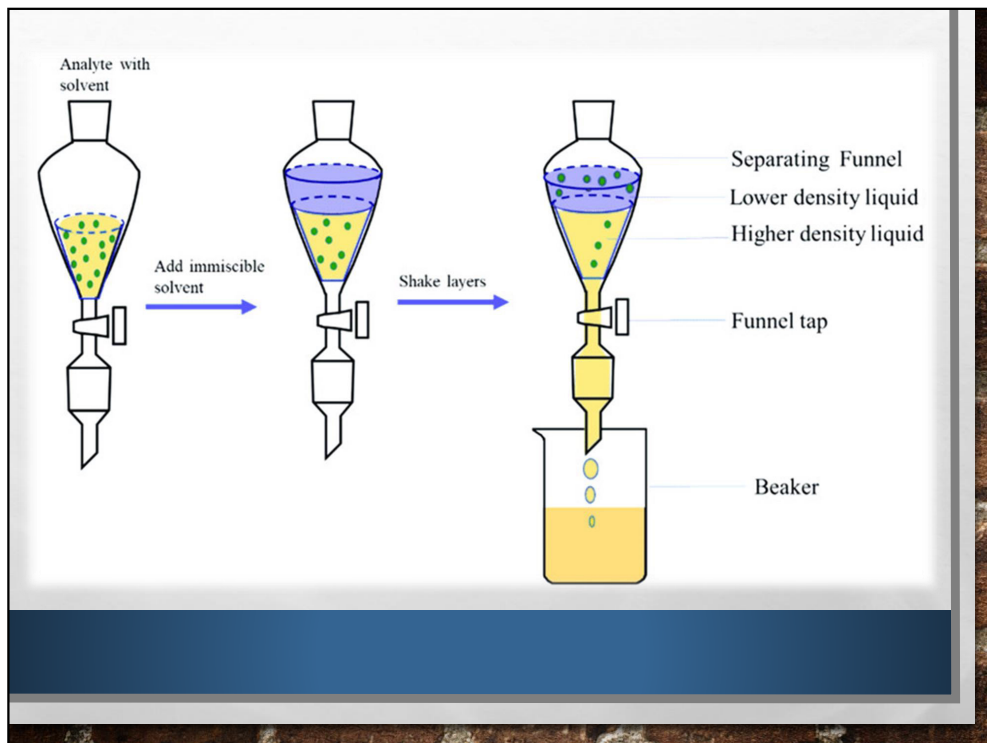
The distribution law can be stated as follows:

$$C_2/C_1=K_d$$

where:

- C₁ is the concentration of the solute in the first solvent (organic phase, for example),
- C₂ is the concentration of the solute in the second solvent (aqueous phase, for example),
- K_d is the distribution coefficient or partition coefficient.

The distribution coefficient (K_d) is a constant at a given temperature and for a specific solute and pair of immiscible solvents. It indicates how the solute is distributed between the two phases when the system reaches equilibrium.



Applications of Distribution Law

This law is pivotal in **separating and purifying** substances, finding applications in pharmaceutical development, analytical chemistry, environmental studies, and metallurgy. It plays a crucial role in predicting the distribution of solutes between different phases, aiding in the optimization of processes such as drug formulation, environmental risk assessment, and the extraction of valuable compounds from complex mixtures.

Lower Osmotic Pressure: The osmotic pressure of a dilute solution is lower compared to a concentrated solution. Osmotic pressure is directly proportional to the concentration of solute particles in the solution, according to van't Hoff's law.

Lower Boiling Point Elevation: Dilute solutions exhibit a smaller increase in boiling point compared to pure solvent due to the presence of solute particles. This phenomenon is governed by Raoult's law and is less pronounced in dilute solutions.

Lower Freezing Point Depression: Similarly, dilute solutions experience a smaller decrease in freezing point compared to pure solvent due to the presence of solute particles. This is also governed by **Raoult's law** and is less significant in dilute solutions.

Ideal Behavior: In dilute solutions, the interactions between solute particles are minimal, and the solution may behave more ideally, adhering closely to ideal solution behavior assumptions. This simplifies calculations and predictions in dilute solution systems.

Colligative Properties: Dilute solutions exhibit colligative properties, which depend only on the number of solute particles and not on their nature. These properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

Increased Reactivity: Dilute solutions can sometimes enhance the reactivity of solute particles due to increased accessibility of solvent molecules to the solute particles. This property is particularly relevant in chemical reactions occurring in dilute solution environments.

The vapor pressure of a solution is directly proportional to the mole fraction of each component in the solution. Raoult's Law is typically applied to solutions where the intermolecular forces between different components are similar, and the solution behaves ideally.

The general form of Raoult's Law for a two-component system is given by:

$$P_{\text{total}} = x_A \cdot P_A + x_B \cdot P_B$$

Where:

P_{total} is the total vapor pressure of the solution.

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

P_A and P_B are the vapor pressures of pure A and B, respectively.

Raoult's Law

Key Applications of Raoult's Law

Ideal Solutions: Raoult's Law is most applicable to ideal solutions, where the interactions between unlike molecules are similar to those between like molecules. In such cases, the vapor pressure of the solution follows the ideal behavior predicted by Raoult's Law.

Vapor Pressure Lowering: Raoult's Law explains how the presence of a non-volatile solute in a solvent lowers the vapor pressure of the solution compared to the pure solvent. This phenomenon is crucial in understanding colligative properties such as boiling point elevation and freezing point depression.

Distillation: Raoult's Law is fundamental in distillation processes. It helps predict the composition of the vapor phase at different temperatures during distillation and is used to separate components based on their different vapor pressures.

Azeotropes: Raoult's Law is used to explain the behavior of azeotropic mixtures, where the vapor composition is the same as the liquid composition. Deviations from Raoult's Law can result in the formation of **positive or negative azeotropes**.

Industrial Processes: Raoult's Law is applied in various industrial processes involving liquid-vapor equilibria, such as the production of chemicals, pharmaceuticals, and petrochemicals. Understanding vapor pressures is crucial for designing and optimizing these processes.

Elevation of boiling point

This phenomenon occurs when a non-volatile solute is added to a solvent, resulting in an increase in the boiling point of the solution compared to the pure solvent. The degree of boiling point elevation is proportional to the molality (moles of solute per kilogram of solvent) of the solution and **can be quantitatively described by the equation:**

$$\Delta T_b = i \cdot K_b \cdot m$$

Where:

ΔT_b is the elevation of boiling point,

i is the van't Hoff factor, representing the number of particles into which the solute dissociates or associates in the solution,

K_b is the ebullioscopic constant, a characteristic property of the solvent,

m is the molality of the solution.

The addition of solute disrupts the normal boiling process by impeding solvent vaporization, requiring a higher temperature for the solution to boil. Boiling point elevation has practical implications, such as in **high temperature boiling liquids**, and its understanding **is crucial in various scientific and industrial applications where precise control of solution properties is essential.**

The depression of freezing point is a colligative property observed in solutions, wherein the freezing point of a solvent is lowered upon the addition of a non-volatile solute. This phenomenon is a consequence of the reduction in vapor pressure at the solution's surface, making it more difficult for the solvent to undergo the phase transition from liquid to solid.

The depression of freezing point can be quantitatively expressed by the equation:

$$\Delta T_f = i \cdot K_f \cdot m$$

where ΔT_f is the depression of freezing point, i is the van't Hoff factor representing solute dissociation or association, K_f is the cryoscopic constant specific to the solvent, and m is the molality of the solution.

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Depression of freezing point

The depression of freezing point is exploited in antifreeze formulations for vehicles, preventing engine coolant from solidifying at low temperatures.

Overall, this colligative property provides valuable insights into the thermodynamic behavior of solutions and finds diverse applications across various disciplines.

Osmotic Pressure

Pharmaceutical industries utilize osmotic pressure concepts to design controlled-release drug delivery systems.

In the food industry, osmotic pressure guides preservation techniques like osmotic dehydration.

Chemical engineering relies on osmotic pressure principles in separation processes such as reverse osmosis (RO), while environmental engineering applies them in water treatment methods like ultrafiltration.

Osmotic pressure also informs laboratory techniques like dialysis for solute separation and plays a role in osmometry for chemical analysis.

Also an example of human blood dialysis of kidney failure patients.

In the context of osmotic pressure, Raoult's Law can be related to the ideal behavior of solvents in a solution and their vapor pressures. Osmotic pressure (π) is then given by Van't Hoff's equation:

$$\pi = i \cdot M \cdot R \cdot T$$

Here, i is the Van't Hoff factor, M is the molarity of the solution, R is the ideal gas constant, and T is the temperature.

Raoult's Law directly addresses vapor pressure in ideal solutions, it lays the groundwork for understanding colligative properties such as osmotic pressure when combined with Van't Hoff's equation. Osmotic pressure ultimately depends on the concentration of solute particles and the properties of the solvent.