Chapter from Chemistry 8

Solutions and Colloids. Solubilities. Salts. Solubility rules.Osmosis and dialysis.
Illustration : Decompression Sickness

8.1. Solutions and Colloids

Mixture are either **homogeneous or heterogeneous**. They are always made of two or more nonreacting, pure substances.

Homogeneous mixtures are those in which the tiniest samples are everywhere identical in composition and properties. **Solutions and colloidal dispersions** are two relatively stable homogeneous mixtures. In solutions, the dispersed particless are the smallest, in colloidal dispersions are larger than those in a solution.

Heterogeneous mixture are those which are not homogeneous.

Solutions

Inside a solution, the particles of the solvent and all solutes have sizes of atoms (0.1 – 1 nm). We usually speak about the liquids. But in reality there are more combination producing solutions.

Gaseous solutions :Gas in a gas (air) ; liquid in gas(droplets in a gas – colloidal system)` solid in a gas (colloidal system)

Liquid solutions: Gas in a liquid (CO_2 in water – carbonated beverages)` liquid in liquid (vinegar)` solid in a liquid (sugar in water, sea water)

Solid solutions : Gas in a solid (alloy)` Liquid in a solid (toluene in rubber)` solid in a solid (zinc in copper – brass)

Colloidal dispersions

Dispersed particles have diameters from 1 nm to 1000 nm and consist of very large clusters of ions or molecules. Colloidal dispersions are :

Type	Dispersed phase	Dispersing Medium	Common examples
Foam	Gas	Liquid	whipped cream, suds
Solid foam	Gas	Solid	pumice
Liquid aerosol	Liquid	Gas	fog, clouds
Emulsion	Liquid	Liquid	cream, mayonnaise, milk
Solid emulsion	Liquid	Solid	butter, cheese
Smoke	Solid	Gas	dust in smog
Sol	Solid	Liquid	starch in water, paints
Solis sol	Solid	Solid	pearls, alloys

In colloidal disperssions, the particles are not trapped by ordinary filter paper. They are large enough, however, to reflect and scatter light (Tyndall effect).

Molecules of water are very polar structure with significant partial charges.

8.2. Solubility.

A solution is made of a **solvent** and one or more **solute**s. The solvent is the medium into which the other substances are mixed or dissolved. The solvent is usually liquid, like water (aqueous solutions).

A solute is anything that is dissolved by the solvent. The solute can be a gas, liquid and solid compound.

Solutions can be **saturated** (it is not possible to dissolve more of the solute at the given temperature, dynamic equilibrium exists between the dissolved and undissolved solute), **unsaturated** (is one in which the ratio of solute to solvent is lower than that of the corresponding saturated solution, a solution into which more solute could be dissolved without changing the temperature). Sometimes can be made **supersaturated** solution. This unstable system has the ratio of dissolved solute to solvent actually higher than that of a saturated solution. The separation of a solid from a solution is called **precipitation,** and the solid - the precipitate. Solubility can be expressed as grams of solute per 100g of solvent. It is the maximum quantity of the solute that can dissolve and form stable solution at the given temperature. The solubilities of most solids increase with temperature. The reason is that most solids dissolve endothermically, that means that they require heat to dissolve. Equilibrium for solutes that dissolves endothermically , the equilibrium expression will be:

$$Solid_{undissolved} + solution + heat\ of\ solution \leftrightarrow more\ concentrated\ solution$$

Le Chatelier´s principle, the stress (increasing of temperature) is absorbed by a shift of the equilibrium to the right. The solution becomes even more concentrated. The rate of the reverse reaction also increases, namely the return of solute to the undissolved state. Eventually. the rates again become equal.

The solubilities of some ionic compounds decrease with increasing temperature, especially the salts of the sulfate ion with metal ions (+2,+3, hydrates).

All gases become less and less soluble in water as the temperature increases (under same pressure).

This is because the dissolving of gases in liquids is always exothermic process:

$$gas_{undissolved} + solution \leftrightarrow more\ concentrated\ solution + heat\ of\ solution$$

The stress (increasing of temperature) is absorbed by lifting equilibrium to the left in favor of undissolved gas.

Gases are more soluble under higher partial pressure. The solubilities of oxygen and nitrogen are directly proportional to the applied pressure. The equilibrium expression is :

$$gas_{undissolved} + solvent \leftrightarrow solution$$
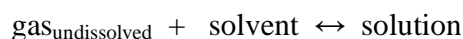
The equilibrium shifts to the right with increasing pressure because only such a change can absorb the volume-squeezing stress of extra pressure.

Similarly, if we reduce the pressure above a liquid that has a dissolved gas, we create a volume-expanding stress and by that we shift the equilibrium to the left. Dissolve gas now leaves the solution.

William Henry noticed that gas solubility is directly proportional to gas pressure - **Henry´s law** (pressure-solubility law):

The concentration of a gas in a liquid at any given temperature is directly proportional to the partial pressure of the gas on the solution.

$$c_g = k_g P_g$$

c_g concentration of gas, k_g constant of proporcionality, P_g partial pressure of gas above the solution

The reference is to **partial** pressure because each gas in a mixture of Gates, like air, dissolves individually according to its own partial pressure and different constant of proporcionality (but for given gas is constant proporcionality the same).

8.3. Salts. Solubility rules.

Salts are ionic compounds whose cations are any except H^+ and whose anions are any except OH^- or O^{2-} . All are crystalline solids at room temperature, because forces of attraction between ions in crystals are very strong. The simple salt is formed of two kinds of oppositely charged ions. Mixed salts have free or more different ions ($KAl(SO_4)_2$ x 12 H_2O).

Formation of salts.

Acid + metal hydroxide $\rightarrow$ a salt + H_2O

Acid + metal bicarbonate $\rightarrow$ a salt + H_2O + CO_2

Acid + metal carbonate $\rightarrow$ a salt + H_2O + CO_2

Acid + metal $\rightarrow$ a salt + H_2

Many salts are insoluble in water. To predict when the salts will be soluble, we use **solubility rules** for ionic compounds:

1. All lithium, sodium, potassium, and ammonium salts are soluble regardless of the counter ion.
2. All nitrates and acetates are soluble, regardless of the counter ions.
3. All chlorides, bromides, and iodides are soluble, except when the counter ion is lead, silver, or mercury(I).
4. All sulfates are soluble except those of lead, kalcium, stroncium, mercury (I), and barium.
5. All hydroxides and metal oxides are insoluble except those of the group IA cations and those of kalcium, stroncium, and barium.
6. All phosphates, carbonates, sulfites, and sulfides are insoluble except those of the group IA cations and NH_4^+.

8.4. Osmosis and dialysis

Colligative properties are properties of solutions or colloidal dispersions that depend only on the number of particles per unit volume of solvent , not on the chemical identities of their solutes (the depression of freezing point,, elevation of the boiling point, migration ability of a solution to pass through semipermeable membranes)

Diffusion is physical process whereby particles intermingle and spread out so as to erase concentration gradient.

Osmosis is the diffusion of solvent molecules through membranes
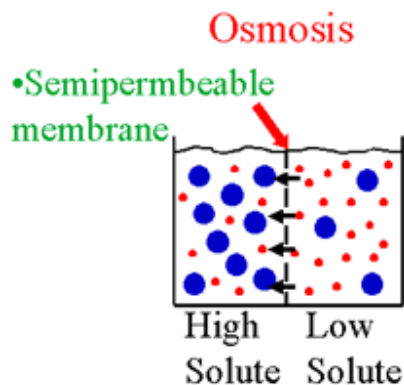
Osmotic membrane. Some membranes have pores so small that only molecules of water can get through them. This membrane is called osmotic membrane. Osmosis is the net migration of the solvent from a solution with the lower concentration of solute into the solution with the higher concentration through the osmotic membrane. The pressure necessary to prevent osmosis is called the **osmotic pressure** of the solution. The value of osmotic pressure is directly proportional to the molar concentration of particles in the solution.

$$\pi = n/V \times RT = c. RT$$

Dialysis is osmosis when the membrane is more permeable. There not only molecules of solvent pass through the membrane but also small molecules move through the membrane.

When the osmotic pressure of blood varies too much, the result can be shock or damage of red blood cells. The body has mechanisms to keep the concentrations of all substances circulating in blood within narrow limits. Consequences of not keeping of relevant concentrations can lead to the life threatening state. When the permeability of the blood capillaries is increased, the colloidal-sized particles leave the blood, the colloidal osmotic pressure of blood decreases. It is less able to take up water from surrounding space. A total blood volume decreases and the brain will not get enough nutrients →shock.

The membranes of red blood cells behave as dialyzing membranes. Within each red cell is an aqueous fluid with dissolved and colloidally dispersed substances. When red cells are placed in pure water, the fluid inside the red cell is more concentrated than the surrounding liquid.Fluid is going to the red cell. The cell burst - the rupturing is called **hemolysis**.

On the other hand when you place the red cells to the solution with higher osmolarity, dialysis occur in the opposite direction. The cells lose the fluid volume, and shrivel and shrink. The process is called **crenation**.

Isotonic solution - two solutions with equal osmolarity.

Hypotonic solution has a lower osmolarity than the one to which it is compared.

Hypertonic solution is one with higher osmotic pressure than another.

Physiological saline solution is the solution of NaCl (0.9%w/w), is isotonic with respect to the fluid inside a red cell.

Osmolarity is the molar concentration of all solute particles active in osmosis or dialysis. Thus 0.1M NaCl solution has molarity 0:1M but osmolarity of 0.2M (osmol/L). The osmolarity of 0.1M Na_2SO_4 is 0.3M (osmol/L) and in case of 0.1M glucose, the osmolarity is the same as the molarity.