

SOLUTIONS

Solute - the substance being dissolved.

Solvent - the substance doing the dissolving (usually present in the greatest amount).

Types of Solution:

Solute	Solvent	Solution Appearance	Example
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda (CO ₂ in water)
Liquid	Liquid	Liquid	Vinegar (CH ₃ COOH in water)
Solid	Liquid	Liquid	Saline solution (NaCl in H ₂ O)
Solid	Solid	Solid	Brass (Zn in Cu) - an alloy

Measuring Solution Concentration:

Molarity (M) = mol solute / L solution

- Since molarity is dependent on volume, which changes slightly with temperature, it is dependent on temperature.

Percent by mass = (Mass of solute/Mass of solution) x 100

Percent by Volume = (Volume of solute/Volume of Solution) x 100

Mass/Volume Percent = (Mass of solute/volume of solution) often measured in mg/dL

Parts per million (ppm), Parts per billion (ppb) and Parts per trillion (ppt)

- For liquids usually based on mass
- For aqueous solutions 1 ppm = 1 mg/L 1 ppb = 1 µg/L 1 ppt = 1 ng/L
- For gaseous solutions, usually based on molecules or volumes. Ex: 1 ppm of a substance in air is 1 molecule of the substance per 1 x 10⁶ molecules of air.

Molality (m) = mol solute / kg solvent

- Note that molality is independent of temperature.

Mole Fraction (X) = moles of component A / Total moles (n_A/ n_A + n_B . . .)

- note there are no units
- the sum of all the mole fractions for all the components in a solution is 1
- **mole percent = mole fraction x 100**

Solution Formation:

The main factors affecting the dissolution of a solute in a solvent are:

1. solute to solute interactions.
2. solvent-solvent interactions.
3. solvent-solute interactions.

Dissolution (formation of a homogeneous mixture) will be favored when the first two are relatively small and the third is relatively large.

Like dissolves like

- in general, if two substances have similar intermolecular forces, the solute should have high solubility in the solvent.

Enthalpy of Solution (heat of solution).

- Consider solution formation as three hypothetical steps, each with an enthalpy change associated with it.

Step 1: Separating solvent molecules ΔH_1

Step 2: Separating solute molecules ΔH_2

Step 3: Allowing the solute and solvent to interact to form the solution ΔH_3

$$\Delta H_{\text{solution}} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

1. Ideal solutions $\Delta H_{\text{solution}} = 0$ $\Delta V_{\text{solution}} = 0$

- All intermolecular forces are of comparable strength.
- No net energy change on solution formation.
- Volume of solution is the sum of the volumes of the solute and solvent.
- Mixtures of ideal gases
- Some liquids come close to forming ideal solutions. Ex: toluene in benzene.

2. $\Delta H_3 > \Delta H_1 + \Delta H_2$ ($\Delta H_{\text{solution}} < 0$ $\Delta V_{\text{solution}} < 0$)

- IE: intermolecular forces between solute and solvent particles are greater than other intermolecular forces.
- Solution formation is exothermic
- Volume of the solution is less than the sum of the solute and solvent volumes.
- Nonideal.

3. $\Delta H_3 < \Delta H_1 + \Delta H_2$ ($\Delta H_{\text{solution}} > 0$)

- IE: intermolecular forces between solute and solvent particles are weaker than other intermolecular forces.
- Solution formation is endothermic.
- May occur due to an increase in entropy (disorder).
- Nonideal.

4. $\Delta H_3 < < \Delta H_1 + \Delta H_2$

- The intermolecular forces between solute and solvent are so much smaller than the sum of the other intermolecular forces that the solution may not form - the two remain as separate (heterogeneous) phases.

For ionic compounds, whether they dissolve in water depends on whether the ion-dipole interactions between ions and water molecules are stronger than the interionic interactions.

Examples: Predict whether the following are likely to be solutions or heterogeneous mixtures.

A) methanol and water B) pentane and octane C) sodium chloride and carbon tetrachloride
D) 1-decanol and water.

Key Terms:

Miscible liquids will mix together in all proportions.

Immiscible liquids do not mix and will form heterogeneous mixtures.

Solvation - the process of dissolving a solute in a solvent

Hydration - the same as solvation, but the term is used when the solvent is water.

Hydrophobic - (water hating) is often used to describe nonpolar substances.

Hydrophilic - (water loving) is often used to describe polar substances.

Factors Affecting Solubility of Solutions:

Equilibrium:

- Dynamic equilibrium occurs when the rate of crystallization is equal to the rate of dissolving.
- **Saturated** solutions have reached dynamic equilibrium. (The solution contains all the dissolved solute it can hold at that temperature).
- The **solubility** of a solute is the concentration of the solution at equilibrium.
- **Unsaturated** solutions contain less solute than can be held at equilibrium.
- **Supersaturated** solutions contain more solute than can be held at equilibrium.

Temperature Effects:

- Solubility changes with temperature.
- If a dissolution process is exothermic, the solubility of a solute will decrease with an increase in temperature.
- If a dissolution process is endothermic, the solubility of the solute will increase with an increase in temperature.
- Most ionic compounds have aqueous solubilities which increase with increased temperature.
- Solubility curves graph solubility v's temperature.

- Supersaturated solutions may be formed by dissolving the solute at a higher temperature then cooling the solution slowly. (Supersaturated solutions are very unstable).

For Gases:

- Most gases are less soluble in water as temperature increases but often become more soluble in organic solvents as temperature increases.

Pressure Effects:

- Changes in pressure have very little effect on the solubilities of solids or liquids in liquid solutions.
- Gas solubilities in all solvents are affected by changes in pressure.
- Henry's Law** $S = k P$ (S=solubility, k=constant depending on the particular gas and temp P=pressure of the gas).

Colligative Properties:

- Physical properties of solutions that depend on the number of solute particles present in the solution rather than the type of solute particles.

A. Colligative Properties for Non-Electrolytes:

- Electrolytes** - ionize in solution.
- Strong electrolytes** - solutes that exist in solution almost exclusively as ions. Most ionic compounds, and a few molecular compounds (mainly acids such as HCl).
- Weak electrolytes** - solutes that are only partially ionized in solution. Some molecular compounds, carboxylic acids, and amines.
- Nonelectrolytes** - solutes that exist in solution almost exclusively as molecules. Most molecular and organic compounds.

1. Lowering of Vapor Pressure.

- The presence of a nonvolatile solute in a solution lowers its vapor pressure.
- Raoult's Law** (for ideal solutions):

$$P_{\text{solv}} = X_{\text{solv}} P_{\text{solv}}^0$$

$$P_{\text{solv}} = \text{Pressure of the solvent, } X_{\text{solv}} = \text{mole fraction of solvent, } P_{\text{solv}}^0 = \text{Pressure of pure solvent}$$
- For solutions containing two volatile liquids, a modified version of Raoult's Law applies:

$$P_{\text{Total}} = P_A + P_B = X_A P_A^0 + X_B P_B^0$$
- Liquid-liquid solutions that obey Raoult's Law are **ideal solutions**.
- Ideal solutions occur where solute and solvent are similar.
- Where solute-solvent interactions are strong (such as hydrogen bonding), a lower vapor pressure than expected will be observed Ex: an acetone-water solution.
- Where solute-solvent interactions are weaker than the interactions in the pure liquids (endothermic), vapor pressure will be higher than expected Ex: ethanol and hexane (polar and nonpolar).

- The vapor in equilibrium with an ideal solution of two volatile components has a higher mole fraction of the more volatile component than is found in the liquid. This can be utilized in fractional distillation.

2. Boiling Point Elevation.

- Since the vapor pressure of a solvent decreases with the addition of a non-volatile solute, it must be heated to a higher temperature than the pure solvent to cause its vapor pressure to equal atmospheric pressure (boiling point). IE: addition of a non-volatile solute causes an increase in the boiling point.
- For a nonvolatile, nonelectrolyte solute:
 $\Delta T_b = K_b m$ $\Delta T_b = bp_{(soln)} - bp_{(pure\ solvent)} = bp\ elevation$ $K_b = \text{molal boiling point constant (different for each solvent)}$ $m = \text{molality of solute.}$
- Boiling points will not be constant since the concentration of solute in the solution will change as the solvent evaporates. As the solution boils, the boiling point will continue to rise as the concentration increases.

3. Freezing Point Depression.

- The freezing point of a solution will be lower than the freezing point of the pure solvent.
- $\Delta T_f = -K_f m$
 $\Delta T_f = fp_{(soln)} - fp_{(solvent)} = fp\ depression$ $K_f = \text{molal freezing point constant (different for each solvent)}$ $m = \text{molality of solute.}$
- Freezing points will not be constant since only the solvent freezes - the solute concentration increases as freezing occurs.
- Note: Make sure you can calculate molar mass from freezing point or boiling point data.
- Make sure you read the examples such as salt on icy sidewalks and antifreeze.

4. Osmotic Pressure (π).

- **Semipermeable membranes** allow small solvent molecules to pass through but not larger solute particles.
- **Osmosis** is the net flow of pure solvent particles through the semipermeable membrane into a solution.
- **Osmotic pressure** is the pressure required to stop osmosis from pure solvent into solution.
 $\pi = MRT$ $\pi = \text{osmotic pressure (in atm or Torr)}$, $V = \text{volume of solution (L)}$,
 $M = \text{molarity of solution}$, $R = \text{gas constant}$, $T = \text{temp (K)}$
- **Isotonic** solutions have identical osmotic pressures.
- **Hypertonic** solutions are more concentrated (have higher osmotic pressure).
- **Hypotonic** solutions are less concentrated (have lower osmotic pressure).
- Note: Be sure you can calculate molar masses from osmotic pressure data.

B. Colligative Properties for Solutions of Electrolytes.

- Colligative properties depend on the number of particles. For nonionizing nonelectrolytes, one mole of solute contains 1 mole of particles. For ionizing electrolytes, 1 mole of solute gives more than 1 mole of particles.
- The colligative property equations can be used after modification with the **van't Hoff Factor (i)**.
- **$i = \text{moles of particles in solution} / \text{moles of solute dissolved}$**
- Although the van't Hoff factor should be easily predictable for strong electrolytes. it is often lower than expected, possibly due to ion pairing.

$$\Delta T_b = i K_b m$$

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

$$\pi = i MRT$$

Colloids:

Solution - Homogeneous mixture with particles less than 1 nm.

Suspension - heterogeneous mixture where solute is temporarily dispersed through the solvent. Particle sizes are $> 1000 \text{ nm}$.

Colloid - dispersion of particles with sizes from 1 nm to 1000 nm. Colloidal particles cannot generally be seen through an ordinary microscope.

Properties of colloids:

- Often appear milky or cloudy
- **Tyndall effect** - a beam of light passed through a colloid will scatter.
- **Coagulation** (destruction of a colloid) can be attained by heating the solution or by adding an electrolyte.

Types of Colloids

Dispersed Phase	Dispersion Medium	Colloid Type	Examples
Solid	Liquid	Sol	Paint, clays, gelatin
Liquid	Liquid	Emulsion	Milk, mayonnaise,
Gas	Liquid	Foam	Whipped cream, soap suds, meringue
Solid	Gas	Aerosol	Smoke, dust, airborne bacteria
Liquid	Gas	Aerosol	Fog, hair sprays
Solid	Solid	Solid sol	Ruby glass, black diamond
Liquid	Solid	Solid emulsion	Butter, pearl, opal
Gas	Solid	Solid foam	Marshmallow, pumice