Solutions

A solution is a homogeneous mixture

The components of a solution are:

- solute: substance (or substances) present in lesser amount being dissolved
- solvent: substance present in greater amount doing the dissolving

Types of solution:

On the basis of nature of solvent and solute

On the basis of amount of solute in solvent/solubility

Types of solutions
(on the basis of the nature of solute and solvent)

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 platinum
solid	solid	liquid	mercury in silver
solid	solid	solid	silver in gold (alloys)

Terms describing how liquids mix

miscible: substances that are soluble in each other in any proportion.

immiscible: substances that do not mix, forms two layers

Types of solutions
(On the basis of the amount of solute and solvent/solubility)

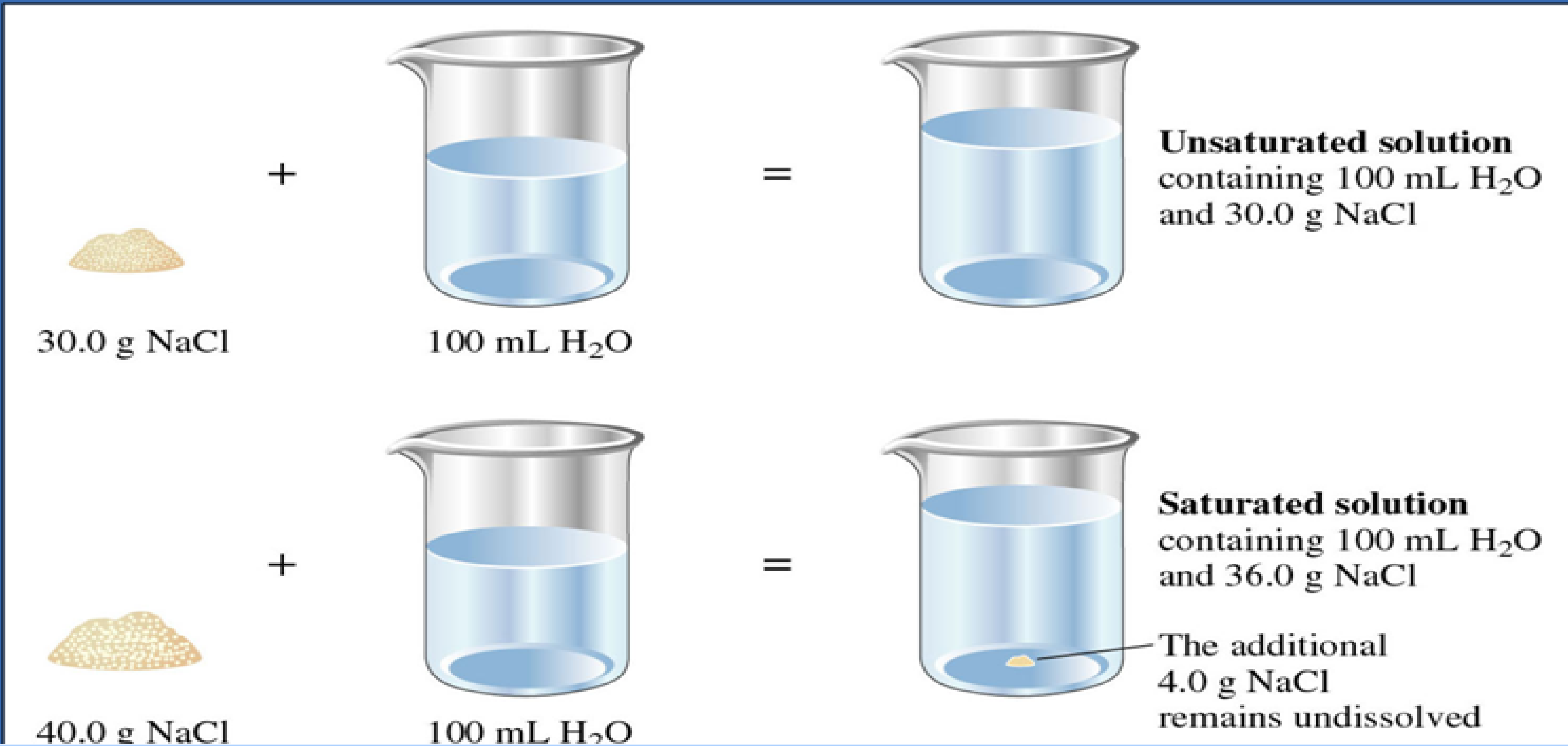
1) **saturated solution**
   - contains the maximum amount of solute in a given solvent at a specific temperature
   - a solution in equilibrium with undissolved solute

2) **unsaturated solution**
   - contains less solute than it has the capacity to dissolve
   - a solution not in equilibrium with dissolved solute
   - more solute can be dissolved

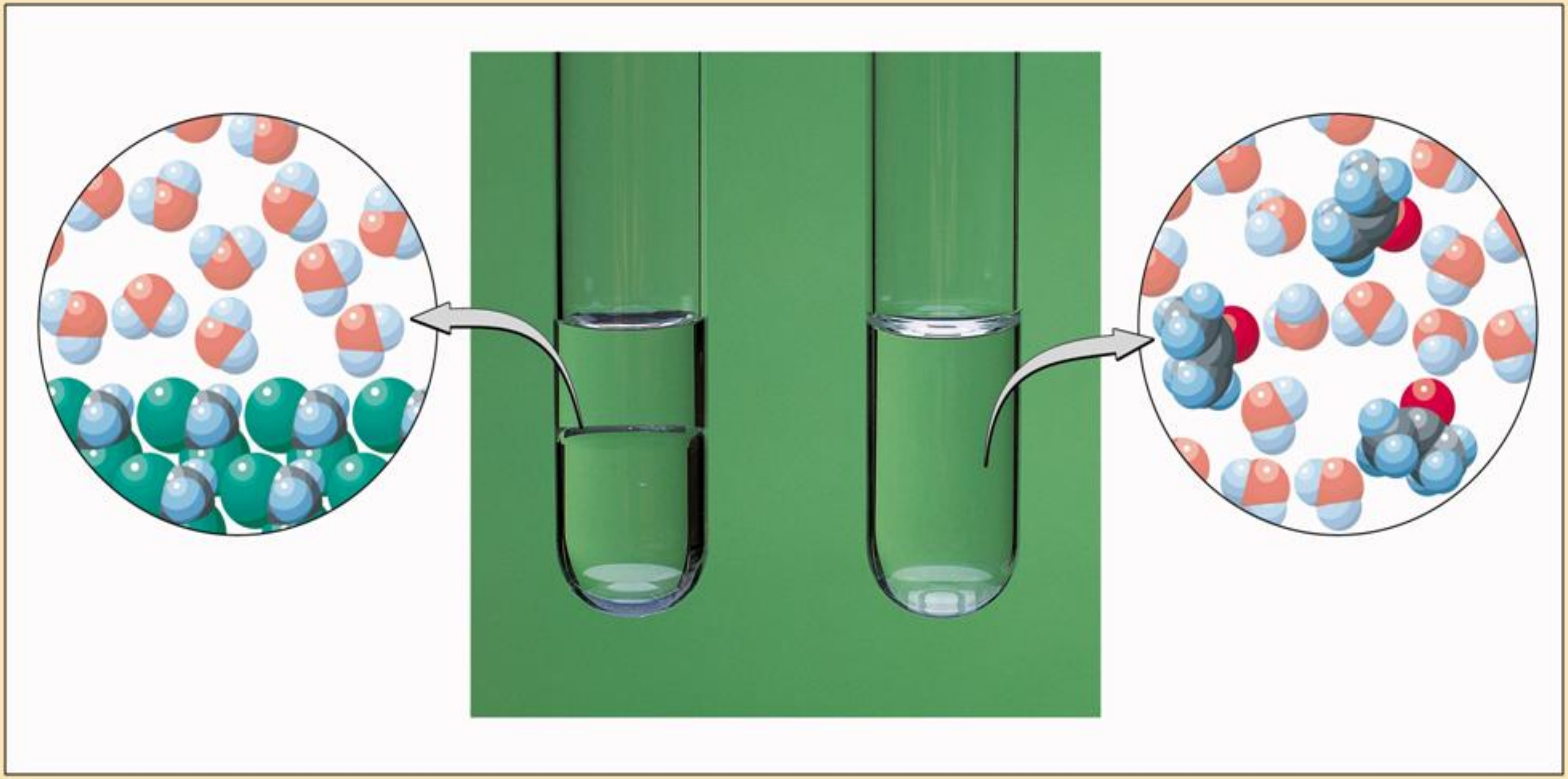
3) **supersaturated solution**
   - contains more solute than is present in a saturated solution
   - unstable solution
   - prep: heat solution to high temperature, then slowly cool

Recognition of unsaturated, saturated and supersaturated solutions.
Hints: Addition of solute to the respected solution

Figure 12.3: Comparison of unsaturated and saturated solutions.



Miscible and immiscible



Liquid Solutions

- **Liquid solutions** are the most common types of solutions found in the chemistry lab.
  - Many inorganic compounds are soluble in water or other suitable solvents.
  - Rates of chemical reactions increase when the likelihood of molecular collisions increases.
  - This increase in molecular collisions is enhanced when molecules move freely in solution.

Solid Solutions

- **Solid solutions** of metals are referred to as alloys.
  - Brass is an alloy composed of copper and zinc.
  - Bronze is an alloy of copper and tin.
  - Pewter is an alloy of zinc and tin.

Solubility and the Solution Process

- The amount of a substance that will dissolve in a solvent is referred to as its **solubility**.
  - Many factors affect solubility, such as **temperature** and, in some cases, **pressure**.
  - There is a limit as to how much of a given solute will dissolve at a given temperature.
  - A **saturated solution** is one holding as much solute as is allowed at a stated temperature.

Analytically, solubility can be expressed as grams of solute/100 grams of solvent

Examples

The solubility of NaCl in water at 80^{o}C is 40g/100 g of water

How many grams of NaCl should dissolve in 250 grams of water at 80^{o}C?

= 100 g NaCl will dissolve

Solubility: Saturated Solutions

- Sometimes it is possible to obtain a **supersaturated solution**, that is, one that contains more solute than is allowed at a given temperature.
  - Supersaturated solutions are unstable.
  - If a small crystal of the solute is added to a supersaturated solution, the excess immediately crystallizes out

Figure 12.4: Crystallization from a supersaturated solution of sodium acetate.

Salting out process of NaCl

Equilibrium:

$$NaCl\,(s) \xrightleftharpoons{H_2O} Na^+\,(aq) + Cl^-\,(aq)$$

↓ **Addition of concentrated HCl**

The above equilibrium will be shifted to left side

↓

Salting out of NaCl crystal

Factors in Explaining Solubility

- In most cases, **"like dissolves like."**
  - This means that polar solvents dissolve polar (or ionic) solutes and nonpolar solvents dissolve nonpolar solutes.
  - The relative **force of attraction** of the solute for the solvent is a major factor in their solubility.

Intermolecular forces in liquid molecules

- Ion-dipole (very much strong) 10-50 kJ/mol
- Dipole-dipole (weak) 3-4 kJ/mol
- Hydrogen bonding (strong) 10-40 kJ/mol
- London dispersion force (weak) 1-10 kJ/mol

Explaining Solubility

Solubility depends on

1) natural tendency for solute and solvent to mix
2) tendency for system to have lowest energy possible

"Like dissolves like" i.e.:

- substances with similar IMF tend to dissolve in one another
- nonpolar solutes dissolve in nonpolar solvents (IMF: London forces)
- polar & ionic solutes are soluble in polar solvents
- macromolecules are not soluble in either polar or nonpolar solvents

Molecular solutions

"like dissolves like"

- nonpolar solutes with nonpolar solvents:
  mixing and London dispersion forces
- polar solutes with polar solvents
  mixing and dipole/dipole interactions

Ionic solutions

- ionic compound with polar solvent
  energy of attraction between an ion and water (**ion-dipole force**)
- attraction of ions for water molecules (**hydration**)
  must be stronger that the attraction for ion in the crystal (lattice energy)

Effects on Solubility

Temperature

1) solid in water
   in most cases (not all)
   solubility of solid ↑ as temp ↑

   examples of exceptions where solubility ↓ with a temp ↑
   $CaSO_4$, $Ca(OH)_2$, $Ce_2(SeO_4)_3$

2) gas in water
   solubility of gas ↓ as temp ↑

Pressure

1) Solid and liquid: Not much effect with pressure

2) Gases
   are greatly affected by pressure

Relationship between gas solubility and pressure

Henry's Law: solubility of a gas in a liquid is directly proportional to the partial pressure of a gas over the solution

$S = k_H P$

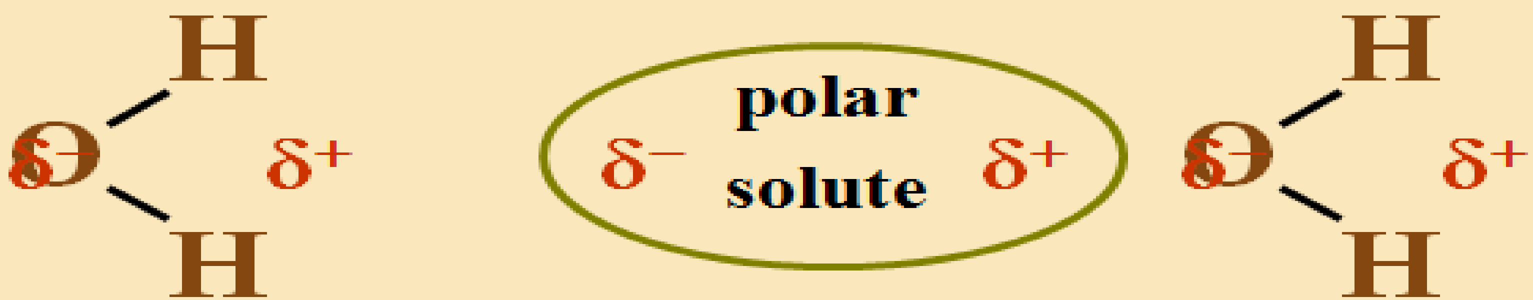
S - solubility
P - partial pressure of gas
k_H - Henry's Law constant for a given gas & temperature

- system at equilibrium
- ↑ partial pressure of gas - more molecules will dissolve in liquid
- why? more molecules are striking the surface of the liquid

Molecular Solutions

Polar molecules interact well with polar solvents such as water.

- The dipole-dipole interactions of water with a polar solvent can be easily explained as electrostatic attraction.



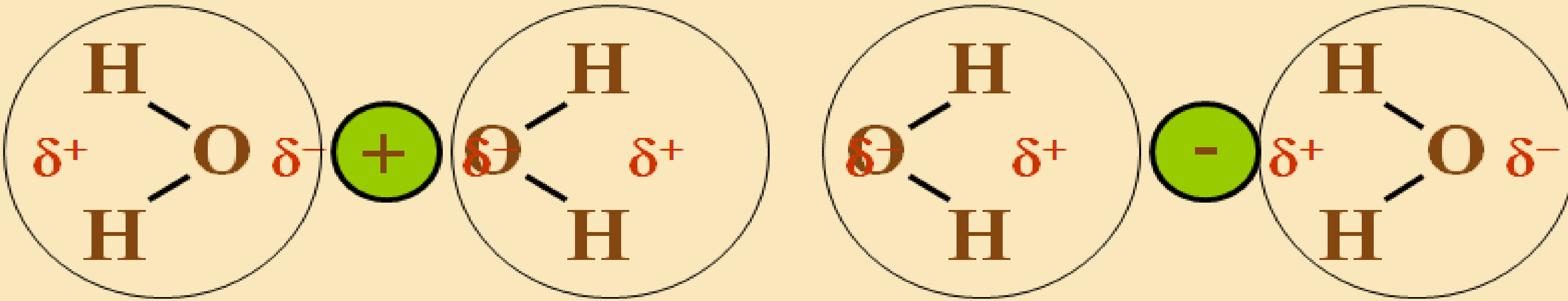
Molecular Solutions

Nonpolar solutes interact with nonpolar solvents primarily due to London forces.

- Heptane, C_7H_{16}, and octane, C_8H_{18}, are both nonpolar components of gasoline and are completely miscible liquids.
- However, for water to mix with gasoline, hydrogen bonds must be broken and replaced with weaker London forces between water and the gasoline.
- Therefore gasoline and water are nearly immiscible.

Ionic Solutions

- **Polar solvents,** such as water, also interact well with **ionic solutes**.
  - Since ionic compounds are the extreme in polarity, we can illustrate the electrostatic attractions of water for cations and anions.



Effects of Temperature and Pressure on Solubility

- The solubility of solutes is very **temperature dependent**.
  - For **gases dissolved in liquids**, as temperature increases, solubility decreases.
  - On the other hand, for most **solids dissolved in liquids**, solubility increases as temperature increases

Temperature Change

- Heat can be evolved or absorbed when an ionic compound dissolves in water.
  - This **heat of solution** can be quite noticeable.
  - When NaOH dissolves in water, it gets very warm (the solution process is exothermic).
  - On the other hand, when ammonium nitrate dissolves in water, it becomes very cold (the solution process is endothermic).

Pressure Change

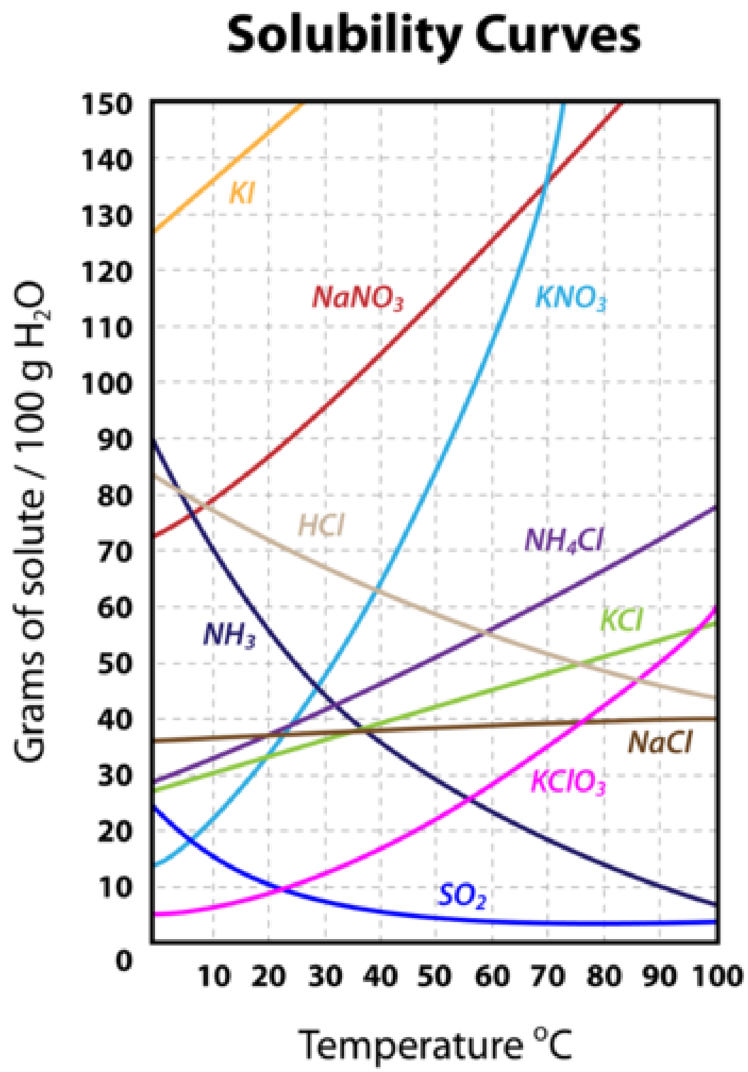
Henry's Law states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas in direct contact with the liquid.

- Expressed mathematically, the law is

$$S = k_H P$$

where **S** is the solubility of the gas, k_H is the Henry's law constant characteristic of the solution, and **P** is the partial pressure of the gas.

Discuss the comparative solubility of gas and solid in a solvent (For example water)



Ways of Expressing Concentration

- **Concentration expressions** are a ratio of the amount of solute to the amount of solvent or solution.
  - The quantity of solute, solvent, or solution can be expressed in **volumes** or in **molar** or **mass** amounts.
  - Thus, there are several ways to express the concentration of a solution.

Molarity

- The molarity of a solution is the moles of solute in a liter of solution.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

  - For example, 0.20 mol of ethylene glycol dissolved in enough water to give 2.0 L of solution has a molarity of

$$\frac{0.20\ \text{mol ethylene glycol}}{2.0\ \text{L solution}} = 0.10\ \text{M ethylene glycol}$$

Mass Percentage of Solute

- The mass percentage of solute is defined as:

$$\text{Mass percentage of solute} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

  - For example, a 3.5% sodium chloride solution contains 3.5 grams NaCl in 100.0 grams of solution.

Molality

- The molality of a solution is the moles of solute per kilogram of solvent.

$$\text{molality}\,(m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

  - For example, 0.20 mol of ethylene glycol dissolved in 2.0 x 10^3 g (= 2.0 kg) of water has a molality of

$$\frac{0.20\ \text{mol ethylene glycol}}{2.0\ \text{kg solvent}} = 0.10\ m\ \text{ethylene glycol}$$

A Problem to Consider

- What is the molality of a solution containing 5.67 g of glucose, C6H12O6, dissolved in 25.2 g of water?
  - First, convert the mass of glucose to moles.

$$5.67\ \mathrm{g\ C_6H_{12}O_6} \times \frac{1\ \mathrm{mol\ C_6H_{12}O_6}}{180.2\ \mathrm{g\ C_6H_{12}O_6}} = 0.0315\ \mathrm{mol\ C_6H_{12}O_6}$$

  - Then, divide it by the kilograms of solvent (water).

$$\text{Molality} = \frac{0.0315\ \mathrm{mol\ C_6H_{12}O_6}}{25.2 \times 10^{-3}\ \mathrm{kg\ solvent}} = 1.25\ m\ \mathrm{C_6H_{12}O_6}$$

Mole Fraction

The mole fraction of a component "A" (χ_A) in a solution is defined as the moles of the component substance divided by the total moles of solution (that is, moles of solute and solvent).

$$\chi_A = \frac{\text{moles of substance A}}{\text{total moles of solution}}$$

- For example, 1 mol ethylene glycol in 9 mol water gives a mole fraction for the ethylene glycol of 1/10 = 0.10.

mass percentage of solute

$$mass\ \%\ of\ solute = \frac{mass\ solute}{mass\ sol'n} \times 100$$

volume percentage of solute

$$volume\ \%\ of\ solute = \frac{volume\ solute}{volume\ sol'n} \times 100$$

mass to volume percentage of solute (g/mL)

$$mass\ to\ volume\ \%\ of\ solute = \frac{mass\ solute}{volume\ sol'n} \times 100$$

Comparison of Concentration Units

molarity

- easier to measure volume of solution, than to weigh solution

molality

- independent of temperature
  (volume of solution usually ↑ with ↑ temp.)

mol fraction

- good for calculations of partial pressures of gas
- dealing with vapor pressures of solutions

percent by mass

- independent of temperature
- do not need to know molar mass of solute

Colligative Properties

- depend on number of solute particles **NOT** the nature of the solute particles
- particles are atoms, ions or molecules

i - van't Hoff factor

- show the number of particles a substance produces in solution

Substance		Particles in Solution	van't Hoff factor
$C_6H_{12}O_6$ (s) glucose	$\xrightarrow{H_2O}$	$C_6H_{12}O_6$ (aq) nonelectrolyte	i = 1
NaCl (s)	$\xrightarrow{H_2O}$	Na^+ (aq) + Cl^- (aq)	i = 2
$Mg(NO_3)_2$ (s)	$\xrightarrow{H_2O}$	Mg^{2+} (aq) + 2 NO^{3-} (aq)	i = 3

Colligative Properties

1) Vapor Pressure Lowering $\quad \Delta P = iP^o{}_A\chi_B$
2) Boiling Pt Elevation $\quad \Delta T_b = ik_bc_m$
3) Freezing Pt Depression $\quad \Delta T_f = ik_fc_m$
4) Osmotic Pressure $\quad \Pi = iMRT$