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

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

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Types of solutions

solubility: max. amount of solute that will dissolve in a given amount of solvent at a specific temperature.

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

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

Types of Concentration Units

Concentration - amount of solute present in a given amount of solution (solution)

1) **Molarity (M)**

$$Molarity = \frac{moles\ of\ solute}{liters\ of\ sol'n}$$ units of mol/L

2) **Molality (m)**

number of moles of solute dissolved in 1 kg of solvent

$$molarity = \frac{mol\ of\ solute}{kg\ of\ solvent}$$

3) **Mole Fraction (χ, greek symbol chi)**

mole fraction of component A = $\chi_A = \frac{mol\ of\ A}{total\ mol\ of\ sol'n}$

total mol of solution = mol solute + mol solvent

4) **Percent Concentration** (*not in text)

a) mass percentage of solute

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

*b) volume percentage of solute

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

*c) 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

Examples using various Concentration Units:

1) An experiment calls for 35.0 g of hydrochloric acid solution that is 20.2% HCl by mass. a) How many grams of HCl is in this solution? b) How many grams of water?

given: solution conc. given in % mass $20.2\%\ \mathrm{HCl} = \dfrac{0.202\ g\ HCl}{1\ g\ sol'n}$

35.0 g of HCl solution

a) $35.0\ g\ HCl\ sol'n\left(\dfrac{0.202\ g\ HCl}{1\ g\ sol'n}\right) = 7.07\ g\ HCl$

b) g solution = g solvent + g solute
g H_2O = g solution - g HCl = 35.0 g solution - 7.07 g HCl = 27.9 g H_2O

2) What mass of glucose ($C_6H_{12}O_6$) must be dissolved in 563 g ethanol (C_2H_5OH) to prepare a 2.40×10^{-2} *m* solution?

3) Hydrogen peroxide is a powerful oxidizing agent that is used in concentrated solution in rocket fuel systems and in dilute solution as a hair bleach. an aqueous solution of H_2O_2 is 30.0% by mass and has a density of 1.11 g/mL. Calculates its

a) molarity

b) molality

c) mole fraction of H_2O_2

Colligative Properties

1) Vapor Pressure Lowering	$\Delta P = iP^{o}_{A}\chi_{B}$
2) Boiling Pt Elevation	$\Delta T_b = ik_b c_m$
3) Freezing Pt Depression	$\Delta T_f = ik_f c_m$
4) Osmotic Pressure	$\Pi = iMRT$

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

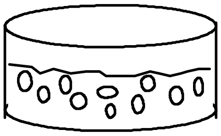
i.e.

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

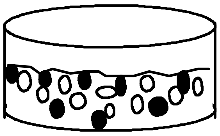
1) **Vapor Pressure (v.p.) Lowering**

a) If a nonvolatile solute is present in solution, the vapor pressure of the solution is always less than pure solvent (less solvent at the surface of the solution to escape into the vapor phase)

b) The relationship between solution vapor pressure and solvent vapor pressure depends on concentration of solute in solution



pure solvent



solution

vapor pressure lowering $\Delta P = iP_A^o\chi_B$

$\Delta\text{P} - P_A^o - P_A$

P_A^o - partial pressure of pure solvent

χ_B - mol fraction of solute

Note: vapor pressure lowering is directly proportional to χ_B

Vapor pressure lowering can be derived from Raoult's Law

$$P_A = P_A^o\chi_A$$

partial pressure of a solvent over a solution, P_A, is given by vapor pressure of pure solvent P_A^o times mol fraction of solvent in the solution, χ_A

if solution contains one solute: $\chi_A + \chi_B = 1$

substitute into Raoult's law: $\chi_A = (1- \chi_B)$

distribute P_A^o : $P_A = P_A^o(1-\chi_B)$

solve for $P_A^o\chi_B$: $P_A = P_A^o - P_A^o\chi_B$

simplify: $P_A^o\chi_B = P_A^o - P_A$

vapor pressure lowering equation: $P_A^o\chi_B = \Delta P$

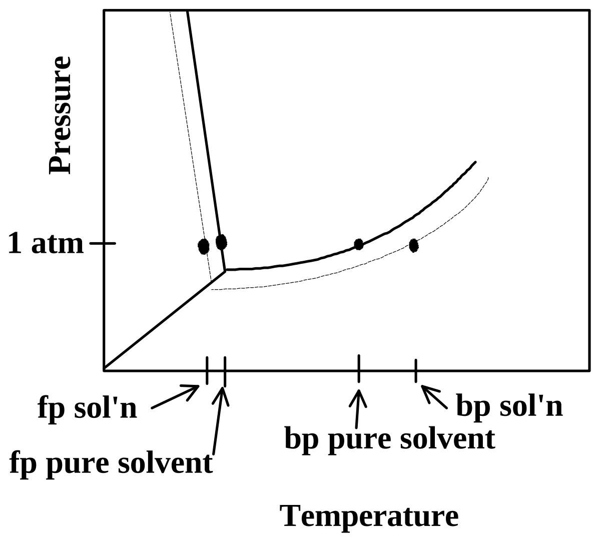
Example of vapor pressure lowering:

Eugenol, $C_{10}H_{12}O_2$ is the chief constituent of oil of clove. It is a pale yellow liquid that dissolves in ethanol, C_2H_5OH; it has a bp of 255°C (relatively low vapor pressure). What is the vapor pressure lowering at 20.0°C of ethanol containing 8.56 g of eugenol in 50.0 g of ethanol? The vapor pressure of ethanol at 20.0°C is 4.6 mmHg.

2) Boiling point (bp) elevation

a) nonvolatile solute lowers vapor pressure of the solution which also affects boiling point of the solution

b) recall: boiling point - temperature at which vapor pressure of solution equals external atmospheric pressure



The solid line represents phase boundaries for the pure solvent and the dotted line represents phase boundaries for the solution.
The normal freezing point (fp) and boiling point (bp) of the pure solvent and the solution are shown by the dots.
At 1 atm pressure the bp of the solution is higher than the pure solvent.

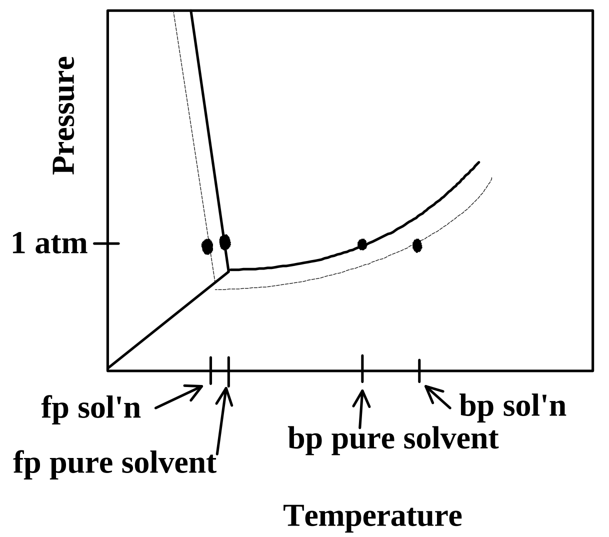
bp elevation ΔT_b

$\Delta T_b = T_b - T_b^o$: T_b - bp solution; T_b^o - bp pure solvent

$\Delta T_b = i\,k_b c_m$:
i - van't Hoff factor
k_b - bp elevation constant (depends on solvent)
c_m - molality concentration

3) **Freezing point (fp) depression**

usually pure solvent freezes out of solution



The solid line represents phase boundaries for the pure solvent and the dotted line represents phase boundaries for the solution.
The normal freezing point (fp) and boiling point (bp) of the pure solvent and the solution are shown by the dots.
At 1 atm pressure the fp of the solution is lower than the pure solvent.

fp depression ΔT_f

$\Delta T_f = T_f - T_f^o$: T_f - fp solution; T_f^o - fp pure solvent

$\Delta T_f = i\, k_f c_m$: i - van't Hoff factor
k_f - fp elevation constant (depends on solvent)
c_m - molality concentration

Note: solute must be nonvolatile for bp elevation; No restriction for fp depression

Ethylene glycol (EG) CH_2OHCH_2OH, is a common automobile antifreeze. It is water soluble and fairly nonvolatile (bp = 197°C). Calculate the freezing pt. of a solution containing 651 g of EG in 2505 g of water. Would you keep this substance in your car during the summer? (EG, M_m = 62.01 g/mol)

Determining Molar mass from bp elevation or fp depression experiments

Derive an equation relating bp to the M_m of solute

bp elevation equation. $\Delta T_b = ik_b c_m$

define c_m $$c_m = \frac{mol\ solute}{kg\ solvent}$$

sub c_m into bp elevation equation. $$\Delta T_b = ik_b \frac{(mol\ solute)}{(kg\ solvent)}$$

define mol solute $$mol\ solute = \frac{g\ solute}{M_m\ solute}$$

sub mol solute $$\Delta T_b = \frac{ik_b \quad (g\ solute)}{(kg\ solvent)(M_m\ solute)}$$

rearrange and solve for M_m

$$M_m = \frac{ik_b \quad (g\ solute)}{\Delta T_b \quad (kg\ solvent)}$$

Example:

An 11.2 g sample of sulfur was dissolved in 40.0 g of carbon disulfide (CS_2). the bp elevation of CS_2 was found to be 2.63 ^oC. What is the M_m of the sulfur in solution?

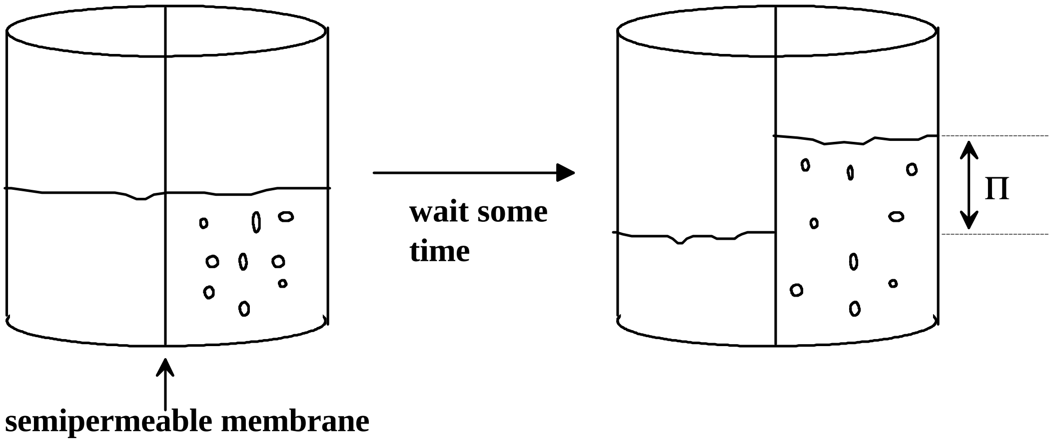
What is the formula of molecular sulfur?

4) Osmotic pressure (P)

Osmosis

is the net movement of solvent molecules through a semipermeable membrane from a pure solvent or from a dilute solution to a more concentrated solution.

semipermeable membrane - allows solvent molecules to pass through but blocks passage of solute molecules.



Osmotic pressure

pressure required to stop osmosis

P = i*M*RT

M - molar concentration
R - gas law constant = 0.0821 L-atm/mol-K
T - absolute temperature
i - van't Hoff factor

Osmosis

- an important biological process

A cell is an aqueous solution surrounded by semipermeable membrane and outside the membrane is an aqueous solution

- If a solution has the same osmotic pressure as then cell, the solution is isotonic
- If the solution has a greater osmotic pressure than the cell, the solution is hypertonic
- If the solution has a smaller osmotic pressure than the cell, the solution is hypotonic

If a cell is place in a

- hypertonic solution, the solvent flows out of cell and the cell shrinks → crenation
- hypotonic solution, the solvent flows into the cell and the cell swells and burst → hemolysis

Colloids

- Dispersion of particle of one substance throughout another substance
- Dividing line between solutions and heterogeneous mixtures
- Dispersed particles are larger than molecules in a solution but not so large that the compounds of the mixture separate under the influence of gravity
- Examples: Fog, smoke, whipped cream, milk, paint, marshmallow, butter, ruby glass
- Particle size: 10Å to 2000Å (can not see without microscope)
- Particles are large enough to scatter light
  Tyndall effect - scattering of light by colloidal particles

Types of Colloids

aerosols - liquid droplets (fog) or solid particles dispersed throughout a gas
emulsion - liquid droplets dispersed throughout another liquid (milk)
sol - solid particles dispersed in a liquid (biological fluids)
foam - gas in a liquid or solid (whipping cream)

Two major classes of colloids

1) hydrophilic colloid
   - water-loving
   - strong attraction between dispersed phase and water

2) hydrophobic colloid
   - water-hating
   - lack of attraction between dispersed phase and water
   - unstable, aggregates into larger particles

Coagulation - process by which dispersed phase of a colloid is made to aggregate and separate from the continuos phase

examples of coagulation
curdling of milk - lactose ferments to lactic acid
delta formation - soil in river water meets concentrated ionic solution in the ocean

Colloids can have both major classes present in one substance

- one end is hydrophilic and the other end is hydrophobic
- in water the hydrophobic ends associate together while the hydrophilic ends associate with water molecules.

This process is known as forming micelles

micelles - a colloidal-sized particle formed in water by the association of molecules or ions that have a hydrophobic end and a hydrophilic end.

Soaps are colloids which form micelles- refer to your text book