NOTES: Solutions & Solubility

Solution:

The solvent is the medium in which the solutes are dissolved. (the solvent is usually the most abundant substance).

Example: saltwater (solution)
salt (solute)
water (solvent)

The process of dissolution is favored by:

1) A decrease in the energy of the system (exothermic)
2) An increase in the disorder of the system (entropy)

Liquids Dissolving in Liquids:

- Liquids that are soluble in one another ("mix") are MISCIBLE.
- Polar liquids are generally soluble in other polar liquids.
- Nonpolar liquids are soluble in other nonpolar liquids.

Factors affecting the rate of dissolution:

1) Surface area/particle size
   - Greater surface area, faster it dissolves
2) Temperature
   - Most solids dissolve faster @ higher temperatures
3) Agitation
   - Stirring/shaking will speed up dissolution

SATURATION:

A solid solute dissolves in a solvent until the solution is SATURATED (equilibrium exists between dissolved & undissolved solute)

- **Unsaturated solution** - is able to dissolve more solute
- **Saturated solution** - has dissolved the maximum amount of solute
- **Supersaturated solution** - has dissolved excess solute (at a higher temperature). Solid crystals generally form when this solution is cooled.

Solubility:

The amount of solute that will dissolve in a given solvent at a given temperature and pressure

Factors affecting solubility:

- **The nature of the solute and solvent**: different substances have different solubilities
- Temperature: many solids substances become more soluble as the temp of a solvent increases; however, gases are less soluble in liquids at higher temps.
- Pressure: Only affects the solubility of gases. As pressure increases, the solubility of gases increases.

Notes: Concentration of Solutions

mass % = $\frac{\text{mass solute}}{\text{Total mass}} \times 100$

ppt (‰) = $\frac{\text{mass solute}}{\text{total mass}} \times 1000$

ppm = $\frac{\text{mass solute}}{\text{total mass}} \times 10^6$

ppb = $\frac{\text{mass solute}}{\text{total mass}} \times 10^9$

Concentration refers to amount of solute dissolved in a solution.

MOLARITY (M) = $\frac{\text{mol solute}}{\text{L solution}}$

- Molarity is the # moles solute per liter of solution.
- The symbol [M] is used to denote molarity of a substance

Dilution:

When you dilute a solution, you can use this equation:

$$M_1 \times V_1 = M_2 \times V_2$$

This is on the TEST *

Example: Describe how you would prepare 2.50 L of 0.665 M Na_2SO_4 solution starting with:

a) solid Na_2SO_4

$0.665\ M = \frac{x}{2.50\ L}$

$x = 1.6625$ mol

1.6625 mol × 142.1 g / 1 mol = 236 g

Dissolve 236 g of Na_2SO_4 in enough water to create 2.50 L of solution

b) 5.00 M Na_2SO_4 solution

$M_1 \times V_1 = M_2 \times V_2$

$(5.00\ M)(V_1) = (0.665\ M)(2.50\ L)$

$V_1 = 0.333\ L = 333\ mL$

Add 333 mL of Na_2SO_4 and dilute with water until you have 2.50 L of solution

ppt (‰) = $\frac{\text{mass solute}}{\text{total mass}} \times 1000$

MASS PERCENT = mass % = $\frac{\text{mass solute}}{\text{total mass of solution}} \times 100$

Example: What is the percent of NaCl in a solution made by dissolving 24 g of NaCl in 152 g of water?

$\frac{24\ g}{176\ g} \times 100 = 14\%$

MOLALITY = $\frac{\text{mol solute}}{\text{Kg solvent}}$

Molality = m = the number of moles of solute per kilogram of solvent.

Example: What is the molaloty of a solution that contains 12.8 g $C_6H_{12}O_6$ in 187.5 g of water?

$12.8\ g \times \frac{1\ mol}{180.18\ g} = 0.7104$ mol

$m = \frac{0.07104\ mol}{.1875\ Kg} = 0.37888 \approx 0.379\ m$

Example: How many grams of H_2O must be used to dissolve 50.0 g of sucrose to prepare a 1.25 m solution of sucrose, $C_{12}H_{22}O_{11}$?

$50.0\ g\ C_{12}H_{22}O_{11} \times \frac{1\ mol}{342.0\ g} = 0.1462$ mol

$1.25\ m = \frac{0.1462\ mol}{x}$

$x = 0.11696 \approx 117\ g\ H_2O$

NORMALITY = M x total positive ox

Example: Find the normality of a 3.0 M solution of H_2SO_4?

$N = (3.0\ M)(+2) = 6.0\ N$

2.0 M $Mg_3(PO_4)_2$

Concentration of each ion?

$[Mg^{2+}] = 3(2.0\ M) = 6.0\ M$

$[PO_4^{3-}] = 2(2.0\ M) = 4.0\ M$

Colligative Properties of Solutions

Colligative properties = physical properties of solutions that depend on the # of particles dissolved, not the kind of particle.

1) lowering vapor pressure
2) raising boiling point
3) lowering freezing point
4) generating an osmotic pressure

Boiling Point Elevation

a solution that contains a nonvolatile solute has a higher boiling pt than the pure solvent; the boiling pt elevation is proportional to the # of moles of solute dissolved in a given mass of solvent.

$$\Delta T_b = K_b m$$

where: ΔT_b = elevation of boiling point
m = molality of solute
K_b = the molal boiling point elevation constant

K_b values are constants; see table 10.2, p. 294

K_b for water = 0.52 °C/m

Example: What is the normal boiling pt of a 2.50 m glucose, $C_6H_{12}O_6$, solution?

$\Delta T_b = K_b m$
$= (0.52°C/m)(2.50\ m)$
$= 1.3°C$

$T_b = 100.0°C + 1.3°C = \boxed{101.3°C}$

Example: How many grams of glucose, $C_6H_{12}O_6$, would need to be dissolved in 535.5 g of water to produce a solution that boils at 101.5°C?

$1.5\ °C = (0.52°C/m)(m)$
$m = 2.885$

$2.885\ m = \frac{x}{0.5355\ kg}$
$x = 1.5449\ mol$
$1.5449\ mol \times \frac{180\ g}{1\ mol} = 278.1\ g = \boxed{280\ g}$

Freezing/Melting Point Depression

the freezing pt of a solution is always lower than that of the pure solvent.

$$\Delta T_f = K_f m$$

where: ΔT_f = lowering of freezing point
m = molality of solute
K_f = the freezing point depression constant

K_f for water = 1.86°C/m

Ex: Calculate the freezing pt of a 2.50 m glucose solution.

$\Delta T_f = K_f m$
$= (1.86°C/m)(2.50 m)$
$= 4.65°C$

$T_f = 0.00°C - 4.65°C = -4.65°C$

Ex: When 15.0 g of ethyl alcohol, C_2H_5OH, is dissolved in 750 grams of formic acid, the freezing pt of the solution is 7.20°C. The freezing pt of pure formic acid is 8.40°C. Determine K_f for formic acid.

15.0 g C_2H_5OH × 1 mol / 46.0 g = 0.3255 mol

m = 0.3255 mol / 0.750 kg = 0.434 m

$\Delta T_f = K_f m$

1.20°C = (K_f)(0.434 m) / 0.434 m 0.434 m = 2.764977 ≈ 2.8°C/m

Ex: An antifreeze solution is prepared containing 50.0 cm^3 of ethylene glycol, $C_2H_6O_2$, (d = 1.12 g/cm^3), in 50.0 g water. Calculate the freezing point of this 50-50 mixture. Would this antifreeze protect a car in Chicago on a day when the temperature gets as low as -10° F?
(-10°F = -23.3°C)

50.0 L $C_2H_6O_2$ × 1.12 g / cm^3 × 1 mol / 62.0g = 0.90323 mol

0.90323 mol / 0.050 kg = 18.1 m

$\Delta T_f = K_f m$
$\Delta T_f = (1.86°C/m)(18.1 m)$
$\Delta T_f = 33.7°C$
$T_f = 0°C - 33.7°C = -33.7°C$

yes it would

Electrolytes and Colligative Properties

- Colligative properties depend on the # of particles present in solution.

- Because ionic solutes dissociate into ions, they have a greater effect on freezing pt and boiling pt than molecular solids of the same molal concentration

- For example, the freezing pt of water is lowered by 1.86°C with the addition of any molecular solute at a concentration of 1 m.

- However, a 1 m NaCl solution contains 2 molal conc. of IONS. Thus, the freezing pt depression for NaCl is 3.72°C...double that of a molecular solute.

- The relationships are given by the following equation:

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

and

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

where: ΔT_f = freezing pt depression and ΔT_b =boiling point elevation

K_f = the molal freezing pt constant
K_b = the molal boiling pt constant
m = the molality of the solute
n = the # of particles formed from the dissociation of each formula unit of the solute.

Examples: What is the freezing pt of:

a) a 1.15 m sodium chloride solution?

$NaCl \rightarrow Na^+ + Cl^-$ n=2

$\Delta T_f = K_f \cdot m \cdot n$
$= (1.86°C/m)(1.15 m)(2)$
$= 4.28°C$

$T_f = 0.00°C - 4.28°C = \boxed{-4.28°C}$

b) a 1.15 m calcium chloride solution?

$CaCl_2 \rightarrow Ca^{2+} + Cl^-$ n=3

$\Delta T_f = (1.86°C/m)(1.15 m)(3)$
$= 6.417°C$

$T_f = 0.00°C - 6.417°C = \boxed{-6.42°C}$

c) a 1.15 calcium phosphate solution?

$Ca_3(PO_4)_2 \rightarrow Ca^{2+} + PO_4^{3-}$ n=5

$\Delta T_f = (1.86°C/m)(1.15 m)(5)$
$= 10.7°C$

$T_f = 0.00°C - 10.7°C = \boxed{-10.7°C}$

Determining Molecular Weights by Freezing Point Depression

Ex: A 1.20 g sample of an unknown compound is dissolved in 50.0 g of benzene. The solution freezes at 4.92°C. Determine the molecular weight of the compound. The freezing pt of pure benzene is 5.48°C, and the K_f for benzene is 5.12°C/m.

$5.18\,°C - 4.92\,°C = 0.26\,°C$

$0.26\,°C = (5.12\,°C/m)(m)$

$m = 0.05078$

$0.05078\,m = \frac{x}{0.050\,kg}$

$x = 0.002539\ mol$

$\frac{1.20\,g}{0.002539\,mol} = \boxed{473\ g/mol}$

Ex: A 37.0 g sample of a new covalent compound was dissolved in 200.0 g of water. The resulting solution froze at -5.58°C. What is the molecular weight of the compound?