CHEM 1035 – Lecture 43

Solution Properties

What happens when 2 or more compounds are mixed? A mixture has 2 characteristics – its composition is variable and it retains some of the properties of its components

2 possibilities –

1. One "dissolves" in the other and they form a solution
2. The are mixed, but retain the individual properties of the 2 compounds

Solutions are homogeneous mixtures (homogeneous – the same throughout). Mixtures that are not solutions can be heterogeneous (different throughout)

Effect of Pressure on Solubility

- For solids and liquids, pressure has no significant influence on solubility. Why?
- For gases, solubility increases with increasing partial pressure of the gas above the liquid. Why?
  - This is quantitatively given by Henry's law:

$$S_{gas} = k_H \times P_{gas}$$

Gases and solids are not compressible. Pressure, therefore, has no influence on these phases of matter.

With gases, the solubility increases as the partial pressure of the gas above the liquid increases. This is due to the increase in the number of collisions that occur between the liquid and the gas (from the kinetic molecular theory of gases, increasing pressure is due either to an increase in the number of collisions, or to an increase in the velocity of the collisions)

Solution Concentration

There are several different ways of expressing the concentration of solutions:

- Molarity – Moles of solute/Liter of solution
- Molality – Moles of solute/Kilogram of Solvent
- Mass Percentage – (Mass of solute/Mass of solution)*100%
- Volume Percentage – (Vol. of solute/Volume of solution)*100%
- Mole Fraction – Mole of solute/Moles of solution

Molarity is symbolized by a capitol M that is underlined (M)
Molality is symbolized by a script, lower-case m (*m*)

Example Problem

An 8.00 mass-% solution of ammonia has a density of 0.9651 gm/mL. What are:

1. The Molality,
2. The Molarity,
3. The mole fraction

Of this solution.

Colligative properties of Solutions

Changes in the physical properties of liquids due to the presence of solutes.

The colligative properties refers to changes in boiling point, freezing point, vapor pressure, and osmotic pressure.

These properties are independent of the chemical identity of the solute, they are only a function of the # of solute particles

Consider a solution that is 1.0L of 0.1 M glucose and one that is 1.0L of 0.1M NaCl. Which has more solute “particles”

NaCl when dissolved gives 2 “particles” for each NaCl molecule. The NaCl, therefore, will have 2x the effect on the colligative properties than the undissociated glucose.

Vapor Pressure Lowering

- Raoult's Law

$$P_{solvent} = X_{solvent}P^{o}_{solvent}$$

The Raoult's law is a limiting law and it holds for dilute solutions. Why does the presence of a solute lower the vapor pressure?

Entropy.

Boiling point elevation

$$\Delta T_b = K_b m$$

Where *m* is the Molality of the solute particles. A 1.00*m* solution of NaCl has 2.0*m* of solute particles.

What is the boiling temperature? The temperature at which the vapor pressure of the liquid is equal to the external pressure. Because a solute lowers the vapor pressure of the solvent, this says that additional energy (heat) is required to increase the vapor pressure to the value that it would be if a solute were not present.

We have seen in Raoult's Law that increasing the solute in a solution will depress the vapor pressure. This would result in having to increase the temperature even higher so that the depressed vapor pressure might become equal to the external pressure. In other words from a molecular view we might expect the boiling point to be elevated when solute is increased in a solution.

Freezing Point Depression

$$\Delta T_m = K_m m$$

The presence of a solute lowers the freezing point of the solvent. Why? When freezing the solvent, the solute has to “undissolve” this process goes against the natural driving force for increased Entropy; hence, we need to remove additional energy from the system (in the form of lower temperatures) to effect this change.

Practical application of this? Melting the ice on the roads after snow or ice.