Chapter 12 – Properties of Solutions

Section 12–1: The Nature of Aqueous Solutions

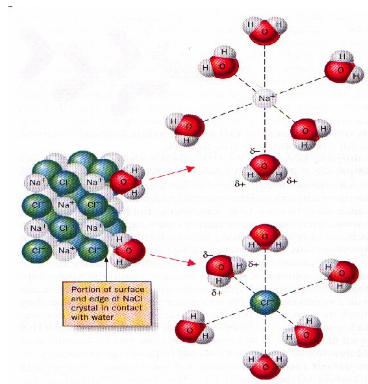
1) Sec 12–1.1 – Mixtures of Two Liquids

   When two liquids are mixed and they form a homogeneous mixture, the liquids are said to be **miscible**.

   When to liquids are mixed and they form a heterogeneous mixture, the liquids are said to be **immiscible**.

   Alcohol and water are miscible.
   Oil and water are immiscible.

2) Sec 12–1.2 – The Formation of Aqueous Solutions of Ionic Compounds



   a) Stirring or shaking the solution – The solute dissolves more rapidly because the agitation brings fresh solvent into contact with solute. This does not change the actual amount of solute that dissolves.

   b) Heating the solution usually causes more solute to dissolve and it also causes it to dissolve faster.

   c) The particle size affects the rate at which a solute dissolves. The greater the surface area exposed (smaller particles) to the solvent, the faster the solute will dissolve. This does not change the amount of solute that dissolves.

Section 12–2: The Effects of temperature and Pressure on Solubility

3) Solubility
   a) Solution – A homogeneous mixture in which all the ions or molecules are fully intermingled.

   b) Solvent – The medium into which material (solute) is mixed or dissolved.

   c) Solute – Any substance that is dissolved in a solvent.

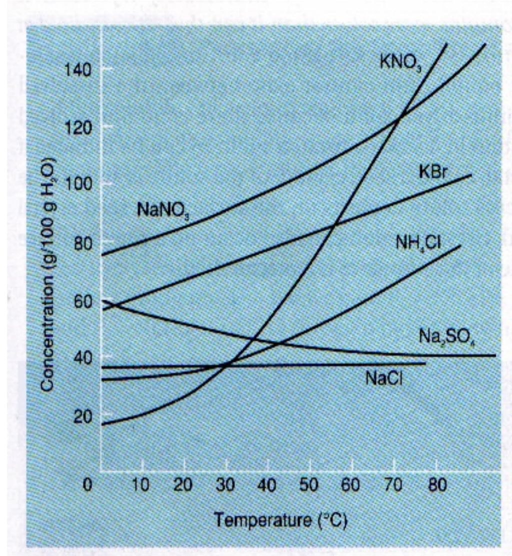
   d) Saturated Solution – The solvent contains all the solute it can hold permanently at a given temperature.

   e) The solubility of a substance is the amount of substance that dissolves in a given quantity of a solvent at a given temperature.

   Example – How much NaCl can be dissolved in 750 g of water at 25 °C. (Assume: 35 g / 100g)

f) Unsaturated Solution – A solution that contains less solute than a saturated solution.

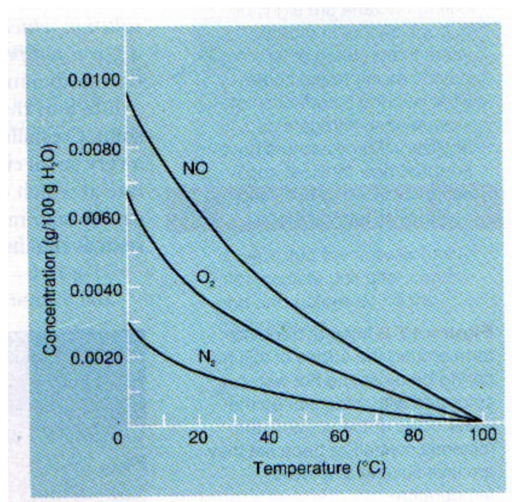
g) Supersaturated Solution – A solution which contains more solute than it can permanently hold at a given temperature. Eventually the excess solute will settle out.



4) Factors Affecting Solubility

a) The solubility of most solids in water increases as the temperature is increased.

Example – How much KNO_3 will dissolve in 100 g of water at 20 °C and how much will dissolve at 60 °C?

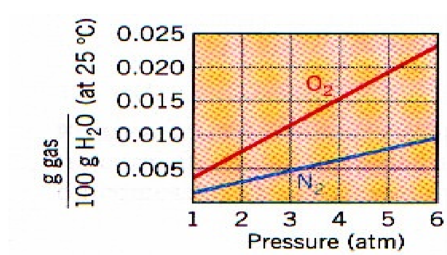


b) The solubility of gases is greater in cold water than in hot water. (This generalization should not be extended to nonaqueous solvents, where the situation is not so simple.)

Example – How much oxygen will dissolve in 100 g of water at 20 °C and how much will dissolve at 60 °C?

c) The solubility of gases in a liquid can be increased by putting the gas under pressure. Henry's Law states that, at a given temperature, the solubility of a gas in a liquid (S) is directly proportional to the pressure of the gas above the liquid (P).

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$



Example – A gas has a solubility at 0 °C of 3.6 g/L at a pressure of 100 kPa. What pressure is needed to produce an aqueous solution containing 9.5 g/L at 0 °C?

Section 12–3: Concentration: Percent By Mass

5) Mass Percent
One way of describing the concentration of a solution is **mass percent** or percent by mass. This is sometimes referred to as "weight percent", weight-weight percent" or "mass-mass percent." It is often indicated by the symbols **% (w/w)**.

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

Example: A solution is prepared by mixing 20.0 g ethanol, C_2H_5OH, 100.0 g of water. Calculate the mass percent of alcohol in this solution.

6) Sec 12–3.1 – Parts per Million and Parts per Billion

Percent by mass is equivalent to parts per hundred. Sometimes, however, concentrations are very low and "percent by mass" would give a very small number. In these cases, two related units become convenient to use: **parts per million** and **parts per billion**.

The basic calculation is done the same way (mass of solute ÷ mass of solution) but the answer is multiplied by 10^6 instead of 100% for part per million, ppm.

Similarly, if you multiply the quotient by 10^9 you get parts per billion, ppb.

$$\frac{2.5 \times 10^{-3}\ \text{g (solute)}}{1.0 \times 10^{3}\ \text{g (solvent)}} \times 100\% = 2.5 \times 10^{-4}\ \%$$

$$\frac{2.5 \times 10^{-3}\ \text{g (solute)}}{1.0 \times 10^{3}\ \text{g (solvent)}} \times 10^{6} = 2.5\ \text{ppm}$$

$$\frac{2.5 \times 10^{-3}\ \text{g (solute)}}{1.0 \times 10^{3}\ \text{g (solvent)}} \times 10^{9} = 2.5 \times 10^{3}\ \text{ppb}$$

7) Volume Percent

A similar way to describe concentration is to use **volume percent** or **percent by volume**. The calculation is like that for mass percent.

$$\text{Volume percent} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

Example: A solution contains 25.0 mL of acetone in water with a total volume of 80.0 mL. What is the percent by volume of acetone in this solution?

[In some instances people also talk about concentration in terms of "mass-volume percent." This does not show up often but do not be confused by it. The idea is the same as with mass percent.]

Section 12–4: Concentration: Molarity

8) Sec 12--4.1 – Calculations Involving Molarity

a) The Concentration of a solution is the measure of the amount of solute that is dissolved in a given amount of the solvent.

b) A dilute solution contains only a small amount of solute.

c) A concentrated solution contains a large amount of solute.

d) The most important unit of concentration is molarity. It is the number of moles of solute dissolved in 1 liter of solution. It is expressed with the symbol M or *M*.

$$M = \frac{\text{moles of solute}}{\text{volume of solution (L)}}$$

A 6 M HCl solution means: $\dfrac{\text{6 moles of HCl}}{\text{1 L of solution}} = \dfrac{\text{6 moles of HCl}}{\text{1000 mL of solution}}$

Examples

1) Calculate the molarity of a solution made by dissolving 80 g of NaOH in enough water to give 400 mL of final solution.

2) How many grams of silver nitrate, $AgNO_3$, are needed to make 2 liters of a 0.20 M silver nitrate solution?

3) Calculate the molarity of a concentrated HCl solution if it has a density of 1.18 g/mL and contains 37 g of HCl per 100 g of solution.

9) Sec 12–4.2 – Dilution of Concentrated Solutions

Dilution – The process in which more solvent is added to a solution in order to lower its concentration. The number of moles of solute does not change when a solution is diluted.

$$M_1V_1 = M_2V_2$$

Example – How many milliliters of 1.0 M NaCl solution are needed to make 250 mL of an 0.20 M NaCl solution?

Section 12–5: Stoichiometry Involving Solutions

10) Sec 12--5.1 – Stoichiometry and Molarity
Just like other reactions, we need to have the balanced chemical equation to know how the reaction goes. In the case of acid-base reactions, we need to include the mole-to-mole ratio in the solution calculation

a A + b B ----> c C + d D where A, B, C and D are compounds and a, b, c, and d are the coefficients

$M_A \times L_A ==$ moles A

moles A $\times \dfrac{\text{b mol B}}{\text{a mol A}} =$ moles B

moles B $= M_B \times L_B$

If you combine all three of these equations you can come up with:

$M_b \times V_b = \dfrac{c_2}{c_1} \times M_a \times V_a$ where c_2/c_1 is the ratio of coeeficients
M_b and M_a are the molarities of the base and acid
V_b and V_a are the volumes of the base and acid

Sometimes you need to do additional calculations. Say, if you calculate the unknown molarity or volume you can calculate the moles and then ...

Examples
1. How many grams of NaOH are needed to neutralize 100 mL of 6.0 M H_2SO_4?

2 What volume of carbon dioxide at 20 °C and 740 mm of Hg will be produced by the reaction of 55 g of Na_2CO_3 with 400 mL of 2.0 M HCl?

11) Neutralization Reactions
Neutralization reactions (remember Chapter 8?) involve acid and bases almost always in solutions. These occur when just enough strong base reacts exactly with the strong acid in a solution. These can also happen when weak acids and weak bases react, but the reaction happens more slowly.

The calculations involved are the same as those mentioned before.

12) Sec 12–5.2 – Titrations

Laboratory experiments often require that we find the molarity of a solution from stoichiometric data.

Titrations are volume-to-volume stoichiometry problems where the concentration is the unknown variable. We can use an **indicator** (a compound that produces a visual clue) or some other means to identify when the reaction is complete.

Example: What volume of 0.250 M NaOH is required to react completely with 13.5 mL of 0.450 M H_2SO_4?

[Remember that you have to have a balanced chemical equation!]

13) Normality {This is NOT in your book!}

In order to make calculations even easier, chemists sometimes use the term **normality** as another unit of concentration. This term is most often used with acids and bases, and focuses mainly on the H^+ and OH^- available in the reaction.

Normality introduces another term: "equivalence." One equivalent of an acid is the amount of that acid that can produce 1 mole of H^+ ions. Likewise, one equivalent of a base is defined as the amount of that base that can furnish 1 mole of OH^- ions.

1 mole HCl = 1 equiv HCl
1 mole HNO_3 = 1 equiv HNO_3
but 1 mole H_2SO_4 = 2 equiv H_2SO_4
1 mole H_3PO_4 = 3 equiv H_3PO_4

Normality (N) is defined as the number of equivalents of solute per liter of solution.

$$\text{Normality} = \frac{\text{number of equivalents}}{\text{1 liter of solution}} = \frac{\text{equivalents}}{\text{liter}} = \frac{\text{equiv}}{\text{L}}$$

This makes calculations even easier since $\mathbf{N_a \times V_a = N_b \times V_b}$ and the concept of equivalents gets rid of the need to use the mole ratio in the calculation. The normality calculation has already taken that into consideration.

Examples:

a. A solution of sulfuric acid contains 86.0 g of H_2SO_4 per liter of solution. Calculate the normality of this solution.

b. Calculate the normality of a solution containing 23.6 g of KOH is 755 mL of solution.

Normality is not always accepted as a unit of concentration because it is not "universal" in its application, and is best used with acids and bases. I mention it because, like the "classical" names of Chapter 4, it does appear in literature and problems and you should be aware of it.