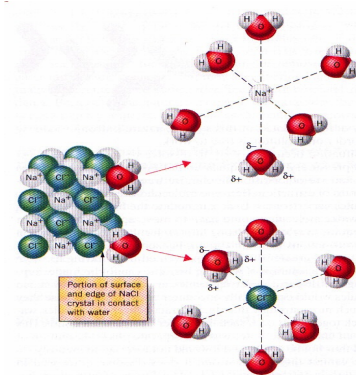


Chapter 15 – Properties of Solutions

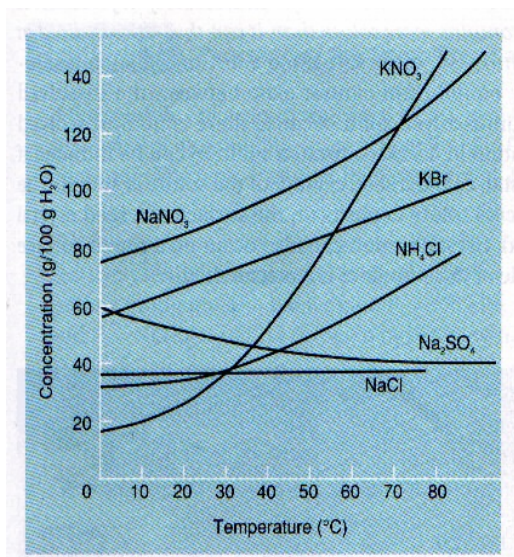
1) Sec 15.1 – Solution Formation

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



2) Sec 15.2 – Solubility

- Solution** – A homogeneous mixture in which all the ions or molecules are fully intermingled.
- Solvent** – The medium into which material (solute) is mixed or dissolved.
- Solute** – Any substance that is dissolved in a solvent.
- Saturated Solution** – The solvent contains all the solute it can hold permanently at a given temperature.
- 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)

- Unsaturated Solution** – A solution that contains less solute than a saturated solution.
- Supersaturated Solution** – A solution which contains more solute than it can permanently hold at a given temperature. Eventually the excess solute will settle out.

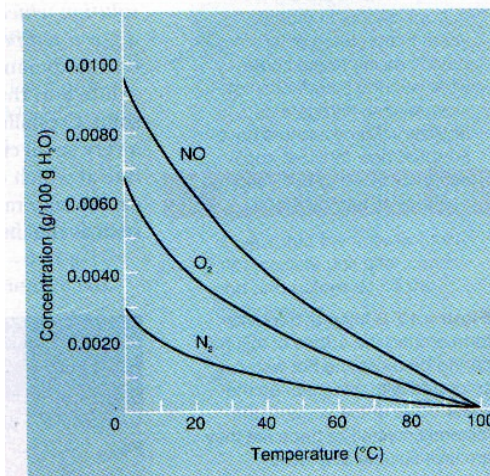
3) Sec 15.3 – Factors Affecting Solubility

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

Example – How much KNO₃ 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.)

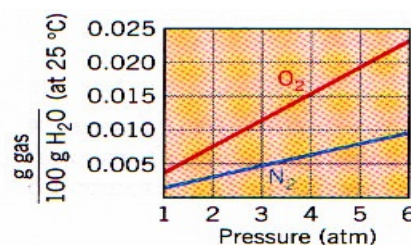
Example – How much oxygen will dissolve in 100 g of water at 20 °C and how much will dissolve at 60 °C?



- 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?



4) Sec 15.4 – Mass Percent

One way of describing the concentration of a solution is **mass percent**. This is sometimes referred to as “weight percent” or even “mass-mass percent.”

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

Example: A solution is prepared by mixing 20.0 g ethanol, C₂H₅OH, 100.0 g of water. Calculate the mass percent of alcohol 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.]

5) Sec 15.4 – Molarity

- The Concentration of a solution is the measure of the amount of solute that is dissolved in a given amount of the solvent.
- A dilute solution contains only a small amount of solute.
- A concentrated solution contains a large amount of solute.
- 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.

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

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

Examples

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

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

- 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.

6) Sec 15.6 – Making Dilutions

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?

7) Sec 15.7 – Solution Stoichiometry

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

$$M_b * V_b = n * M_a * V_a$$

where n is the mole-to-mole ratio
 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

Examples

- How many grams of NaOH are needed to neutralize 100 mL of 6.0 M H_2SO_4 ?
- 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?

8) Sec 15.8 – 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.

9) Sec 15.9 – Normality

In order to make calculations even easier, chemists 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.

$$\begin{array}{l} 1 \text{ mole HCl} = 1 \text{ equiv HCl} \\ 1 \text{ mole HNO}_3 = 1 \text{ equiv HNO}_3 \\ \text{but } 1 \text{ mole H}_2\text{SO}_4 = 2 \text{ equiv H}_2\text{SO}_4 \\ 1 \text{ mole H}_3\text{PO}_4 = 3 \text{ equiv H}_3\text{PO}_4 \end{array}$$

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

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

This makes calculations even easier since $N_a \times V_a = N_b \times V_b$ and the concept of equivalents gets rid of the need to use the mole ratio in the calculation.

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 in 755 mL of solution.

10) Sec 15.10 – Colligative Properties – Boiling Point and Freezing Point

A solution property that depends upon the concentration of a solute (ie., the number of solute particles) is called a colligative property.

Examples of these are boiling point elevation (raising the boiling point of a liquid by adding a solute) and freezing point depression (lowering the freezing point by adding a solute). The antifreeze in your car has ethylene glycol added to raise the boiling point of the water in the radiator. Salt spread on the road lowers the freezing point of water (and, thus, melts the ice).