Section 14.1 The Nature of Acids and Bases

Arrhenius concept:

Acid – produces hydrogen ions in aqueous solutions.
Bases – produces hydroxide ions in aqueous solutions.

Bronsted – Lowery model

Acid – a proton (H+) donor.
Base – a proton acceptor

Hydronium ion the ion that forms when water accepts a proton (H+)

General reaction that occurs when an acid is dissolved in water.

$$\underset{\text{Acid}}{HA(aq)} + \underset{\text{Base}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate Acid}}{H_3O^+(aq)} + \underset{\text{Conjugate Base}}{A^-(aq)}$$

Conjugate Base everything that remains of acid molecule after a proton is lost.

Conjugate Acid the molecule which is formed when the proton is transferred to the base.

Conjugate Acid - Base pair consists of two substances related to each other by the donating and accepting of a single proton.

Equilibrium expression for a general acid dissolved in water reaction would be

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{[H^+][A^-]}{[HA]}$$

K_a is called the Acid Dissociation Constant.

Section 14.2 Acid Strength

Strong Acid - one in which the equilibrium lies far to the right. It has a large value for K_a. Strong acids also yeild weak conjugate bases.

Weak Acid - one in which the equilibrium lies far to the left. It has a small value for K_a. Weak acids yeild relatively strong conjugate bases. See Table 14.1 p.651

Diprotic Acid - an acid having two acidic protons. ex. - H_2SO_4

Oxyacids - the acidic proton is attached to an oxygen atom.

Organic Acids - those acids with a carbon atom backbone, and commonly contain the carboxyl group.

Monoprotic Acid - acids having only one acidic proton.

Amphoteric - can behave either as an acid or as a base. ex. - water

Autoionization - involves the transfer of a proton from one water molecule to another to produce a hydroxide ion and a hydronium ion.

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.

Ion-Product Constant K_w (or dissociation constant for water) - always refers to the autoionization of water. Using the above reaction for the autoionization of water:

$$K_w = [H_3O^+][OH^-] = [H^+][OH^-]$$

at 25°C

$$[H^+] = [OH^-] = 1.0 \times 10^{-7}$$

so at 25°C

$$K_w = (1.0 \times 10^{-7}\ M)^2 = 1.0 \times 10^{-14}\ mol^2/L^2$$

It is important to recognize the meaning of K_w. The product of $[H^+]$ and $[OH^-]$ must always equal 1.0×10^{-14}. This lends itself to three possibilities:

1. A neutral solution, where $[H^+] = [OH^-]$.
2. An acidic solution, where $[H^+] > [OH^-]$.
3. A basic solution, where $[H^+] < [OH^-]$.

Section 14.3 The pH Scale

pH Scale - provides a convenient way to represent solution acidity

$$pH = -\log [H^+]$$

Significant Figures for Logarithms - the number of decimal places in the log is equal to the number of significant figures in the original number.

Similar scales can be formed for pK and pOH.

$$pOH = -\log [OH^-]$$
$$pK = -\log K$$

Many relationships maybe formed between these three equations - useful equations to know

$$K_w = [H^+][OH^-]$$
$$\log K_w = \log [H^+] + \log [OH^-]$$
$$pK_w = pH + pOH$$

and at 25°C

$$pH + pOH = 14.00$$

Section 14.4 Calculating the pH of Strong Acid Solutions

Common Strong Acids - (commit to memory) - HCl (aq), HNO_3(aq), H_2SO_4(aq), and $HClO_4$(aq)

One key to solving acid-base equilibrium, is you must focus on the solution components and their chemistry.

Major Species - are those solution components present in relatively large amounts.

Section 14.5 Calculating the pH of Weak Acid Solution

Step 1 - List the major species in the solution.
Step 2 - Choose the species that can produce H^+, and write a balanced equation for the reaction producing H^+.
Step 3 - Using the values of the equilibrium constant for the reactions you have written, decide which equilibrium will dominate in producing H^+.
Step 4 - Write the equilibrium expression for the dominant equilibrium.

Step 5 - List the initial concentrations of the species participating in the dominant equilibrium.
Step 6 - Define the change needed to achieve equilibrium; that is, define x.
Step 7 - Write the equilibrium concentrations in terms of x.
Step 8 - Substitute the equilibrium concentrations into the equilibrium expression.
Step 9 - Solve for x the "easy" way: that is by assuming $[HA]_0 - x \approx [HA]_0$
Step 10 - Using the 5% rule verify whether the approximation is valid.
Step 11 - Calculate $[H^+]$ and pH

5% Rule - $\frac{x}{[HA]_0} \times 100 < 5\ \%$ then the approximation is acceptable.

Percent Dissociation - $$\text{Percent dissociation} = \frac{\text{amount dissociated } \left(\text{mol}/\text{L}\right)}{\text{initial concentration } \left(\text{mol}/\text{L}\right)} \times 100\%$$

For a given weak acid, the percent dissociation increases as the acid becomes more dilute. For solutions of any weak acid HA, $[H^+]$ decreases as $[HA]_0$ decreases, but the percent dissociation increases as $[HA]_0$ decreases.

Section 14.6 Bases

Strong Base - one which dissociates completely in water and has large K_b values.

General reaction for a base B and water is given by:

$B(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$BH^+(aq)$	+	$OH^-(aq)$
Base		Acid		Conjugate Acid		Conjugate Base

The equilibrium constant for this equation:

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Weak Base - small K_b values.

Section 14.7 Polyprotic Acids

Polyprotic Acids - acids which can furnish more than one proton. Table 14.4 p.677 - important K_a values.

Characteristics of Weak Polyprotic Acids

1 Typically, successive K_a values are so much smaller than the first value that only the first dissociation step makes a significant contribution to the equilibrium concentration of H^+. This means that the calculation of the pH for a solution of a typical weak polyprotic acid is identical to that for a solution of a weak monoprotic acid.

2 Sulfuric acid is unique in being a strong acid in its first dissociation step and a weak acid in its second step. For relatively concentrated solutions of sulfuric acid (1.0 M or higher), the large concentration of H^+ from the first dissociation step represses the second step. Which can be neglected as a contributor of H^+ ions. For dilute solutions of sulfuric acid, the second step does make a significant contribution, and the quadratic equation must be used to obtain the total H^+ concentration.

Section 14.8 Acid - Base Properties of Salts

Salt - another name for an ionic compound.

Salts that produce neutral solutions - salts that consist of the cations of strong bases and the anions of strong acids have no effect on $[H^+]$ when dissolved in water.

Salts that produce basic solutions - any salt whose cation has neutral properties and whose anion is the conjugate base of a weak acid, the aqueous solution will be basic. Ex:

$$\underset{\text{Conjugate base of the weak acid}}{C_2H_3O_2^-(aq)} + H_2O(l) \rightleftharpoons \underset{\text{Weak acid}}{HC_2H_3O_2(aq)} + OH^-(aq)$$

Salts that produce acidic solutions - salts in which the anion is not a base and the cation is the conjugate acid of a weak base produce acidic solutions. Ex:

$$\underset{\text{Conjugate acid of the weak base}}{NH_4^+(aq)} \rightleftharpoons \underset{\text{Weak base}}{NH_3(aq)} + H^+(aq)$$

Highly charged metal ions - salts which contain highly charged metal ions produce an acidic solution. The metal ion becomes hydrated which then causes the solution to become acidic. See p. 687 - 689 (and Table 14.6)

Section 14.9 The Effect of Structure on Acid-Base Properties

There are two main factors that determine whether a molecule containing an X - H bond will behave as a Bronsted-Lowry acid: the strength of the bond and the polarity of the bond. Increased polarity and high electron density typically lends to large K_a values (strong acids).

Section 14.10 Acid-Base Properties of Oxides

A compound containing the H - O - X group will produce an acidic solution in water if the O - X bond is strong and covalent. If the O - X bond is ionic, the compound will produce a basic solution in water.

Acidic Oxides - a covalent oxide dissolves in water, and an acidic solution forms.

Basic Oxides - an ionic oxide dissolves in water, and oxide has a great affinity for H^+ causing basic solutions.

Section 14.11 The Lewis Acid-Base Model

Lewis acid - is an electron-pair acceptor.

Lewis base - is an electron-pair donor.

The Lewis model encompasses the Bronsted-Lowry model, but the reverse is not true.

Section 14.12 Strategy for Solving Acid-Base Problems: A Summary

Read through pages 695 - 696

**Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.