Chapter 16

Section 16.1 You should know that acids and bases have distinctive physical properties, such as the slippery feel of bases and the sour taste of acids. Some of the acids present in our environment occur naturally. Others, including H_2SO_4 and HNO_3, are also the result of pollution released by our industrialized society and contribute to the formation of acid rain.

Section 16.2 You should know that the **Brønsted-Lowry model** of acids and bases defines acids as H^+ ion donors and bases as H^+ ion acceptors. Strong acids include binary (HX where X is a group 17 element other than fluorine) acids and oxoacids formed when nonmetal oxides combine with water. It is very important that you understand that all strong acids are completely ionized (close to 100%) in water. The H^+ ions that acids release combine with water molecules to form hydronium (H_3O^+) ions. Strong bases are those group1 and 2 hydroxides that are soluble in water and dissociate completely when they dissolve. Most acids, including the organic acids in biological samples, are weak acids that are only partially ionized in water. When weak acid "HA" undergoes ionization, it forms its **conjugate base,** A^-. When base "B" acquires a H^+ ion, it forms its **conjugate acid,** HB^+. The strongest acid in water is the H_3O^+ ions; the strongest base is the OH^- ion.

$$\text{Weak acid (only partially ionized)}: HA + H_2O \rightarrow H_3O^+ + A^-$$
$$\text{Weak base (only partially ionized)}: B + H_2O \Leftrightarrow BH^+ + OH^-$$

Section 16.3 You should know that water is an **amphoteric** substance in that it is capable of behaving both as an acid and as a base. This behavior is evident in the autoionization of water in which strong hydrogen bonding between water molecules results in formation of a hydronium ion and a hydroxide ion by one pair of water molecules out of every 10 million at 25ºC:

$$2\ H_2O(\ell) \leftrightarrow H_3O^+(aq) + OH^-(aq) \qquad K_w\ (25^{\circ}C) = 1.00E\text{-}14$$

In a neutral solution, $[H_3O^+ \text{ or } H^+] = [OH^-] = 1.00 \times 10^{-7}$. **pH** is a logarithmic scale for expressing the acidic or basic strength of solutions. Acidic solutions have pH values less than 7.00; basic solutions have pH values greater than 7.00. An increase in one pH unit represents a decrease in $[H^+]$ to $1/10^{th}$ of its initial value.

$$pH = -\log[H_3O^+ \text{ or } H^+]$$

Section 16.4 You should know that calculations of the pH of weak acids and bases employ ICE tables listing the initial concentrations of reactants and products, the changes in concentration that they undergo, and their concentrations at equilibrium. The **degree of ionization** of a weak acid (HA) is the ratio of $\{A^-]$ to the initial (total) acid concentration and is usually expressed as a percentage.

Section 16.5 You should know that weak **polyprotic acids** can undergo more than one acid-ionization reaction, but the first one is always the one that controls pH.

Section 16.6 You should know that the strength of an oxoacid is related to the stability of the anion that is formed as a result of losing a H^+ ion. This stability increases as the electronegativity of the central atom increases and as the number of oxygen atoms double bonded to the central atom increases.

Section 16.7 You should know that a salt is an acidic salt if its formula contains a cation that is the conjugate acid of a weak base and an anion that is the conjugate base of a strong acid. A salt is a basic salt if its formula includes an anion that is the conjugate base of a weak acid, and a cation that is the ion of a group1 or 2 element.

Section 16.8 You should know that adding a salt of a weak acid (HA) to a solution of the acid provides a second source of the conjugate base (A^-). As predicted by LeChâtelier's principle, part of the added A^- is consumed as the acid-ionization reaction

$$HA + H_2O \leftrightarrow H_3O^+ + A^-$$

runs in reverse to consume H_3O^+ and raise pH. Similarly, adding a salt of a weak base to a solution of the base provides a second source of its conjugate acid, which lowers pH. These shifts are examples of the **common-ion effect.**

Section 16.9 You should know that a **pH buffer** is a solution that contains either a weak acid and a salt of its conjugate base or a weak base and a salt of its conjugate acid. Please remember that buffer solutions resist pH change when strong acids or bases are added. Additions of acid are neutralized by the basic component of a buffer; additions of a base are neutralized by the acidic component. The result is a change in the concentration ratio in the **Henderson-Hasselbach** equation, but the changes in the logarithm of the ratio, and the impact on pH, tend to be small as long as neither component is completely consumed.

$$\text{Henderson - Hasselbach Equation}: \text{pH} = \text{pK}_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

Section 16.10 You should know that a **pH indicator** is a weak acid or base that has a color that is different from that of its conjugate base or acid. Indicators can be used to determine the pH (±pH units) of a solution that is within 1.0 pH unit of the pK_a or pK_b of the indicator. You should also be aware that these indicators are used to detect the equivalence points in pH titrations, which are highly precise methods for determining the concentration of a weak or strong acid or base, called the analyte, in an aqueous sample. A known volume of the sample is titrated with a known concentration of a solution, called the titrant, of strong base or acid. Titrant is added until all the analyte has been consumer – a point called the equivalence point in the titration.