Chapter 10 - **Monoprotic Acid-Base Equilibria**

Homework: Due Monday, March 6

Problems 10-2, 10-3, 10-5, 10-8, 10-11, 10-19, 10-26, 10-27, 10-28, 10-29, 10-32, 10-36

ACID – BASE CHEMISTRY

Protic Acids and Bases (Bronsted-Lowry)

- Acid => increases [H_3O^+] or [H^+]
- Base => increases [OH^-]
- Protic => transfer of H^+
- Bronsted acid => proton donor (like HCl)
- Bronsted base => proton acceptor (like NaOH)
- Salt => ionic solid made from reaction of acid and base (like $CaCl_2$ – R_x of $2HCl + Ca(OH)_2 = CaCl_2 + 2HOH$)
- Conjugate => acid-base pair (like acetic acid $H_3C\text{-}C(=O)\text{-}OH$ and acetate ion $H_3C\text{-}C(=O)\text{-}O^-$)

Equilibrium of Water

$H_2O = H^+ + OH^-$ $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ (25°C)

- K_w allows you to calculate $[H^+]$ and $[OH^-]$
- $[H^+]$ and $[OH^-]$ describes the acid/base content of solution
- Since K_w is small, $[H^+][OH^-]$ will also be small.
- pH ($-\log[H^+]$) or pOH ($-\log[OH^-]$) is a useful way to report $[H^+]$ or $[OH^-]$
- $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ So….

$$pK_w = pH + pOH = 14$$

Acid/Base Strength

Strong acids and bases **COMPLETELY** dissociates

$$HCl \rightarrow H^+ + Cl^-$$

(No reverse arrow as no HCl present in solution)

TABLE 6-2: Strong acids and Strong bases

Be able to recognize them (memorize)

Examples: Acids – HCl, HNO_3, H_2SO_4

Bases – NaOH, KOH, NH_4OH

Determining pH and pOH of strong acid/base

Example:

0.1 M HCl means $[H^+] = 0.1$ or 1×10^{-1}

$pH = -\log(1 \times 10^{-1}) = 1$

So…

$[H^+]$	$[H^+]$	pH
0.01	1×10^{-2}	2
0.001	1×10^{-3}	3
0.0001	1×10^{-4}	4

0.035 M HCl $[H^+] = 3.5 \times 10^{-2}$ M $pH = -\log(3.5 \times 10^{-2}) = 1.46$

Given pH 5.67, what is the $[H^+]$?

$[H^+] =$ antilog $(-5.67) \Rightarrow 10^{-pH} = 10^{-5.67} = 2.14 \times 10^{-6}$ M

Example: What is the pH of a 0.0045 M NaOH solution?

Example: What is the $[OH^-]$ of a solution at pH 4.3?

If concentration of strong acid or strong base is very small (< 10^{-6} M), then you cannot simply use the concentration of the strong acid or strong base to determine pH or pOH. The reason is that at this low concentration, the dissociation of H_2O will also contribute to the $[H^+]$ and $[OH^-]$. ($H_20 \Leftrightarrow H^+ + OH^-$)

In the case of extremely small strong acid or strong base concentrations (< 10^{-8}), then the main contribution of $[H^+]$ or $[OH^-]$ is from the dissociation of H_2O so the solution has a pH 7 and pOH 7.

See pg. 180 and 181 of Harris

Weak acids and bases:
PARTIAL dissociation

Acid: $HA \Leftrightarrow H^+ + A^-$ $\quad K_a = [H^+][A^-] / [HA]$

acid dissociation constant

Base: $B + H_2O \Leftrightarrow BH^+ + OH^-$ $\quad K_b = [BH^+] [OH^-] / [B]$

base hydrolysis constant

Know the conjugate pairs:

$$H_3C - CH_2 - \overset{\overset{O}{\|}}{C} - OH \quad = \quad H_3C - CH_2 - \overset{\overset{O}{\|}}{C} - O^- \quad + \quad H^+$$

acid — base (acid – H^-)

$$H_3C - N^+H_3 \quad = \quad H_3C\text{-}NH_2 \quad + \quad H^+$$

acid (base + H^+) — base

NOTE:

- The conjugate base of an acid has a negative charge (loss of H^+).
- The conjugate acid of a base has a positive charge (gain of H^+).

Relationship of K_a and K_b

The relationship between K_a and K_b for a conjugate pair is:

$$K_a \times K_b = K_w = 1 \times 10^{-14}$$

Also $$pK_w = pK_a + pK_b$$

$$H_2A = H^+ + HA^- \qquad K_a$$

(acid) ***(conjugate base)***

$$HA^- + H_2O = H_2A + OH^- \qquad K_b$$

(conjugate base) ***(acid)***

multiply the K's for the added R_x's

$$H_2O = H^+ + OH^- \qquad K_w$$