CHEM 20B Week 8

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Acid-Base Equilibria

Arrhenius Definition of Acids and Bases

- Acids: when dissolved in water increases the concentration of H_3O^+
- Bases: when dissolved in water increases the concentration of OH^-

Brønsted-Lowry Definition of Acids and Bases

- Acids: Proton donor
- Bases: Proton acceptor

Lewis Definition of Acids and Bases

- Electron pair acceptor
- Electron pair donor

Properties of Acids and Bases in Aqueous Solutions: The Brønsted-Lowry Scheme

Autoionization of Water

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

K_w, the acid-base constant of water is determined by

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

This constant varies by temperature.

Temperature Dependence of K_w

Temperature ($^\circ$ C)	K_w	**pH of Water**
0	0.114×10^{-14}	7.47
10	0.292×10^{-14}	7.27
20	0.681×10^{-14}	7.08
25	1.01×10^{-14}	7.00
30	1.47×10^{-14}	6.92
40	2.92×10^{-14}	6.77
50	5.47×10^{-14}	6.63
60	9.61×10^{-14}	6.51

Strong Acids

Strong Acids ionize completely in aqueous solution.

$$H_2O(l) + HA(aq) \rightarrow H_3O^+(aq) + A^-(aq)$$

Common Strong Acids

- HBr (aq)
- HCl (aq)
- HI (aq)
- HNO_3 (aq)
- $HClO_4$ (aq)
- $HClO_3$ (aq)
- H_2SO_4 (aq)

Strong Bases

Strong Bases react completely to give OH^- when put in water.

$$H_2O(l) + NH_2^-(aq) \rightarrow OH^-(aq) + H_2(aq)$$

$$H_2O(l) + H^-(aq) \rightarrow OH^-(aq) + H_2(aq)$$

Common Strong Bases

- Group 1 hydroxides
- Alkaline earth metal hydroxides
- Group 1 and Group 2 oxides

The pH Function

$$pH = -\log_{10}[H_3O^+]$$

At 25°C,

- pH $<$ 7 $=$ Acidic solution
- pH $=$ 7 $=$ Neutral solution
- pH $>$ 7 $=$ Basic solution

Similar to how $[H_3O^+][OH^-] = K_w$, $pH + pOH = pK_w$.
As a consequence,

- $[H^+] = 10^{-pH}$
- $[OH^-] = 10^{-pOH}$

Acid and Base Strength

Acid strength is based on the extent to which they are ionized in solution.

$$HA(aq) \leftrightharpoons H^+(aq) + A^-(aq)$$

The Acid Ionization Constant, K_a is a quantitative measure of the strength of the acid.

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

If:

- $K_a \gg 1 \rightarrow$ HA is a strong acid.
- $K_a \ll 1 \rightarrow$ HA is a weak acid.

Convenient characterization of strength of acid is pK_a.

$$pK_a = -\log_{10}(K_a)$$

For Example:
$K_a = 10^7 \rightarrow pK_a = -7$ (strong acid)
$K_a = 10^-5 \rightarrow pK_a = 5$ (weak acid)

Similarly to acids, base strength is represented by K_b, which is inversely related to the strength of its conjugate acid.

$$H_2O(l) + B(aq) \leftrightharpoons OH^-(aq) + BH^+(aq)$$

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

Similarly, a convenient characterization of strength of base is pK_b.

$$pK_b = -\log_{10}(K_b)$$

If:

- $K_b \gg 1 \rightarrow$ strong base, many OH^- produced, little $[B]$ left.
- $K_b \ll 1 \rightarrow$ weak base, most $[B]$ remains.

Importantly,

$$K_b K_a = K_w$$

$$pK_b + pK_a = pK_w$$

Equilibrium Calculations for Weak Acids and Bases

- Calculate K_a or K_b from initial concentration(s) and measured pH
- Calculate equilibrium concentrations and pH from initial concentrations and known K_a or K_b

Do the following steps:

1. write the ionization equation.
2. write the equilibrium expression.
3. construct an ICE table
4. solve for unknown (may or may not be quadratic)
    - if quadratic, choose the value that does not result in negative concentrations.

Buffer Solutions

- Resist changes in pH from the addition of strong acid (H^+) or base (OH^-).
- Can be a weak acid and a salt of its conjugate base.
- Example:
  - H_2CO_3 / HCO_3^- (in Blood Plasma)