

# CHEM 20B Week 8

Aidan Jan

March 1, 2024

## Acid-Base Equilibria

### Arrhenius Definition of Acids and Bases

- Acids: when dissolved in water increases the concentration of  $\text{H}_3\text{O}^+$
- Bases: when dissolved in water increases the concentration of  $\text{OH}^-$

### Brønsted-Lowry Definition of Acids and Bases

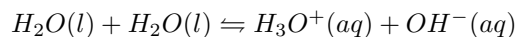
- Acids: Proton donor
- Bases: Proton acceptor

### 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



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

$$K_w[\text{H}_3\text{O}^+][\text{OH}^-]$$

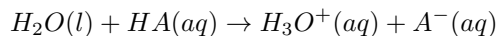
This constant varies by temperature.

Temperature Dependence of  $K_w$

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

## Strong Acids

Strong Acids ionize completely in aqueous solution.

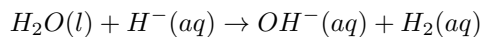
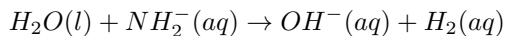


### Common Strong Acids

- HBr (aq)
- HCl (aq)
- HI (aq)
- HNO<sub>3</sub> (aq)
- HClO<sub>4</sub> (aq)
- HClO<sub>3</sub> (aq)
- H<sub>2</sub>SO<sub>4</sub> (aq)

## Strong Bases

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



### 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

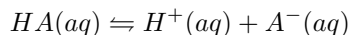
Similar to how  $[H_3O^+][OH^-] = K_w$ ,  $pH + pOH = pK_w$ .

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.



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  $pK_a$ .

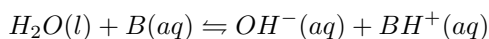
$$pK_a = -\log_{10}(K_a)$$

**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.



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

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 ( $\text{H}^+$ ) or base ( $\text{OH}^-$ ).
- Can be a weak acid and a salt of its conjugate base.
- Example:
  - $\text{H}_2\text{CO}_3$  /  $\text{HCO}_3^-$  (in Blood Plasma)