Acids and Bases

1. Acid-Base Theories

Arrhenius: Acids – Dissociates completely to produce H^+
Bases – Dissociates completely to produce OH^-

Bronsted-Lowry: Acids – H^+ (proton) donors
Bases – H^+ (proton) acceptors

$HA + H_2O \leftrightarrow H_3O^+ + A^-$
A B CA CB (CA = Conjugate Acid, CB = Conjugate Base)

$B + H_2O \leftrightarrow BH^+ + OH^-$
B A CA CB

- **Example:**
  Write a balanced chemical equation for the reaction between N_2H_4 and H_2O that explains why a solution of N_2H_4 in water has a pH greater than 7.

$$N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^-$$

H^+

2. Strong and Weak Acids/Bases

Strong Acids (H^+) - HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4
Strong Bases (OH^-) – LiOH, NaOH, KOH, RbOH, CsOH, $Ba(OH)_2$, $Sr(OH)_2$

Strong Acids and Bases are strong electrolytes, conduct electricity, and dissociate completely. $HClO_4$ is stronger than $HClO_3$ which is stronger than $HClO_2$ due to the fact that the increasing number of oxygen atoms that are attached to the central atom weakens the attraction that the central atom has for the H^+ ion. Therefore, more oxygen atoms attached to a central atom assists in greater dissociation of H^+.

Salts made from the conjugate of strong acids and strong bases are neutral salts when dissolved in water. Therefore, NaCl would be neutral.

Weak Acids - All others including HF and $HC_2H_3O_2$ (can abbreviate them as HA)
Weak Bases - All others including NH_3 (can abbreviate them as B)

Weak Acids and Bases are weak electrolytes and do not conduct electricity well and dissociate less than approximately 5%.
Salts made from conjugates of weak acids and weak bases are basic and acidic when dissolved in water. Therefore, NaF would be a base (contains the conjugate base F^-) and NH_4Cl would be an acid (contains the conjugate acid NH_4^+).

- **Example:**
  Determine whether the pH of the following salts above, below, or equal to 7:
  (a) $NaC_2H_3O_2$
  pH > 7 (contains the conjugate base $C_2H_3O_2^-$)
  (b) KBr
  pH = 7 (contains conjugates of strong acids and strong bases)
  (c) NH_4Br
  pH < 7 (contains the conjugate acid NH_4^+)

3. pH Calculations

pH + pOH = 14
$[H^+]\,[OH^-] = 1 \times 10^{-14}$
pH = - log $[H^+]$ $\quad\quad$ $[H^+] = 10^{\wedge[-pH]}$
pOH = - log $[OH^-]$

- **Example:**
  Calculate the pH of the following solutions.
  (a) 0.0015 *M* HNO_3

$$pH = -\log\,[H^+]$$
$$pH = -\log(0.0015) = 2.82$$

  (b) 2.50 g of NaOH is dissolved in 500 mL of water.

$$moles = \frac{mass}{molar\ mass} = \frac{2.50\ g}{40\ g\ mol^{-1}} = 0.0625\ moles\ OH^-$$
$$Molarity = \frac{moles}{Liter} = \frac{0.0625\ moles\ OH^-}{0.500\ L} = 0.125\ M\ OH^-$$
$$pOH = -\log\,[OH^-] = -\log(0.125) = 0.90$$
$$pH = 14 - pOH = 14 - 0.90 = 13.10$$

4. Dilutions or Titrations

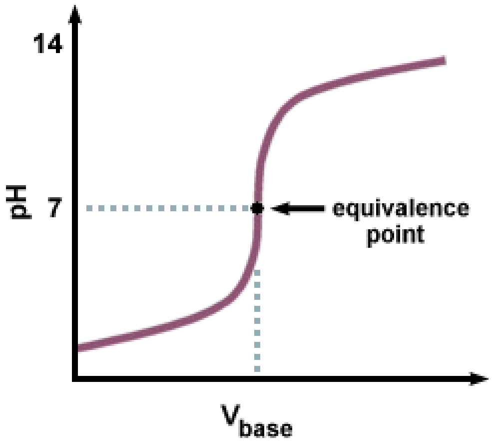
$M_1V_1 = M_2V_2$
Acid $\quad$ Base
Equivalence Point - when the moles of acid = moles of base
M * V (in L) = $\quad$ moles H^+
M * V (in L) = $\quad$ moles OH^-

$$\frac{\text{Subtract to find moles}}{\text{Total new volume in L}} = X \quad pH = -\log\,[X]$$

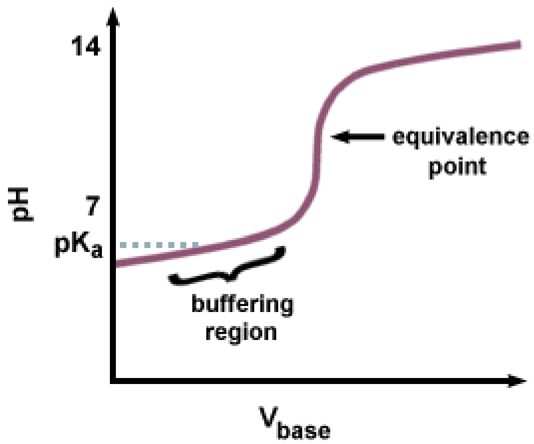
(Note: If adding two strong acids, you will add the moles together and divide by total new volume.

5. Titration Curves

Strong Acid Titrated with a Strong Base (Equivalence Point at pH = 7.00)



Weak Acid Titrated with a Strong Base (Equivalence Point above pH = 7.00)



Weak Base Titrated with a Strong Acid (Equivalence Point below pH = 7.00)