Chapter Summary: Acid Base Equilibria

Acid ionization constant

CH_3COOH + H_2O $\rightleftharpoons$ H_3O^+ + CH_3COO^-

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

$pK_a = -\log K_a$ $pK_b = -\log K_b$ $pK_w = pK_a + pK_b$

lower value of pK_a = strong acid; higher value of pK_a = weak acid

Polyprotic acids: more than one ionizable proton e.g. H_2SO_4 is diprotic and H_3PO_4 is triprotic. In a solution there will be more than just the one conjugate acid/base pair.

Acidic and Basic Ions

E.g. Na_2CO_3 $\longrightarrow$ Na^+ + CO_3^{2-}

The carbonate ion can react with water molecules as follows:

CO_3^{2-} + H_2O $\longrightarrow$ HCO_3^- + OH^-

Na^+ do not react with water. Therefore the carbonate ion undergoes hydrolysis reaction producing hydroxide ions – resulting in a basic ion.

Salts of strong acid + strong base give neutral solutions e.g. NaCl, KNO_3
Salts of weak acid + strong base give basic solutions e.g. Na_2CO_3
Salts of strong acid + weak bases give acidic solutions e.g. NH_4Cl

Common Ion Effect: The common ion in an equilibrium can shift the equilibrium to the opposite side. E.g. adding sodium acetate CH_3COONa to will increase the concentration of acetate ion thus moving the equilibrium to the left.

CH_3COOH + H_2O $\rightleftharpoons$ H_3O^+ + CH_3COO^-

Buffer Solutions: a solution to which addition of a little acid or base changes the pH only slightly.

Henderson-Hasselbalch Equation: used to calculate new pH after addition of acid or base.

$$pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{weak acid}]}$$

Acid Base Indicators: chemicals that change color with pH.

Titration Curves: a neutralization graph to calculate end point of a titration.