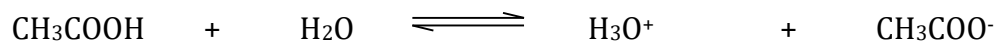


Chapter Summary: Acid Base Equilibrium

Acid ionization constant



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_b = -\log K_b$$

$$\text{p}K_w = \text{p}K_a + \text{p}K_b$$

lower value of $\text{p}K_a$ = strong acid; higher value of $\text{p}K_a$ = weak acid

Polyprotic acids: Acids that have 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 one conjugate acid/base pair.

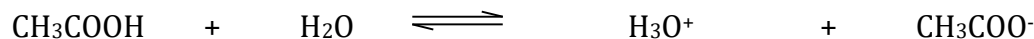
Acidic and Basic Ions

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 will increase the concentration of acetate ion thus moving the equilibrium to the left.



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.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

Acid Base Indicators: Chemicals that change color with pH.

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

Key Words

Equilibrium constant of acids and bases	Acidic and basic salts	Common ion effect
Buffers	Hendersen-Hasselbalch equation	Indicators
Titrations	End point	Equivalent point