COMMON ION EFFECT IN THE IONIZATION OF ACIDS AND BASES

Factors Affecting Acidic Strength

We understand that a weak electrolyte undergoes partial dissociation in a solution, establishing an equilibrium between the undissociated electrolyte and the ions formed in the solution. When a strong electrolyte, sharing a common ion with the weak electrolyte, is introduced, the dissociation of the weak electrolyte is further restrained or diminished. This phenomenon is termed the common ion effect.

For instance, consider the ionization of acetic acid (CH_3COOH) and the impact of adding a small amount of acetate ion (CH_3COO^-):

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

$$CH_3COONa(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq)$$

Due to the presence of common CH_3COO^- ions, according to Le Chatelier's principle, the equilibrium is shifted to the left, suppressing the dissociation of acetic acid.

To determine the pH of the solution resulting from the addition of 0.05 M acetate ion to a 0.05 M acetic acid solution, let's reconsider the acetic acid dissociation equilibrium:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

The initial concentrations are 0.05 M for acetic acid, 0.05 M for acetate ion, and 0 for hydrogen ions. The equilibrium concentrations are $0.05 - x$ for acetic acid, $0.05 + x$ for acetate ion, and x for hydrogen ions (where a is the extent of ionization of acetic acid).

The equilibrium constant Ka is given by:

$$Ka = \frac{(CH_3COO^-)(H^+)}{CH_3COOH}$$

Since the dissociation of the acid is now influenced by the non-common ion, we can express hydrogen ion concentration as:

$$[H^+] = \frac{Ka(CH_3COOH)}{CH_3CO\ \ ^-}$$

After adding acetate ion, the concentration of $[CH_3COO^-]$ becomes $0.05 + x$.

For a weak electrolyte, $x \ll 0.05$, simplifying to $0.05 + x \approx 0.05x = 0.05$.

Therefore,

$$[H^+] = \frac{Ka \times 0.05}{0.05}$$

$$[H^+] = Ka = 1.8 \times 10^{-5}\ M$$

The pH can be determined as:

$$[H^+] = -\log(1.8 \times 10^{-5}) = 4.74$$

Relationship between pK and pK_b

The term pK_a represents the ionization constant of a weak acid, while pK_b corresponds to the ionization constant of its conjugate base. The equilibrium for the dissociation of a weak acid (HA) can be expressed as:

$$HA \rightleftharpoons H^+ + A^-$$

The ionization constant Ka is given by the expression:

$$K_a = [H^+][A^-]/[HA] \qquad (i)$$

Additionally, considering the ionization of the conjugate base (A^-) with water:

$$A^- + H_2O \rightleftharpoons OH^- + HA$$

The ionization constant Kb is expressed as:

$$K_b = [OH^-][HA]/[A-] \qquad (ii)$$

By multiplying equations (i) and (ii), the relation $K_a \times K_b = [H^+][OH^-] = K_w$ is obtained, indicating the equilibrium constant for the dissociation of water.

This can be further expressed as: $\log K_a + \log K_b = \log K_w$

Therefore, combining $pK_a = -\log K_a$ and $pK_b = -\log K_b$,

The relationship $pK_a + pK_b = -\log K_w$ is derived.

This equation signifies the relationship between the ionization constants of a weak acid and its conjugate base in terms of the ion product constant for water (K_w).