FIRST STAGE II

Volumetric Analysis

The Third Lecture

Lecture Objective

- Weak Acids & weak Bases
- Ionisation constant of weak Acids & weak Bases
- Weak acids pH equation
- Weak bases pH equation
- The relationship between ph, pOH, K_a, and K_b
- Examples

Weak Acids & weak Bases

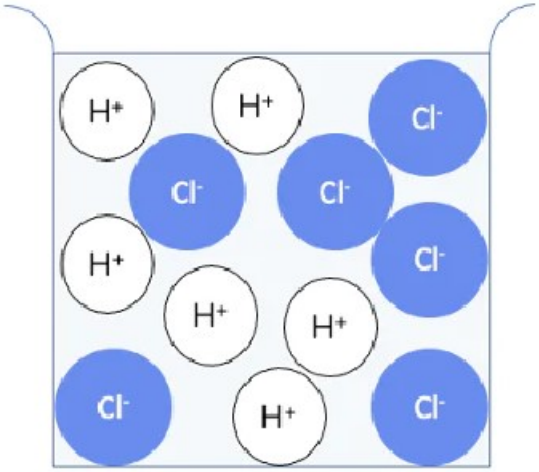
Weak Acids & weak Bases are **partially** ionized such as of thire:

Weak acids		Weak bases	
Acetic acid	CH_3COOH	Ammonia	NH_3
Carbonic acid	H_2CO_3	Diethylamine	$(CH_3CH_2)_2NH$
Formic acid	CHOOH	Methylamine	CH_3NH_2
Hydrocyanic acid	HCN	Sodium bicarbonate	$NaHCO_3$
Hydrofluoric acid	HF		
Phosphoric acid	H_3PO_4		

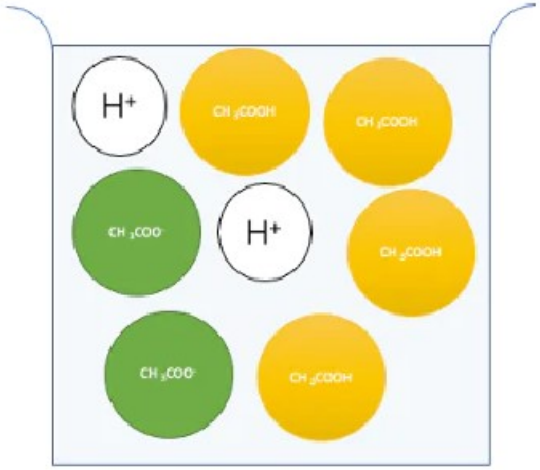
Weak Acids & weak Bases

By contrast from strong acid & bases, weak acids and bases ionize only partially, and the ionization reaction is reversible. Thus, weak acid and base solutions contain multiple charged and uncharged species in dynamic equilibrium

Hydrochloric acid fully dissociates in water and so is a strong acid

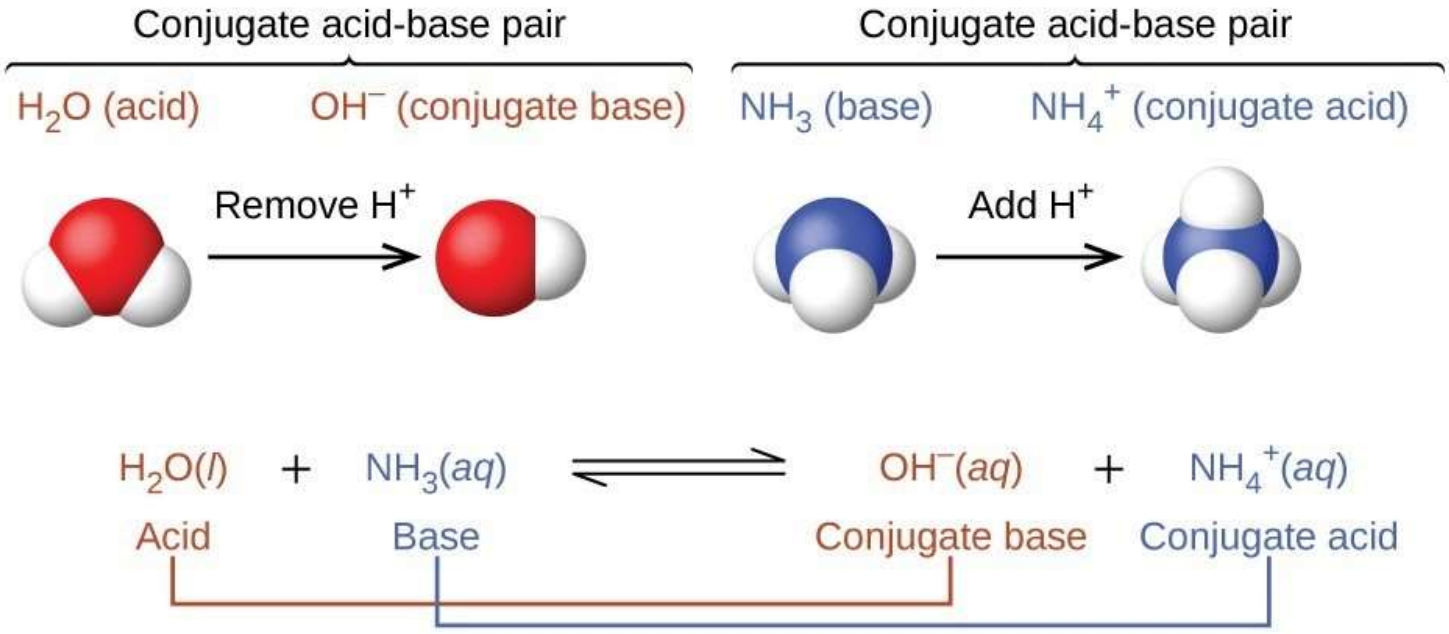


Acetic acid only partially dissociates in water and so is a weak acid



Weak Acids & weak Bases

In this connection, you probably realize that conjugate acids of weak bases are weak acids and conjugate bases of weak acids are weak bases.



K_a and K_b

K_a is an acid dissociation constant (also known as acidity constant, or acid-ionization constant.

K_a is a quantitative measure of the strength of an acid in solution. It is the equilibrium constant for a chemical reaction.

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+].[A^-]}{[AH]}$$

very small of water is ionised, **H_2O is constant**

K_a and K_b

K_b is an base dissociation constant (also known as basicity constant, or base-ionization constant.

K_b is a quantitative measure of the strength of an base in solution. It is the equilibrium constant for a chemical reaction.

$$B + H_2O \rightleftharpoons HB^+ + OH^-$$

$$K_b = \frac{[HB^+].[OH^-]}{[B]}$$ very small of water is ionised, **H_2O is constant**

Weak acids , pH equation

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

$$K_a = \frac{[H^+].[CH_3COO^-]}{[CH_3COOH]} \quad \text{.....(1)}$$

From above equation, $[H^+] = [CH_3COO^-]$

$$K_a = \frac{[H^+]^2}{[CH_3COOH]} \quad \text{.....(2)}$$

And $[CH_3COOH] = Ca$

$$K_a = \frac{[H^+]^2}{[C_a]} \quad \text{.....(3)}$$

Weak acids , pH equation

$$K_a = \frac{[H^+]^2}{[C_a]} \text{ .....(3)}$$

$$[H^+]^2 = K_a .[C_a] \text{ .....(4)}$$

$$[H^+] = K_a .[C_a]^{1/2} \text{ .....(5)}$$

Taking loge to both sides:

$$\mp \log [H^+] = \mp \frac{1}{2} \log K_a \mp \frac{1}{2} \log [C_a]$$

$$pH = ½ \ pK_a – ½ \log C_a$$

Weak base , pOH equation

$$NH_4OH + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+].[HO^-]}{[NH_4OH]} \text{.....(1)}$$

From above equation, $[OH^-] = [NH_4^+]$

$$K_b = \frac{[OH^-]^2}{[NH_4OH]} \text{.....(2)}$$

And $[[NH_4OH]] = C_b$

$$K_b = \frac{[OH^-]^2}{[C_b]} \text{.....(3)}$$

Weak base , pOH equation

$$K_b = \frac{[OH^-]^2}{[C_b]} \quad \text{.....(3)}$$

$$[OH^-]^2 = K_b \,.[C_b] \quad \text{.....(4)}$$

$$[OH^-] = K_b \,.[C_b]^{1/2} \quad \text{.....(5)}$$

Taking loge to both sides:

$$\mp \log [OH^-] = \mp \frac{1}{2} \log K_b \mp \frac{1}{2} \log [C_b]$$

$$pOH = ½ \; pK_b - ½ \log C_b$$

The Relationship pH, pOH, K_a, K_b and K_w

$$\cancel{NH_4^+} + H_2O \rightleftharpoons \cancel{NH_3} + H_3O^+ \qquad K_a = 5.6 \times 10^{-10}$$

$$\cancel{NH_3} + H_2O \rightleftharpoons \cancel{NH_4^+} + OH^- \qquad K_b = 1.8 \times 10^{-5}$$

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \qquad K_w = 1.01 \times 10^{-14}$$

$$K_a \times K_b = K_w$$

$$(5.6 \times 10^{-10}).(1.8 \times 10^{-5}) = 1.01 \times 10^{-14}$$

$$K_a = \frac{[\cancel{NH_3}].[H_3O^+]}{[\cancel{NH_4^+}]} \times K_b = \frac{[\cancel{NH_4^+}].[HO^-]}{[\cancel{NH_3}]} = K_W$$

The Relationship pH, pOH, K_a, K_b and K_w

$$H_3O^+ \times OH^- = K_w = 1.01 \times 10^{-14}$$ log to both side

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$$\log([H_3O^+] \times [OH^-]) = \log K_w =$$

$$\text{Log}[H_3O^+] \pm \log [OH^-] = \log 1.01 \times 10^{-14}$$

$$pH + pOH = 14$$

Examples

Ex. Calculate the pH of 0.01M CH_3COOH, where it's the ionization constant of acid = 1.85×10^{-5} at 25℃

Solution

$$\underset{0.01}{CH_3COOH} \rightleftharpoons \underset{x}{H^+} + \underset{x}{CH_3COO^-}$$

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

$$K_a = \frac{[H^+]^2}{[C_a]}$$

$$[H^+] = K_a .[C_a]^{1/2}$$

$$\mp \log [H^+] = \mp \frac{1}{2} \log K_a \mp \frac{1}{2} \log [C_a]$$

pH = ½ pK_a – ½ logC_a

pH = ½ 4.74 – ½ log10^{-2}

pH = 2.37+ 1 **pH= 3.37**

pKa = -log Ka
= - log 1.85×10^{-5}
= -(5+0.026)
= -(-4.74) === 4.74

Examples

Ex. Calculate the pH of 0.01M NH_4OH, where it's the ionization constant of base = 1.85×10^{-5} at 25℃

Solution

$$NH_4OH + H_2O \rightleftharpoons NH_4^+ + OH^-$$
$$0.01 \qquad\qquad x \qquad\qquad x$$

$$K_b = \frac{[NH_4^+].[HO^-]}{[NH_4OH]}$$

$$K_b = \frac{[OH^-]^2}{[C_b]}$$

$$[OH^-] = K_b .[C_b]^{1/2}$$

$$\mp \log [OH^-] = \mp \frac{1}{2} \log K_b \mp \frac{1}{2} \log [C_b]$$

pOH = ½ pK_b – ½ logC_b

pOH = ½ 4.74 – ½ log10^{-2}

pOH = 2.37+ 1

pOH= 3.37 pH = 14 – 3.37 = 10.63

pK_b = -log K_b
= - log 1.85×10^{-5}
= -(5+0.026)
= -(-4.74) === 4.74

Examples

Ex. K_b methylamine 3.7×10^{-4} calculate the K_a value of methylammonium ion ?

Solution

$$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$$

$$CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$$

$$K_a \times K_b = K_w$$

$$K_a \times 3.7\times 10^{-4} = 1.01\times 10^{-14}$$

$$K_a = 2.7\times 10^{-11}$$

$K_a \times K_b = K_w$ x (log)

$\log (K_a \times K_b) = \log K_w$

$-\log K_a \mp \log K_b = -\log K_w$

$-\log 2.7\times 10^{-11} \mp \log K_b = -\log K_w$

$10.57 + pK_b = 14$

$pK_b = 3.43$

pKa = -log Ka

= - log 2.7×10^{-11}

= -(-11+0.43)

= -(-10.57) === 10.57

Third Lecture

REVIEW

- How to weak acids and bases un ionization?
- Equilibrium reaction of weak Acid & base
- Conjugate acids bases?
- K_a and K_b
- Derive pH equation
- pH calculation of weak acids and bases equation
- The relationship between pH, pOH, K_a, and K_b
- Examples

Home Work

1. **Calculate the pH and pOH of the following acidic solutions?**
   (a) 0.02 M HCN
   (b) 0.8 g of HCN dissolve to 100 ml water $K_{a\ HCN} = 6.2 \times 10^{-10}$

2. **Calculate the pH and pOH of the following basic solutions?**
   (a) 0.075 M NH_3
   (b) 1 mL of 0.1M of NH_3 add to 100 ml water $K_{b\ NH_3} = 1.8 \times 10^{-5}$