Chemistry 3.6
Aqueous Systems

Acids and Bases

From Level 2

Water dissociation

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

$$K_w = [H_3O^+][OH^-] = 10^{-14}$$

"p" = -log

$$pH = -\log [H_3O^+]$$

$$10^{-pH} = [H_3O^+]$$

$$pH + pOH = 14$$

Dissociating

- Once a solid dissolve, each species can further under goes **dissociation** with **water**.
- If it is an **acid** species, it will **donates** a proton (H^+) forming a conjugate base and **H_3O^+ ions**
- If it is a **base** species, it will **accepts** a proton (H^+) forming conjugate acid and **OH^- ions**

Strong acid and weak acid

- Strong acid fully dissociate when it is dissolved

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

- Weak acid **first dissolve then partially dissociate**

$$NH_4Cl_{(s)} \rightarrow NH_4^+{}_{(aq)} + Cl^-$$

$$H_2O_{(l)} + NH_4^+{}_{(aq)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$$

Strong acid Calculation

- Because strong acid fully dissociate, the concentration of H_3O^+ (H^+) ions are the same as displayed in the formula of acid

Example

$[H^+]$ in 0.1 $molL^{-1}$ of HCl = 0.1 $molL^{-1}$

$[H^+]$ in 0.1 $molL^{-1}$ of H_2SO_4 = 0.2 $molL^{-1}$

$pH = -\log[H^+]$

Weak acid calculation

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

- K_a is usually given

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

- The assumption is that most of the acid molecules do not dissociate, therefore [HA] is very similar to the original concentration cHA
- Another assumption is $[H_3O^+] = [A^-]$. Rearranging the formula gives:

$$[H_3O^+] = \sqrt{K_a \times cHA}$$

Example

Calculate the pH of 0.05 molL^{-1} of ethanoic acid solution. $K_a = 1.74 \times 10^{-5}$

$CH_3COOH(aq) + H_2O(l) \qquad CH_3COO^- + H_3O^+$

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.74 \times 10^{-5}$$

$$[H_3O^+] = \sqrt{(1.74 \times 10^{-5}) \times 0.05}$$

$$[H_3O^+] = 9.33 \times 10^{-4}$$

$$pH = -\log (9.33 \times 10^{-4}) = 3.03$$

Strong base and weak base

- Strong base fully dissociate when it dissolve.

$$NaOH_{(s)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

- Weak base first dissolve then dissociate partially.

$$NH_{3(g)} \rightarrow NH_{3(aq)}$$

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^{+}{}_{(aq)} + OH^{-}_{(aq)}$$

Strong base calculation

- Calculation the concentration of OH^-
- $10^{-14} \div [OH^-] = [H_3O^+]$
- $-\log [H_3O^+] = pH$

Weak base calculation

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

- K_b is given in terms of K_a because…

$$K_b = \frac{K_w}{K_a} = \frac{[HA][OH^-]}{[A^-]}$$

- Same assumption with weak acid that the concentration of weak base $[A^-] \approx cA^-$
- $[OH^-] = [HA]$ therefore

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

$$[OH^-] = \sqrt{K_b \times cA^-}$$

Example

Calculate the pH of 0.05 molL^{-1} of ethanoic acid solution. $K_a = 1.74 \times 10^{-5}$

$CH_3COOH(aq) + H_2O(l) \qquad CH_3COO^- + H_3O^+$

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.74 \times 10^{-5}$$

$$[H_3O^+] = \sqrt{(1.74 \times 10^{-5}) \times 0.05}$$

$$[H_3O^+] = 9.33 \times 10^{-4}$$

$$pH = -\log(9.33 \times 10^{-4}) = 3.03$$