<u>Acid & Base Equilibrium Constants</u>

- weak acids and weak bases can be represented as **equilibrium** systems because they do not completely ionization
  - > the **acid ionization** reaction of a **weak acid** (such as CH_3COOH) with water is shown by:

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

- An equilibrium expression can be written for the ionization as:

$$Ka = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

- the value of Ka is the **<u>acid ionization constan</u>**t
- the <u>greater the value of Ka, the stronger the acid</u>
- larger Ka value indicates the equilibrium favours the products; they are therefore stronger because of greater ionization
- Ka values for STRONG acids are not listed since these acids are 100% ionized and the concentration of the denominator is therefore ZERO

- the base ionization reaction of a weak base such as NH_3 with water is shown by:

$$NH_3(aq) + H_2O(\ell) \rightleftarrows NH_4^+(aq) + OH^-(aq)$$

- the equilibrium expression for the ionization is:

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

- the value of Kb is called the **base ionization constant**
- the greater the Kb value, the stronger the base
- Kb values must be calculated using the Ka values of the conjugate acids

- there is an important relationship that exists between Ka and Kb for conjugate acid-base pairs; consider the following
  - > the acid NH_4^+ has the <u>acid ionization equation:</u>

$$NH_4^+ + H_2O \rightleftarrows NH_3 + H_3O^+$$

  - > the acid ionization expression:

$$Ka = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

  - > the conjugate base NH_3 has the <u>base ionization equation:</u>

$$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$$

  - > and the base ionization expression

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

  - > when Ka and Kb are multiplied together we get:

$$Ka \times Kb = \frac{[NH_3][H_3O^+]}{[NH_4^+]} \times \frac{[NH_4^+][OH^-]}{[NH_3]} = [H_3O^+][OH^-]$$

  - > and since $[H_3O^+][OH^-] = Kw$

$Ka \times Kb = Kw$
(at 25°C $Kw = 1.00 \times 10^{-14}$)

Calculate the Kb for $C_2O_4^{2-}$.

◦need Ka for conj. acid

$$Kb = \frac{Kw}{Ka} = \frac{1\times10^{-14}}{6.4\times10^{-5}} = 1.6\times10^{-10}$$

*Note that all Brønsted-Lowry reactions in Chem 12 will only involve transfer of a single proton

- if solutions containing amphiprotic ions are mixed, the stronger of the two acids will donate a proton to the other
- ex. HCO_3^- (Ka = 5.6 x 10^{-11}) & $H_2PO_4^-$ (Ka = 6.2 x 10^{-8}), since $H_2PO_4^-$ has a larger Ka, it will donate a proton to HCO_3^-

$$HCO_3^- + H_2PO_4^- \rightleftarrows H_2CO_3 + HPO_4^{2-}$$

- the position of an equilibrium depends on the strengths of the acids
- ex. H_2CO_3 (Ka = 4.3 x 10^{-7}) is a stronger acid than HSO_3^- (Ka = 1.0 x 10^{-7})

  > H_2CO_3 has a greater tendency to donate protons; there will be more products than reactants (products are favoured)

$$H_2CO_3 + SO_3^{2-} \rightleftarrows HCO_3^- + HSO_3^-$$

Weaker acids and bases are favoured in an equilibrium.