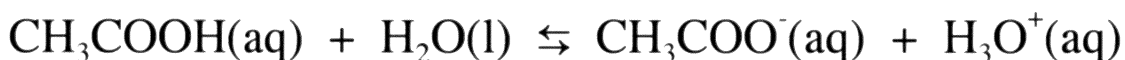


## Acid & Base Equilibrium Constants

- weak acids and weak bases can be represented as **equilibrium** systems because they do not completely ionize

> the **acid ionization** reaction of a **weak acid** (such as CH<sub>3</sub>COOH) with water is shown by:

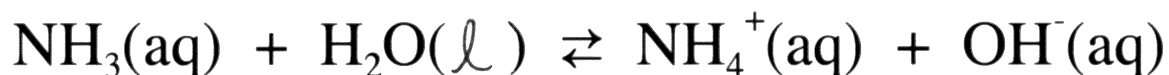


- An equilibrium expression can be written for the ionization as:

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

- the value of  $K_a$  is the **acid ionization constant**
- the **greater the value of  $K_a$ , the stronger the acid**
- larger  $K_a$  value indicates the equilibrium favours the products; they are therefore stronger because of greater ionization
- $K_a$  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  $\text{NH}_3$  with water is shown by:



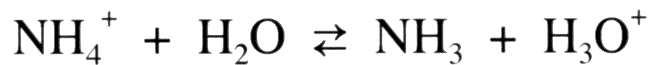
- the equilibrium expression for the ionization is:

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

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

- there is an important relationship that exists between  $K_a$  and  $K_b$  for conjugate acid-base pairs; consider the following

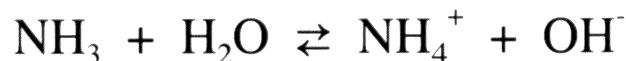
> the acid  $\text{NH}_4^+$  has the acid ionization equation:



> the acid ionization expression:

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

> the conjugate base  $\text{NH}_3$  has the base ionization equation:



> and the base ionization expression

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

> when  $K_a$  and  $K_b$  are multiplied together we get:

$$K_a \times K_b = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} \times \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

> and since  $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$

$K_a \times K_b = K_w$ <p>(at <math>25^\circ\text{C}</math> <math>K_w = 1.00 \times 10^{-14}</math>)</p>
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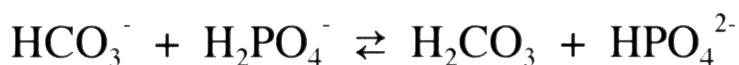
Calculate the  $K_b$  for  $C_2O_4^{2-}$ .

• need  $K_a$  for conj. acid

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{0.4 \times 10^{-5}} = 1.6 \times 10^{-10}$$

\*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.  $HCO_3^-$  ( $K_a = 5.6 \times 10^{-11}$ ) &  $H_2PO_4^-$  ( $K_a = 6.2 \times 10^{-8}$ ), since  $H_2PO_4^-$  has a larger  $K_a$ , it will donate a proton to  $HCO_3^-$



- the position of an equilibrium depends on the strengths of the acids
- ex.  $H_2CO_3$  ( $K_a = 4.3 \times 10^{-7}$ ) is a stronger acid than  $HSO_3^-$  ( $K_a = 1.0 \times 10^{-7}$ )
  - >  $H_2CO_3$  has a greater tendency to donate protons; there will be more products than reactants (products are favoured)



Weaker acids and bases are favoured in an equilibrium.