4.7-9 Equilibrium Dissociation for Water, Acids and Bases

a) What is Dissociation of Water?

i) Liquid water is in equilibrium with its aqueous ions:

$$2H_2O_{(l)} + 59kJ \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

ii) $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

iii) Water is neutral so: $[H_3O^+] = [OH^-]$

b) What Does K_w Tell Us Immediately?

i) There will always be a wee bit of **acid** in an aqueous solution. Even in a strong base! Since the value is so small, if a strong base is present, we can ignore the value!

ii) What is $[H_3O^+]$ in water?

$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

Since $[H_3O^+] = [OH^-]$...........$[x][x] = 1.00 \times 10^{-14}$ **$x = 1.00 \times 10^{-7}$ M**

iii)There will always be a wee bit of **base** in an aqueous solution. Even in a strong acid! Since the value is so small, if a strong acid is present, we can ignore the value!

iv) What is $[OH^-]$ in water?

$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

Since $[H_3O^+] = [OH^-]$...........$[x][x] = 1.00 \times 10^{-14}$ **$x = 1.00 \times 10^{-7}$ M**

c) Use of K_w Expression

i) We can find the $[OH^-]$ if we know the $[H_3O^+]$ and vice versa!

ii) Example: What is $[H_3O^+]$ and $[OH^-]$ in 0.015M HCl?

① HCl is a strong acid, so is 100% dissociated

② Therefore [HCl] = $[H_3O^+]$ = **0.015M**

③ $[OH^-] = \frac{Kw}{[H_3O^+]} = \frac{1.00 \times 10^{-14}}{0.015M} = \mathbf{6.7 \times 10^{-13}\ M}$

iii) Example: What is $[H_3O^+]$ and $[OH^-]$ in 1.5×10^{-6} M NaOH?

① NaOH is a metal hydroxide, so is a strong base, so is 100% dissociated

② Therefore [NaOH] = $[OH^-]$ = **1.5×10^{-6} M**

③ $[H_3O^+] = \frac{Kw}{[OH^-]} = \frac{1.00 \times 10^{-14}}{1.5 \times 10^{-6}\ M} = \mathbf{6.7 \times 10^{-9}\ M}$

d) Acid Dissociation

i) Recall, only weak acids will form an equilibrium.

$$H_2S_{(aq)} + H_2O_{(l)} \rightleftharpoons HS^-_{(aq)} + H_3O^+_{(aq)}$$

ii) $K_a = \frac{[HS^-][H_3O^+]}{[H_2S]} = 9.1 \times 10^{-8}$

iii) K_a is called the Acid Ionization Constant

iv) The greater the value of K_a, the stronger the acid (see Table of Relative Strengths of Acids p.334)

v) Example: Write the reaction and the K_a expression of H_2SO_3 with water

$$H_2SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HSO_3^-{}_{(aq)} + H_3O^+_{(aq)}$$

$$K_a = \frac{[HSO_3^-][H_3O^+]}{[H_2SO_3]} = 1.5 \times 10^{-2}$$

d) Base Dissociation

i) Recall, only weak bases will form an equilibrium.

$$CO_3^{-2}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + OH^-{}_{(aq)}$$

ii) $K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{-2}]} = 1.8 \times 10^{-4}$

iii) K_b is called the Base Ionization Constant

iv) The greater the value of K_b, the stronger the base.

e) Finding K_b

i) Theory

There is a relationship between K_a and K_b

①**Acid** dissociation: $\mathbf{H_2S}_{(aq)} + H_2O_{(l)} \rightleftharpoons HS^-{}_{(aq)} + H_3O^+{}_{(aq)}$

$$K_a = \frac{[HS^-][H_3O^+]}{[H_2S]} = 9.1 \times 10^{-8}$$

②**Base** dissociation: $\mathbf{HS^-}(aq) + H_2O_{(l)} \rightleftharpoons H_2S_{(aq)} + OH^-{}_{(aq)}$

$$K_b = \frac{[H_2S][OH^-]}{[HS^-]} = 1.1 \times 10^{-7}$$

③$K_a \times K_b = \frac{[HS^-][H_3O^+]}{[H_2S]} \times \frac{[H_2S][OH^-]}{[HS^-]} = [H_3O^+][OH^-] = K_w$

ii) Procedure:

①Look on *Base side* of Table of Relative Strengths to find your base species

②Find the Ka for the reverse reaction.

③Calculate K_b from $K_a \times K_b = K_w$

iii) Example: Find the K_b for SO_4^{-2}

On the table we find: $HSO_4^- \rightleftharpoons H^+ + SO_4^{-2}$ $K_a = 1.2 \times 10^{-2}$

$K_b = K_w / K_a = 1.00 \times 10^{-14} / 1.2 \times 10^{-2} = \mathbf{8.3 \times 10^{-13}}$

Do questions: #29-30 page 127; #31-34 page 128; #35-37 page 130