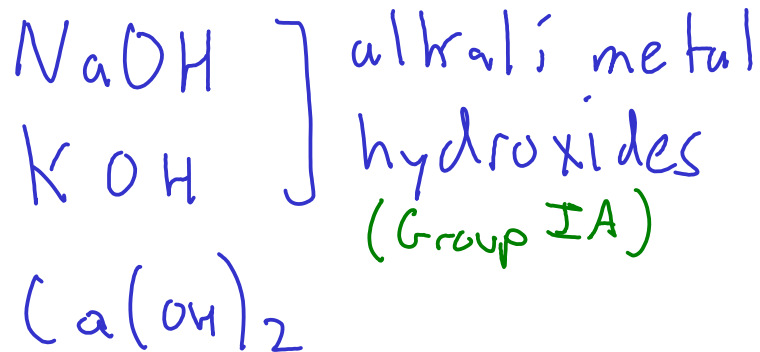# ACID-BASE EQUILIBRIUM IN WATER

- Like other ELECTROLYTES, acids and bases IONIZE to some extent in water

- STRONG electrolytes ionize completely. Acids and bases that ionize completely in water are called STRONG ACIDS and STRONG BASES

- WEAK electrolytes ionize partially, remaining mostly non-ionized. Acids and bases that ionize only partially in solution are called WEAK ACIDS and WEAK BASES.

 - Most acids and bases are WEAK!

Common strong acids

HCl

HNO_3

H_2SO_4 (only 1st proton)

HBr

HI

Common strong bases

$NaOH$ ⎤ alkali metal
KOH ⎦ hydroxides
(Group IA)

$Ca(OH)_2$

SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

- With strong acids and bases, the acid or base completely ionizes in water. So, we only have to worry about the effect of the acid or base on the water equilibrium itself.

- Since the equilibrium constant for the self-ionization of water is so small, the strong acid or base will overpower the hydronium (for acids) or hydroxide (for bases) produced by the water.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \quad ; \quad k_w = 1.0 \times 10^{-14}$$
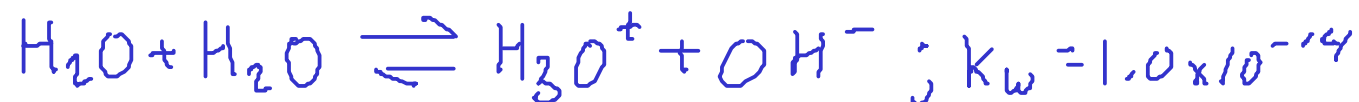
Consider a solution of 0.025 M nitric acid (a strong acid):

$$HNO_3 (aq) + H_2O (l) \longrightarrow NO_3^- (aq) + H_3O^+ (aq)$$

Assume that all the acid ionizes (strong acid), and we'l assume that all the hydronium ion in solution comes from the acid. The presence of hydronium from the acid (a lot of it, relatively) will suppress the self-ionization of water.

$$So, [H_3O^+] = 0.025 M \quad ([HNO_3]_{nominal})$$

$$\boxed{pH = 1.60} \quad *$$

* For logarithms, the places AFTER the decimal point are significant figures, while the places in front are not - they're effectively placeholders!

What would the HYDROXIDE ion concentration be under these conditions?

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$(0.025)[OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = 4.0 \times 10^{-13}$$

... this also equals the concentration of HYDRONIUM produced by the self-ionization of water - and it's extremely small!

Consider a solution of 0.0125 M sodium hydroxide (a strong base):

$$NaOH\,(aq) \longrightarrow Na^+\,(aq) + OH^-\,(aq)$$

Like before, we'll assume that all the hydroxide ion in solution comes from the base, since the base's hydroxide ion should suppress water's self-ionization.

$$[OH^-] = 0.0125\ M \quad \left([NaOH]_{nominal}\right)$$

We want to know pH. Taking the negative log of hydroxide concentration gives us pOH ...

$$pOH = 1.90$$

Remember that pOH and pH are related:

$$pH + pOH = 14.00$$

$$pH + 1.90 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's check the concentration of hydronium ion produced by the water wquilibrium:

$$[H_3O^+] = 10^{-pH} = 8.0 \times 10^{-13}\ M = [H_3O^+]$$

... since this also equals the amount of hydroxide ion produced by the water equilibrium, we can safely conclude the presence of the strong base effectively shuts down the self-ionization of water.

(A) What is the concentration of hydronium ion in an aqueous solution whose pH is 10.50? (B) What is the hydroxide ion concentration? (C) What molar concentration of sodium hydroxide solution would provide this pH?

A) $pH = 10.50$, $[H_3O^+] = ?$

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

$[H_3O^+] = 10^{-10.50} = \boxed{3.2 \times 10^{-11} \text{ M } H_3O^+}$

B) $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

$(3.2 \times 10^{-11})[OH^-] = 1.0 \times 10^{-14}$
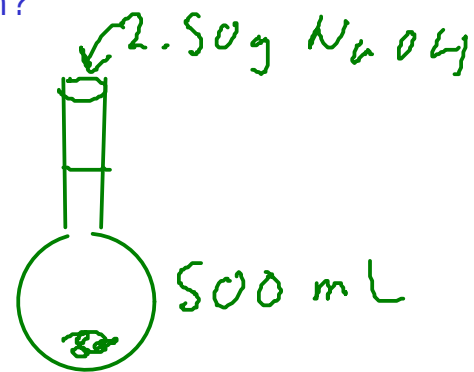
$[OH^-] = \boxed{3.2 \times 10^{-4} \text{ M } OH^-}$

C) $NaOH \longrightarrow Na^+ + OH^-$

1:1 ratio of $NaOH : OH^-$, so (0.00034 m !)

$[NaOH]_{nominal} = \boxed{3.2 \times 10^{-4} \text{ M } NaOH}$

What is the pH of a sodium hydroxide solution made from dissolving 2.50 g of sodium hydroxide in enough water to make 500.0 mL of solution?

NaOH: 40.00 g/mol

$$M_{NaOH} = \frac{mol \ NaOH}{L \ solution} \leftarrow 500 \ mL = 0.5000 \ L$$

← 2.50 g NaOH

500 mL

Find moles NaOH to find concentration:

$$2.50 \text{ g } NaOH \times \frac{mol \ NaOH}{40.00 \text{ g } NaOH} = 0.0625 \text{ mol } NaOH$$

$$M = \frac{0.0625 \text{ mol } NaOH}{0.5000 \ L} = 0.125 \ M \ NaOH$$

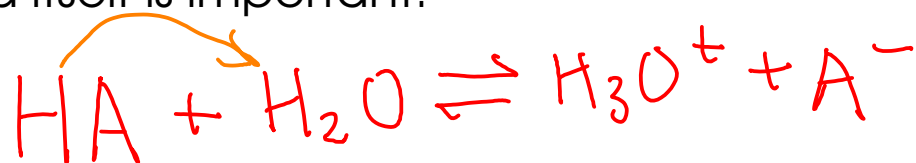Since sodium hydroxide is a strong base, it will completely ionize and control the amount of hydroxide ion present:

$$NaOH \rightarrow Na^+ + OH^- ; \ [OH^-] = [NaOH] = 0.125 \ M$$

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H_3O^+](0.125) = 1.0 \times 10^{-14}$$

$$[H_3O^+] = 8.0 \times 10^{-14}$$

$$\boxed{pH = 13.10}$$

For a WEAK ACID, equilibrium does not lie far to the right. The ionization equilibrium of the acid itself is important!
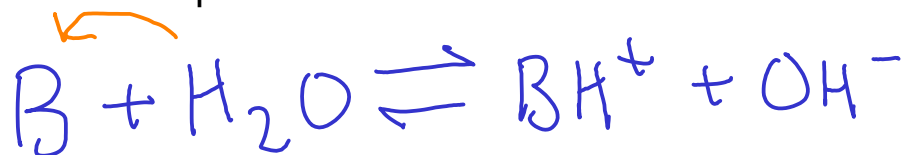
$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

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

acid ionization constant

Again, water's concentration will not change significantly, so it is folded into the ionization constant

(HA) = concentration of undissociated acid

For a WEAK BASE, equilibrium does not lie far to the right. The ionization equilibrium of the base itself is important!

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

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

base ionization constant

Values for Ka and Kb can often be found in data books / tables / or on the web.

In Ebbing, this data is in the appendices, on pages A-13 and A-14

WEAK ELECTROLYTES

- In solutions of weak acids or bases, the UNDISSOCIATED form is present in significantly high concentration.

- The pH of a solution of weak acid will be HIGHER than the pH of a strong acid solution with the same nominal concentration!

H^+ Cl^-

H^+ Cl^-

H^+ Cl^-

HCl

strong acid

VS

$HC_2H_3O_2$

$HC_2H_3O_2$

H^+ $C_2H_3O_2^-$

$HC_2H_3O_2$
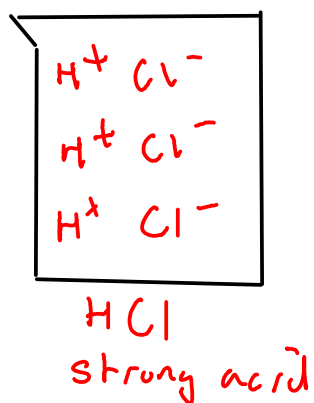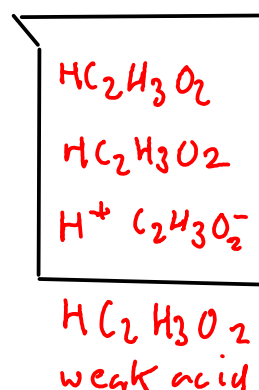
weak acid

Fewer molecules of the weak acid ionize, so the concnetration of hydrogen/hydronium ion is lower, meaning a higher pH!

- The pH of a solution of weak base will be LOWER than the pH of a strong base solution with the same nominal concentration!