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

$\left.\begin{array}{l} NaOH \\ KOH \end{array}\right]$ alkali metal 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.

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \; ; K_w = 1.0 \times 10^{-14}$$

Consider a solution of 0.025 M nitric acid (a strong acid):

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

Since nitric acid is strong, the above reaction goes fully to the right, meaning that all the nitric acid molecules produce hydronium ion! Since the presence of hydronium ion from the acid suppresses an already small amount produced by the water, we'll assume that essentially all hydronium in solution comes from the acid.

$$[H_3O^+] = [HNO_3]_{nominal} = 0.025 \text{ M } H_3O^+$$

$$pH = -\log_{10}[H_3O^+] = -\log(0.025) = \boxed{1.60}$$

* For logarithms, the significant figures are the digits BEHIND the decimal!

Let's calculate the HYDROXIDE concentration at this pH ...

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

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

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

Since the hydroxide is produced only by the water equilibrium, we know that it will produce the same amount of hydronium (1:1 ratio!). So the amount of hydronium produced by water is FAR smaller than 0.025 M!

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

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

Sodium hydroxide is a soluble ionic compound, so it completely ionizes in water. Similar to what we did with hydronium for the acid, we will assume that all the hydroxide ion in the solution comes from NaOH, since hydroxide from NaOH should suppress water's own ionization - which is already small!

$$[OH^-] = [NaOH]_{nominal} = 0.0125\ M\ OH^-$$

$$pOH = -\log_{10}[OH^-] = -\log_{10}(0.0125) = 1.90$$

Since $pH + pOH = 14.00$, $pH + 1.90 = 14.00$

$$\boxed{pH = 12.10}$$

Like before, we can check to see what amount of hydroxide ion water itself would make at this pH ... It will be equal to the amount of HYDRONIUM produced, since the water equilibrium must produce a hydroxide ion each time it produces a hydronium ion.

$$10^{-pH} = [H_3O^+], \quad 10^{-12.10} = [H_3O^+] = 8.0 \times 10^{-13}\ M\ H_3O^+$$

Based on the amount of hydronium produced by the water equilibrium here, we can also conclude that the amount of hydroxide ion produced is also really small (same numerical value as hydronium) compared to the 0.0125 M produced by NaOH.

(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

$[H_3O^+] = 10^{-10.50}$ $(10^{-pH} = H_3O^+)$

$= \boxed{3.2 \times 10^{-11}\ M\ H_3O^+}$

B) $(3.2 \times 10^{-11})[OH^-] = 1.0 \times 10^{-14}$ $([H_3O^+][OH^-] = 1.0 \times 10^{-14})$

$[OH^-] = \boxed{3.2 \times 10^{-4}\ M\ OH^-}$

C) Since NaOH is strong $(NaOH \rightarrow Na^+ + OH^-)$,

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