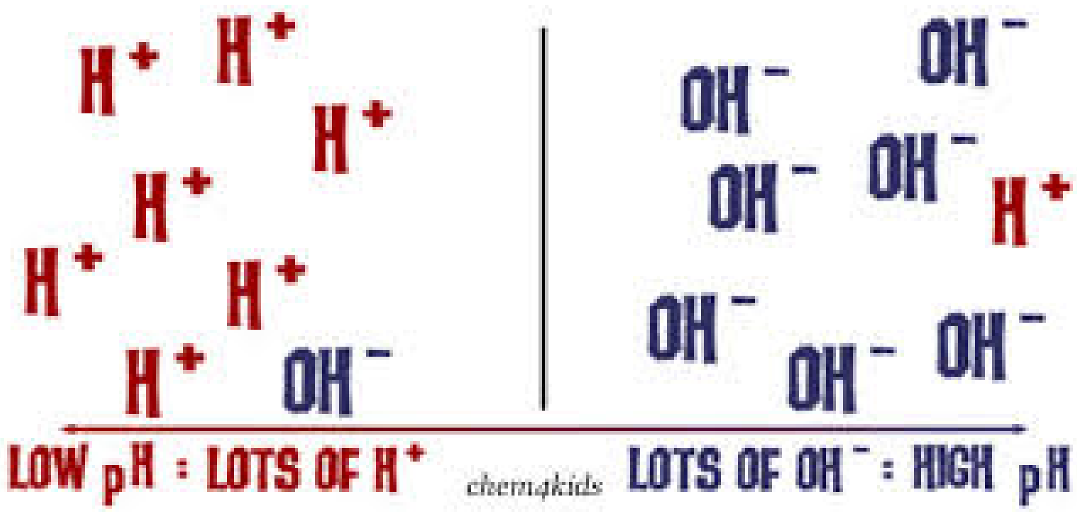
pH of Weak/Strong Acids & Bases…



Outcomes:

Write the equilibrium expression (K_a or K_b) from a balanced chemical equation.

Use K_a or K_b to solve problems for pH, percent dissociation, and concentration.

pH of Strong & Weak Ac ids/Bases:

Recall that pH is the **NEGATIVE** **LOGARITHM** of the **HYDRONIUM** or **HYDROXIDE** ion **CONCENTRATION** in a solution.

pH of Strong Acids/Bases:

- Since strong acids/bases **IONIZE** **COMPLETELY**, we can simply use **STOICHIOMETRY** to find pH.

Example:

strong

Find the pH of a 0.08 M solution of $Ba(OH)_2$.

$$Ba(OH)_2 \longrightarrow Ba^{2+} + \boxed{2OH^-}$$

$$0.08 \longrightarrow \boxed{0.16 \frac{mol}{L}}$$

$$pOH = -\log 0.16 = 0.79$$

$$pH = 14 - 0.79 = 13.21$$

pH of Strong & Weak Ac ids/Bases:

pH of Weak Acids/Bases:

- Since **WEAK** acids/bases do **NOT** ionize **COMPLETELY**, we **CANNOT** simply use stoichiometry.
- We must be given either the **DISSOCIATION CONSTANT**, or **PERCENT IONIZATION**.
- We must solve for the **[H_3O^+]** or **[OH^-]** as before, and then calculate **pH**.

pH of Strong & Weak Ac ids/Bases:

pH of Weak Acids/Bases Examples:

1. Given K_a or K_b:

Calculate the pH of a 0.10M solution of hydrogen sulfide (K_a=1.0x10^{-7})

$$H_2S + H_2O \rightleftharpoons H_3O^+ + HS^-$$

	H_2S	H_2O	H_3O^+	HS^-
	0.1	X	0	0
	$-x$		$+x$	$+x$
	$0.1-x$		x	x

$$K_a = \frac{[H_3O^+][HS^-]}{[H_2S]}$$

$$1\times10^{-7} = \frac{(x)^2}{0.1-x}$$ assume x is small

$$x = 1\times10^{-4}\ \frac{mol}{L} = [H_3O^+]$$

$$pH = -\log 1\times10^{-4} = 4$$

******When determining pH, we only do so for the first donated proton, since the K_a for the second proton is very small (has no real effect on pH).***

pH of Strong & Weak Ac ids/Bases:

pH of Weak Acids/Bases Examples:

2. Given the percent dissociation:

Calculate the pH of a 0.03M solution of sulfurous acid if 0.02% is ionized

$$H_2SO_3 + H_2O \rightleftharpoons H_3O^+ + HSO_3^-$$

H_2SO_3	H_2O	H_3O^+	HSO_3^-
0.03		0	0
-6×10^{-6}		$+6\times10^{-6}$	$+6\times10^{-6}$
		6×10^{-6}	

$$0.03 \times \frac{0.02\%}{100} = 6\times10^{-6}\ \frac{mol}{L} = [H_3O^+]$$

$$pH = -\log 6\times10^{-6} = 5.22$$