Part 2 - Ka, Kb and pH and pOH calculations

- the acid ionization constant, K_a, is the equilibrium constant for the reaction of an acid with water

$$HA_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)} \qquad K_a = \frac{\text{products}}{\text{reactants}}$$

- K_a values are found in the right hand column of your Acid-base chart on pg 8-9 of your data booklet.

- Note that in this table, the first six acids have Ka values given only as "very large."

- These acids are collectively called the strong acids because they all react quantitatively (>99.9%) with water to form hydronium ions

Example

$$HF(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + F^-(aq)$$

The equilibrium law expression is

$$K_a = \frac{[H_3O^+(aq)][F^-(aq)]}{[HF(aq)]} = 6.3 \times 10^{-4}$$

- Two calculations involving the K_a constant are common for weak acid solutions:

1. calculating a K_a value from measured (empirical) amount concentration data

2. using a K_a value to predict a concentration of hydronium ions for an aqueous solution where the initial weak acid amount concentration is known

Calculating K_a from Amount Concentrations

Example 1:

@ equilibrium

The pH of a 1.00 mol/L solution of acetic acid is carefully measured to be $\boxed{2.38}$ at SATP. What is the value of Ka for acetic acid?

Step 1: Write the reaction of the acid with water and write the Ka equilibrium law expression

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Step 2: find the equilibrium concentration of aqueous hydronium ion from the pH

Step 3: Use an ICE table to calculate the equilibrium concentration of the weak acid

	$1\,CH_3COOH$	$+\ H_2O_{(l)} \rightleftharpoons$	$1\,CH_3COO^-$	$+1\,H_3O^+$
I	1.0 mol/L		0	0
C	−0.004168		+0.004168	+0.004168
E	0.9958... mol/L		0.004168...	0.004168... mol/L

Step 4: Use your equilibrium concentrations to calculate Ka

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = \frac{(0.004168...)(0.004168...)}{0.9958....}$$

$$K_a = \boxed{1.7 \times 10^{-5}}$$

Example 2:

A student measures the pH of a 0.25 mol/L solution of carbonic acid to be 3.48. Calculate the Ka for carbonic acid from this evidence.

$$H_2CO_3 + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCO_3^-{}_{(aq)}$$

	H_2CO_3	H_2O	H_3O^+	HCO_3^-
I	0.25		0	0
C	-3.311×10^{-4}		3.311×10^{-4}	$+3.311\times10^{-4}$
E	0.24966...		3.311×10^{-4}	3.311×10^{-4}

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} = \boxed{4.4\times10^{-7}}$$

Example 3:

The pH of a 0.400 mol/L solution of sulfurous acid is measured to be 1.17. Calculate the Ka for sulfurous acid from this evidence.

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

	H_2SO_3	H_2O	H_3O^+	HSO_3^-
I	0.400		0	0
C	-0.0676..		+0.0676...	+0.0676...
E	0.3323...		0.0676..	0.0676....

$$K_a = \frac{[H_3O^+][HSO_3^-]}{[H_2SO_3]}$$

$$\boxed{= 0.014}$$

Calculating $[H_3O^+_{(aq)}]$ from K_a

Example 1:

Predict the $[H_3O^+_{(aq)}]$ and pH for a 0.200 mol/L aqueous solution of methanoic acid

Use the balanced equation to write the equilibrium law expression.

$$HCOOH + H_2O \rightleftharpoons H_3O^+ + HCOO^-_{(aq)}$$

I 0.200 mol/L   0   0

C

E

$$1.8\times10^{-4} = \frac{x^2}{0.200} \qquad x = \sqrt{(1.8\times10^{-4})(0.200)}$$

$$\boxed{[H_3O^+] = x = 0.006 \text{ mol/L}}$$

$$pH = -\log(\checkmark) = \boxed{2.222}$$

Example 2:

Predict the $[H_3O^+_{(aq)}]$ and pH for a 0.500 mol/L aqueous solution of hydrocyanic acid.

$$HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CN^-_{(aq)}$$

0.500   0   0

$-x$ $+x$ $+x$

0.500 x x

$$k_a = \frac{[H_3O^+][CN^-]}{[HCN]} \qquad 6.2\times10^{-10} = \frac{x^2}{0.500}$$

$$\boxed{\begin{array}{c}[H_3O^+] = x = 1.76\times10^{-5} \text{ mol/L} \\ pH = 4.754\end{array}}$$