

Period 42

- Quiz on Bronsted-Lowry Acids and Bases
- Go over Ka worksheet #1

① Calculation of Ka values

If we are given the [H+] concentration and the concentration of the acid solution, we can easily calculate the Ka value for the acid.

✓ Example #1 :

An unknown weak acid, HW, is found to have a [H+] of 1.5 x 10⁻⁵ M when we mix a 2.5 M solution of the acid. Calculate Ka for this acid.

Answer #1 :

$$\begin{aligned}
 & \text{HW}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{W}^-_{(aq)} \\
 \text{Ka} &= \frac{[\text{H}^+][\text{W}^-]}{[\text{HW}]} = \frac{(1.5 \times 10^{-5})(1.5 \times 10^{-5})}{(2.5 - (1.5 \times 10^{-5}))} \\
 &= \frac{(1.5 \times 10^{-5})^2}{(2.5)} \\
 &= 9.0 \times 10^{-11}
 \end{aligned}$$

Example #2 :

Another weak acid, HZ, has a 1.8 x 10⁻⁸ M [OH⁻] in a 0.025 M solution of the acid. Calculate the Ka of HZ.

Answer #2 :

$$\begin{aligned}
 & \text{HZ}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{Z}^-_{(aq)} \\
 \text{If } [\text{OH}^-] &= 1.8 \times 10^{-8} \text{ M the } [\text{H}^+] = 5.6 \times 10^{-7} \text{ M} \\
 \text{Ka} &= \frac{[\text{H}^+][\text{Z}^-]}{[\text{HZ}]} = \frac{(5.6 \times 10^{-7})(5.6 \times 10^{-7})}{(0.025 - (5.6 \times 10^{-7}))} \\
 &= \frac{(5.6 \times 10^{-7})^2}{(0.025)} \\
 &= 1.2 \times 10^{-11}
 \end{aligned}$$

- Assignment :

- Ka worksheet question #2
- Read sections 20-6, 20-7, 20-8

Period 41

p.554 (not calculation)

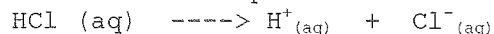
- Go over worksheet

- Quiz tomorrow on worksheet

② Ka Calculations

Q1 : What is the $[H^+]$ in a 0.50 M HCl solution ?

A1 : Because HCl is a strong acid, it completely dissociates as per the following equation :



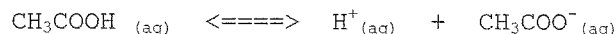
The ratio between HCl and H^+ is one to one therefore the $[H^+]$ will be 0.50 M.

Q2 : What is the $[H_3O^+]$ in a 15.0 M HNO_3 solution ?

A2 : $[H_3O^+] = 15.0$ M

✓ Q3 : What is the $[H^+]$ in a 0.50 M CH_3COOH solution ?

A3 : Because CH_3COOH is a weak acid, it will not completely dissociate but will set up an equilibrium as follows :



According to our chart, the equilibrium constant for this reaction is 1.8×10^{-5} .

$$K_{eq} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

Like with solubility, the equilibrium expression involving an acid is used so frequently, we give it its own name : the ACID CONSTANT and call it K_a . K_a in fact is nothing more than a specialized K_{eq} expression. Now, back to our question. When we dissolve an acid into water, we are setting up an 'ICE' problem.

	CH_3COOH	\rightleftharpoons	H^+	+	CH_3COO^-
I	0.50		0		0
C	-X		+X		+X
E	0.50-X		X		X

$$K_{eq} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{(X)(X)}{(0.50-X)}$$

As we can see, this will work out quite quickly into an absolutely messy quadratic equation to solve for X. However, there is an easier method.

Look at the denominator : $(0.50-X)$. Because the K_a value is very low, the amount of dissociation is insignificant compared to the original concentration. We can therefore ignore the X in the denominator.

$$1.8 \times 10^{-5} = \frac{(X)(X)}{(0.50-X)} \quad \text{assume } X \text{ is insignificant}$$

$$1.8 \times 10^{-5} = \frac{(X)(X)}{(0.50)}$$

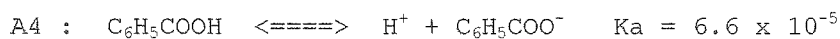
$$9.0 \times 10^{-6} = X^2$$

$$3.0 \times 10^{-3} = X$$

$$\text{The } [H^+] = 3.0 \times 10^{-3} \text{ M}$$

NOTE : if we were to have calculated the question via the quadratic formula, we would have got 2.99×10^{-3} M. Either way, when we make a subtraction from 0.50 M (keeping in mind the rules for significant figures) the amount of dissociation is indeed insignificant compared to the original concentration.

✓ Q4 : What is the $[H_3O^+]$ in a 3.0 M solution of benzoic acid ?



$$K_a = \frac{[H^+][C_6H_5COO^-]}{[C_6H_5COOH]}$$

$$6.6 \times 10^{-5} = \frac{(X)(X)}{(3.0 - X)} \quad \text{X is insignificant}$$

$$6.6 \times 10^{-5} = \frac{(X)(X)}{(3.0)}$$

$$1.98 \times 10^{-4} = X^2$$

$$1.4 \times 10^{-2} = X$$

$$[H_3O^+] = 1.4 \times 10^{-2} \text{ M}$$

Note:
Similar
calculations
for K_b (K_b
must be
calculated)