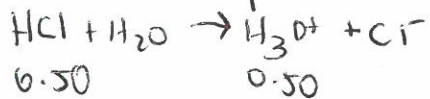


Ex A.

What is the pH of 0.50M HCl?

Name: Key

Blk: \_\_\_\_\_ Date: \_\_\_\_\_



CHEMISTRY 12  
ACID BASES UNIT

Lesson #12

pH =  $-\log[0.50] = 0.30$  **Calculations Involving  $K_a$  and pH**

IMPT: The pH scale is NOT LINEAR!!! When the pH is increased by 1, the  $[\text{H}_3\text{O}^+]$  is DECREASED by 10!!!

Recall that Strong Acids IONIZE 100 %, therefore the concentration of the strong acid will equal the concentration of  $\text{H}_3\text{O}^+$  in solution!!

HOWEVER, Weak acids DO NOT IONIZE 100% in water, therefore we must use an **ICE TABLE** to determine the  $[\text{H}_3\text{O}^+]$  that is actually present in solution!

Generic Equation for a WEAK ACID in water:



There are THREE types of problems that you can solve associated with a weak acid:

**Type 1.** Given the  $K_a$  and the concentration of the weak acid, solve for the pH

**Example 1.** What is the pH for a 0.500 M solution of  $\text{CH}_3\text{COOH}$ ?

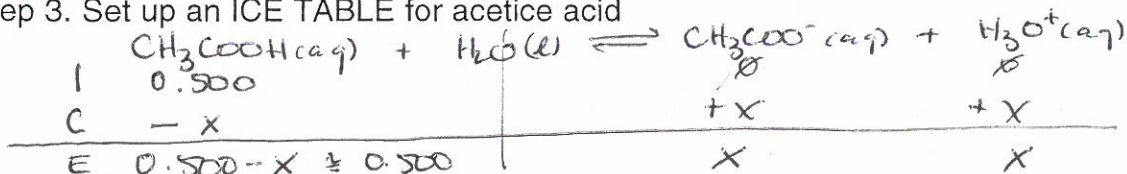
Step 1. Write out the ionization equation with water



Step 2. Write out the  $K_a$  expression, identify the  $K_a$  value

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

Step 3. Set up an ICE TABLE for acetic acid



Step 4. Plug values into  $K_a$  and solve for x

$$1.8 \times 10^{-5} = \frac{[x][x]}{0.500} = \frac{x^2}{0.500} \rightarrow \sqrt{1.8 \times 10^{-5} \times 0.500} = \sqrt{x^2}$$

$$\underline{3.0 \times 10^{-3}} = x$$

Step 5. Use x to solve for pH

$$\therefore [\text{H}_3\text{O}^+] = 3.0 \times 10^{-3} \quad \text{pH} = -\log[\text{H}_3\text{O}^+] = \boxed{2.52}$$

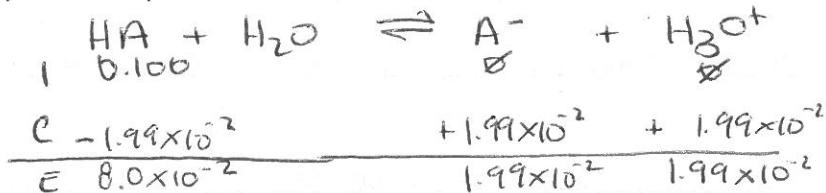
**Type 2.** Given the pH and the initial concentration of the weak acid, solve for the  $K_a$  value.

**Example 2.** If the pH of a 0.100 M solution of an unknown weak acid is 1.70, determine the  $K_a$  and identify the weak acid.

Step 1 . Write out the generic ionization with water



Step 2. Set up an ICE TABLE for the weak acid



Step 3. Use the pH to fill in the values of the ICE TABLE

$$pH = 1.70 \quad \therefore [H_3O^+]_E = \text{antilog}(-1.70) \Rightarrow 1.99 \times 10^{-2}$$

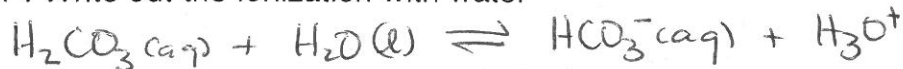
Step 4. Use the equilibrium values to calculate the  $K_a$

$$K_a = \frac{[1.99 \times 10^{-2}][1.99 \times 10^{-2}]}{8.0 \times 10^{-2}} \rightarrow 5.0 \times 10^{-3} \quad \therefore \text{either Citric or Iron III}$$

**Type 3.** Given pH and  $K_a$ , determine the initial concentration of the weak acid.

**Example 3.** What concentration of  $H_2CO_3$  is required to produce a pH of 4.18?

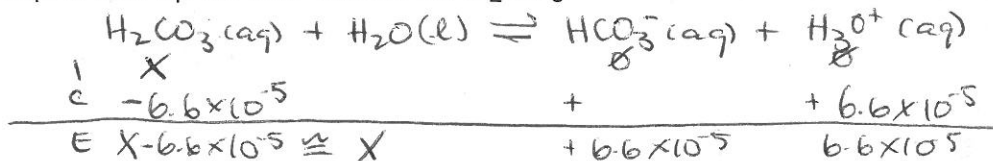
Step 1 . Write out the ionization with water



Step 2. Write out the  $K_a$  expression, look up  $K_a$  value

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} = 4.3 \times 10^{-7}$$

Step 3. Set up an ICE TABLE for  $H_2CO_3$



Step 4. Use pH to fill in the table

$$pH = 4.18 \quad \therefore [H_3O^+]_E = \text{antilog}(-4.18) = 6.6 \times 10^{-5}$$

Step 5. Use  $K_a$  to solve for initial concentration

$$\rightarrow 4.3 \times 10^{-7} = \frac{[6.6 \times 10^{-5}]^2}{X}$$

**Seatwork/Homework:** Exercises 74-78, 80, 82-83  
**PLO's:** Part of M3 and M5 for  $K_a$

$$X = \frac{[6.6 \times 10^{-5}]^2}{4.3 \times 10^{-7}}$$

$$X = 1.0 \times 10^{-2} \text{ M}$$

$$[H_2CO_3]_I = 1.0 \times 10^{-2} \text{ M}$$