

## Weak Acid Problems:

### Steps to Solving problems:

1. list major species
2. choose species that can produce  $H^+$  ions and write a balanced equation for the reaction producing  $H^+$
3. Use values of equilibrium constants for reaction written and decide which will dominate in producing  $H^+$
4. write equilibrium expression
5. ICE the problem
6. write the equilibrium expression and substitute equilibrium concentrations into the expression
7. Solve the "easy" way: **AXIS (assume x is small)**  
 $[HA]_o - x \approx [HA]_o$
8. calculate  $[H^+]$  and pH

Ex. 1

The hypochlorite ion is a strong oxidizing agent found in bleaches and disinfectants. It also forms when swimming pool water is treated with chlorine. The ion also has a high affinity for protons (it's a stronger base than the  $\text{Cl}^-$ ) and forms weakly acidic hypochlorous acid ( $\text{HOCl}$ ,  $K_a = 3.5 \times 10^{-8}$ ). Calculate the pH of a 0.100 M aqueous solution of hypochlorous acid.

See yesterday's work

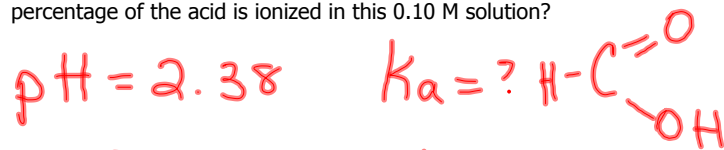
Ex. 2

Calculate the pH of a solution that contains 1.00 M HCN ( $K_a = 6.2 \times 10^{-10}$ ) and 5.00 M  $\text{HNO}_2$  ( $K_a = 4.0 \times 10^{-4}$ ). Also calculate the concentration of cyanide ion ( $\text{CN}^-$ ) in this solution at equilibrium.

See yesterday's  
work

Ex. 3

A student prepared a 0.10 M solution of formic acid ( $\text{HCHO}_2$ ) and measured its pH using a pH meter. The pH at  $25^\circ\text{C}$  was found to be 2.38. (a) Calculate  $K_a$  for formic acid at this temperature. (b) What percentage of the acid is ionized in this 0.10 M solution?



I    0.100            0            0

C    -x                +x            +x

E    0.100-x            ~~x~~            ~~x~~

$\text{pH} = 2.38$      $4.17 \times 10^{-3}$      $4.17 \times 10^{-3}$

$$[\text{H}^+] = 10^{-2.38}$$

$$= 4.17 \times 10^{-3} \text{ M}$$

↑ This is value of x.

$$K_a = \frac{[\text{H}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]}$$

$$= \frac{(4.17 \times 10^{-3})(4.17 \times 10^{-3})}{0.100 - x}$$

~~0.100 - x~~    **AXIS**

$$K_a = 1.74 \times 10^{-4}$$

$0.100 - x \approx 0.100$

$$\begin{aligned} \% \text{ dissociation} &= \frac{[\text{H}^+]}{[\text{HCHO}_2]} \times 100 \\ &= \frac{4.17 \times 10^{-3}}{0.100} \times 100 \\ &= 4.17\% \end{aligned}$$

Ex. 4

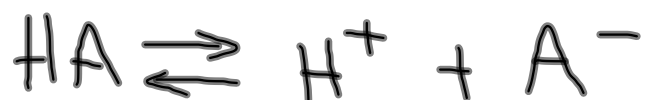
Calculate the pH of a 0.20 M solution of HCN. ( $K_a = 6.2 \times 10^{-10}$ )

See yesterday's  
work

Ex. 5



The  $K_a$  for niacin is  $1.6 \times 10^{-5}$ . What is the pH of a 0.010 M solution of niacin? What is the percentage of ionization? Calculate the percent dissociation of niacin.

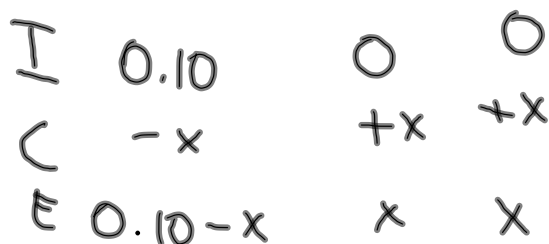


Ex. 6



Calculate the percentage of HF molecules ionized in (a) a 0.10 M HF;  
(b) a 0.0010 M HF.

$$K_a = 7.2 \times 10^{-4}$$



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

$$7.2 \times 10^{-4} = \frac{(x)(x)}{0.100}$$

AXIS

$$0.100 - x \approx 0.10$$

$$7.2 \times 10^{-5} = x^2$$

$$\sqrt{7.2 \times 10^{-5}} = x$$

$$.00849 \text{ M} = x = [\text{H}^+] = [\text{F}^-]$$

$$\% \text{ ionization} = \frac{[\text{H}^+]}{[\text{HF}]} \times 100$$

$$= \frac{0.00849}{0.100} \times 100$$

$$= 8.49\%$$