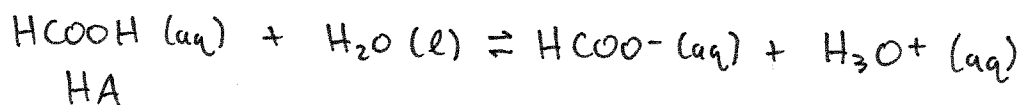


For a weak acid, the equation for its dissociation is



$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

pH of 0.1M soln of HCOOH @ 25°C is 2.38. Calculate K_a ← 2 sig figs (after decimal)



$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$$

- Assume that $[H_3O^+] = [HCOO^-]$
- $[H_3O^+] = 10^{-2.38} = 4.2 \times 10^{-3} M$

	$[HCOOH], M$	$[H_3O^+], M$	$[HCOO^-], M$
I	0.10	0	0
C	-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
E	$0.1 - 4.2 \times 10^{-3} = 0.0958$	4.2×10^{-3}	4.2×10^{-3}

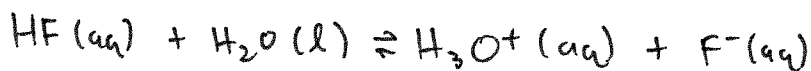
$$K_a = \frac{[4.2 \times 10^{-3}][4.2 \times 10^{-3}]}{[0.0958]} = 1.8 \times 10^{-4}$$

Calculating Percent Ionization

$$\text{Percent Ionization} = \frac{[H_3O^+]_{eq}}{[HA]_{initial}} \times 100$$

$$\text{In this example, } \frac{4.2 \times 10^{-3}}{0.10} \times 100 = 4.2\%$$

0.100M HF, the percent ionization is 8.1%. Calculate K_a



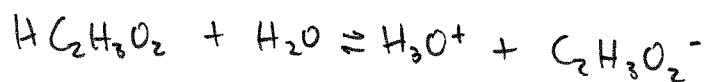
$$K_a = \frac{[H_3O^+][F^-]_{eq}}{[HF]_{eq}} = \frac{[8.1 \times 10^{-3}]^2}{0.100 - 8.1 \times 10^{-3}} = 7.14 \times 10^{-4}$$

$$\text{amount dissociated} = \frac{8.1}{100} \times 0.100 M = 8.1 \times 10^{-3} M = [F^-] = [H_3O^+]$$

Calculating pH Using K_a

- ① Write equilibrium constant expression from ionization equilibrium
- ② Set up ICE table to determine equilibrium concentrations as a function of x
- ③ Substitute equilibrium concentrations into the expressions for K_a and solve for x

Calculate pH of 0.30 M soln of acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$ @ 25°C



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.8 \times 10^{-5}$$

	$\text{CH}_3\text{COOH (M)}$	$\text{H}_3\text{O}^+ \text{ (M)}$	$\text{CH}_3\text{COO}^- \text{ (M)}$
I	0.30	0	0
C	-x	+x	+x
E	0.3-x	x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{x^2}{0.3-x} = 1.8 \times 10^{-5}$$

Option 1: Use quadratic equation to solve for x , then calculate pH

$$x = 0.0023$$

$$\text{pH} = 2.63$$

Option 2: Assume $x \ll 0.30$ and solve the following

$$\frac{x^2}{0.30} = 1.8 \times 10^{-5} \quad x = 2.3 \times 10^{-3}$$

$$\text{pH} = 2.63$$

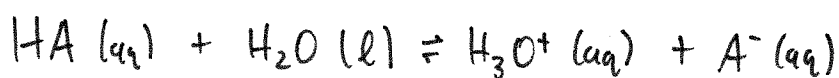
When is it OK to ignore x in the denominator?

- Rule of thumb, If $[\text{HA}]_0 \geq 100 K_a$, assume $[\text{HA}]_{\text{eq}} = [\text{HA}]_0$

- Check with "5% rule"

$$\frac{x}{[\text{HA}]_0} \times 100 < 5\%$$

What is pH of 0.00100 M acetic acid? $K_a = 1.8 \times 10^{-5}$



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

	HA	H_3O^+	A^-
I	0.00100	0	0
C	-x	x	x
E	0.00100 - x	x	x

$$K_a = \frac{x^2}{0.00100 - x}$$

Assume $0.00100 - x = 0.001$

$$x = 1.34 \times 10^{-4} \text{ M}$$

$$\text{pH} = 3.873$$

$$\frac{1.34 \times 10^{-3}}{0.001} \Rightarrow 13\%$$

Successive Approximations (Technique)

$$1.8 \times 10^{-5} = \frac{x^2}{0.001 - x}$$

First approximation, $x = 1.34 \times 10^{-4}$

$$\text{Solve } \frac{x^2}{0.001}$$

Second approximation,

$$\text{Solve } \frac{x^2}{0.001 - 1.34 \times 10^{-4}}$$

$$x = 1.25 \times 10^{-4}$$

Third approximation,

$$\text{Solve } \frac{x^2}{0.001 - 1.25 \times 10^{-4}}$$

$$x = 1.26 \times 10^{-4}$$

$$\text{pH} = 3.900$$

Polyprotic Acids

- Polyprotic Acids have more than one acidic proton
- Generally, $K_{a1} \gg K_{a2} \gg K_{a3}$ so, you can ignore 2nd (and 3rd) ionizations