Acids and Bases III Chapter 16

For a weak acid, the equation for its dissociation is

$$HA\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + A^-\,(aq)$$

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

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

$$HCOOH\,(aq) + H_2O\,(l) \rightleftharpoons HCOO^-\,(aq) + H_3O^+\,(aq)$$

HA

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

- 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

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

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

$[HA]_0$

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

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

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

Calculating pH Using Ka

1. Write equilibrium constant expression from ionization equilibrium
2. Set up ICE table to determine equilibrium concentrations as a function of x
3. Sustitute equilibrim concentrations into the expressions for Ka and solve for x

Calculate pH of 0.30 M sln of acetic acid, $HC_2H_3O_2$ @ 25°C

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.8 \times 10^{-5}$$

	CH_3COOH (M)	H_3O^+ (M)	CH_3COO^- (M)
I	0.30	0	0
C	$-x$	$+x$	$+x$
E	$0.3-x$	x	x

$$K_a = \frac{[H_3O^+][HC_2H_3O_2]}{[HC_2H_3O_2]} = \frac{x^2}{0.3-x} = 1.8 \times 10^{-5}$$

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

$x = 0.0023$

$pH = 2.63$

Option 2: Assume $x << 0.30$ and solve the following

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

$pH = 2.63$

When is it ok to ignore x in the demoniator?

- Rule of thumb, If $[HA]_0 \geq 100\ K_a$, assume $[HA]_{eq} = [HA]_0$
- Check with "5% rule"

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

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

$$HA\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + A^-\,(aq)$$

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

	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 $0.00100 - x = 0.001$

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

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

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

Second approximation,

Solve $\frac{x^2}{0.001 - 1.34 \times 10^{-4}}$ $x = 1.25 \times 10^{-4}$

third approximation,

Solve $\frac{x^2}{0.001 - 1.25 \times 10^{-4}}$ $x = 1.26 \times 10^{-4}$

pH = 3.900

Polyprotic Acids

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