March 15, 2002

- No office hour today
- No class next week!
- Exam solution key will be online sometime this weekend
- *Demo Today!*

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Successive Approximations

$$\frac{x^2}{0.10 - x} = 1.613 \times 10^{-5}$$

Assume: $x << 0.10$

$$\frac{(x')^2}{0.10} = 1.613 \times 10^{-5}$$

First Approximation

$x' = 1.27 \times 10^{-3}$

$$\frac{(x'')^2}{0.10 - 1.27 \times 10^{-3}} = 1.613 \times 10^{-5}$$

2nd Approximation

$x'' = \underline{\mathbf{1.26 \times 10^{-3}}}$

Less than 1% change - ***Stop!***

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pH Please!

$[OH^-] = x = 1.26 \times 10^{-3}$ *M*

pOH = -Log(1.26×10^{-3}) = *2.899*

pH = 14.00 - pOH = ***__11.10__***

- *Anions* of Weak Acids are Weak Bases
  - ✓ KCN in water will be basic
- **Cations** of Weak Bases are Weak Acids
  - ✓ NH_4NO_3 in water will be acidic

Common Ion Effect

- Suppose we add a *common ion* to a weak acid system at equilibrium?

<u>Example:</u> add *Sodium Acetate* to a solution of Acetic Acid

$$NaCH_3COO \rightarrow Na^+ (aq) + CH_3COO^- (aq)$$

$$CH_3COOH (aq) + H_2O (l) \leftrightarrows CH_3COO^- (aq) + H_3O^+ (aq)$$

- Adding *acetate ion* (a product) will shift equilibrium to the **__LEFT__**
  - ✓ $[H_3O^+]$ will *decrease*
  - ✓ pH will **increase**

Quantitatively

- Example: What happens to the pH when we add 10.00 mL of 0.100 *M* $NaCH_3COO$ to 10.00 mL of 0.100 *M* CH_3COOH?

First: Find intial pH (of the CH_3COOH alone)

	CH_3COOH (aq) + H_2O (l) ⇆	CH_3COO^- (aq) +	H_3O^+ (aq)
I	0.100 *M*	0	0
C	-x	+x	+x
E	0.100 - x	x	x

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pH of CH_3COOH alone

- Plug into K_a expression:

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.76 \times 10^{-5}$$

$$K_a = \frac{(x)(x)}{0.100 - x} = 1.76 \times 10^{-5}$$

- By *successive approximations* (or **quadratic formula)**:

$x = [H_3O^+] = 1.318 \times 10^{-3}$ *M*

pH = -Log(1.318×10^{-3}) = ***2.88***

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Now, add the $NaCH_3COO$

- We need to re-calculate concentrations when we change the volume of the solution

Trick: use *mmol* instead of **mol**

mL x M = *mmol*

10.00 mL x 0.100 M CH_3COOH = ***1.00 mmol*** CH_3COOH

10.00 mL x 0.100 M $NaCH_3COO$ = ***1.00 mmol*** $NaCH_3COO$

✓Divide by new solution volume to get initial concentrations:

$\frac{1.00 \text{ mmol}}{20.00 \text{ mL}} = 5.00 \times 10^{-2}\ M$ (for both CH_3COOH and $NaCH_3COO$)

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ICE Time

$CH_3COOH\ (aq) + H_2O\ (l) \leftrightarrows CH_3COO^-\ (aq) + H_3O^+\ (aq)$

	CH_3COOH	CH_3COO^-	H_3O^+
I	0.0500 M	0.0500 M	0
C	-x	+x	+x
E	0.0500 - x	0.0500 + x	x

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{(x)(0.0500 + x)}{0.0500 - x} = 1.76 \times 10^{-5}$$

Using Successive Approximations: $x' = 1.76 \times 10^{-5}$

So: $[H_3O^+] = 1.76 \times 10^{-5}\ M$ and pH = ***4.75***

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