

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: $x \ll 0.10$

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

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

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

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

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pH Please!

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

$$\text{pOH} = -\text{Log}(1.26 \times 10^{-3}) = 2.899$$

$$\text{pH} = 14.00 - \text{pOH} = \underline{\underline{11.10}}$$

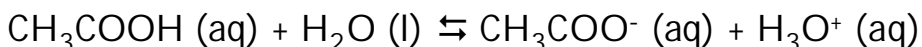
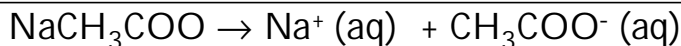
- **Anions** of Weak Acids are Weak Bases
 - ✓ KCN in water will be basic
- **Cations** of Weak Bases are Weak Acids
 - ✓ NH_4NO_3 in water will be acidic

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Common Ion Effect

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

Example: add *Sodium Acetate* to a solution of *Acetic Acid*



- Adding *acetate ion* (a product) will shift equilibrium to the **LEFT**

- ✓ $[\text{H}_3\text{O}^+]$ will *decrease*

- ✓ pH will **increase**

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Quantitatively

- Example: What happens to the pH when we add 10.00 mL of 0.100 M NaCH₃COO to 10.00 mL of 0.100 M CH₃COOH?

First: Find initial pH (of the CH₃COOH alone)

$$\text{CH}_3\text{COOH (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{CH}_3\text{COO}^- \text{ (aq)} + \text{H}_3\text{O}^+ \text{ (aq)}$$

<i>I</i>	0.100 M	0	0
<i>C</i>	-x	+x	+x
<hr/>			
<i>E</i>	0.100 - x	x	x

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pH of CH₃COOH alone

- Plug into K_a expression:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1.76 \times 10^{-5}$$

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

- By successive approximations (or quadratic formula):

$$x = [\text{H}_3\text{O}^+] = 1.318 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\text{Log}(1.318 \times 10^{-3}) = \underline{\underline{2.88}}$$

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Now, add the NaCH₃COO

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

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

$$\text{mL} \times M = \text{mmol}$$

$$10.00 \text{ mL} \times 0.100 \text{ M CH}_3\text{COOH} = \underline{1.00 \text{ mmol}} \text{ CH}_3\text{COOH}$$

$$10.00 \text{ mL} \times 0.100 \text{ M NaCH}_3\text{COO} = \underline{1.00 \text{ mmol}} \text{ NaCH}_3\text{COO}$$

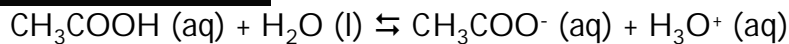
✓ Divide by new solution volume to get initial concentrations:

$$\underline{1.00 \text{ mmol}} = 5.00 \times 10^{-2} \text{ M} \text{ (for both CH}_3\text{COOH and NaCH}_3\text{COO)}$$

20.00 mL

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ICE Time



I	0.0500 M	0.0500 M	0
C	-x	+x	+x
E	0.0500 - x	0.0500 + x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{(x)(0.0500 + x)}{0.0500 - x} = 1.76 \times 10^{-5}$$

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

So: $[\text{H}_3\text{O}^+] = 1.76 \times 10^{-5} \text{ M}$ and $\text{pH} = \underline{4.75}$

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