March 27, 2002

- ➢Friday's office hour cancelled
- ➢Registration for Fall Semester
  - ✓Chem 141 versus Chem 143
  - ✓Chem 121 (Quantitative Analysis)?
  - ✓Math? Physics?
- ➢***Quiz on Friday***

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Back to Equilibrium

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

	CH_3COOH	CH_3COO^-	H_3O^+
I	0.90 mmol/30.00 mL	1.10 mmol/30.00 mL	-
C	-x	+x	+x
E	***3.0×10^{-2} -x***	***3.67×10^{-2} + x***	***x***

Substituting:

$$1.76 \times 10^{-5} = \frac{(3.67 \times 10^{-2} - x)(x)}{(3.0 \times 10^{-2} + x)}$$

Assume: $x << 10^{-2}$

$$\frac{(3.67 \times 10^{-2})x'}{(3.0 \times 10^{-2})} = 1.76 \times 10^{-5}$$

$$x' = \boxed{\mathbf{1.44 \times 10^{-5}}\ M = [H_3O^+]}$$ (assumption holds: $10^{-5} << 10^{-2}$)

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Effect on pH?

- pH = -Log[H_3O^+] = ***4.84***
- So: ΔpH = 4.75 - 4.84 = ***-0.09***

Very little change in pH!

<u>Compare with adding NaOH to water:</u>

<u>Add:</u> 10.00 mL 0.0100 *M* NaOH

to *20.00 mL H_2O*

[OH^-] = $\dfrac{0.100 \text{ mmol } OH^-}{30.00 \text{ mL}}$ = ***3.33 x 10^{-3} M***

pOH = *2.48*

pH = 14.00 - 2.48

pH = ***11.52***

So: ΔpH = 7.00 - 11.52 = ***-4.52*** (Huge!)

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Buffers

A mixture of a *weak acid* and its *conjugate base* **buffers** pH:

+ H^+

$$CH_3COOH + H_2O \leftrightarrows CH_3COO^- + H_3O^+$$

+ OH^-

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The Henderson-Hasselbalch Equation

➢Log formulation of the K_a expression:

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

Solve for $[H_3O^+]$:

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

$$\boxed{\mathbf{pH = pK_a + Log([A^-]/[HA])}}$$

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Titrations

➢If K_{rxn} is *large*, then it is useful for titrimetric determinations

➢We can *predict the pH* of the solution resulting from each addition of titrant in an **acid/base** titration, if we know:

✓Initial concentrations

✓Equilibrium constants

The Process:
-convert to *mmol*
-*complete* reaction
-convert to *concentrations*
-*ICE* to get $[H_3O^+]$
-convert to pH

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Example Titration

Let's titrate!

- **Analyte:**
  50.00 mL 0.100 *M* Acetic Acid (HAc)
  $K_a = 1.76 \times 10^{-5}$
- **Titrant:** 0.100 *M* NaOH

The Chemistry:

$HAc + OH^- \rightarrow Ac^- + H_2O$ (titration reaction)

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Initial pH (0.00 mL NaOH)

Just HAc (weak acid) alone in solution:

	$HAc + H_2O$	$\leftrightarrows$	Ac^-	$+ H_3O^+$
I	0.100 *M*		-	-
C	-x		+x	+x
E	0.100 - x		x	x

$$K_a = \frac{x^2}{0.100 - x}$$

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

$\boxed{pH = \underline{\mathbf{2.88}}}$

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