March 29, 2002

- ➢No office hour today
- ➢***Quiz Today!***

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5.00 mL NaOH added

Convert to mmol:

- 5.00 mL (0.100 mol OH^-/L) = 0.500 mmol OH^-
- 50.00 mL (0.100 mol HAc/L) = 5.00 mmol HAc

$$HAc + OH^- \rightarrow Ac^- + H_2O$$

Complete reaction:

	HAc	OH^-	Ac^-
I	5.00 mmol	0.500 mmol	-
C	-0.50 mmol	-0.500 mmol	+0.500 mmol
F	*4.50 mmol*	-	*0.500 mmol*

Convert to concentrations:

$$C_{HAc} = \frac{4.50\ \text{mmol}}{55.00\ \text{mL}} = 8.182 \times 10^{-2}\ M$$

$$C_{Ac^-} = \frac{0.50\ \text{mmol}}{55.00\ \text{mL}} = 9.091 \times 10^{-3}\ M$$

It's a Buffer!

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On to Equilibrium!

$$HAc + H_2O \leftrightarrows Ac^- + H_3O^+$$

	HAc		Ac⁻	H₃O⁺
I	8.182×10^{-2} *M*		9.091×10^{-3} *M*	-
C	-x		+x	+x
E	*$8.182 \times 10^{-2} - x$*		*$9.091 \times 10^{-3} + x$*	*x*

$pH = pK_a + Log([Ac^-]/[HAc])$

$pH = 4.754 + Log\ [(9.091 \times 10^{-3} + x)/(8.182 \times 10^{-2} - x)]$

<u>Assume:</u> $x \ll 10^{-2}$

$pH = 4.754 + Log(9.091 \times 10^{-3}/8.182 \times 10^{-2}) = 3.7998 =$ ***<u>3.80</u>***

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25.00 mL NaOH added

It's a Buffer!

Convert to mmol:

- $25.00\ mL\ (0.100\ mol\ OH^-/L) = 2.50\ mmol\ OH^-$
- $50.00\ mL\ (0.100\ mol\ HAc/L) = 5.00\ mmol\ HAc$

$$HAc + OH^- \rightarrow Ac^- + H_2O$$

Complete reaction:

	HAc	OH⁻	Ac⁻
I	5.00 mmol	2.50 mmol	-
C	- 2.50 mmol	- 2.50 mmol	+2.50 mmol
F	*2.50 mmol*	-	*2.50 mmol*

$pH = pK_a + Log([Ac^-]/[HAc])$

$pH = pK_a + Log(1) = 4.754 + 0 =$ ***<u>4.75</u>***

What are we assuming when we do this?

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50.00 mL NaOH added

Equiv. Point

Convert to mmol:

- 50.00 mL (0.100 mol OH^-/L) = 5.00 mmol OH^-
- 50.00 mL (0.100 mol HAc/L) = 5.00 mmol HAc

$$HAc + OH^- \rightarrow Ac^- + H_2O$$

Complete reaction:

	HAc	OH^-	Ac^-
I	5.00 mmol	5.00 mmol	-
C	- 5.00 mmol	-5.00 mmol	+5.00 mmol
F	-	-	*5.00 mmol*

Convert to concentrations:

$$C_{Ac^-} = \frac{5.00\text{ mmol}}{100.00\text{ mL}} = 5.00 \times 10^{-2}\ M$$

Equilibrium: Weak Base

$$Ac^- + H_2O \leftrightarrows HAc + OH^-$$

	Ac^-	HAc	OH^-
I	$5.00 \times 10^{-2}\ M$	-	-
C	-x	+x	+x
E	$5.00 \times 10^{-2}\ M - x$	x	*x*

$K_b = K_a/K_w = 5.68 \times 10^{-10}$

$$K_b = \frac{x^2}{0.0500 - x}$$

$x = 5.33 \times 10^{-6} = [OH^-]$

pOH = *5.27*

pH = 14.00 - *5.27* = $\boxed{\mathbf{8.73}}$

60.00 mL NaOH added

Convert to mmol:

- 60.00 mL (0.100 mol OH^-/L) = 6.00 mmol OH^-
- 50.00 mL (0.100 mol HAc/L) = 5.00 mmol HAc

$$HAc + OH^- \rightarrow Ac^- + H_2O$$

Complete reaction:

	HAc	OH^-	Ac^-
I	5.00 mmol	6.00 mmol	-
C	- 5.00 mmol	-5.00 mmol	+5.00 mmol
F	-	*1.00 mmol*	*5.00 mmol*

Two bases - which one will control pH?

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Excess OH^- Rules!

$[OH^-] = [OH^-]_{NaOH} + [OH^-]_{Ac-} + [OH^-]_{H2O}$

$[OH^-] = \dfrac{1.00 \text{ mmol}}{110.00 \text{ mL}} + (<10^{-6}) + (<10^{-7})$

$[OH^-] = 9.091 \times 10^{-3}$ *M*

pOH = ***2.04***

pH = 14.00 - *2.04* = ***<u>11.96</u>***

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Equivalence Point Detection

- ***Inflection Point* of titration curve**
  - Plot *pH versus mLs titrant*
- **Indicator**

$$\underset{\text{acid form}}{HIn} \leftrightarrows H^+ + \underset{\text{base form}}{In^-}$$

  - Color change observed when: $[HIn] = [In^-]$
    (when $pH = pK_a$ of indicator)
  - Choose indicator so that:
    Indicator pK_a » Equiv Pt pH

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