

March 29, 2002

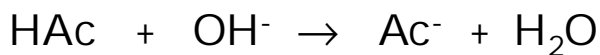
- No office hour today
- **Quiz Today!**

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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



Complete reaction:

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:

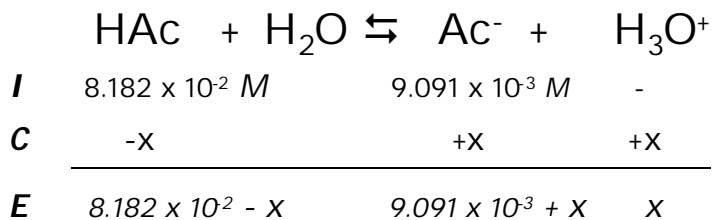
$$C_{\text{HAc}} = \frac{4.50 \text{ mmol}}{55.00 \text{ mL}} = 8.182 \times 10^{-2} \text{ M}$$

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

It's a Buffer!

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On to Equilibrium!



$$\text{pH} = \text{pK}_a + \text{Log}([\text{Ac}^-]/[\text{HAc}])$$

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

Assume: $x \ll 10^{-2}$

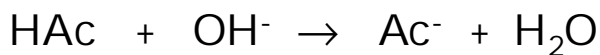
$$\text{pH} = 4.754 + \text{Log}(9.091 \times 10^{-3}/8.182 \times 10^{-2}) = 3.7998 = \underline{\underline{3.80}}$$

25.00 mL NaOH added

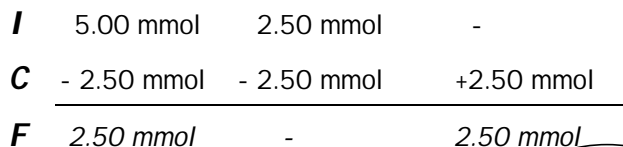
It's a Buffer!

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



Complete reaction:



$$\text{pH} = \text{pK}_a + \text{Log}([\text{Ac}^-]/[\text{HAc}])$$

$$\text{pH} = \text{pK}_a + \text{Log}(1) = 4.754 + 0 = \underline{\underline{4.75}}$$

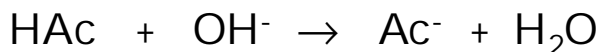
What are we assuming when we do this?

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



Complete
reaction:

I	5.00 mmol	5.00 mmol	-
C	- 5.00 mmol	- 5.00 mmol	+ 5.00 mmol
F	-	-	5.00 mmol

Convert to
concentrations:

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

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Equilibrium: Weak Base



I	5.00 x 10 ⁻² M	-	-
C	-x	+x	+x
E	5.00 x 10 ⁻² M - x	x	x

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

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

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

$$\text{pOH} = 5.27$$

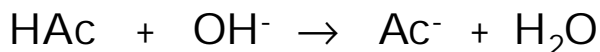
$$\text{pH} = 14.00 - 5.27 = \underline{\underline{8.73}}$$

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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



Complete reaction:

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!

$$[\text{OH}^-] = [\text{OH}^-]_{\text{NaOH}} + [\text{OH}^-]_{\text{Ac}^-} + [\text{OH}^-]_{\text{H}_2\text{O}}$$

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

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

$$\text{pOH} = \mathbf{2.04}$$

$$\text{pH} = 14.00 - 2.04 = \mathbf{11.96}$$

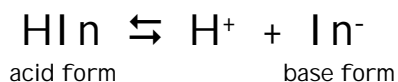
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Equivalence Point Detection

➤ Inflection Point of titration curve

- Plot *pH versus mLs titrant*

➤ Indicator



- Color change observed when: $[\text{HIn}] = [\text{In}^-]$
(when $\text{pH} = \text{pK}_a$ of indicator)
- Choose indicator so that:
Indicator $\text{pK}_a \gg \text{Equiv Pt pH}$