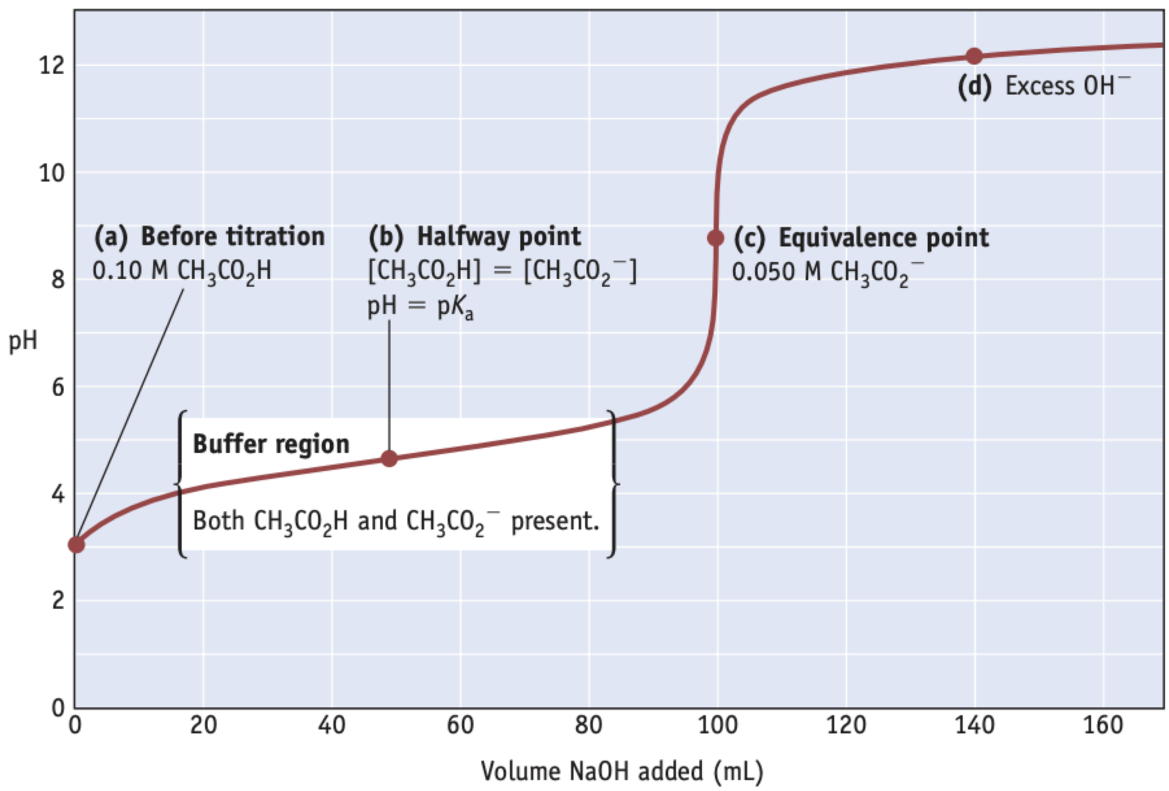
Extra Titration Problems Answer Key
7/25/24

1) Consider the titration of 100.0 mL of 0.100 M acetic acid (K_a = 1.8 × 10^{-5}) with 0.100 M NaOH.

$$CH_3CO_2H(aq) + OH^-(aq) \rightarrow CH_3CO_2^-(aq) + H2O(\ell)$$



a) What is the pH of the solution when 90.0 mL of 0.100 M NaOH has been added to 100.0 mL of 0.100 M acetic acid?

Step 1. First calculate the amounts of reactants before reaction (= concentration × volume) and then use the principles of stoichiometry to calculate the amounts of reactants and products after reaction. The limiting reactant is NaOH, so some CH_3CO_2H remains along with the product, $CH_3CO_2^-$.

Equation	CH_3CO_2H	+	OH^-	$\rightleftharpoons$	$CH_3CO_2^-$	+	H_2O
Initial (mol)	0.01000		0.009000		0		
Change (mol)	−0.009000		−0.009000		+0.009000		
After rxn (mol)	0.00100		0		0.009000		

Step 2. The ratio of amounts (moles) of acid to conjugate base is the same as the ratio of their concentrations. Therefore, you can use the amounts of weak acid remaining and conjugate base formed to find the pH from Equation 17.3.

$$[H_3O^+] = \frac{\text{mol } CH_3CO_2H}{\text{mol } CH_3CO_2^-} \times K_a = \left(\frac{0.00100 \text{ mol}}{0.009000 \text{ mol}}\right)(1.8 \times 10^{-5}) = 2.00 \times 10^{-6} \text{ M}$$

$$pH = -\log(2.00 \times 10^{-6}) = 5.70$$

b) What is the pH at the equivalence point?

Step 1. To reach the equivalence point, 0.0100 mol of NaOH was added to 0.0100 mol of CH_3CO_2H and 0.0100 mol of $CH_3CO_2^-$ was formed.

Equation	CH_3CO_2H	+	OH^-	$\rightarrow$	$CH_3CO_2^-$	+	H_2O
Initial (mol)	0.01000		0.01000		0		
Change (mol)	−0.01000		−0.01000		+0.01000		
After rxn (mol)	0		0		0.01000		

Step 2. Combining the two solutions, each with a volume of 100.0 mL, results in a total volume of 200.0 mL, so the concentration of $CH_3CO_2^-$ at the equivalence point is (0.01000 mol/0.200 L) = 0.05000 M. Next, set up an ICE table for the reaction of this weak base with water,

Equation	$CH_3CO_2^-$	+	H_2O	$\rightleftharpoons$	CH_3CO_2H	+	OH^-
Initial (M)	0.05000				0		0
Change (M)	$-x$				$+x$		$+x$
Equilibrium (M)	$0.05000 - x$				x		x

and calculate the concentration of OH^- ion using K_b for the weak base.

$$K_b \text{ for } CH_3CO_2^- = \frac{K_w}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} = \frac{[CH_3CO_2H][OH^-]}{[CH_3CO_2^-]} = \frac{(x)(x)}{0.05000 - x}$$

Assume that x is small with respect to 0.05000 M,

$$x = [OH^-] = 5.27 \times 10^{-6} \text{ M and so the pOH} = 5.278$$

$$\text{pH} = 14.00 - 5.278 = 8.72$$

c) What is the pH after 110.0 mL of NaOH is added?

Step 1. Following the equivalence point, CH_3CO_2H is the limiting reactant. After the reaction, excess hydroxide ion remains as well as the product acetate ion.

Equation	CH_3CO_2H	+	OH^-	$\rightarrow$	$CH_3CO_2^-$	+	H_2O
Initial (mol)	0.01000		0.01100		0		
Change (mol)	−0.01000		−0.01000		+0.01000		
After rxn (mol)	0		0.00100		0.01000		

Step 2. The solution contains excess OH^- from the unused NaOH. Additional OH^- produced by $CH_3CO_2^-$ hydrolysis is very small [see part (b)], so the pH of the solution after the equivalence point is determined by the excess NaOH (in 210.0 mL of solution).

$$[OH^-] = 0.00100 \text{ mol}/0.2100 \text{ L} = 4.76 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log[OH^-] = -\log[4.76 \times 10^{-3}] = 2.322$$

$$\text{pH} = 14.00 - 2.322 = 11.68$$

2) Phenol, C_6H_5OH, is a weak organic acid. Suppose 0.515 g of the compound is dissolved in enough water to make 125 mL of solution. The resulting solution is titrated with 0.123 M NaOH. (Assume K_a for phenol = 1.3×10^{-10}).

$$C_6H_5OH(aq) + OH^-(aq) \rightleftharpoons C_6H_5O^-(aq) + H2O(\ell)$$

a) What is the pH of the original solution of phenol?

b) What are the concentrations of all of the following ions at the equivalence point: Na^+, H_3O^+, OH^-, and $C_6H_5O^-$?

c) What is the pH of the solution at the equivalence point?

a) $C_6H_5OH\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+(aq) + C_6H_5O^-\,(aq)$

	C_6H_5OH	H_3O^+	$C_6H_5O^-$
I (M)	0.044	0	0
C (M)	$-x$	$+x$	$+x$
E (M)	$0.044-x$	x	x

$$0.515\text{ g} \times \frac{1\text{ mol}}{94.11\text{ g}} \times \frac{1}{0.125\text{ L}} = 0.044\text{ M}$$

$$K_a = \frac{[H_3O^+][C_6H_5O^-]}{[C_6H_5OH]} \Rightarrow 1.3\times10^{-10} = \frac{x^2}{0.044}$$

$x = 2.4\times10^{-6}$ $\qquad [H_3O^+]_E = 2.4\times10^{-6}$ M

$pH = -\log(2.4\times10^{-6}) = 5.62$

b) $C_6H_5OH(aq) + OH^-(aq) \rightarrow C_6H_5O^-(aq) + H_2O\,(\ell)$

	C_6H_5OH	OH^-	$C_6H_5O^-$
I (moles)	5.48×10^{-3}	5.48×10^{-13}	0
C (moles)	-5.48×10^{-3}	-5.48×10^{-3}	$+5.48\times10^{-3}$
F (moles)	0	0	5.48×10^{-3}

$$5.48\times10^{-13}\text{ moles NaOH} \times \frac{1\text{ L}}{0.123\text{ moles}} = 0.0445\text{ L}$$

$$[C_6H_5O^-]_i = \frac{5.48\times10^{-3}\text{ moles}}{0.1695\text{ L}} = 0.0323\text{ M}$$

Total Volume = 0.125 L + 0.0445 L = 0.1695 L

$C_6H_5O^-\,(aq) + H_2O\,(\ell) \rightleftharpoons C_6H_5OH\,(aq) + OH^-\,(aq)$

	$C_6H_5O^-$	C_6H_5OH	OH^-
I (M)	0.0323	0	0
C (M)	$-x$	$+x$	$+x$
E (M)	$0.0323-x$	x	x

$$K_B = \frac{K_w}{K_A} = \frac{1 \times 10^{-14}}{1.3 \times 10^{-10}} = 7.7 \times 10^{-5}$$

$$7.7 \times 10^{-5} = \frac{x^2}{0.0323 - x}$$

$x = 0.00154$ (Using quadratic formula)

$x = 0.00158$ (Using approximation)

$x \approx 0.0015$

$$[Na^+] = \frac{5.48 \times 10^{-3} \text{ moles}}{0.1695 \text{ L}} = 0.0323 \text{ M}$$

$$[OH^-] = 0.0015 \text{ M}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{0.00154} = 6.5 \times 10^{-12} \text{ M}$$

$$[C_6H_5O^-] = 0.0323 \text{ M} - 0.00154 \text{ M} = 0.0308 \text{ M}$$

c)

$$pH = -\log(6.5 \times 10^{-12}) = 11.19$$

3) You require 36.78 mL of 0.0105 M HCl to reach the equivalence point in the titration of 25.0 mL of aqueous ammonia. (K_a of NH_4^+ = 5.6 x 10^{-10})

a) What was the concentration of NH_3 in the original ammonia solution?

b) What are the concentrations of H_3O^+, OH^-, and NH_4^+ at the equivalence point?

c) What is the pH of the solution at the equivalence point?

a)

$$36.78\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{0.0105\ \text{moles}}{1\ \text{L}} = 3.8619 \times 10^{-4}\ \text{moles HCl}$$

$$\frac{3.8619 \times 10^{-4}\ \text{moles } NH_3}{25 \times 10^{-3}\ \text{L}} = 0.0154\ \text{M}$$

$[NH_3] = 0.0154$ M

b)

	$NH_3(aq)$ +	H_3O^+ ⟶	$NH_4^+(aq)$ + $H_2O(l)$
I (moles)	3.8619×10^{-4}	3.8619×10^{-4}	0
C (moles)	-3.8619×10^{-4}	-3.8619×10^{-4}	$+3.8619 \times 10^{-4}$
F (moles)	0	0	3.8619×10^{-4}

$$[NH_4^+]_i = \frac{3.8619 \times 10^{-4}\ \text{moles}}{61.78 \times 10^{-3}\ \text{L}} = 6.25 \times 10^{-3}\ \text{M}$$

	$NH_4^+(aq) + H_2O(l)$ ⇌	$H_3O^+(aq)$ +	$NH_3(aq)$
I (M)	6.25×10^{-3}	0	0
C (M)	$-x$	$+x$	$+x$
E (M)	$6.25 \times 10^{-3} - x$	x	x

$$5.6 \times 10^{-10} = \frac{x^2}{6.25 \times 10^{-3} - x} \qquad x = 1.9 \times 10^{-6}$$

$[H_3O^+] = 1.9 \times 10^{-6}$ M

$$[OH^-] = \frac{1 \times 10^{-14}}{1.9 \times 10^{-6}} = 5.3 \times 10^{-9}\ \text{M}$$

$[NH_4^+] = 6.25 \times 10^{-3} - 1.9 \times 10^{-6} = 6.25 \times 10^{-3}$ M

c) $pH = -\log(1.9 \times 10^{-6})$

pH = 5.72