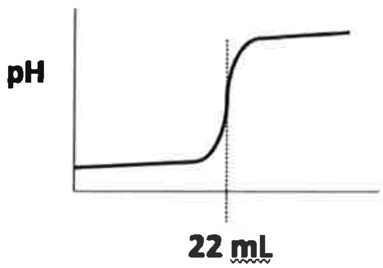
Chapter 16 Number 2 (16.5 – 16.7) (Unit 3) 26 February 2018

1. This curve represents the titration of a strong acid with strong base. In this case, 25.00 mL of HNO_3 was titrated with 0.1184 M NaOH.



1a. How many moles of NaOH were added at the equivalence point?

$n_{OH} = M \cdot V = 0.1184 \frac{mol}{L} \times .022 L$

2.60×10^{-3} mol OH^-

1b. Suppose the 25.00 mL acid sample came from a very large container. What is the concentration of the acid in the container?

$[HX] = \frac{2.60 \times 10^{-3} \text{ mol } H_3O^+}{.025 L}$

$= 0.1042$ mol/L

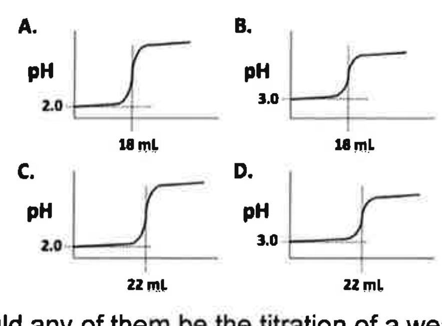
1c. What is the pH of the solution after 18.00 mL NaOH(aq) has been added?

25 mL
+ 18 mL
43 mL

n	H_3O^+	+ OH^-	$\longrightarrow$ 2 H_2O
I	2.60×10^{-3}	2.13×10^{-3} ← $n = MV$	
C	-2.13×10^{-3}	-2.13×10^{-3}	$= 0.1184 \times .018$
E	4.74×10^{-4}	~ 0	$= 2.13 \times 10^{-3}$ mol

$[H_3O^+] = 4.74 \times 10^{-4}$ mol$/0.043$ L → pH $= 1.96$

2. Compare these four titration curves.

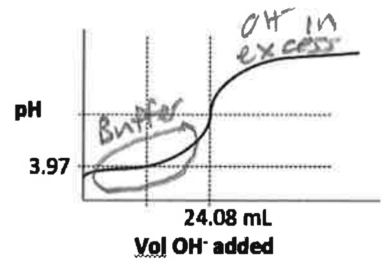


2a. Could any of them be the titration of a weak acid with OH^-? **Circle: Yes or No**

2b. Which two represent the least concentrated base? **Circle two: A B C D**

2c. Which two represent the least concentrated acid? **Circle two: A B C D**

3. This curve represents the titration of a weak acid with 0.1184 M NaOH. Suppose that 0.5050 g of a solid weak acid, HA, was dissolved in water to make 50.00 mL of solution and titrated to give the curve as shown.



3a. How many moles of NaOH were added at the equivalence point?

$n = MV$
$= 0.1184 \cdot 0.02408 = 2.85 \times 10^{-3}$ mol OH^-

3b. How many moles of HA were originally present?

2.85×10^{-3} mol HA

3c. What is the molar mass of the acid?

$MM = \frac{0.5050 g}{2.85 \times 10^{-3} mol} = 177$ g/mol

3d. What is the K_a of the acid?

$pK_a = 3.97$ from picture
$K_a = 10^{-3.97} = 1.07 \times 10^{-4}$

3e. Indicate the "buffer region" in the titration curve. ✓

3f. Indicate the region in which OH^- is in excess. ✓

3g. What is the pH of the acid solution before any OH^- was added? (Hint: Chapter 15 weak acid K_a calculation, MICE table, 400 Rule, etc.)

$[HA] = 2.85 \times 10^{-3}$ mol$/.05$ L
$= 5.7 \times 10^{-2}$ M

400 Rule = 533

M	HA	+ H_2O ⇌	H_3O^+	+ A^-
I	0.057			
C				
E	0.057-x		x	x

$x = 2.47 \times 10^{-3}$
pH = 2.61

Now try these problems from the book:
Section 16.5 – 16.7. Problems 14 – 17, 44, 46, 84 – 88 even, 92, and 94

Chapter 16 Number 1 (16.1 – 16.4) (Unit 3) 23 February 2018

1. Determine the numerical value for each of these equilibria, given K_a for HA = 2.2×10^{-5}. Add arrows of appropriate length to each equilibria. Add an appropriate subscript to K (as I did in (a). Your choices include: K_a, K_b, K_n, K_w, K_a, and K_c.

(a) $HA + H_2O \rightleftarrows H_3O^+ + A^-$ $K_a = 2.2\times10^{-5}$

(b) $HA + OH^- \rightleftarrows H_2O + A^-$ $K = \frac{1}{K_b} = \frac{K_a^{HA}}{K_w} = 2.2\times10^{9}$

(c) $H_3O^+ + OH^-$ $2\ H_2O$ $K = \frac{1}{K_w} = 1\times10^{+14}$

(d) $A^- + H_2O \rightleftarrows OH^- + HA$ $K = K_b = 4.5\times10^{-10}$

(e) $2\ H_2O \rightleftarrows H_3O^+ + OH^-$ $K = K_w = 1\times10^{-14}$

(f) $A^- + H_3O^+$ $H_2O + HA$ $K = \frac{1}{K_a} = 4.5\times10^{+4}$

2. Consider the two acids, HA and HB, for which $K_a^{HA} = 2.2 \times 10^{-5}$ and $K_a^{HB} = 4.1 \times 10^{-6}$. Calculate K for the reaction:

$$HA + B^- \rightleftharpoons A^- + HB$$

$HA + H_2O \rightleftarrows H_3O^+ + A^-$ 2.2×10^{-5}
$B^- + H_3O^+ \rightleftarrows H_2O + HB$ $1/4.1\times10^{-6}$
$HA + B^- \rightleftarrows A^- + HB$ $K = 5.4$

3. Calculate K_c for the reaction:

$$NH_4^+ + F^- \rightleftharpoons NH_3 + HF$$

K_b for NH_3 = 1.8×10^{-5} and K_a for HF = 3.5×10^{-4}

$F^- + H_3O^+ \rightleftarrows H_2O + HF$ $1/3.5\times10^{-4}$
$NH_4^+ + OH^- \rightleftarrows H_2O + NH_3$ $1/1.8\times10^{-5}$
$2H_2O \rightleftarrows H_3O^+ + OH^-$ 1×10^{-14}
$NH_4^+ + F^- \rightleftarrows NH_3 + HF$ 1.6×10^{-6}

4a. What is the pH of a buffer prepared by dissolving 0.0100 mol $HC_2H_3O_2$ (pK_a = 4.74) and 0.0150 mol $C_2H_3O_2^-$ in enough water to dissolve the substances.

$pH = 4.74 + \log\left(\frac{0.015}{0.01}\right)$
$= 4.92$

4b. What is the pH of the previous buffer after 0.0030 mol NaOH were added?

$pH = 4.74 + \log\left(\frac{0.015+.003}{0.010-0.003}\right)$
$= 5.15$

4c. Instead of adding NaOH as in Question 4b, suppose 0.0022 mol HCl were added instead to the buffer described in Question 4a. What is the pH?

$pH = 4.74 + \log\left(\frac{0.015-.0022}{0.010+.0022}\right)$
$= 4.76$

5. What is the pH of a buffer prepared by reacting 0.0155 mol NH_4Cl with 0.0080 mol NaOH?

$pH = 9.26 + \log\left(\frac{0+0.0080}{0.0155-0.0080}\right)$
$= 9.29$
$K_b^{NH_3} = 1.8\times10^{-5}$
$pK_a^{NH_4^+} = 9.26$

6. How many grams of NaOH must be added to 500.0 mL 0.100 M $HC_2H_3O_2$ in order to prepare a buffer with a pH = 5.00? $pK_a = 4.74$

$5.00 = 4.74 + \log\left(\frac{0+x}{0.0500\text{ mol}-x}\right)$
$\frac{x}{0.05-x} = 1.82$ $x = 0.0323$ mol
$x = 1.82(0.05-x)$ $m = 1.29$ g NaOH

7. What will happen to the pH of any of the buffers in the previous problems if water is added? no change

8. What will happen to the pH of any of the buffers in the previous problems if strong acid is added? pH ↓

Now try these problems from the book:
Section 16.1 – 16.4. Problems 1 – 13, 40, 42, 50 – 82 (even) Note: (You can use HHBE for any of these that are buffers)