

Acid - Base Theory

I. Review of Acids & Bases

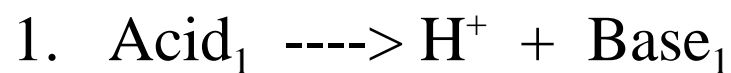
A. Characteristics of Acids

B. Characteristics of Bases

C. Three Concepts of Acids-Bases

<p>A R R H E N I U S</p>	<p>Acid</p> <p>Substance that yields H^+ (H_3O^+) in Solution.</p>	<p>Base</p> <p>Substance that Yields OH^- in Solution.</p>	<p>Neutralization</p> <p>$HCl + NaOH$ \downarrow is actually $H^+ + OH^- \rightarrow H_2O$</p>
<p>B R ö N S T E D</p>	<p>Proton Donor</p>	<p>Proton Acceptor</p>	<p>Transfer of proton from donor to acceptor</p>
<p>L E W I S</p>	<p>Electron Pair Acceptor</p>	<p>Electron Pair Donor</p>	<p>Formation of a Coordinate Covalent Bond</p>

D. Expansion of Brønsted-Lowry Concept



Hence:



II. pH

A. Def-

$$pH = -\log[H^+] = \log \frac{1}{[H^+]}$$

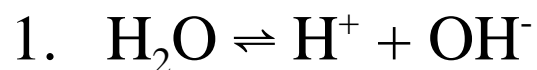
$$= -\log [H_3O^+]$$

B. Calculate pH for

[H⁺]

pH

C. Ionization of Water



$$K_{eq} = \frac{[H^+][OH^-]}{[H_2O]}$$

$$K_{eq} [H_2O] = [H^+][OH^-]$$

$$K_w =$$

2. At 25°C $[H^+] = [OH^-] = 1 \times 10^{-7}M$

$$K_w =$$

3. Hence from $K_w = [H^+][OH^-]$
 $[H^+] =$ $[OH^-] =$

4. Eg. What is the pH of 0.003 M NaOH?

5. Eq. What is the pH of a 0.0005 M $\text{Ba}(\text{OH})_2$ Solution?

6. Additional relationships resulting from the ionization of water

- a. Calculate the pH of a 0.00003333 m solution of $\text{Ce}(\text{OH})_3$ in water.

7. Solvents other than Water

III. Ionization of Weak Acids and Weak Bases

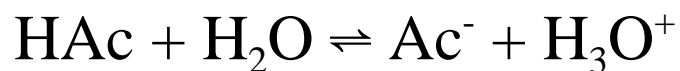
-Calculation of pH

A. No Common Ion

B. Common Ion present from external source

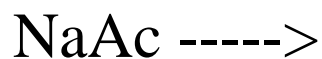
IV. Hydrolysis (or Rxn of Weak Acid's Conjugate Base or Weak Base's Conjugate Acid with water)

A. Concept



$$K_a = \underline{\hspace{2cm}}$$

NaAc in H₂O- Hydrolysis



$$K_b = K_h = \underline{\hspace{2cm}}$$

Note:

$$K_a \cdot K_h = \underline{\hspace{2cm}} \times \underline{\hspace{2cm}}$$

B. Other relationships derived from the concept

1. If $K_a \cdot K_h = K_w$

$$K_a =$$

$$K_h =$$

2. If $K_a \cdot K_h = K_w$

$$pK_a + pK_h = pK_w$$

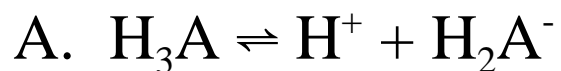
$$pK_a + pK_h = 14$$

C. Working of Hydrolysis Problem

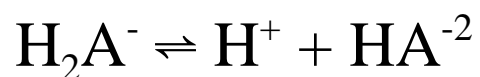
1. Recognize it as hydrolysis
2. Write hydrolysis Rxn
3. Find K_h & write expression for K_h
4. Use conditions in the problem to work out the equilibrium calculations

D. What is a solution's pH if it is 0.20 molar in NH_4Cl (aq) ? The K_b for NH_3 is 1.8×10^{-5} .

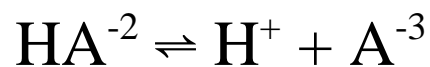
V. Ionization of Polyprotic Acids



$$K_{a1} = \frac{[\text{H}_2\text{A}^-][\text{H}^+]}{[\text{H}_3\text{A}]}$$



$$K_{a2} = \frac{[\text{HA}^{-2}][\text{H}^+]}{[\text{H}_2\text{A}^-]}$$



$$K_{a3} = \frac{[\text{A}^{-3}][\text{H}^+]}{[\text{HA}^{-2}]}$$

VI. Henderson-Hasselbach Equation- To calculate the pH of a solution containing a mixture of a Weak Acid and its conjugate base (salt) when the K_a (or pK_a) of the (weak) acid is known.

A. Acid Equation: $HA \rightleftharpoons H^+ + A^-$

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

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \frac{[Salt]}{[Acid]}$$

B. Base Equation: $B + H_2O \rightleftharpoons BH^+ + OH^-$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

$$pOH = pK_b + \log \frac{[BH^+]}{[B]}$$

$$pOH = pK_b + \log \frac{[Salt]}{[Base]}$$

$$pH = pK_w - pOH = pK_w - \left(pK_b + \log \frac{[BH^+]}{[B^+]} \right)$$

$$pH = 14 - pK_b - \log \frac{[BH^+]}{[B]}$$

- C. Suppose you have a mixture of $5.00 \times 10^{-5} \text{ M HA}_c$ and $1.00 \times 10^{-5} \text{ M Ac}^-$. Calculate the pH of the mixture. The K_a for HAc is 1.80×10^{-5} .

D. Note: A mixture of a weak acid and its conjugate base

or

Weak base and its conjugate acid is called a Buffer {resists all pH changes}

VII. Calculate pH during Titration

A. Approach

1. Calculate the # moles of acid & base present
2. See which is in excess
3. Use volume of solution to calculate the conc. of excess acid or base
4. Use concentration of excess & whether it is strong or weak to calc. the pH at that point.
 - a. strong- calc. conc. of H^+ or OH^- (may need to use $K_w = [H^+][OH^-]$)
 - b. Weak - Equil Calculation
5. At Equivalence Point- it may be a hydrolysis calculation assuming all the weak acid (base) is present as its salt- Calc. the conc. of the salt and let x moles/l hydrolyze

6. Beyond the Equivalence Point-
Calculate Concent. of (excess)
strong acid or base

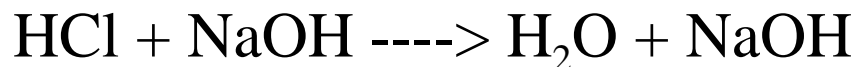
B. Summary

1. Before EP:
2. At EP:
3. Beyond EP:

Titration Curve Calculations

Calculate and graph the titration curve obtained when 50.00 ml of 0.1000 M HCl are titrated with 0.1000 M NaOH. Calculate the pH values when 0 ml, 10.00 ml, 20.00 ml, 30.00 ml, 40.00 ml, 45.00 ml, 48.00 ml, 49.00 ml, 49.99 ml, 50.00 ml, 50.02 ml, 50.06 ml, 50.20 ml, 50.90 ml, 52.00 ml, 55.00 ml, 60.00 ml, 70.00 ml, 80.00 ml, and 90.00 ml of 0.1000 M NaOH are added.

$$\# \text{ mmoles HCl} = 0.1000 \frac{\text{mmoles}}{\text{ml}} \times 50.00 \text{ ml} = 5.000 \text{ mmoles HCl}$$



a. $V_b = 0 \text{ ml}$
 $[\text{H}^+] = 0.1000 \text{ M} = 1 \times 10^{-1}$
 $\text{pH} = 1.0$

b. $V_b = 10.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 10.00 \text{ ml} = 1.000 \text{ mmoles}$
mmoles H^+ left = 4.000 mmoles
Total Vol = $50.00 \text{ ml} + 10.00 \text{ ml} = 60.00 \text{ ml}$
 $[\text{H}^+] = \frac{4.000 \text{ mmoles}}{60.00 \text{ ml}} = .06667 \text{ M}$
 $\text{pH} = -\log (6.667 \times 10^{-2}) = 1.18$

c. $V_b = 20.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 20.00 \text{ ml} = 2.000 \text{ mmoles}$
mmoles H^+ left = $5.000 - 2.000 = 3.000 \text{ mmoles}$
Total Vol = $50.00 \text{ ml} + 20.00 \text{ ml} = 70.00 \text{ ml}$
 $[\text{H}^+] = \frac{3.000 \text{ mmoles}}{70.00 \text{ ml}} = .04286 \text{ M}$
 $\text{pH} = -\log (4.286 \times 10^{-2}) = 1.37$

d. $V_b = 50.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 30.00 \text{ ml} = 3.000 \text{ mmoles}$
mmoles H^+ left = $5.00 - 3.000 = 2.000 \text{ mmoles}$

$$\text{Total Vol} = 50.00 \text{ ml} + 30.00 \text{ ml} = 80.00 \text{ ml}$$

$$[\text{H}^+] = \frac{2.000 \text{ mmoles}}{80.00 \text{ ml}} = .02500$$

$$\text{pH} = -\log (2.500 \times 10^{-2}) = 1.60$$

e. $V_b = 40.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 40.00 \text{ ml} = 4.000 \text{ mmoles}$
mmoles H^+ left = $5.000 - 4.000 = 1.000 \text{ mmoles}$

$$\text{Total Vol} = 50.00 \text{ ml} + 40.00 \text{ ml} = 90.00 \text{ ml}$$

$$[\text{H}^+] = \frac{1.000 \text{ mmoles}}{90.00 \text{ ml}} = .01111 \text{ M}$$

$$\text{pH} = -\log (1.111 \times 10^{-2}) = 1.95$$

f. $V_b = 45.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 45.00$
 $= 4.500 \text{ mmoles}$
mmoles H^+ left = $5.000 - 4.500 =$
 0.500 mmoles
Total Vol = $50.00 \text{ ml} + 45.00 =$
 95.00 ml
 $[\text{H}^+] = \frac{0.500 \text{ mmoles}}{95.00 \text{ ml}} = .005263 \text{ M}$
 $\text{pH} = -\log (5.263 \times 10^{-3}) = 2.28$

g. $V_b = 48.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 48.00$
 $\text{ml} = 4.800 \text{ mmoles}$
mmoles H^+ left = $5.000 - 4.800 =$
 0.200 mmoles
Total Vol = $50.00 \text{ ml} + 48.00 \text{ ml} =$
 98.00 ml
 $[\text{H}^+] = \frac{0.200 \text{ mmoles}}{98.00 \text{ ml}} = .00204 \text{ M}$
 $\text{pH} = -\log (2.041 \times 10^{-3}) = 2.69$

- h. $V_b = 49.00 \text{ ml}$
mmoles NaOH = $.1000 \text{ M} \times 49.00$
= 4.900 mmoles
mmoles H^+ left = $5.000 - 4.900 =$
 0.100 mmoles
Total Vol = $50.00 \text{ ml} + 49.00 \text{ ml} =$
 99.00 ml

$$[\text{H}^+] = \frac{.100 \text{ mmoles}}{99.00 \text{ ml}} = .001010 \text{ M}$$

$$\text{pH} = -\log (1.0101 \times 10^{-3}) = 3.00$$

- i. $V_b = 49.99 \text{ ml}$
mmoles NaOH = $.1000 \text{ m} \times 49.99$
= 4.999 mmoles
mmoles H^+ left = $5.000 - 4.999 =$
 $.001 \text{ mmoles}$
Total Vol = $50.00 \text{ ml} + 49.99 \text{ ml} =$
 99.99 ml

$$[\text{H}^+] = \frac{0.001 \text{ mmoles}}{99.99 \text{ ml}} = 1 \times 10^{-5} \text{ M}$$

$$\text{pH} = -\log (1 \times 10^{-5}) = 5.00$$

j. $V_b = 50.00 \text{ ml}$
 # mmoles NaOH = $.1000 \text{ M} \times 50.00$
 $= 5.000 \text{ mmoles}$
 # mmoles H^+ left = $5.000 - 5.000 =$
 0.000 mmoles

Source of H^+ from $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$

$$1 \times 10^{-14} = K_w = [\text{H}^+] [\text{OH}^-]$$

$$x^2 = 1 \times 10^{-14}$$

$$x = [\text{H}^+] = 1 \times 10^{-7}$$

$$\text{pH} = -\log (1 \times 10^{-7}) = 7.00$$

k. $V_b = 50.02 \text{ ml}$
 # mmoles excess NaOH = $.1000 \text{ M} \times$
 $.02 \text{ (excess)} = .002 \text{ mmoles}$
 Total Vol = 100.02 ml

$$[\text{OH}^-] = \frac{.002 \text{ mmoles}}{100.02 \text{ ml}} = 1.990 \times 10^{-5} \text{ M}$$

$$\text{pOH} = -\log (1.990 \times 10^{-5}) = 4.70$$

$$\text{pH} = 14 - \text{pOH} = 9.30$$

l. $V_b = 50.06 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times .06 \text{ ml} = .006 \text{ mmoles}$
Total Vol = 100.06 ml
 $[OH^-] = \frac{.006 \text{ mmoles}}{100.06 \text{ ml}} = 5.99 \times 10^{-5} \text{ M}$
 $pOH = -\log (5.99 \times 10^{-5}) = 4.22$
 $pH = 14 - pOH = 9.78$

m. $V_b = 50.20 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times .20 = .020 \text{ mmoles}$
Total Vol = 100.20 ml
 $[OH^-] = \frac{.020 \text{ mmoles}}{100.20 \text{ ml}} = 1.996 \times 10^{-4} \text{ M}$
 $pOH = -\log (1.996 \times 10^{-4}) = 3.70$
 $pH = 14 - pOH = 10.30$

n. $V_b = 50.90 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times .90 \text{ ml} = .090 \text{ mmoles}$
Total Vol = 100.90 ml
 $[OH^-] = \frac{.090 \text{ mmoles}}{100.90 \text{ ml}} = 8.92 \times 10^{-4} \text{ M}$
 $pOH = -\log (8.92 \times 10^{-4}) = 3.05$
 $pH = 14 - pOH = 10.95$

o. $V_b = 52.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 2.00 \text{ ml} = .200 \text{ mmoles}$
Total Vol = 102.0 ml
 $[OH^-] = \frac{.200 \text{ mmoles}}{102.0 \text{ ml}} = 1.96 \times 10^{-3} \text{ M}$
 $pOH = -\log (1.96 \times 10^{-3}) = 2.71$
 $pH = 14 - pOH = 11.29$

p. $V_b = 55.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 5.00 \text{ ml} = .500 \text{ mmoles}$
Total Vol = 105.0 ml
 $[OH^-] = \frac{.500 \text{ mmoles}}{105.0 \text{ ml}} = 4.76 \times 10^{-3} \text{ M}$
 $pOH = -\log (4.76 \times 10^{-3}) = 2.32$
 $pH = 14 - pOH = 11.68$

q. $V_b = 60.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 10.00 \text{ ml} = 1.000 \text{ mmoles}$
Total Vol = 110.00 ml
 $[OH^-] = \frac{1.000 \text{ mmoles}}{110.00 \text{ ml}} = 9.09 \times 10^{-3} \text{ M}$
 $pOH = -\log (9.09 \times 10^{-3}) = 2.04$
 $pH = 14 - pOH = 11.96$

r. $V_b = 70.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 20.00 \text{ ml} = 2.000 \text{ mmoles}$
Total Vol = 120.00 ml
 $[OH^-] = \frac{2.000 \text{ mmoles}}{120.00 \text{ ml}} = 1.667 \times 10^{-2} \text{ M}$

$$pOH = -\log (1.667 \times 10^{-2}) = 1.78$$
$$pH = 14 - pOH = 12.22$$

s. $V_b = 80.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 30.00 \text{ ml} = 3.000 \text{ mmoles}$
Total Vol = 130.00 ml
 $[OH^-] = \frac{3.000 \text{ mmoles}}{130.00 \text{ ml}} = 2.31 \times 10^{-2} \text{ M}$

$$pOH = -\log (2.31 \times 10^{-2}) = 1.64$$
$$pH = 14 - pOH = 12.36$$

- t. $V_b = 90.00 \text{ ml}$
mmoles excess NaOH = $.1000 \text{ M} \times 40.00 \text{ ml} = 4.000 \text{ mmoles}$
Total Vol = 140.0 ml

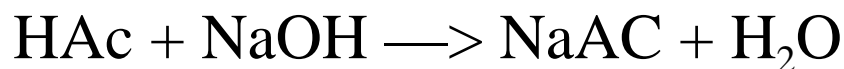
$$[OH^-] = \frac{4.000 \text{ mmoles}}{140.00 \text{ ml}} = 2.86 \times 10^{-2} \text{ M}$$

$$pOH = -\log(2.86 \times 10^{-2}) = 1.54$$

$$pH = 14 - pOH = 12.46$$

2. Calculate and graph titration curve obtained when 50.00 ml of 0.1000 m HAc are titrated with 0.1000 m NaOH. Calculate the pH values when 0.00 ml, 10.00 ml, 20.00 ml, 30.00 ml, 40.00 ml, 45.00 ml, 48.00 ml, 49.00 ml, 49.99 ml, 50.00 ml, 50.02 ml, 50.06 ml, 50.20 ml, 50.90 ml, 52.00 ml, 55.00 ml, 60.00 ml, 70.00 ml, 80.00 ml, 90.00 ml of 0.1000 m NaOH are added. $K_a = 1.8 \times 10^{-5}$.

mmoles HAc = 0.1000 mmoles/ml x 50.00 ml = 5.00 mmoles



a. $V_b = 0$ ml

$[\text{H}^+]$ due to breakdown of HAc



$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{Ac}^-]}{[\text{HA}]}$$

List x moles HAc react

$$[\text{H}^+] = x$$

$$[\text{Ac}^-] = x$$

$$[\text{HAc}] = .1000 - x \approx .1000$$

$$1.8 \times 10^{-5} = \frac{x^2}{.1000}$$

$$1.8 \times 10^{-6} = x^2$$

$$x = [\text{H}^+] = 1.34 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log(1.34 \times 10^{-3}) = 2.87$$

→ when base is added

$$[H^+] = K_a \frac{[HA]}{[Ac^-]}$$

b. $V_b = 10.00$ ml NaOH

mmoles NaOH added = $.1000$ M x
 10.00 ml = 1.000 mmoles

mmoles Ac^- produced = 1.000
mmoles

mmoles HAc left = $5.000 - 1.000 =$
 4.000 mmoles

Total Vol = $50.00 + 10.00 = 60.00$ ml

$$[HAc] = \frac{4.000 \text{ mmoles}}{60 \text{ ml}} = 6.667 \times 10^{-2} \text{ M}$$

$$[Ac^-] = \frac{1.000 \text{ mmoles}}{60 \text{ ml}} = 1.667 \times 10^{-2} \text{ M}$$

$$[H^+] = K_a \frac{[HAc]}{[Ac^-]} = (1.5 \times 10^{-5}) \frac{(6.667 \times 10^{-2})}{(1.667 \times 10^{-2})} = 7.20 \times 10^{-5}$$

$$\text{pH} = -\log (7.20 \times 10^{-5}) = 4.14$$

c. $V_b = 20.00 \text{ ml NaOH}$

$$\# \text{ mmoles NaOH added} = .1000 \text{ M} \times 20.00 \text{ ml} = 2.000 \text{ mmoles}$$

$$\# \text{ mmoles NaAc}^- \text{ produced} = 2.000 \text{ mmoles}$$

$$\# \text{ mmoles HAc left} = 5.000 - 2.000 = 3.000 \text{ mmoles}$$

$$\text{Total Vol} = 50.00 + 20.00 = 70.00 \text{ ml}$$

$$[\text{HAc}] = \frac{3.000 \text{ mmoles}}{70.00 \text{ ml}} = 4.28 \times 10^{-2} \text{ M}$$

$$[\text{NaAc}^-] = \frac{2.000 \text{ mmoles}}{70.00 \text{ ml}} = 2.86 \times 10^{-2} \text{ M}$$

$$[\text{H}^+] = (1.8 \times 10^{-5}) \frac{(4.28 \times 10^{-2})}{(2.80 \times 10^{-2})} = 2.69 \times 10^{-5}$$

$$\text{pH} = -\log (2.69 \times 10^{-5}) = 4.57$$

d. $V_b = 30.00 \text{ ml NaOH}$

mmoles NaOH added = $.1000 \text{ M} \times 30.00 \text{ ml} = 3.000 \text{ mmoles}$

mmoles Ac^- produced = 3.000 mmoles

mmoles HAc left = $5.000 - 3.000 = 2.000 \text{ mmoles}$

Total Vol = 80.00 ml

$$[\text{HAc}] = \frac{2.000 \text{ mmoles}}{80.00 \text{ ml}} = 2.50 \times 10^{-2} \text{ M}$$

$$[\text{Ac}^-] = \frac{3.000 \text{ mmoles}}{80.00 \text{ ml}} = 3.75 \times 10^{-2} \text{ M}$$

$$[\text{H}^+] = K_a \frac{[\text{HAc}]}{[\text{Ac}^-]} = (1.8 \times 10^{-5}) \frac{(2.50 \times 10^{-2})}{(3.75 \times 10^{-2})} = 1.20 \times 10^{-5}$$

$$\text{pH} = -\log [\text{H}^+] = -\log(1.20 \times 10^{-5}) = 4.92$$

e. $V_o = 40.00$ ml NaOH
mmoles NaOH added = $.1000$ M x
 40.00 ml = 4.000 mmoles

mmoles Ac^- produced = 4.000
mmoles

mmoles HAc left = $5.000 - 4.000 =$
 1.000 mmoles

Total Vol = 90.00 ml

$$[hAc] = \frac{1.000 \text{ mmoles}}{90.00 \text{ ml}} = 1.11 \times 10^{-2} \text{ M}$$

$$[2c^-] = \frac{4.000 \text{ mmoles}}{90.00 \text{ ml}} = 4.44 \times 10^{-2} \text{ M}$$

$$[H^+] = Xa \frac{[HA]}{[Ac]^-} = 1.8 \times 10^{-5} \frac{(1.11 \times 10^{-2})}{(4.44 \times 10^{-2})}$$

$$= 450 \times 10^{-5} = 4.50 \times 10^{-6}$$

$$\text{pH} = -\log (4.50 \times 10^{-6}) = 5.35$$

$$f. \quad V_b = 45.00 \text{ ml}$$

$$\# \text{ mmoles NaOH} = 4.500 \text{ mmoles}$$

$$\# \text{ mmoles NaAc} = 4.500 \text{ mmoles}$$

$$\# \text{ mmoles Hac} = 0.500 \text{ mmoles}$$

$$\text{Total Vol} = 95.0 \text{ ml}$$

$$[\text{HAc}] = 0.500 \text{ mmoles} / 95.00 \text{ ml} =$$

$$5.26 \times 10^{-3} = .526 \times 10^{-2}$$

$$[\text{Ac}^-] = 4.500 \text{ mmoles} / 95.00 \text{ ml} =$$
$$4.74 \times 10^{-2}$$

$$[\text{H}^+] = 1.8 \times 10^{-5} \frac{(.526 \times 10^{-2})}{(4.76 \times 10^{-2})}$$
$$= .200 \times 10^{-5} = 2.00 \times 10^{-6} \text{ M}$$

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

g. $V_b = 48.00 \text{ ml}$
mmoles NaOH = 4.800 mmoles

mmoles Ac^- = 4.800 mmoles

mmoles HAc = 0.200 mmoles

Total Vol = 98.00 ml

$[\text{HAc}] = 0.200 \text{ mmoles} / 98.00 \text{ ml} =$

$$2.04 \times 10^{-3} = .204 \times 10^{-2}$$

$[\text{Ac}^-] = 4.800 \text{ mmoles} / 98.00 \text{ ml} =$
 4.90×10^{-2}

$$[\text{H}^+] = 1.8 \times 10^{-5} \frac{(.204 \times 10^{-2})}{(4.90 \times 10^{-2})}$$

$$= 7.49 \times 10^{-7} \text{ M}$$

$$\text{pH} = -\log(7.49 \times 10^{-7}) = 6.12$$

h. $V_b = 49.00 \text{ ml}$

mmoles NaOH = 4.900 mmoles =

mmoles $\text{Ac}^- = 4.900 \text{ mmoles}$

mmoles HAc = 0.100 mmoles

Total Vol = 99.00 ml

$$[\text{HAc}] = 0.100 \text{ mmoles} / 99.00 \text{ ml} = .101 \times 10^{-2} \text{ M}$$

$$[\text{Ac}^-] = 4.900 \text{ mmoles} / 99.00 \text{ ml} = 4.95 \times 10^{-2} \text{ M}$$

$$\begin{aligned} [H^+] &= 1.8 \times 10^{-5} \frac{(.101 \times 10^{-2} \text{ M})}{(4.95 \times 10^{-2} \text{ M})} \\ &= 3.67 \times 10^{-7} \text{ M} \end{aligned}$$

$$\text{pH} = -\log(3.67 \times 10^{-7}) = 6.44$$

i. $V_b = 49.99 \text{ ml}$

$$\# \text{ mmoles NaOH} = 4.999 \text{ mmoles} =$$

$$\text{mmoles Ac}^- = 4.999 \text{ mmoles}$$

$$\# \text{ mmoles HAc} = 0.001 \text{ mmoles}$$

$$\text{Total Vol} = 99.99 \text{ ml}$$

$$[\text{HAc}] = 1 \times 10^{-3} \text{ mmoles} / 99.99 \text{ ml}$$

$$= 1.000 \times 10^{-5} = .001000 \times 10^{-2}$$

$$[\text{Ac}^-] = 4.999 \text{ mmoles} / 99.00 \text{ ml} =$$

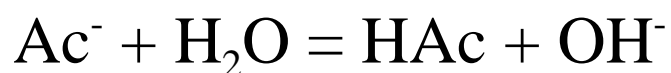
$$5.00 \times 10^{-2}$$

$$[H^+] = 1.8 \times 10^{-5} \frac{(.001000 \times 10^{-2})}{5.00 \times 10^{-2}} = 3.6 \times 10^{-9}$$

$$\text{pH} = -\log(3.6 \times 10^{-9}) = 8.44$$

j. $V_b = 50.00 \text{ ml}$

“All” HAc has been neutralized - The $[H^+]$ depends upon the hydrolysis of the Ac^- produced



$$K_h = \frac{K_w}{K_a} = \frac{[HAc][OH^-]}{[Ac^-]}$$

original $[Ac^-] = 5.000 \text{ mmoles} / 100\text{ml} = 0.05000 \text{ M}$

Let x react

$$[OH^-] = [HAc] = x$$

$$[Ac^-] = 0.05000 - x$$

Assume $0.05000 \gg x$

$$\text{So } [Ac^-] \approx 0.05000$$

$$\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{x^2}{5.00 \times 10^{-2}}$$

$$(.556 \times 10^{-9})(5.00 \times 10^{-2}) = x^2$$

$$2.78 \times 10^{-11} = x^2$$

$$27.8 \times 10^{-12} = x^2$$

$$x = [\text{OH}^-] = 5.27 \times 10^{-6}$$

$$\text{pOH} = -\log (5.27 \times 10^{-6}) = 5.28$$

$$\text{pH} = 14 - \text{pOH} = 8.72$$

After the equiv point ($V_b = 50.00$ ml) NaOH in excess and the pH is determined by the $[\text{OH}^-]$ excess. See calculations k) \rightarrow t) in the strong acid-strong base curve for these values.