*Chem 1515* Name____________________________________
Problem Set #13 TA Name________________________________
Fall 2001
Lab Section # ______

ALL work must be shown in all problems for full credit. **Due at the beginning of class on Wednesday, November 21, 2001**

PS13.1. Determine the pH for a solution containing the following substances† .

a) 0.550 M $HC_4H_7O_2$ and 0.550 M $NaC_4H_7O_2$

| | $HC_4H_7O_{2}(aq) \rightleftarrows$ | $H^+(aq)$ + | $C_4H_7O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.550 M** | **~ 0** | **0.550 M** | |
| **C** | **-x** | **+x** | **+x** | $x=[HC_4H_7O_2]_{diss}$ |
| **E** | **0.550 - x** | **+x** | **0.550 + x** | |

$$K_a = \frac{[C_4H_7O_2^-][H^+]}{[HC_4H_7O_2]}$$

$$1.5 \times 10^{-5} = \frac{(0.55 + x)(x)}{(0.55 - x)} \quad x < 0.55$$

$$1.5 \times 10^{-5} = \frac{0.55(x)}{0.55}$$

$$1.5 \times 10^{-5}\ M = x = [H^+] \quad pH = 4.82$$

b) 0.460 M $HC_4H_7O_2$ and 0.460 M $NaC_4H_7O_2$

| | $HC_4H_7O_{2}(aq) \rightleftarrows$ | $H^+(aq)$ + | $C_4H_7O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.460 M** | **~ 0** | **0.460 M** | |
| **C** | **-x** | **+x** | **+x** | $x=[HC_4H_7O_2]_{diss}$ |
| **E** | **0.460 - x** | **+x** | **0.460 + x** | |

$$K_a = \frac{[C_4H_7O_2^-][H^+]}{[HC_4H_7O_2]}$$

$$1.5 \times 10^{-5} = \frac{(0.46 + x)(x)}{(0.46 - x)} \quad x < 0.46$$

$$1.5 \times 10^{-5} = \frac{0.46(x)}{0.46}$$

$$1.5 \times 10^{-5}\ M = x = [H^+] \quad pH = 4.82$$

c) 0.300 M $NH_4Cl$ and 0.520 M $NH_3$

| | $NH_3(aq)$ + | $H_2O(l) \rightleftarrows$ | $NH_4^+(aq)$ + | $OH^-(aq)$ | |
|---|---|---|---|---|---|
| **I** | **0.520 M** | **-** | **0.300 M** | **~ 0** | |
| **C** | **-x** | **-** | **+x** | **+x** | $x=[NH_3]_{reacting}$ |
| **E** | **0.520-x** | **-** | **0.3+x** | **+x** | |

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{(0.3 + x)(x)}{0.520-x} \quad x < .3$$

$$1.8 \times 10^{-5} = \frac{(0.3)(x)}{.520}$$

$$3.12 \times 10^{-5} = x = [OH^-]$$

$$pOH = 4.51 \quad pH = 9.49$$

d) 0.300 M HCl and 0.890 M $HC_3H_5O_2$

**HCl is a strong acid and it will completely dissociate according to the equation**

| | $HCl_{(aq)}$ | $\rightleftarrows$ | $H^+_{(aq)}$ | + | $Cl^-_{(aq)}$ |
|---|---|---|---|---|---|
| I | 0.30 M | | 0 | | 0 |
| F | 0 | | 0.3 M | | 0.3 M |

**The $H^+$ ion formed from the dissociation of the HCl is used as the initial $[H^+]$ in the equilibrium reaction.**

| | $HC_3H_5O_{2(aq)}$ | $\rightleftarrows$ | $H^+_{(aq)}$ | + | $C_3H_5O_2^-{}_{(aq)}$ | |
|---|---|---|---|---|---|---|
| I | 0.890 M | | 0.3 M | | 0 | |
| C | -x | | +x | | +x | $x=[HC_3H_5O_2]_{diss}$ |
| E | 0.890 - x | | 0.3 + x | | 0 + x | |

$$K_a = \frac{[C_3H_5O_2^-][H^+]}{[HC_3H_5O_2]}$$

$$1.3 \times 10^{-5} = \frac{(0.3 + x)(x)}{(0.890 - x)} \qquad x <<< 0.2$$

$$1.3 \times 10^{-5} = \frac{0.3(x)}{0.890}$$

$$3.9 \times 10^{-5}\ M = x$$

$$[H^+] = 0.3 + x = 0.2\ M + 3.9 \times 10^{-5}\ M = 0.300\ M$$

$$pH = 0.52$$

**The amount of $H^+$ ion formed from the dissociation of the propionic acid is negligible compared to the amount of hydrogen ion from the HCl.**

e) 0.375 M $C_6H_5NH_3NO_3$ and 0.565 M $C_6H_5NH_2$

| | $C_6H_5NH_{2(aq)}$ | + $H_2O_{(l)}$ | $\rightleftarrows$ | $C_6H_5NH_3^+{}_{(aq)}$ | + $OH^-_{(aq)}$ | |
|---|---|---|---|---|---|---|
| I | 0.255 | - | | 0.835 | ~0 | |
| C | -x | - | | +x | +x | $x=[C_6H_5NH_2]_{reacting}$ |
| E | 0.255 - x | - | | 0.835 + x | +x | |

$$K_b = \frac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]}$$

$$4.3 \times 10^{-10} = \frac{(0.835 + x)(x)}{0.255 - x} \qquad x < .255$$

$$4.3 \times 10^{-10} = \frac{(0.835)(x)}{0.255}$$

$$1.31 \times 10^{-10} = x = [OH^-]$$

$$pOH = 9.88 \quad pH = 4.12$$

**Notice because the conjugate acid is stronger (has a larger equilibrium constant) than the base, the solution is acidic!**

PS13.2. Determine the magnitude of the equilibrium constant for the following reactions

a) $H^+(aq) + OH^-(aq) \rightleftarrows H_2O(l)$

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

b) $HNO_3(aq) + KOH(aq) \rightleftarrows H_2O(l) + KNO_3(aq)$ (Note: write the net ionic equation)

**$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$ ionic equ.**

**$H^+(aq) + OH^-(aq) \rightleftarrows H_2O(l)$ net ionic equ.**

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

c) $HOCl(aq) + NaOH(aq) \rightleftarrows H_2O(l) + NaOCl(aq)$

**$HOCl(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + OCl^-(aq)$ ionic equ.**

**$HOCl(aq) + OH^-(aq) \rightarrow H_2O(l) + OCl^-(aq)$ net ionic equ.**

**The above ionic equation can be separated into two equations whose K's are known.**

**$HOCl(aq) + \rightarrow H^+(aq) + OCl^-(aq)$ $K_a$**

**$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ $\frac{1}{K_w}$**

$$\mathbf{K = \frac{K_a(HOCl)}{K_w}}$$

$$\mathbf{K = \frac{3.0 \times 10^{-8}}{1.0 \times 10^{-14}} = 3.0 \times 10^6}$$

d) $CH_3NH_2(aq) + HClO_4(aq) \rightleftarrows CH_3NH_3ClO_4(aq)$

**$CH_3NH_2(aq) + H^+(aq) + ClO_4^-(aq) \rightarrow CH_3NH_3^+(aq) + ClO_4^-(aq)$ ionic equ.**

**$CH_3NH_2(aq) + H^+(aq) \rightarrow CH_3NH_3^+(aq)$ net ionic equation**

**The above ionic equation can be separated into two equations whose K's are known.**

**$CH_3NH_2(aq) + H_2O(l) \rightarrow CH_3NH_3^+(aq) + OH^-(aq)$ $K_b$**

**$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ $\frac{1}{K_w}$**

$$\mathbf{K = \frac{K_b}{K_w}}$$

$$\mathbf{K = \frac{4.4 \times 10^{-4}}{1.0 \times 10^{-14}} = 4.4 \times 10^{10}}$$

**Notice the magnitude for the equilibrium constant for any type of neutralization reaction is very large. Because of the large equilibrium constant, anytime an acid and base are mixed, the reaction proceeds to completion forming the product, salt and water.**

PS13.3. A titration is performed by adding 0.250 M KOH to 25.0 mL of 0.350 M $HNO_3$.

a) Calculate the pH before addition of any KOH.

**$HNO_3$ is a strong acid and it will complete dissociate according to the equation**

| | $HNO_3(aq)$ | $\rightarrow$ | $H^+(aq)$ | $+\ NO_3^-(aq)$ |
|---|---|---|---|---|
| **I** | **0.350 M** | | **0** | **0** |
| **F** | **0** | | **0.35** | **0.35** |

**$[H^+] = 0.350$ M**
**pH = 0.456**

b) Calculate the pH after the addition of 4.0, 18.0 and 34.0 mL of the base.(Show your work in detail for <u>one</u> of the volumes.)

**Add 4.0 mL of 0.250 M KOH**

$$4.0\text{ mL}\left(\frac{1\text{ L}}{1000\text{ mL}}\right)\left(\frac{0.250\text{ mol}}{1\text{ L}}\right) = 0.0010\text{ mol KOH}$$

$$25.0\text{ mL}\left(\frac{1\text{ L}}{1000\text{ mL}}\right)\left(\frac{0.350\text{ mol}}{1\text{ L}}\right) = 0.00875\text{ mol } HNO_3$$

| | $HNO_3(aq)$ | + $KOH(aq)$ $\rightarrow$ | $H_2O(l)$ | + $KNO_3(aq)$ |
|---|---|---|---|---|
| **initial** | **0.00875 mol** | **0.001 mol** | **–** | **0** |
| **final** | **0.00775 mol** | **0** | **–** | **0.001 mol** |

$$[HNO_3] = \frac{0.\ 00775\text{ mol}}{0.029\text{ L}} = 0.267\text{ M} \qquad \text{pH} = 0.573$$

**Add 18.0 mL of 0.250 M KOH**

| | $HNO_3(aq)$ | + $KOH(aq)$ $\rightarrow$ | $H_2O(l)$ | + $KNO_3(aq)$ |
|---|---|---|---|---|
| **initial** | **0.00875 mol** | **0.0045 mol** | **–** | **0** |
| **final** | **0.00425 mol** | **0** | **–** | **0.0045 mol** |

$$[HNO_3] = \frac{0.00425\text{ mol}}{0.043\text{ L}} = 9.88 \times 10^{-2}\text{ M}$$

$$\text{pH} = 1.00$$

**Add 34.0 mL of 0.250 M KOH**

| | $HNO_3(aq)$ | + $KOH(aq)$ $\rightarrow$ | $H_2O(l)$ | + $KNO_3(aq)$ |
|---|---|---|---|---|
| **initial** | **0.00875 mol** | **0.0085 mol** | **–** | **0** |
| **final** | **0.00025 mol** | **0** | **–** | **0.0085 mol** |

$$[HNO_3] = \frac{0.00025\text{ mol}}{0.059\text{ L}} = 4.24 \times 10^{-3}\text{ M}$$

$$\text{pH} = 2.37$$

c) Calculate the volume of base needed to reach the equivalence point.

At the equivalence point

$\mathbf{moles_{acid} = moles_{base}}$

$$\mathbf{moles_{acid}\left(\frac{volume_{acid}}{volume_{acid}}\right) = moles_{base}\left(\frac{volume_{base}}{volume_{base}}\right)}$$

$\mathbf{M_{acid}V_{acid} = M_{base}V_{base}}$

$\mathbf{0.350\ M \cdot 25.0\ mL = 0.250\ M \cdot V_{base}}$

$$\mathbf{\frac{0.350\ M \cdot 25.0\ mL}{0.250\ M} = V_{base} = 35.0\ mL}$$

d) Calculate the pH at the equivalence point.

**Add 35.0 mL of 0.250 M KOH**

| | $HNO_3$*(aq)* + | $KOH$*(aq)* → | $H_2O$*(l)* | + $KNO_3$*(aq)* |
|---|---|---|---|---|
| **initial** | **0.00875 mol** | **0.00875 mol** | **–** | **0** |
| **final** | **0** | **0** | **–** | **0.00875 mol** |

**$KNO_3$ is the salt of a strong acid and strong base, so the pH = 7.00 at the equivalence point.**

e) Calculate the pH after adding 5.00 mL of KOH past the endpoint.

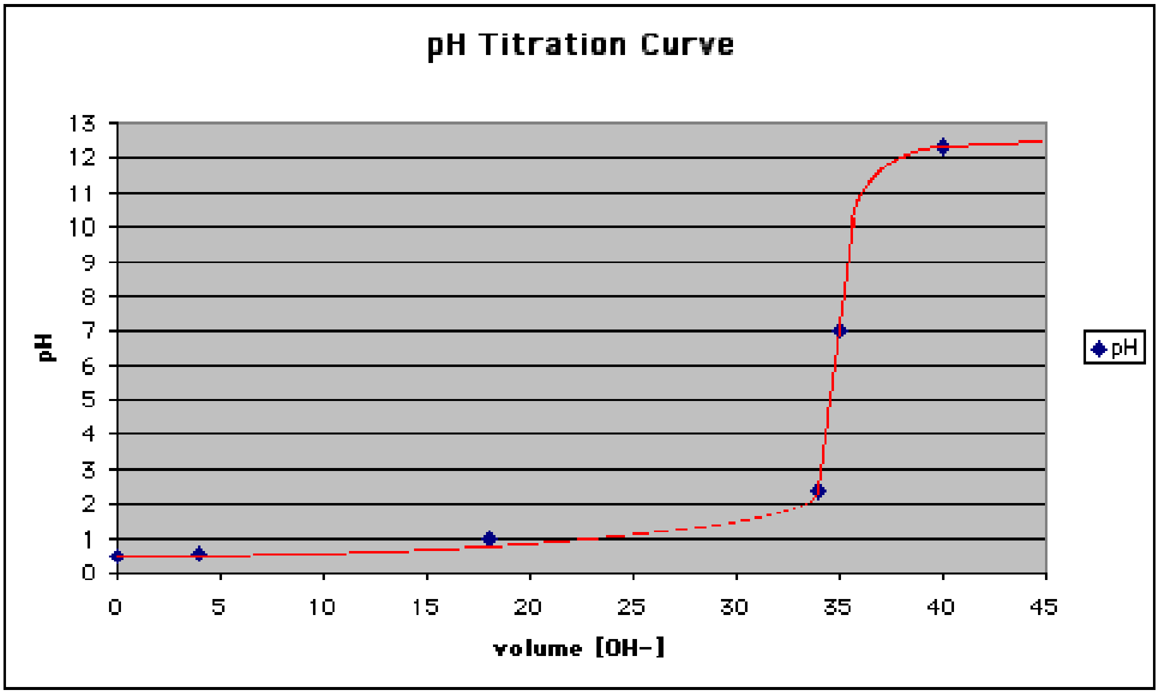
**Add 40.0 mL of 0.250 M KOH**

| | $HNO_3$*(aq)* + | $KOH$*(aq)* → | $H_2O$*(l)* | + $KNO_3$*(aq)* |
|---|---|---|---|---|
| **initial** | **0.00875 mol** | **0.0100 mol** | **–** | **0** |
| **final** | **0** | **0.00125 mol** | **–** | **0.00875 mol** |

$$\mathbf{[KOH] = \frac{0.00125\ mol}{0.065\ L} = 1.92 \times 10^{-2}\ M = [OH^-]}$$

**pOH = 1.72** **pH = 12.3**

f) Plot pH (y axis) versus volume of KOH added (x axis) for each calculation above. Sketch the titration curve.


pH Titration Curve
pH
13
12
11
10
9
8
7
6
5
4
3
2
1
0
0
5
10
15
20
25
30
35
40
45
volume [OH-]
pH

PS13.4. A titration is performed by adding 0.200 M NaOH to 30.0 mL of 0.275 M $HC_3H_5O_2$.

a) Calculate the pH before addition of any NaOH.

| | $HC_3H_5O_2$ | $\rightleftarrows$ | $H^+$ | + | $C_3H_5O_2^-$ | |
|---|---|---|---|---|---|---|
| **I** | **0.275 M** | | **~0** | | **0** | |
| **C** | **-x** | | **+x** | | **+x** | $x = [HC_3H_5O_2]_{diss}$ |
| **E** | **0.275-x** | | **x** | | **x** | |

$$K_a = \frac{[H^+][C_3H_5O_2^-]}{[HC_3H_5O_2]}$$

$$1.3 \times 10^{-5} = \frac{(x)(x)}{0.275 + x} \quad \text{assume } x << 0.275$$

$$1.3 \times 10^{-5} = \frac{(x)(x)}{0.275}$$

$$3.8 \times 10^{-6} = x^2$$

$$1.9 \times 10^{-3} = x = [H^+]$$

$$pH = 2.72$$

b) Calculate the pH after the addition of 1.5, 20.0, and 40.5 mL of the base. (Show your work in detail for <u>one</u> of the volumes.)

**Add 1.5 mL of NaOH**

$$[HC_3H_5O_2] = \frac{0.275 \text{ mol}}{L}(0.030 \text{ L}) = 0.00825 \text{ mol}$$

$$[NaOH] = \frac{0.200 \text{ mol}}{L}(0.0015 \text{ L}) = 0.0003 \text{ mol}$$

| | $HC_2H_3O_2(aq)$ + | $NaOH(aq)$ | $\rightleftarrows$ | $C_3H_5O_2^-(aq)$ | + | $H_2O(l)$ |
|---|---|---|---|---|---|---|
| **I** | **0.00825 mol** | **0.0003 mol** | | **0** | | **-** |
| **E** | **0.00795 mol** | **0** | | **0.0003 mol** | | **-** |

$$[HC_3H_5O_2] = \frac{0.00795 \text{ mol}}{0.0315 \text{ L}} = 0.252 \text{ M} \qquad [C_3H_5O_2^-] = \frac{0.0003 \text{ mol}}{0.0315 \text{ L}} = 0.00952 \text{ M}$$

| | $HC_3H_5O_2$ | $\rightleftarrows$ | $H^+$ | + | $C_3H_5O_2^-$ | |
|---|---|---|---|---|---|---|
| **I** | **0.252 M** | | **~0** | | **0.00952 M** | |
| **C** | **-x** | | **+x** | | **+x** | $x = [HC_3H_5O_2]_{diss}$ |
| **E** | **0.252 - x** | | **x** | | **0.00952 + x** | |

$$K_a = \frac{[H^+][C_3H_5O_2^-]}{[HC_3H_5O_2]}$$

$$1.3 \times 10^{-5} = \frac{(x)(0.00952 + x)}{0.252-x} \quad \text{assume } x << 0.00952$$

$$1.3 \times 10^{-5} = \frac{(x)(0.00952)}{0.252}$$

$$3.44 \times 10^{-4} = x = [H^+]$$

$$pH = 3.46$$

**Add 20.0 mL of NaOH**

$$[HC_3H_5O_2] = \frac{0.275 \text{ mol}}{L}(0.030 \text{ L}) = 0.00825 \text{ mol}$$

$$[NaOH] = \frac{0.200 \text{ mol}}{L}(0.020 \text{ L}) = 0.0040 \text{ mol}$$

| | $HC_3H_5O_2(aq)$ | $+ NaOH(aq)$ | $\rightleftarrows$ | $C_3H_5O_2^-(aq)$ + | $H_2O(l)$ |
|---|---|---|---|---|---|
| **I** | 0.00825 mol | 0.0040 mol | | 0 | - |
| **E** | 0.00425 mol | 0 | | 0.0040 mol | - |

$$[HC_3H_5O_2] = \frac{0.00425\ \text{mol}}{0.050\ \text{L}} = 0.085\ \text{M} \quad [C_3H_5O_2^-] = \frac{0.0040\ \text{mol}}{0.050\ \text{L}} = 0.080\ \text{M}$$

| | $HC_2H_3O_2$ | $\rightleftarrows$ | $H^+$ | + | $C_3H_5O_2^-$ | |
|---|---|---|---|---|---|---|
| **I** | 0.085 M | | ~0 | | 0.080 M | |
| **C** | -x | | +x | | +x | $x = [HC_3H_5O_2]_{diss}$ |
| **E** | 0.085 - x | | x | | 0.080 + x | |

$$K_a = \frac{[H^+][C_3H_5O_2^-]}{[HC_3H_5O_2]}$$

$$1.3 \times 10^{-5} = \frac{(x)(0.080 + x)}{0.085 - x} \qquad \text{assume } x << 0.0694$$

$$1.3 \times 10^{-5} = \frac{(x)(0.080)}{0.085}$$

$$1.38 \times 10^{-5} = x = [H^+]$$

$$pH = 4.86$$

**Add 40.5 mL of NaOH, the pH = 6.62**

c) Calculate the volume of base needed to reach the equivalence point.

$$\text{moles}_{acid} = \text{moles}_{base}$$

$$\text{moles}_{acid}\left(\frac{\text{volume}_{acid}}{\text{volume}_{acid}}\right) = \text{moles}_{base}\left(\frac{\text{volume}_{base}}{\text{volume}_{base}}\right)$$

$$M_{acid}V_{acid} = M_{base}V_{base}$$

$$0.275\ \text{M} \cdot 30.0\ \text{mL} = 0.200\ \text{M} \cdot V_{base}$$

$$\frac{0.275\ \text{M} \cdot 30.0\ \text{mL}}{0.200\ \text{M}} = V_{base} = 41.25\ \text{mL}$$

d) Calculate the pH at the equivalence point.

$$[HC_3H_5O_2] = \frac{0.275\ \text{mol}}{\text{L}}(0.030\ \text{L}) = 0.00825\ \text{mol}$$

$$[NaOH] = \frac{0.200\ \text{mol}}{\text{L}}(0.04125\ \text{L}) = 0.00825\ \text{mol}$$

| | $HC_3H_5O_2\ (aq)$ + | $NaOH(aq)$ | $\rightleftarrows$ | $C_3H_5O_2^-(aq)$ | $+ H_2O(l) + Na^+(aq)$ |
|---|---|---|---|---|---|
| **I** | 0.00825 mol | 0.00825 mol | | 0 | - |
| **E** | 0 | 0 | | 0.00825 mol | - |

$$[C_2H_3O_2^-] = \frac{0.00825\ \text{mol}}{0.07125\ \text{L}} = 0.116\ \text{M}$$

| | $H_2O$ + | $C_3H_5O_2^-$ | $\rightleftarrows$ | $OH^-$ + | $HC_3H_5O_2$ |
|---|---|---|---|---|---|
| **initial** | — | 0.116 | | 0 | 0 |
| **change** | — | –x | | +x | +x $x = [C_3H_5O_2^-]_R$ |
| **final** | — | 0.116–x | | +x | +x |

$$K_b = \frac{1 \times 10^{-14}}{1.3 \times 10^{-5}} = \frac{[HC_3H_5O_2][OH^-]}{[C_3H_5O_2^-]}$$

$$7.7 \times 10^{-10} = \frac{[x][x]}{[0.116 - x]} \qquad x<<<0.127$$

$$9.4 \times 10^{-6} = x = [OH^-]$$

$$pOH = 5.02 \qquad pH = 8.98$$

e) Calculate the pH after adding 5.00 mL of NaOH past the endpoint.

**Add 46.25 mL of NaOH**

$$[HC_3H_5O_2] = \frac{0.275 \text{ mol}}{\text{L}}(0.030 \text{ L}) = 0.00825 \text{ mol}$$

$$[NaOH] = \frac{0.200 \text{ mol}}{\text{L}}(0.04625 \text{ L}) = 0.00925 \text{ mol}$$

| | $HC_3H_5O_2\,(aq)$ + | $NaOH(aq)$ | $\rightleftarrows$ | $C_3H_5O_2^-(aq)$ + | $H_2O(l)$ + $Na^+(aq)$ |
|---|---|---|---|---|---|
| I | 0.00825 mol | 0.00925 mol | | 0 | - |
| E | 0 | 0.00100 | | 0.00825 mol | - |

$$[NaOH] = \frac{0.0010 \text{ mol}}{0.07625 \text{ L}} = 0.0131 \text{ M} \quad [C_2H_3O_2^-] = \frac{0.00825 \text{ mol}}{0.07625 \text{ L}} = 0.108 \text{ M}$$

| | $H_2O$ + | $C_3H_5O_2^-$ $\rightleftarrows$ | $OH^-$ + | $HC_3H_5O_2$ |
|---|---|---|---|---|
| initial | — | 0.108 | 0.0131 | 0 |
| change | — | –x | +x | +x  x = $[C_3H_5O_2^-]_R$ |
| final | — | 0.108–x | 0.0131+x | +x |

$$K_b = \frac{1 \times 10^{-14}}{1.3 \times 10^{-5}} = \frac{[HC_3H_5O_2][OH^-]}{[C_3H_5O_2^-]}$$

$$7.7 \times 10^{-10} = \frac{[x][.0131 + x]}{[0.108 - x]} \quad x<<<0.0131$$

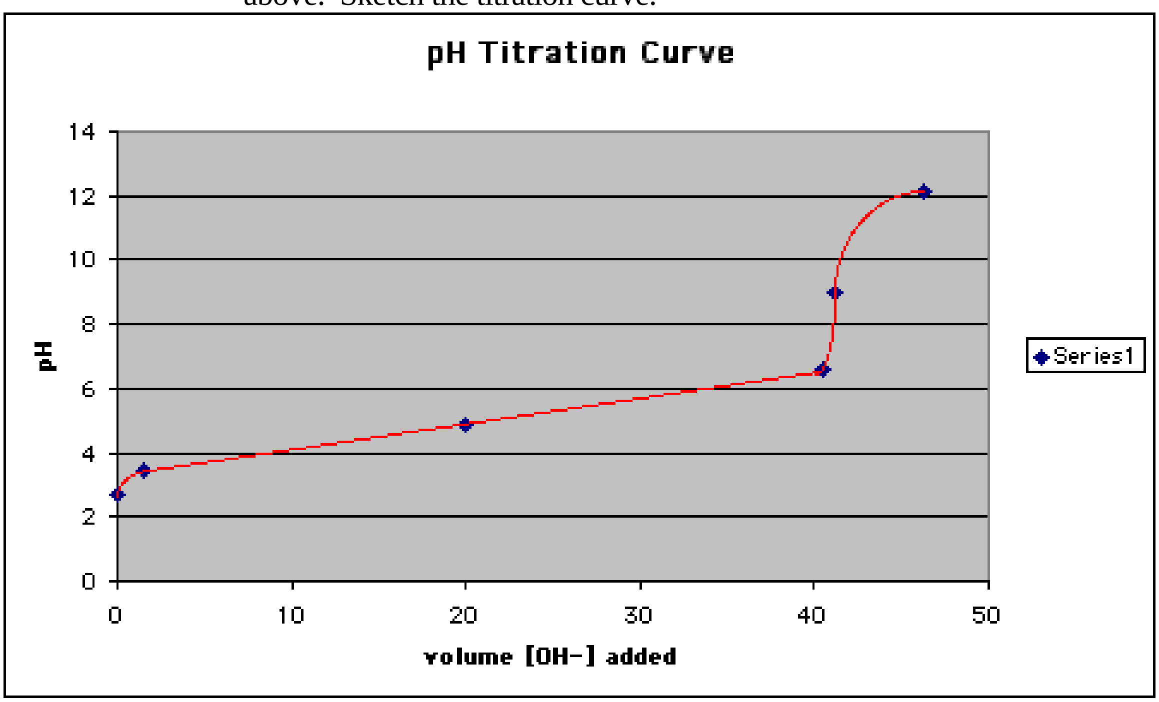
$$6.35 \times 10^{-9} = x$$

$$[OH^-] = 0.0131 + x$$

$$[OH^-] = 0.0131 + 6.35 \times 10^{-9} = \quad 0.0131 \text{ M}$$

$$pOH = 1.88 \text{ and } pH = 12.1$$

f) Plot pH (y axis) verses volume of NaOH added (x axis) for each calculation above. Sketch the titration curve.



PS13.5. Calculate the pH at the equivalence point when 35.0 mL of 0.160 M ethylamine, $CH_3CH_2NH_2$, is titrated with 0.120 M HBr

$$\mathbf{M_{acid}V_{acid} = M_{base}V_{base}}$$

$$\mathbf{0.120\ M \cdot V_{acid} = 0.160\ M \cdot 35.0\ mL}$$

$$\mathbf{V_{acid} = \frac{0.160\ M \cdot 35.0\ mL}{0.120\ M} = 46.7\ mL}$$

**46.7 mL of 0.120 M HBr is needed to reach the end point**

$$\mathbf{moles\ CH_3CH_2NH_2 = \frac{0.160\ mol}{L}(0.35\ L) = 0.0056\ mol}$$

$$\mathbf{moles\ HBr = \frac{0.120\ mol}{L}(0.0467\ L) = 0.0056\ mol}$$

| | $CH_3CH_2NH_2$*(aq)* + | HBr*(aq)* ⇄ | $CH_3CH_2NH_3Br$*(aq)* |
|---|---|---|---|
| **initial** | **0.0056** | **0.0056** | **0** |
| **final** | **0** | **0** | **0.0056** |

$$[\mathbf{CH_3CH_2NH_3}] = \mathbf{\frac{0.0056\ mol}{0.0817\ L} = 0.0686\ M}$$

| | $CH_3CH_2NH_3^+$*(aq)* ⇄ | $H^+$*(aq)* + | $CH_3CH_2NH_2$*(aq)* | |
|---|---|---|---|---|
| **initial** | **0.0686 M** | **0** | **0** | |
| **change** | **–x** | **+x** | **+x** | $\mathbf{x = [CH_3CH_2NH_3^+]_D}$ |
| **equilibrium** | **0.0686 – x** | **+x** | **+x** | |

$$\mathbf{K_a = \frac{1 \times 10^{-14}}{6.4 \times 10^{-4}} = \frac{[CH_3CH_2NH_2][H^+]}{[\ CH_3CH_2NH_3^+]}}$$

$$\mathbf{1.56 \times 10^{-11} = \frac{[x][x]}{[0.0686 - x]}} \quad \mathbf{x<<<0.0686}$$

$$\mathbf{1.0 \times 10^{-6} = x = [H^+]}$$

$$\mathbf{pH = 5.98}$$

PS13.6. Calculate the pH of a solution prepared by mixing
a) 25.0 mL of 0.512 M NaOH and 34.0 mL of 0.187 M HCl

$$\mathbf{moles\ NaOH = \frac{0.512\ mol}{L}(0.025\ L) = 0.0128\ mol}$$

$$\mathbf{moles\ HCl = \frac{0.187\ mol}{L}(0.034\ L) = 0.00636\ mol}$$

| | HCl*(aq)* + | NaOH*(aq)* ⇄ | $H_2O$*(aq)* + | NaCl*(aq)* |
|---|---|---|---|---|
| **initial** | **0.00636 mol** | **0.0128 mol** | **–** | **0** |
| **change** | **-0.00636** | **-0.00636** | **–** | **+0.00636** |
| **final 0.00143 mol** | **0** | **0.00644** | **–** | **0.00636 mol** |

$$\mathbf{[HCl] = \frac{0.00644\ mol}{0.059\ L} = 0.109\ M}$$

$$\mathbf{pH = 0.9623}$$

PS13.6. (CONTINUED)

b) 46.0 mL of 0.235 M KOH and 50.0 mL of 0.420 M $HC_3H_5O_2$

$$\text{moles } HC_3H_5O_2 = \frac{0.420 \text{ mol}}{L}(0.050 \text{ L}) = 0.0210 \text{ mol}$$

$$\text{moles KOH} = \frac{0.235 \text{ mol}}{L}(0.046 \text{ L}) = 0.0108 \text{ mol}$$

| | $HC_3H_5O_2(aq)$ + | $KOH(aq)$ ⇄ | $C_3H_5O_2^-(aq)$ + | $H_2O(l)$ |
|---|---|---|---|---|
| I | 0.0210 mol | 0.0108 mol | 0 | - |
| C | -0.0108 | -0.0108 | +0.0108 | - |
| F | 0.0102 mol | 0 | 0.0108 mol | - |

$$[HC_3H_5O_2] = \frac{0.0102 \text{ mol}}{0.096 \text{ L}} = 0.106 \text{ M} \quad [C_3H_5O_2^-] = \frac{0.0108 \text{ mol}}{0.096 \text{ L}} = 0.113 \text{ M}$$

| | $H_2O$ + | $C_3H_5O_2^-$ ⇄ | $OH^-$ + | $HC_3H_5O_2$ | |
|---|---|---|---|---|---|
| initial | — | 0.113 mol | ~0 | 0.106 mol | |
| change | — | –x | +x | +x | $x = [C_2H_3O_2^-]_R$ |
| final | — | 0.113 –x | +x | 0.106 +x | |

$$K_b = \frac{1 \times 10^{-14}}{1.3 \times 10^{-5}} = \frac{[HC_3H_5O_2][OH^-]}{[C_3H_5O_2^-]}$$

$$7.69 \times 10^{-10} = \frac{[x][0.106 + x]}{[0.113 - x]} \quad x<<<0.106$$

$$8.20 \times 10^{-10} = x = [OH^-]$$

$$pOH = 9.09$$

$$pH = 4.91$$

c) 400 mL of 0.250 M $NH_3$ and 250 mL of 0.120 M HCl

$$\text{moles } NH_3 = \frac{0.250 \text{ mol}}{L}(0.400 \text{ L}) = 0.100 \text{ mol}$$

$$\text{moles HCl} = \frac{0.120 \text{ mol}}{L}(0.250 \text{ L}) = 0.0300 \text{ mol}$$

| | $NH_3(aq)$ + | $HCl(aq)$ ⇄ | $NH_4Cl(aq)$ |
|---|---|---|---|
| initial | 0.10 | 0.03 | 0 |
| change | -0.03 | -0.03 | +0.03 |
| final | 0.07 | 0 | 0.03 |

$$[NH_3] = \frac{0.070 \text{ mol}}{0.650 \text{ L}} = 0.108 \text{ M} \quad [NH_4^+] = \frac{0.0300 \text{ mol}}{0.650 \text{ L}} = 0.0462 \text{ M}$$

| | $H_2O(l)$ + | $NH_3(aq)$ ⇄ | $OH^-(aq)$ + | $NH_4^+(aq)$ |
|---|---|---|---|---|
| initial | — | 0.108 | 0 | 0.0462 |
| change | — | –x | +x | +x |
| final | — | 0.108 –x | +x | 0.0462+x |

$$K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{[x][0.0462+x]}{[0.108–x]} \quad x<<<0.0154$$

$$4.2 \times 10^{-5} = x = [OH^-]$$

$$pOH = 4.38 \qquad pH = 9.62$$