*Chem 1515.001 - 006*
Problem Set #14
Fall 2001

Name________________________________
TA's Name ____________________________
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ALL work must be shown in all problems for full credit. **Due at the beginning of class on Wednesday, December 5, 2001.**

PS14.1. Phosphate buffers are commonly used in soft drinks to help control the pH. What is the pH of a soft drink in which the major buffer components are 8.50 g of $NaH_2PO_4$ and 9.23 g of $Na_2HPO_4$ per 355 mLs of solution?

$$\mathbf{8.50\ g\ NaH_2PO_4\left(\frac{1\ mol}{120\ g}\right)\left(\frac{1}{0.355\ L}\right) = 0.200\ M}$$

$$\mathbf{9.23\ g\ Na_2HPO_4\left(\frac{1\ mol}{142\ g}\right)\left(\frac{1}{0.355\ L}\right) = 0.183\ M}$$

| | $H_2PO_4^-(aq)$ | $\rightleftarrows$ | $H^+(aq)$ | + | $HPO_4^{2-}(aq)$ | |
|---|---|---|---|---|---|---|
| **I** | **0.200 M** | | **~0** | | **0.183 M** | |
| **C** | **-x** | | **+x** | | **+x** | $x=[HC_3H_5O_2]_{diss}$ |
| **E** | **0.200 - x** | | **+ x** | | **0.183 + x** | |

$$K_a = \frac{[HPO_4^{2-}][H^+]}{[H_2PO_4^-]}$$

$$6.2 \times 10^{-8} = \frac{(x)(0.183x)}{(0.200 - x)} \quad x <<< 0.183$$

$$6.2 \times 10^{-8} = \frac{(x)0.183}{0.200}$$

$$6.78 \times 10^{-8}\ M = x = [H^+]$$

$$pH = 7.17$$

PS14.2. Specify the reagents (an acid and its conjugate base or a base and its conjugate acid) and the concentration of each reagent needed to prepare buffer solutions having the listed pH values.

NOTE: The optimum buffer solution is one with equal concentrations of the weak acid (weak base) and its conjugate base (conjugate acid). Under these conditions, the pH of the solution is equal to the $pK_a$ ($pK_b$). So the best reagent for each of the solutions below is one whose pK is equal to the pH. Since the tables in the appendix list K values, each of the pH's must be converted to their corresponding $[H^+]$ and compared to an equilibrium constant in Appendix D.

a) 4.74

**pH = 4.74 : $[H^+] = 1.8 \times 10^{-5}$ Acetic acid has a $K_a = 1.8 \times 10^{-5}$; therefore, a solution of .85 M $HC_2H_3O_2$ and .85 M $NaC_2H_3O_2$ would have a pH of 4.74.**

b) 9.81

**pH = 9.81 : $[H^+] = 1.54 \times 10^{-10}$ M and $[OH^-] = 6.4 \times 10^{-5}$ M. Since the pH is basic, a weak base and its conjugate acid should be considered. Ethylamine has a $K_b = 6.4 \times 10^{-5}$; therefore, a solution of 0.70 M $(CH_3)_2N$ and . M $(CH_3)_2NHCl$ would have a pH of 9.81.**

c) 2.92

**pH = 2.92 : $[H^+] = 1.2 \times 10^{-3}$ Chloroacetic acid has a $K_a = 1.4 \times 10^{-3}$; therefore, a solution of .65 M $HC_2H_2O_2Cl$ and .76 M $NaC_2H_2O_2Cl$ would have a pH of 4.74.**

PS14.3. Write the reaction that occurs when a strong acid is added to each of the buffer solutions below.

a) $HC_2H_3O_2/NaC_2H_3O_2$

$$\mathbf{NaC_2H_3O_2}\textit{(aq)} + \mathbf{HCl}\textit{(aq)} \rightleftarrows \mathbf{HC_2H_3O_2}\textit{(aq)} + \mathbf{NaCl}\textit{(aq)}$$

**or**

$$\mathbf{C_2H_3O_2^-}\textit{(aq)} + \mathbf{H^+}\textit{(aq)} \rightleftarrows \mathbf{HC_2H_3O_2}\textit{(aq)}$$

d) $C_2H_5NH_3$/ $C_2H_5NH_3Cl$

$$\mathbf{CH_3CH_2NH_2}\textit{(aq)} + \mathbf{HCl}\textit{(aq)} \rightleftarrows \mathbf{CH_3CH_2NH_3^+}\textit{(aq)} + \mathbf{Cl^-}\textit{(aq)}$$

**or**

$$\mathbf{CH_3CH_2NH_2}\textit{(aq)} + \mathbf{H^+}\textit{(aq)} \rightleftarrows \mathbf{CH_3CH_2NH_3^+}\textit{(aq)}$$

c) HCN/KCN

$$\mathbf{CN^-}\textit{(aq)} + \mathbf{HCl}\textit{(aq)} \rightleftarrows \mathbf{HCN}\textit{(aq)} + \mathbf{Cl^-}\textit{(aq)}$$

**or**

$$\mathbf{CN^-}\textit{(aq)} + \mathbf{H^+}\textit{(aq)} \rightleftarrows \mathbf{HCN}\textit{(aq)}$$

PS14.4. Write the reaction that occurs when a strong base is added to each of the buffer solutions below.

a) $HC_2H_3O_2/NaC_2H_3O_2$

$$\mathbf{HC_2H_3O_2}\textit{(aq)} + \mathbf{NaOH}\textit{(aq)} \rightleftarrows \mathbf{NaC_2H_3O_2}\textit{(aq)} + \mathbf{H_2O}\textit{(l)}$$

**or**

$$\mathbf{HC_2H_3O_2}\textit{(aq)} + \mathbf{OH^-}\textit{(aq)} \rightleftarrows \mathbf{C_2H_3O_2^-}\textit{(aq)} + \mathbf{H_2O}\textit{(l)}$$

b) $C_2H_5NH_3$/ $C_2H_5NH_3Cl$

$$\mathbf{CH_3CH_2NH_3Cl}\textit{(aq)} + \mathbf{NaOH}\textit{(aq)} \rightleftarrows \mathbf{H_2O}\textit{(l)} + \mathbf{CH_3CH_2NH_2}\textit{(aq)} + \mathbf{NaCl}\textit{(aq)}$$

**or**

$$\mathbf{CH_3CH_2NH_3^+}\textit{(aq)} + \mathbf{OH^-}\textit{(aq)} \rightleftarrows \mathbf{H_2O}\textit{(l)} + \mathbf{CH_3CH_2NH_2}\textit{(aq)}$$

c) HCN/KCN

$$\mathbf{HCN}\textit{(aq)} + \mathbf{OH^-}\textit{(aq)} \rightleftarrows \mathbf{CN^-}\textit{(aq)} + \mathbf{H_2O}\textit{(l)}$$

**or**

$$\mathbf{HCN}\textit{(aq)} + \mathbf{NaOH}\textit{(aq)} \rightleftarrows \mathbf{NaCN}\textit{(aq)} + \mathbf{H_2O}\textit{(l)}$$

PS14.5. Determine the pH of a buffer prepared by mixing 0.675 moles of $HC_2H_3O_2$ and 0.575 moles of $NaC_2H_3O_2$ in enough water to give 1.00 liter of solution.

|   | $HC_2H_3O_{2(aq)}$ ⇄ | $H^+_{(aq)}$ + | $C_2H_3O_2^-{}_{(aq)}$ |   |
|---|---|---|---|---|
| **I** | **0.675 M** | $1 \times 10^{-7}$ | **0.575 M** |   |
| **C** | **-x** | **+x** | **+x** | $x=[HC_2H_3O_2]_{diss}$ |
| **E** | **0.675 - x** | $1 \times 10^{-7}+x$ | **0.575 + x** |   |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.575 + x)(1 \times 10^{-7} + x)}{(0.675 - x)} \qquad x < 0.475 \text{ and } x > 1 \times 10^{-7}$$

$$1.8 \times 10^{-5} = \frac{0.575(x)}{0.675}$$

$$2.1 \times 10^{-5}\ M = x = [H^+] \qquad pH = 4.68$$

Calculate the pH when

a) 0.0500 mol of HCl is added

|   | $H^+_{(aq)}$ + | $C_2H_3O_2^-{}_{(aq)}$ ⇄ | $HC_2H_3O_{2(aq)}$ |
|---|---|---|---|
| **I** | **.050 mol** | **0.575 mol** | **0.675 mol** |
| **F** | **0** | **0.525** | **0.725** |

|   | $HC_2H_3O_{2(aq)}$ ⇄ | $H^+_{(aq)}$ + | $C_2H_3O_2^-{}_{(aq)}$ |   |
|---|---|---|---|---|
| **I** | **0.725 M** | $1 \times 10^{-7}$ | **0.525 M** |   |
| **C** | **-x** | **+x** | **+x** | $x=[HC_2H_3O_2]_{diss}$ |
| **E** | **0.725 - x** | $1 \times 10^{-7}+x$ | **0.525 + x** |   |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.525 + x)(1 \times 10^{-7} + x)}{(0.725 - x)} \qquad x < 0.475 \text{ and } x > 1 \times 10^{-7}$$

$$1.8 \times 10^{-5} = \frac{0.525(x)}{0.725}$$

$$2.5 \times 10^{-5}\ M = x = [H^+] \qquad pH = 4.60$$

b) 0.0500 mol of NaOH is added

|   | $HC_2H_3O_{2(aq)}$ + | $OH^-_{(aq)}$ ⇄ | $C_2H_3O_2^-{}_{(aq)}$ + | $H_2O_{(l)}$ |
|---|---|---|---|---|
| **I** | **0.675** | **0.050** | **0.575** | **-** |
| **E** | **0.625** | **0** | **0.625** | **-** |

|   | $HC_2H_3O_{2(aq)}$ ⇄ | $H^+_{(aq)}$ + | $C_2H_3O_2^-{}_{(aq)}$ |   |
|---|---|---|---|---|
| **I** | **0.625 M** | $1 \times 10^{-7}$ | **0.625 M** |   |
| **C** | **-x** | **+x** | **+x** | $x=[HC_2H_3O_2]_{diss}$ |
| **E** | **0.625 - x** | $1 \times 10^{-7}+x$ | **0.625 + x** |   |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.625 + x)(1 \times 10^{-7} + x)}{(0.625 - x)} \qquad x < 0.625$$

$$1.8 \times 10^{-5} = \frac{0.625\ (x)}{0.625}$$

$$1.8 \times 10^{-5}\ M = x = [H^+] \qquad pH = 4.74$$

PS14.5. (CONTINUED)

c) 1 liter of water is added

$$\mathbf{M_1V_1 = M_1V_2}$$

$$\mathbf{0.675\ M \cdot 1\ L = M_1 \cdot 2\ L}$$

$$\mathbf{M_1 = \frac{0.675\ M \cdot 1\ L}{2\ L} = 0.338\ M\ HC_2H_3O_2}$$

$$\mathbf{M_1V_1 = M_1V_2}$$

$$\mathbf{0.575\ M \cdot 1\ L = M_1 \cdot 2\ L}$$

$$\mathbf{M_1 = \frac{0.575\ M \cdot 1\ L}{2\ L} = 0.288\ M\ C_2H_3O_2^-}$$

$$\mathbf{HC_2H_3O_2}\mathit{(aq)} \rightleftarrows \mathbf{H^+}\mathit{(aq)} + \mathbf{C_2H_3O_2^-}\mathit{(aq)}$$

| | | | | |
|---|---|---|---|---|
| **I** | **0.338 M** | **$1 \times 10^{-7}$** | **0.288 M** | |
| **C** | **-x** | **+x** | **+x** | **$x=[HC_2H_3O_2]_{diss}$** |
| **E** | **0.338 - x** | **$1 \times 10^{-7}$+x** | **0.288 + x** | |

$$\mathbf{K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}}$$

$$\mathbf{1.8 \times 10^{-5} = \frac{(0.288 + x)(1 \times 10^{-7} + x)}{(0.338 - x)}} \quad \mathbf{x < 0.288\ M}$$

$$\mathbf{1.8 \times 10^{-5} = \frac{0.288(x)}{0.338}}$$

$$\mathbf{2.1 \times 10^{-5}\ M = x = [H^+]} \qquad \mathbf{pH = 4.68}$$

**Dilution has no effect on the pH of the buffer. pH = 4.79**

d) 5.00 mL of a 1.00 M $HNO_3$ solution is added

$$\mathbf{\left(1.00\ \frac{mol}{L}\right) 0.005\ L = 0.005\ mol\ H^+}$$

$$\mathbf{H^+}\mathit{(aq)} + \mathbf{C_2H_3O_2^-}\mathit{(aq)} \rightleftarrows \mathbf{HC_2H_3O_2}\mathit{(aq)}$$

| | | | |
|---|---|---|---|
| **I** | **0.0050 mol** | **0.575 mol** | **0.675 mol** |
| **F** | **0** | **0.570** | **0.680** |

$$\mathbf{\frac{0.570\ mol}{1.005\ L} = 0.567\ M} \qquad \mathbf{\frac{0.680\ mol}{1.005\ L} = 0.677}$$

$$\mathbf{HC_2H_3O_2}\mathit{(aq)} \rightleftarrows \mathbf{H^+}\mathit{(aq)} + \mathbf{C_2H_3O_2^-}\mathit{(aq)}$$

| | | | | |
|---|---|---|---|---|
| **I** | **0.677 M** | **$1 \times 10^{-7}$** | **0.567 M** | |
| **C** | **-x** | **+x** | **+x** | **$x=[HC_2H_3O_2]_{diss}$** |
| **E** | **0.677 - x** | **$1 \times 10^{-7}$+x** | **0.567 + x** | |

$$\mathbf{K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}}$$

$$\mathbf{1.8 \times 10^{-5} = \frac{(0.567 + x)(1 \times 10^{-7} + x)}{(0.677 - x)}} \quad \mathbf{x < 0.570}$$

$$\mathbf{1.8 \times 10^{-5} = \frac{0.567(x)}{0.677}}$$

$$\mathbf{2.15 \times 10^{-5}\ M = x = [H^+]}$$

$$\mathbf{pH = 4.67}$$

PS14.6. Calculate the pH **change** produced when 0.100 mol of gaseous HCl is added to each of the following buffer solutions.

a) 500 mL of 0.900 M $NH_3$ and 0.900 M $NH_4Cl$

| | $NH_3(aq)$ + | $H_2O(l)$ ⇌ | $NH_4^+(aq)$ + | $OH^-(aq)$ | |
|---|---|---|---|---|---|
| **I** | **0.900 M** | **-** | **0.900 M** | **~0** | |
| **C** | **-x** | **-** | **+x** | **+x** | $x=[NH_3]_{reacting}$ |
| **E** | **0.900 - x** | **-** | **0.900 + x** | **+x** | |

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

$$1.8 \times 10^{-5} = \frac{(0.900 + x)(+x)}{0.900 - x} \qquad x < 0.90$$

$$1.8 \times 10^{-5} = \frac{(0.90)(x)}{0.90}$$

$$1.8 \times 10^{-5} = x = [OH^-]$$

$$pOH = 4.74 \quad pH = 9.26$$

$$0.900\,\frac{mol}{L} \cdot 0.500\ L = 0.450\ mol\ NH_3$$

$$0.900\,\frac{mol}{L} \cdot 0.500\ L = 0.450\ mol\ NH_4^+$$

| | $NH_3(aq)$ + | $H^+(aq)$ ⇌ | $NH_4^+(aq)$ |
|---|---|---|---|
| **I** | **0.45 mol** | **0.100 mol** | **0.45 mol** |
| **E** | **0.35 mol** | **-** | **0.55 mol** |

$$\frac{0.350\ mol}{0.500\ L} = 0.700\ M \qquad \frac{0.550\ mol}{0.500\ L} = 1.1\ M$$

| | $NH_3(aq)$ + | $H_2O(l)$ ⇌ | $NH_4^+(aq)$ + | $OH^-(aq)$ | |
|---|---|---|---|---|---|
| **I** | **0.70 mol** | **-** | **1.1 mol** | $1 \times 10^{-7}$ | |
| **C** | **-x** | **-** | **+x** | **+x** | $x=[NH_3]_{reacting}$ |
| **E** | **0.70 - x** | **-** | **1.1 + x** | $1 \times 10^{-7}+x$ | |

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

$$1.8 \times 10^{-5} = \frac{(1.1 + x)(1 \times 10^{-7} + x)}{0.70 - x} \qquad x < 0.35 \text{ and } x > 1 \times 10^{-7}$$

$$1.8 \times 10^{-5} = \frac{(1.1)(x)}{0.70}$$

$$1.14 \times 10^{-5} = x = [OH^-]$$

$$pOH = 4.94 \quad pH = 9.06$$

**The pH change is 0.20 pH units.**

b) 500 mL of 0.200 M $NH_3$ and 0.800 M $NH_4Cl$

| | $NH_3(aq)$ | + $H_2O(l)$ | $\rightleftarrows$ $NH_4^+(aq)$ | + $OH^-(aq)$ | |
|---|---|---|---|---|---|
| **I** | **0.200 M** | **-** | **0.800 M** | **~0** | |
| **C** | **-x** | **-** | **+x** | **+x** | $x=[NH_3]_{reacting}$ |
| **E** | **0.200 - x** | **-** | **0.800 + x** | **+x** | |

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

$$1.8 \times 10^{-5} = \frac{(0.800 + x)(1 \times 10^{-7} + x)}{0.200 - x} \quad x < 0.20$$

$$1.8 \times 10^{-5} = \frac{(0.80)(x)}{0.20}$$

$$4.5 \times 10^{-6} = x = [OH^-]$$

$$pOH = 5.35 \quad pH = 8.65$$

$$0.200\ \frac{mol}{L} \cdot 0.500\ L = 0.100\ mol\ NH_3$$

$$0.800\ \frac{mol}{L} \cdot 0.500\ L = 0.400\ mol\ NH_4^+$$

| | $NH_3(aq)$ | + $H^+(aq)$ | $\rightleftarrows$ $NH_4^+(aq)$ |
|---|---|---|---|
| **I** | **0.10 mol** | **0.100 mol** | **0.40 mol** |
| **E** | **0** | **-** | **0.50 mol** |

$$\frac{0.50\ mol\ NH_4^+}{0.500\ L} = 1.00\ M$$

| | $NH_4^+(aq)$ | $\rightleftarrows$ $NH_3(aq)$ | + $H^+(aq)$ | |
|---|---|---|---|---|
| **I** | **1.00 M** | **0** | **~0** | |
| **C** | **-x** | **+x** | **+x** | $x=[NH_4^+]_{reacting}$ |
| **E** | **1.00 - x** | **+x** | **+x** | |

$$K_a = \frac{K_w}{K_b} = \frac{[NH_3][H^+]}{[NH_4^+]}$$

$$\frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{(x)(x)}{1.00 - x} \qquad x < 1.00$$

$$5.56 \times 10^{-10} = \frac{(x)(x)}{1.00}$$

$$5.56 \times 10^{-10} = x^2$$

$$2.36 \times 10^{-5} = x = [H^+]$$

$$pH = 4.63$$

**The pH change is 4.02 pH units.**

c) 500 mL of 0.100 M $NH_3$ and 0.900 M $NH_4Cl$

**$NH_3(aq) + H_2O(l) \rightleftarrows NH_4^+(aq) + OH^-(aq)$**

| | $NH_3(aq)$ | $H_2O(l)$ | $NH_4^+(aq)$ | $OH^-(aq)$ | |
|---|---|---|---|---|---|
| **I** | **0.100** | **-** | **0.900** | **~0** | |
| **C** | **-x** | **-** | **+x** | **+x** | **$x=[NH_3]_{reacting}$** |
| **E** | **0.100 - x** | **-** | **0.900 + x** | **+x** | |

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

$$1.8 \times 10^{-5} = \frac{(0.900 + x)(1 \times 10^{-7} + x)}{0.100 - x} \qquad x < 0.10 \text{ and } x >> 1 \times 10^{-7}$$

$$1.8 \times 10^{-5} = \frac{(0.90)(x)}{0.10}$$

$$2.0 \times 10^{-6} = x = [OH^-]$$

$$pOH = 5.70 \quad pH = 8.30$$

**After the neutralization reaction, the solution contains 0.100 M HCl and 1.0 M $NH_4^+$. The pH of a solution of a strong acid, at a high concentration, and a weak acid is dominated by the strong acid. Therefore, the pH of the solution is;**

**pH = -log(0.100 M) = 1.00 pH change is 7.30 pH units.**

PS14.7. Calculate the pH **change** produced when 0.100 mol of solid NaOH is added to each of the following buffer solutions.

a) 500 mL of 0.600 M $HC_2H_3O_2$ and 0.600 M $NaC_2H_3O_2$

**$HC_2H_3O_2(aq) \rightleftarrows H^+(aq) + C_2H_3O_2^-(aq)$**

| | $HC_2H_3O_2(aq)$ | $H^+(aq)$ | $C_2H_3O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.600 M** | **~0** | **0.600 M** | |
| **C** | **-x** | **+x** | **+x** | **$x=[HC_2H_3O_2]_{diss}$** |
| **E** | **0.600 - x** | **+x** | **0.600 + x** | |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.600 + x)(x)}{(0.600 - x)} \qquad x < 0.600$$

$$1.8 \times 10^{-5} = \frac{0.600(x)}{0.600}$$

$$1.8 \times 10^{-5}\ M = x = [H^+]$$

$$pH = 4.74$$

**When the NaOH is added to the solution we must convert to moles of the weak acid and conjugate base.**

$$0.600 \frac{mol\ HC_2H_3O_2}{L} \cdot 0.500\ L = 0.300\ mol\ HC_2H_3O_2$$

$$0.600 \frac{mol\ HC_2H_3O_2}{L} \cdot 0.500\ L = 0.300\ mol\ C_2H_3O_2^-$$

**$HC_2H_3O_2(aq) + OH^-(aq) \rightleftarrows C_2H_3O_2^-(aq) + H_2O(l)$**

| | $HC_2H_3O_2(aq)$ | $OH^-(aq)$ | $C_2H_3O_2^-(aq)$ | $H_2O(l)$ |
|---|---|---|---|---|
| **I** | **0.300 mol** | **0.100 mol** | **0.300 mol** | **-** |
| **E** | **0.200 mol** | **0** | **0.400 mol** | **-** |

$$\frac{0.200\ mol}{0.500\ L} = 0.400\ M \qquad \frac{0.400\ mol}{0.500\ L} = 0.800\ M$$

| | $HC_2H_3O_2(aq) \rightleftarrows$ | $H^+(aq)$ + | $C_2H_3O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.200 mol** | **~0** | **0.400 mol** | |
| **C** | **-x** | **+x** | **+x** | **x= mol $HC_2H_3O_{2diss}$** |
| **E** | **0.200 - x** | **$1 \times 10^{-7}$+x** | **0.400 + x** | |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.400 + x)(1 \times 10^{-7} + x)}{(0.200 - x)} \quad x < 0.350$$

$$1.8 \times 10^{-5} = \frac{0.400(x)}{0.200}$$

$$9.0 \times 10^{-6}\ M = x = [H^+]$$

$$pH = 5.04$$

**The change is 0.300 pH units**

b) 500 mL of 0.400 M $HC_2H_3O_2$ and 0.600 M $NaC_2H_3O_2$

| | $HC_2H_3O_2(aq) \rightleftarrows$ | $H^+(aq)$ + | $C_2H_3O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.400 M** | **~0** | **0.600 M** | |
| **C** | **-x** | **+x** | **+x** | **$x=[HC_2H_3O_2]_{diss}$** |
| **E** | **0.400 - x** | **+x** | **0.600 + x** | |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.600 + x)(x)}{(0.400 - x)} \quad x < 0.600$$

$$1.8 \times 10^{-5} = \frac{0.600(x)}{0.400}$$

$$1.2 \times 10^{-5}\ M = x = [H^+]$$

$$pH = 4.92$$

**When the NaOH is added to the solution we must convert to moles of the weak acid and conjugate base.**

$$0.400 \frac{mol\ HC_2H_3O_2}{L} \cdot 0.500\ L = 0.200\ mol\ HC_2H_3O_2$$

$$0.600 \frac{mol\ HC_2H_3O_2}{L} \cdot 0.500\ L = 0.300\ mol\ C_2H_3O_2^-$$

| | $HC_2H_3O_2(aq)$ + | $OH^-(aq) \rightleftarrows$ | $C_2H_3O_2^-(aq)$ + | $H_2O(l)$ |
|---|---|---|---|---|
| **I** | **0.200 mol** | **0.100 mol** | **0.300 mol** | **-** |
| **E** | **0.100 mol** | **0** | **0.400 mol** | **-** |

$$\frac{0.100\ mol}{0.500\ L} = 0.200\ M \qquad \frac{0.400\ mol}{0.500\ L} = 0.800\ M$$

| | $HC_2H_3O_2(aq) \rightleftarrows$ | $H^+(aq)$ + | $C_2H_3O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.200 M** | **~0** | **0.800 M** | |
| **C** | **-x** | **+x** | **+x** | **x= mol $HC_2H_3O_{2diss}$** |
| **E** | **0.200 - x** | **+x** | **0.800 + x** | |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.800 + x)(1 \times 10^{-7} + x)}{(0.200 - x)} \quad x < 0.350$$

$$1.8 \times 10^{-5} = \frac{0.800(x)}{0.200}$$

$$4.5 \times 10^{-6}\ M = x = [H^+]$$

$$pH = 5.35$$

**The change is 0.43 pH units**

c) 500 mL of 0.100 M $HC_2H_3O_2$ and 0.700 M $NaC_2H_3O_2$

| | $HC_2H_3O_2(aq) \rightleftarrows$ | $H^+(aq)$ | $+\ C_2H_3O_2^-(aq)$ | |
|---|---|---|---|---|
| **I** | **0.100 M** | **~0** | **0.700 M** | |
| **C** | **-x** | **+x** | **+x** | $x = [HC_2H_3O_2]_{diss}$ |
| **E** | **0.100 - x** | **+x** | **0.700 + x** | |

$$K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.700 + x)(x)}{(0.100 - x)} \quad x < 0.600$$

$$1.8 \times 10^{-5} = \frac{0.700(x)}{0.100}$$

$$2.6 \times 10^{-6}\ M = x = [H^+]$$

$$pH = 5.59$$

**When the NaOH is added to the solution we must convert to moles of the weak acid and conjugate base.**

$$0.100\ \frac{\text{mol } HC_2H_3O_2}{L} \cdot 0.500\ L = 0.050\ \text{mol } HC_2H_3O_2$$

$$0.700\ \frac{\text{mol } HC_2H_3O_2}{L} \cdot 0.500\ L = 0.350\ \text{mol } C_2H_3O_2^-$$

| | $HC_2H_3O_2(aq)$ | $+\ OH^-(aq) \rightleftarrows$ | $C_2H_3O_2^-(aq)$ | $+\ H_2O(l)$ |
|---|---|---|---|---|
| **I** | **0.050 mol** | **0.100 mol** | **0.350 mol** | **-** |
| **E** | **0 mol** | **0.05 mol** | **0.400 mol** | **-** |

$$\frac{0.050\ \text{mol}}{0.500\ L} = 0.100\ M\ [OH^-] \qquad \frac{0.400\ \text{mol}}{0.500\ L} = 0.800\ M$$

| | $C_2H_3O_2^-(aq)$ | $+\ H_2O(l) \rightleftarrows$ | $HC_2H_3O_2(aq)$ | $+\ OH^-(aq)$ |
|---|---|---|---|---|
| **I** | **0.400 M** | **-** | **0** | **0.100 M** |
| **C** | **-x** | **-** | **+x** | **+x** |
| **E** | **0.400 - x** | **-** | **0** | **0.100 + x** |

$$K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{[HC_3H_5O_2][OH^-]}{[C_3H_5O_2^-]}$$

$$5.5 \times 10^{-10} = \frac{[x][.100 + x]}{[0.400 - x]} \quad x<<<0.100$$

$$2.2 \times 10^{-9} = x$$

$$[OH^-] = 0.100 + x = 0.100 + 2.2 \times 10^{-9} = 0.100\ M$$

$$pOH = 1.00 \text{ and } pH = 13$$

**The change is 7.41 pH units**