

# UNIT 3 - CHEMICAL SYSTEMS AND EQUILIBRIUM

- ① Factors that can affect the equilibrium of a reaction are temperature, enthalpy, and entropy.
- ② A  $\downarrow$  in T will shift the equilibrium to the right. The lower temperature relieves the heat produced in the exothermic reaction. A decrease in volume will have no effect on the equilibrium because there are the same numbers of molecules on both sides of the reaction.
- ③ An  $\uparrow$  in pressure will result in a decrease in volume. Since the reactant side of the equation has fewer molecules than the product side, the equilibrium will shift to the left.

④

	$2 \text{SO}_2(\text{g})$	$+\text{O}_2(\text{g})$	$\rightleftharpoons$	$2 \text{SO}_3(\text{g})$
I	1.7	1.7		0
C	$-2x$	$-x$	$+2x$	
E	$1.7 - 2x$	$1.7 - x$	$2x$	

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]}$$

$$K_c = 3.6 \times 10^{-3}$$

$$\frac{1.7}{3.6} \times 10^{-3} > 100$$

$$K_c = \frac{(2x)^2}{(1.7)^2 (1.7)}$$

$$\frac{(2x)^2}{(1.7)^3} = 3.6 \times 10^{-3}$$

$$x = 0.066$$

$$\begin{aligned} [\text{SO}_3] &= 2(0.066) \\ &= 0.13 \text{ mol/L} \end{aligned}$$

$$\textcircled{5} \quad [\text{CO}] = [\text{H}_2\text{O}] = 0.50 \text{ mol/L} \quad K_c = 8.3$$

	$\text{CO(g)} + \text{H}_2\text{O(g)} \rightleftharpoons \text{H}_2\text{(g)} + \text{CO}_2\text{(g)}$			
I	0.05	0.05	0	0
C	$-x$	$-x$	$+x$	$+x$
E	$0.05-x$	$0.05-x$	$x$	$x$

$$K_c = \frac{[\text{H}_2][\text{CO}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

$$8.3 = \frac{x^2}{(0.05-x)^2}$$

$$2.88 = \frac{x}{0.05-x}$$

$$x = 0.037$$

$$[\text{H}_2] = [\text{CO}_2] = 0.037 \text{ mol/L}$$

$$[\text{CO}] = [\text{H}_2\text{O}] = 0.05 - 0.037 \\ = 0.013 \text{ mol/L}$$

\textcircled{6} For experiment 1

$$K_c = \frac{[\text{SO}_3]^2}{[\text{O}_2][\text{SO}_2]^2} = \frac{(3.50)^2}{(1.25)(1.5)^2} \\ = 4.36$$

For experiment 2

$$K_c = \frac{[\text{SO}_3]^2}{[\text{O}_2][\text{SO}_2]^2} = \frac{(0.260)^2}{(0.0450)(0.590)^2} \\ = 4.32$$

Within experimental error, these values may be considered equal.

(7)

- a) The graph should show the concentration of  $H_2$  decreasing with time, and the concentration of  $HI$  increasing with time.
- b) The graph of  $[I_2]$  and  $[HI]$  with time should look similar to the graph in a. The concentrations of  $I_2$  and  $H_2$  should decrease at the same rate, and that of  $HI$  should increase.

(8)

- a)  $HCl$ ,  $H_2SO_4$   
 b)  $NaOH$ ,  $KOH$   
 c)  $NH_3$ ,  $C_5H_5N$

(9)

- a)  $O^{2-}$   
 b)  $NH_4^+$   
 c)  $H_2CO_3$   
 d)  $CO_3^{2-}$

(10)

- a)  $SO_4^{2-}(\text{aq})$   
 b)  $S^{2-}(\text{aq})$   
 c)  $HPO_4^{2-}(\text{aq})$   
 d)  $CO_3^{2-}(\text{aq})$

(11)

- $HCOOH$  is a weak acid; most of its molecules are not dissociated.  $NaHCOO$  is a good electrolyte; the concentration of  $HCOO^-$  is relatively high. If an acid is added to this buffer solution, the hydronium ions are removed by the  $HCOO^-$  ions. If a base is added, the  $OH^-$  ions react with the  $H_3O^+$  ions formed by the dissociation of  $HCOOH$ .

(21)

$$\textcircled{12} \quad \text{Initial } [\text{HAsc}] = \frac{5.0\text{g}}{(0.250\text{L})(176\text{g/mol})} = 0.11 \text{ mol/L}$$

	$\text{HAsc}(\text{aq})$	$\rightleftharpoons$	$\text{HAsc}^-(\text{aq})$	$\text{H}_3\text{O}^+(\text{aq})$
I	0.11		0	0
C	-x		+x	+x
E	$0.11 - x$		x	x

$$K_a = \frac{[\text{HAsc}^-][\text{H}_3\text{O}^+]}{[\text{HAsc}]}$$

$$= \frac{(x)(x)}{0.11 - x}$$

$$= 8.0 \times 10^{-5}$$

$$\boxed{\frac{0.11}{8.0 \times 10^{-5}} > 100}$$

$$\frac{x^2}{0.11} = 8.0 \times 10^{-5}$$

$$x = 3.0 \times 10^{-3} \text{ mol/L}$$

$$[\text{H}_3\text{O}^+] = 3.0 \times 10^{-3} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log(3.0 \times 10^{-3}) \\ &= 2.52 \end{aligned}$$

$$\textcircled{13} \quad \begin{aligned} V_{\text{HNO}_3} &= 25.0 \text{ mL} \\ [\text{HNO}_3] &= 1.40 \text{ mol/L} \\ V_{\text{NaOH}} &= 15.0 \text{ mL} \\ [\text{NaOH}] &= 2.00 \text{ mol/L} \end{aligned}$$



$$n = cV$$

$$n_{\text{HNO}_3} = 1.40 \frac{\text{mol}}{\text{L}} \times 0.0250 \text{ L}$$

$$= 0.0350 \text{ mol}$$

$$n_{\text{NaOH}} = 2.00 \text{ mol/L} \times 0.0150 \text{ L} \\ = 0.0300 \text{ mol}$$

ratio is 1:1; the amount of NaOH is less, so this is the limiting reactant

$$\text{excess HNO}_3(\text{aq}) = 0.0350 \text{ mol} - 0.0300 \text{ mol} \\ = 0.0050 \text{ mol}$$

$$\therefore [\text{H}_3\text{O}^+(\text{aq})] = 5.0 \times 10^{-3} \text{ mol}$$

$$\text{Total volume} = (25.0 + 15.0) \text{ mL} = 400 \text{ mL}$$

$$[\text{H}_3\text{O}^+] = \frac{5.0 \times 10^{-3} \text{ mol}}{0.400 \text{ L}} \\ = 0.0125 \text{ mol/L}$$

The solution is acidic.

(14) pH = 5.53

$$\text{pOH} = 14.00 - 5.53 \\ = 8.47$$

$$[\text{H}_3\text{O}^+] = 10^{-5.53} \\ = 3.0 \times 10^{-6} \text{ mol/L}$$

$$[\text{OH}^-] = 10^{-8.47} \\ = 3.4 \times 10^{-9} \text{ mol/L}$$

(15)

$$K_a = ? \quad \% \text{ dissociation} = ?$$

$$[\text{CH}_3\text{CH}_2\text{COOH}]_{\text{init.}} = 0.10 \text{ mol/L}$$

$$\text{pH} = 2.96$$

	$\text{CH}_3\text{CH}_2\text{COOH}(\text{aq})$	$\text{H}_2\text{O}(l)$	$\text{CH}_3\text{CH}_2\text{COO}^-(\text{aq})$	$\text{H}_3\text{O}^+(\text{aq})$
I	0.10		0	0
C	$-x$		$+x$	$+x$
E	$0.10 - x$		$x$	$x$

$$K_a = \frac{[\text{CH}_3\text{CH}_2\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CH}_2\text{COOH}]}$$

$$= \frac{(x)(x)}{(0.10 - x)}$$

$$[\text{H}_3\text{O}^+] = 10^{-2.96}$$

$$= 1.1 \times 10^{-3} \text{ mol/L}$$

$$K_a = \frac{(1.1 \times 10^{-3})^2}{0.10 - (1.1 \times 10^{-3})}$$

$$= 1.2 \times 10^{-5}$$

$$\% \text{ dissociation} = \frac{1.1 \times 10^{-3} \text{ mol/L}}{0.10 \text{ mol/L}} \times 100$$

$$= 1.1 \%$$

(24)

$$(16) \quad [\text{HCOOH}] = 0.025 \text{ mol/L}$$

$$K_a = 1.8 \times 10^{-4}$$

	$\text{HCOOH}_{(\text{aq})} + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCOO}^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$		
I	0.025	0	0
C	$-x$	$+x$	$+x$
E	$0.025 - x$	$x$	$x$

$$\begin{aligned} K_a &= \frac{[\text{HCOO}^-][\text{H}_3\text{O}^+]}{[\text{HCOOH}]} \\ &= \frac{(x)(x)}{0.025 - x} \\ &= 1.8 \times 10^{-4} \end{aligned}$$

$$\frac{0.025}{1.8 \times 10^{-4}} = 139$$

relatively close to 100; solve quadratic equation

$$\frac{x^2}{0.025 - x} = 1.8 \times 10^{-4}$$

$$x^2 + (1.8 \times 10^{-4})x - (4.5 \times 10^{-6}) = 0$$

$$\begin{aligned} x &= \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \\ &= \frac{-1.8 \times 10^{-4} \pm \sqrt{(1.8 \times 10^{-4})^2 - 4 \times 1 \times (-4.5 \times 10^{-6})}}{2} \end{aligned}$$

$$x = 0.0020 \quad \text{or} \quad x = -0.002$$

$$x = 0.0020 \text{ mol/L} \quad (= [\text{H}_3\text{O}^+] = [\text{HCOO}^-])$$

$$\begin{aligned} \text{pH} &= -\log 0.0020 \\ &= 2.70 \end{aligned}$$

$$(17) \quad K_b = ?$$

$$[C_5H_5N] = 0.125 \text{ mol/L}$$

$$pH = 9.10$$

	$C_5H_5N(aq)$	$+ H_2O(l)$	$\rightleftharpoons$	$C_5H_5NH^+(aq)$	$+ OH^-(aq)$
I	0.125			0	0
C	-x			+x	+x
E	$0.125 - x$			x	x

$$K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5N]}$$

$$= \frac{(x)(x)}{0.125 - x}$$

$$pOH = 14.00 - 9.10$$

$$= 4.90$$

$$[OH^-] = 10^{-4.90}$$

$$= 1.3 \times 10^{-5} \text{ mol/L}$$

$$(0.125 - (1.3 \times 10^{-5}) = 0.125)$$

$$K_b = \frac{(1.3 \times 10^{-5})^2}{0.125}$$

$$= 1.4 \times 10^{-9}$$

(18)  $k_b = ?$   
 $pH = ?$

$$K_a = 1.81 \times 10^{-5}$$

12.5 g of  $\text{CH}_3\text{COONa}$  is dissolved in 1.00 L of water

$$k_b = \frac{K_w}{K_a}$$

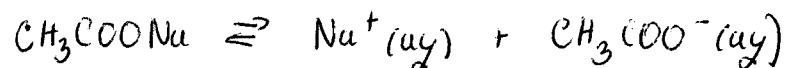
$$k_b = \frac{1.0 \times 10^{-14}}{1.81 \times 10^{-5}}$$

$$= 5.6 \times 10^{-10}$$

$$M(\text{CH}_3\text{COONa}) = 82.0 \text{ g/mol}$$

$$n = \frac{12.5 \text{ g}}{82.0 \text{ g/mol}}$$

$$= 0.152 \text{ mol/L}$$



$$\therefore [\text{CH}_3\text{COO}^-] = 0.152 \text{ mol/L}$$

	$\text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O(l)}$	$\rightleftharpoons$	$\text{CH}_3\text{COOH(aq)} + \text{OH}^-(\text{aq})$	
I	0.152		0	0
C	-x		+x	+x
E	$0.152 - x$		x	x

$$k_b = \frac{[\text{CH}_3\text{COOH}] [\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

$$= \frac{(x)(x)}{0.152 - x}$$

$$\frac{0.152}{5.6 \times 10^{-10}} \gg 100$$

$$\frac{x^2}{0.152} = 5.6 \times 10^{-10}$$

$$x = 9.2 \times 10^{-6} \quad \therefore [\text{OH}^-] = 9.2 \times 10^{-6}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]}$$

$$= \frac{1.0 \times 10^{-14}}{9.2 \times 10^{-6}}$$

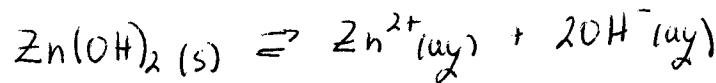
$$[\text{H}_3\text{O}^+] = 1.1 \times 10^{-9} \text{ mol/L}$$

$$\text{pH} = -\log(1.1 \times 10^{-9})$$

$$= 8.96$$

19

$$K_{sp} = 7.7 \times 10^{-17}$$



$$K_{sp} = [Zn^{2+}(aq)] [OH^-(aq)]^2$$

$$= 7.7 \times 10^{-17}$$

$$[OH^-] = 2 [Zn^{2+}]$$

$$K_{sp} = [Zn^{2+}] (2 [Zn^{2+}])^2$$

$$7.7 \times 10^{-17} = 4 [Zn^{2+}]^3$$

$$[Zn^{2+}] = \sqrt[3]{\frac{7.7 \times 10^{-17}}{4}}$$

$$= 2.7 \times 10^{-6} \text{ mol/L}$$

The molar solubility of zinc hydroxide is  $2.7 \times 10^{-6} \text{ mol/L}$ .

OR

	$Zn(OH)_2(s)$	$\rightleftharpoons$	$Zn^{2+}(aq)$	$+ 2OH^-(aq)$
I	-		0	0
C	-		$+x$	$+2x$
E	-		x	$2x$

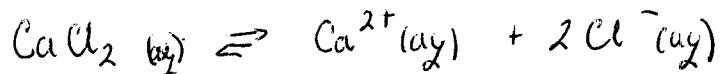
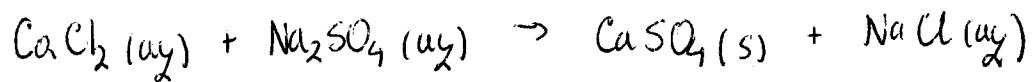
$$K_{sp} = (x)(2x)^2$$

$$4x^3 = 7.7 \times 10^{-17}$$

;

28

(20)



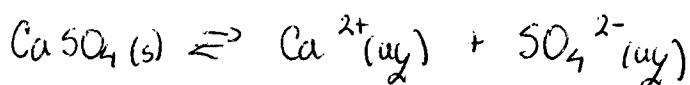
$$[\text{Ca}^{2+}(\text{aq})] = [\text{CaCl}_2(\text{aq})] = 0.200 \text{ mol/L} \quad \text{before mixing}$$

$$[\text{SO}_4^{2-}(\text{aq})] = [\text{Na}_2\text{SO}_4(\text{aq})] = 0.05 \text{ mol/L} \quad \text{before mixing}$$

After mixing total volume  $\uparrow$  (so initial concentrations  $\downarrow$ )

$$[\text{Ca}^{2+}(\text{aq})] = 0.200 \text{ mol/L} \times \frac{150 \text{ mL}}{300 \text{ mL}} \\ = 0.100 \text{ mol/L}$$

$$[\text{SO}_4^{2-}(\text{aq})] = 0.05 \text{ mol/L} \times \frac{150 \text{ mL}}{300 \text{ mL}} \\ = 0.025 \text{ mol/L}$$



$$Q = [\text{Ca}^{2+}(\text{aq})][\text{SO}_4^{2-}(\text{aq})] \\ = (0.100)(0.025) \\ = 2.5 \times 10^{-3}$$

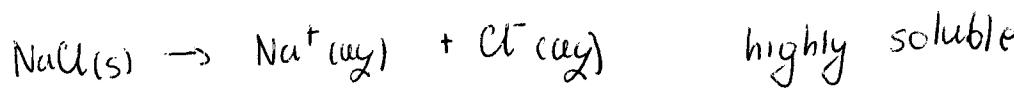
$Q \gg K_{\text{sp}} (3.6 \times 10^{-5})$  - more ions are present than would be present at equilibrium, so the reaction must shift toward the solid

$\therefore$  precipitate forms

(21)

(21)

The 2 salts have a common ion  $\text{Cl}^-$  (aq)



$\therefore$  before the addition of  $\text{PbCl}_2$

$$[\text{Cl}^-]_{\text{aq}} = [\text{NaCl(aq)}] = 0.4 \text{ mol/L}$$

$K_{\text{sp}}$  for  $\text{PbCl}_2$  indicates low solubility ( $K_{\text{sp}} = 1.7 \times 10^{-5}$ )

$\text{PbCl}_2(\text{s}) \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$			
I	-	0	0.4
C	-	$+x$	$+2x$
E	-	$x$	$0.4 + 2x$

$$\begin{aligned} K_{\text{sp}} &= [\text{Pb}^{2+}(\text{aq})][\text{Cl}^-(\text{aq})]^2 \\ &= (x)(0.4 + 2x)^2 = 1.7 \times 10^{-5} \end{aligned} \quad \frac{0.4}{1.7 \times 10^{-5}} \gg 100$$

$$\therefore K_{\text{sp}} = x(0.4)^2$$

$$x = \frac{1.7 \times 10^{-5}}{(0.4)^2}$$

$$x = 1.0625 \times 10^{-4}$$

The molar solubility of  $\text{PbCl}_2$  in 0.4 mol/L  $\text{NaCl}$  (aq) is  $1.0625 \times 10^{-4}$  mol/L

(22)

a) Before titration

The major entities are  $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$  and  $\text{H}_2\text{O}(\text{l})$

	$\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$	$\rightleftharpoons \text{H}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$		
I	0.400	0.00	0.00	
C	-x	+x	+x	
E	$0.400 - x$	x	x	

$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

$$\frac{x^2}{0.400 - x} = 1.8 \times 10^{-5} \quad \frac{0.4}{1.8 \times 10^{-5}} \gg 100$$

$$\therefore \frac{x^2}{0.400} = 1.8 \times 10^{-5}$$

$$x = 2.7 \times 10^{-3}$$

$$x = 2.7 \times 10^{-3} \text{ mol/L}$$

$$[\text{H}^+(\text{aq})] = 2.7 \times 10^{-3} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}^+] \\ &= -\log(2.7 \times 10^{-3}) \\ &= 2.62 \end{aligned}$$

Before titration begins, the pH of the sample is 2.62

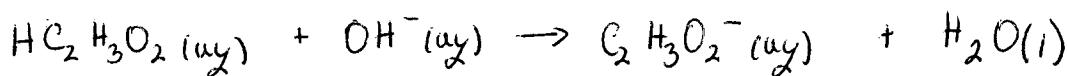
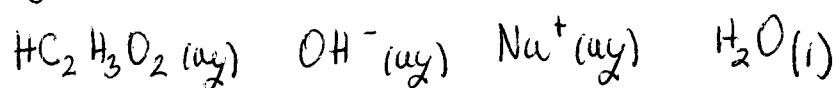
$$V_{\text{HC}_2\text{H}_3\text{O}_2} = 30.00 \text{ mL}$$

$$C_{\text{HC}_2\text{H}_3\text{O}_2} = 0.400 \text{ mol/L}$$

$$\begin{aligned} n &= CV \\ &= 30.00 \text{ mL} \times 0.400 \text{ mol/L} \\ &= 12.00 \text{ mmol} \end{aligned}$$

b) During titration

Major entities (before reaction)



$$V_{\text{NaOH}} = 15.00 \text{ mL}$$

$$C_{\text{NaOH}} = 0.400 \text{ mol/L}$$

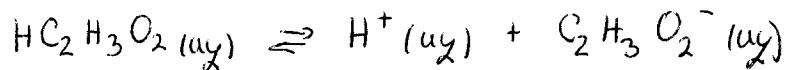
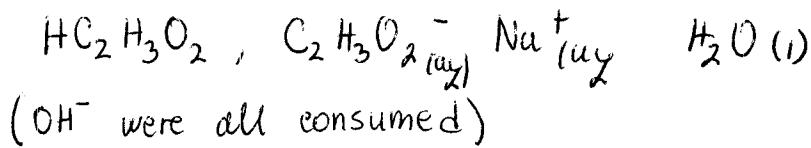
$$\begin{aligned} n &= C V \\ &= 15.00 \text{ mL} \times 0.400 \text{ mol/L} \\ &= 6.00 \text{ mmol} \end{aligned}$$

6.00 mmol of NaOH will react with 6.00 mmol of  $\text{H}^+$  leaving 6.00 mmol of acetic acid unreacted.

The total volume of the solution is

$$\begin{array}{ccc} 30.00 \text{ mL} & + & 15.00 \text{ mL} \\ \text{original solution} & & \text{added} \\ & & \text{NaOH} \end{array} = \begin{array}{c} 45.00 \\ \text{total} \\ \text{volume} \end{array}$$

After the neutralization the major entities are



$$\begin{aligned} [\text{HC}_2\text{H}_3\text{O}_2(\text{aq})] &= \frac{6.00 \text{ mmol}}{45.0 \text{ mL}} \\ &= 0.13 \text{ mol/L} \end{aligned}$$

	$\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$	$\rightleftharpoons$	$\text{H}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$
I	0.13	0	0.13
C	-x	+x	0.13 + x
E	0.13 - x	x	0.13 + x

$$K_a = \frac{[\text{H}^+(\text{aq})][\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})]}{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})]}$$

$$= \frac{x(0.13+x)}{0.13-x}$$

$$\frac{x(0.13)}{0.13} = \frac{1.8 \times 10^{-5}}{1.8 \times 10^{-5} + 100} \Rightarrow 100$$

$$= 1.8 \times 10^{-5}$$

$$\frac{x(0.13)}{0.13} = 1.8 \times 10^{-5}$$

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

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

$$\text{pH} = -\log(1.8 \times 10^{-5})$$

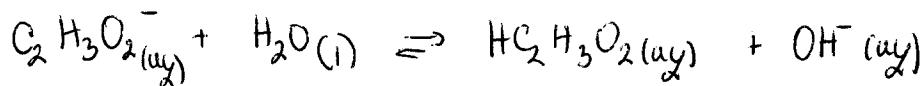
$$\text{pH} = 4.74$$

c) At the equivalence point 30.00 mL of 0.400 mol/L  $\text{NaOH}(\text{aq})$  has been added

$\text{H}^+$  will be neutralized by  $\text{OH}^-$

The remaining entities are  $\text{Na}^+(\text{aq})$   $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$   $\text{H}_2\text{O}(l)$

$\text{Na}^+$  does not affect the pH



$$K_b = \frac{[\text{OH}^-][\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$$

$$K_a = 1.8 \times 10^{-5}$$

$$K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$$

$$K_b = 5.6 \times 10^{-10}$$

$$\frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.6 \times 10^{-10}$$

At the equivalence point total volume is 30.00 + 30.00 mL  
= 60.00 mL.

Initial amount of acetic acid

$$V = 30.00 \text{ mL}$$

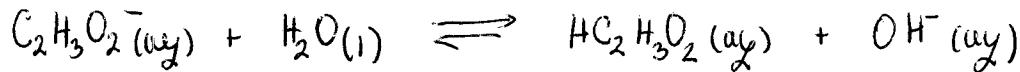
$$C = 0.400 \text{ mol/L}$$

$$n = 30.00 \text{ mL} \times 0.400 \text{ mol/L}$$

$$n = 12 \text{ mmol}$$

Since we began with 12.00 mmol of  $\text{HC}_2\text{H}_3\text{O}_2$  we will end up with 12.00 mmol of  $\text{C}_2\text{H}_3\text{O}_2^-$  (aq) at the equivalence point.

$$[\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})] = \frac{12.00 \text{ mmol}}{60.00 \text{ mL}} \\ = 0.200 \text{ mol/L}$$



	$\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$	$\text{H}_2\text{O}(l)$	$\rightleftharpoons$	$\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$	$\text{OH}^-(\text{aq})$
I	0.200			0	0
C	-x			+x	+x
E	0.200 - x			x	x

$$K_b = \frac{[HC_2H_3O_2(aq)] [OH^-]_{aq}}{[C_2H_3O_2^-(aq)]}$$

$$\frac{x^2}{(0.200-x)} = 5.6 \times 10^{-10} \quad \frac{0.200}{5.6 \times 10^{-10}} \gg 100$$

$$\frac{x^2}{0.200} = 5.6 \times 10^{-10}$$

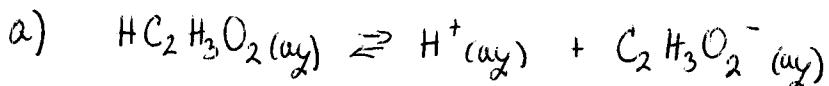
$$x = 1.06 \times 10^{-5}$$

$$[OH^-] = 1.06 \times 10^{-5}$$

$$\begin{aligned} pOH &= -\log (1.06 \times 10^{-5}) \\ &= 4.97 \end{aligned}$$

$$\begin{aligned} pH &= 14.00 - 4.97 \\ &= 9.02 \end{aligned}$$

(23)



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

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

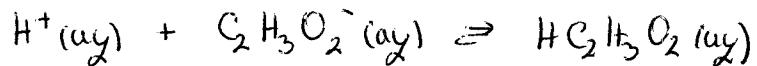
$$[H^+(aq)] = K_a \frac{0.40 \text{ mol/L}}{0.40 \text{ mol/L}}$$

$$[H^+] = 1.8 \times 10^{-5} \text{ mol/L}$$

$$pH = 4.74$$

(35)

b) When  $H^+$  ions are produced in the buffer they react with acetate ions



adding 0.1 mol of  $H^+$   $\rightarrow$  acetate ions will decrease by 0.1 (to 0.30)  
acetic acid will  $\uparrow$  by 0.1 (to 0.50)

$$\begin{aligned} [H^+(aq)] &= K_a \cdot \frac{[HC_2H_3O_2(aq)]}{[C_2H_3O_2^-(aq)]} \\ &= 1.8 \times 10^{-5} \cdot \frac{0.50}{0.30} \\ &= 3 \times 10^{-5} \text{ mol/L} \end{aligned}$$

$$pH = -\log(3 \times 10^{-5})$$

$$= 4.52$$

pH dropped from 4.74 to 4.52, a rather small decrease.

c) In the water case

$$[H^+] = 0.10 \text{ mol/L}$$

$$pH = -\log(0.10)$$

$$pH = 1.00$$

The pH of water drops from 7.00 to 1.00