REVIEW QUESTIONS
Chapter 16

1. A 0.10 M solution of lactic acid ($HC_3H_5O_3$) has a pH of 2.44. Calculate K_a for lactic acid.

$$[H^+] = 10^{-2.44} = 3.6x10^{-3}\ M$$
$$[C_3H_5O_3^-] = [H^+] = 3.6x10^{-3}\ M$$
$$[HC_3H_5O_3] = 0.10 - (3.6x10^{-3}) \approx 0.10$$
$$K_a = \frac{[H^+][C_3H_5O_3^-]}{[HC_3H_5O_3]} = \frac{(3.6x10^{-3})^2}{0.10} = 1.3x10^{-4}$$

2. A 0.200 M solution of a weak acid HX is 9.4% ionized. Calculate the pH and K_a for this acid.

$$[H^+] = [A^-] = 0.200 \text{ x } 0.094 = 0.019\ M$$
$$pH = -\log(0.019) = 1.72$$
$$[HA] = 0.200 - 0.019 = 0.181$$
$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(0.019)^2}{0.181} = 2.0x10^{-3}$$

3. Calculate the pH of a 0.050 M solution of ethylamine ($C_2H_5NH_2$, K_b= 6.4 x10^{-4}).

$$C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3^+ + OH^-$$

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{x^2}{0.050 - \cancel{x}} = 6.4x10^{-4}$$
$$x = [OH^-] = \sqrt{(0.050)(6.4x10^{-4})} = 5.7x10^{-3}\ M$$
$$pOH = -\log(5.7x10^{-3}) = 2.24$$
$$pH = 14.00 - 2.24 = 11.76$$

4. The K_a for hydrocyanic acid, HCN, is 5.0 x10^{-10}. What is the K_b for CN^-?

$$K_b = \frac{K_w}{K_a} = \frac{1.0x10^{-14}}{5.0x10^{-10}} = 2.0x10^{-5}$$

5. Hydrosulfuric acid is a polyprotic acid with the following equilibria:

$$H_2S\,(aq) \rightleftharpoons H^+\,(aq) + HS^-\,(aq) \qquad Ka_1 = 1.1 \times 10^{-7}$$
$$HS^-\,(aq) \rightleftharpoons H^+\,(aq) + S^{2-}\,(aq) \qquad Ka_2 = 1.2 \times 10^{-13}$$

a) Calculate the pH of a 0.100 M H_2S solution.

$$K_{a1} = \frac{[H^+][HS^-]}{[H_2S]} = \frac{x^2}{0.100 - \cancel{x}} = 1.1\text{x}10^{-7}$$
$$x = [H^+] = [HS^-] = \sqrt{(0.100)(1.1\text{x}10^{-7})} = 1.0\text{x}10^{-4}\ M$$
$$pH = -\log(1.0\text{x}10^{-4}) = 4.00$$

b) Calculate the $[S^{2-}]$ for the solution above.

$$K_{a2} = \frac{[H^+][S^{2-}]}{[HS^-]} = \frac{(1.0\text{x}10^{-4} + \cancel{x})(x)}{(1.0\text{x}10^{-4}) - \cancel{x}} = 1.2\text{x}10^{-13}$$
$$x = [S^{2-}] = 1.2\text{x}10^{-13}\ M$$

6. Sodium benzoate, $C_6H_5CO_2Na$, is the salt of the weak acid, benzoic acid ($C_6H_5CO_2H$). A 0.10 M solution of sodium benzoate has a pH of 8.60 at room temperature.
   a) Calculate the K_b value for benzoate ion ($C_6H_5CO_2^-$).

$$C_6H_5CO_2^- + H_2O \rightleftharpoons C_6H_5CO_2H + OH^-$$

$$pOH = 14.00 - 8.60 = 5.40$$
$$[OH^-] = [C_6H_5CO_2^-] = 10^{-5.40} = 4.0\text{x}10^{-6}$$
$$[C_6H_5CO_2H] = 0.10 - (4.0\text{x}10^{-6}) \approx 0.10$$
$$K_b = \frac{[C_6H_5CO_2H][OH^-]}{[C_6H_5CO_2^-]} = \frac{(4.0\text{x}10^{-6})^2}{0.10} = 1.6\text{x}10^{-10}$$

   b) Calculate the K_a value for benzoic acid.

$$C_6H_5CO_2H + H_2O \rightleftharpoons H_3O^+ + C_6H_5CO_2^-$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0\text{x}10^{-14}}{1.6\text{x}10^{-10}} = 6.3\text{x}10^{-5}$$

7. Potassium sorbate ($KC_6H_7O_2$) is the salt of the weak acid, sorbic acid ($HC_6H_7O_2$, Ka = 1.7×10^{-5}), and is commonly added to cheese to prevent mold. What is the pH of a solution containing 4.93 g of potassium sorbate in 500 mL of solution?

$$\mathbf{4.93\ \cancel{g} \times \frac{1\ mol}{150.1\ \cancel{g}} \times \frac{1}{0.500\ L} = 0.0657\ M}$$

$$\mathbf{C_6H_7O_2^- + H_2O \rightleftharpoons HC_6H_7O_2 + OH^-}$$

$$\mathbf{K_b = \frac{K_w}{K_a} = \frac{1.0x10^{-14}}{1.7x10^{-5}} = 5.9x10^{-10}}$$

$$\mathbf{K_b = \frac{[HC_6H_7O_2][OH^-]}{[C_6H_7O_2^-]} = \frac{x^2}{0.0657 - x} = 5.9x10^{-10}}$$

$$\mathbf{x = [OH^-] = \sqrt{(0.0657)(5.9x10^{-10})} = 6.2x10^{-6}\ M}$$

$$\mathbf{pOH = -log(6.2x10^{-6}) = 5.21}$$

$$\mathbf{pH = 14.00 - 5.21 = 8.79}$$

8. A buffer is prepared by adding 20.0 g of acetic acid ($HC_2H_3O_2$) and 20.0 g of sodium acetate ($NaC_2H_3O_2$) in enough water to prepare 2.00 L of solution. Calculate the pH of this buffer? ($K_a = 1.8 \times 10^{-5}$)

$$\mathbf{20.0\ \cancel{g}\ HAc \times \frac{1\ mol}{60.0\ \cancel{g}} \times \frac{1}{2.00\ L} = 0.167\ M}$$

$$\mathbf{20.0\ g\ NaAc \times \frac{1\ mol}{82.0\ g} \times \frac{1}{2.00\ L} = 0.122\ M}$$

	$HC_2H_3O_2$	+ H_2O	$\rightleftharpoons$ H_3O^+ +	$C_2H_3O_2^-$
Initial	**0.167**	**-----**	**0**	**0.122**
Δ	**–x**	**-----**	**+x**	**+x**
Equil.	**0.167 – x**	**-----**	**x**	**0.122 + x**

$$\mathbf{K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{(x)(0.122 + \cancel{x})}{0.167 - \cancel{x}} = 1.8x10^{-5}}$$

$$\mathbf{x = \frac{(0.167)(1.8x10^{-5})}{0.122} = 2.\underline{4}6x10^{-5}}$$

$$\mathbf{pH = -log(2.\underline{4}6x10^{-5}) = 4.61}$$

9. What is the ratio of HCO_3^- to H_2CO_3 in blood of pH 7.4? (K_a for $H_2CO_3 = 4.3 \times 10^{-7}$)

$$\mathbf{H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^-}$$

$$\mathbf{pH = pK_a + \log \frac{[HCO_3^-]}{[H_2CO_3]} \qquad 7.4 = 6.37 + \log \frac{[HCO_3^-]}{[H_2CO_3]}}$$

$$\mathbf{\frac{[HCO_3^-]}{[H_2CO_3]} = antilog\ (7.4 - 6.37) = 10^{1.03} = 11}$$

10. How many grams of NaBrO should be added to 1.00 L of 0.200 M HBrO to form a buffer with a pH of 8.80? (K_a for HBrO = 2.5×10^{-9})

$$\mathbf{HBrO + H_2O \rightleftharpoons H_3O^+ + BrO^-}$$

$$\mathbf{pH = pK_a + \log \frac{[BrO^-]}{[HBrO]} \qquad 8.80 = 8.60 + \log \frac{[BrO^-]}{[HBrO]}}$$

$$\mathbf{\frac{[BrO^-]}{[HBrO]} = antilog\ (8.80 - 8.60) = 10^{0.20} = 1.6}$$

$$\mathbf{[BrO^-] = 1.6\ (0.200\ M) = 0.32\ M}$$

$$\mathbf{1.00\ \cancel{L} \times \frac{0.32\ mol}{1\ \cancel{L}} \times \frac{118.9\ g}{1\ mol} = 38\ g}$$

11. Acetylsalicylic acid (aspirin, $HC_9H_7O_4$) is a weak acid with Ka = 2.75×10^{-5} at 25°C. 3.00 g of sodium acetylsalicylate ($NaC_9H_7O_4$) is added to 200.0 mL of 0.100 M solution of this acid. Calculate the pH of the resulting solution at 25°C.

$$\mathbf{Molarity\ of\ NaC_9H_7O_4 = 3.00\ \cancel{g} \times \frac{1\ mol}{202\ \cancel{g}} \times \frac{1}{0.200\ L} = 0.0743\ M}$$

	$HC_9H_7O_4$ +	H_2O $\rightleftharpoons$		H_3O^+ +	$C_9H_7O_4^-$
Initial	0.100	-----	0	0.0743
Δ	–x	-----	+x	+x
Equil.	0.100 – x	-----	x	0.0743 + x

$$\mathbf{K_a = \frac{[H_3O^+][C_9H_7O_4^-]}{[HC_9H_7O_4]} = \frac{(x)(0.0743 + \cancel{x})}{0.100 - \cancel{x}} = 2.75 \times 10^{-5}}$$

$$\mathbf{x = \frac{(0.100)(2.75 \times 10^{-5})}{0.0743} = 3.70 \times 10^{-5}}$$

$$\mathbf{pH = -\log(3.70 \times 10^{-5}) = 4.432}$$

12. The equations and dissociation constants for three different acids are given below:

$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	$K_a = 4.2x10^{-7}$	**pKa = 6.4**
$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$	$K_a = 6.2x10^{-8}$	**pKa = 7.2**
$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$	$K_a = 1.3x10^{-2}$	**pKa = 1.9**

Identify the conjugate pair that is best for preparing a buffer with a pH of 7.2. Clearly explain your choice.

The best conjugate pair would be $H_2PO_4^-$ and HPO_4^{2-}
The pH = pKa = 7.2 for this buffer when $[H_2PO_4^-] = [HPO_4^{2-}]$

$$pH = pK_a + \log \frac{[HPO_4^{2-}]}{[H_2PO_4^-]}$$

13. A buffer solution is prepared by adding 0.10 L of 2.0 M acetic acid solution to 0.10 L of 1.0 M NaOH solution.
a) Calculate the pH of this buffer solution.

$$0.10\ \not{L} \times \frac{2.0\ mol}{1\ \not{L}} = 0.20\ mol\ HC_2H_3O_2 \qquad 0.10\ \not{L} \times \frac{1.0\ mol}{1\ \not{L}} = 0.10\ mol\ NaOH$$

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$ +	H_2O
Initial	0.20	0.10	0	----
Δ	–0.10	–0.10	+0.10	----
Final	0.10	0	0.10	----

$$[C_2H_3O_2^-] = \frac{0.10\ mol}{0.20\ L} = 0.50\ M \qquad [HC_2H_3O_2] = \frac{0.10\ mol}{0.2\ L} = 0.50\ M$$

$$pH = pKa + \log \frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

From textbook Ka = $1.7x10^{-5}$ pKa = 4.77

$$pH = 4.77 + \log \frac{0.50}{0.50} = 4.77$$

b) 0.10 L of 0.20 M HCl is added to 0.40 L of the buffer solution above. What is the pH of the resulting solution?

The H_3O^+ ions provided by HCl react with the acetate ions in the buffer.

$$[H_3O^+] = (0.10L)(0.20\ M) = 0.020\ mol$$

$$[C_2H_3O_2^-] = [HC_2H_3O_2] = (0.40\ L)(0.50\ M) = 0.20\ mol$$

	$C_2H_3O_2^-$ +	H_3O^+ →	$HC_2H_3O_2$ +	H_2O
Initial	0.20	0.020	0.20	----
Δ	–0.020	–0.020	+0.020	----
Final	0.18	0	0.22	----

$$[C_2H_3O_2^-] = \frac{0.18\ mol}{0.50\ L} = 0.36\ M \qquad [HC_2H_3O_2] = \frac{0.22\ mol}{0.50\ L} = 0.44\ M$$

$$pH = pKa + \log\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

From textbook $Ka = 1.7x10^{-5}$ $\qquad pKa = 4.77$

$$pH = 4.77 + \log\frac{0.36}{0.44} = 4.68$$

14. A 10.0 mL solution of 0.100 M NH_3 ($K_b = 1.8\ x10^{-5}$) is titrated with a 0.100 M HCl solution. Calculate the pH of this solution at equivalence point.

At equivalence point $\quad 10.0\ mL\ NH_3\ x\ \frac{0.100\ mol}{1\ L}\ x\ \frac{1\ HCl}{1\ NH_3}\ x\ \frac{1\ L}{0.100\ mol} = 10.0\ mL\ of\ HCl$

$$[NH_3] = \frac{(0.100\ M)(10.0\ mL)}{(20.0\ mL)} = 0.0500\ M \qquad [HCl] = \frac{(0.100\ M)(10.0\ mL)}{(20.0\ mL)} = 0.0500\ M$$

First assume all of the HCl and NH_3 react to form NH_4Cl, then some of the NH_4^+ hydrolyzes back to ammonia.

	NH_3 +	HCl →	NH_4^+ +	Cl^-
Initial	0.0500	0.0500	0	0
Δ	–0.0500	–0.0500	+0.0500	+0.0500
Final	0	0	0.0500	0.0500

	NH_4^+ +	H_2O ⇌	NH_3 +	H_3O^+
Initial	0.0500	----	0	0
Δ	–x	----	+x	+x
Equil.	0.0500–x	----	x	x

$$K_a = \frac{K_w}{K_b} = \frac{1.0x10^{-14}}{1.8x10^{-5}} = 5.6x10^{-10}$$

$$K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]} = \frac{(x)(x)}{0.050 - \cancel{x}} = 5.6x10^{-10}$$

$$[H_3O^+] = x = \sqrt{(0.050)(5.6x10^{-10})} = 5.3x10^{-6}$$

$$pH = -\log(5.3x10^{-6}) = 5.28$$

15. A sample of 25.0 mL of 0.100 M solution of HBr is titrated with 0.200 M NaOH. Calculate the pH of solution after 10.0 mL of the base is added.

	HBr	+ NaOH	→ NaBr	+ H_2(
Initi	2.50 mı	2.00 mm	0	----
Δ	–2.00 m	–2.00 mn	+2.00 mn	----
Fin	0.50 mı	0	2.00 mm	----

$$[H^+]=[HBr]= \frac{0.50\ mmol}{35.0\ mL} = 0.01\underline{4}3\ M$$

$$pH= -\log(0.01\underline{4}3) = 1.85$$

16. A 10.0-mL solution of 0.300 M NH_3 is titrated with a 0.100 M HCl solution. Calculate the pH after the following additions of the HCl solution:
(a) 0.0 mL, (b) 10.0 mL, (c) 30.0 mL

a) Since no acid has been added, NH_3 ionizes as shown below:

	NH_3	+ H_2O	⇌ NH_4^+	+ OH
Initi	0.30(	----	0	0
Δ	–x	----	+x	+x
Equ	0.300-	----	x	x

From textbook, K_b= 1.8 x 10^{-5}

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{0.300 - \cancel{x}} = 1.8x10^{-5}$$

$$x =[OH^-] = \sqrt{(0.300)(1.8x10^{-5})} = 2.\underline{3}2x10^{-3}$$

$$pOH = -\log(2.\underline{3}2x10^{-3}) = 2.63$$

$$pH = 14.00-2.63 = 11.37$$

b) Addition of 10.0 mL of acid neutralizes some of the ammonia, as shown below:

	NH_3 +	HCl →	NH_4^+ +	Cl^-
Initial	3.00 mmol	1.00 mmol	0	----
Δ	–1.00 mmol	–1.00 mmol	+1.00 mmol	----
Final	2.00 mmol	0	1.00 mmol	----

$$[NH_3] = \frac{2.00 \text{ mmol}}{20.0 \text{ mL}} = 0.100 \text{ M} \qquad [NH_4^+] = \frac{1.00 \text{ mmol}}{20.0 \text{ mL}} = 0.0500 \text{ M}$$

$$K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \qquad pK_a = \log K_a = 9.25$$

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]} = 9.25 + \log \frac{0.100}{0.0500} = 9.55$$

c) After addition of 30.0 mL of HCl equivalence point is reached and pH of solution is based on hydrolysis of the salt formed.

	NH_3 +	HCl →	NH_4^+ +	Cl^-
Initial	3.00 mmol	3.00 mmol	0	----
Δ	–3.00 mmol	–3.00 mmol	+3.00 mmol	----
Final	0	0	3.00 mmol	----

$$[NH_4^+] = \frac{3.00 \text{ mmol}}{40.0 \text{ mL}} = 0.0750 \text{ M}$$

	NH_4^+ +	H_2O ⇌	NH_3 +	H_3O^+
Initial	0.0750	----	0	0
Δ	–x	----	+x	+x
Equil.	0.0750–x	----	x	x

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = \frac{(x)(x)}{(0.0750 - \cancel{x})} = 5.56 \times 10^{-10}$$

$$x = [H_3O^+] = \sqrt{(0.0750)(5.56 \times 10^{-10})} = 6.46 \times 10^{-6}$$

$$pH = -\log(6.46 \times 10^{-6}) = 5.19$$

17. A 45.0-mL sample of 0.200 M acetic acid is titrated with 0.180 M NaOH. Calculate the pH of the solution (a) before addition of NaOH, (b) after addition of 20.0 mL of NaOH and (c) at the equivalence point.

a) Since no base has been added, acetic acid ionizes as shown below:

	$HC_2H_3O_2$ +	H_2O $\rightleftharpoons$	$C_2H_3O_2^-$ +	H_3
Initi	0.20	----	0	0
Δ	–x	----	+x	+x
Equ	0.200–	----	x	x

From textbook, $K_a = 1.7 \times 10^{-5}$

$$K_a = \frac{[CHO_2^-][H_3O^+]}{[HC_2H_3O_2]} = \frac{(x)(x)}{0.200 - \cancel{x}} = 1.7x10^{-5}$$

$$x = [H_3O^+] = \sqrt{(0.200)(1.7x10^{-5})} = 1.\underline{8}4x10^{-3}$$

$$pH = -\log(1.\underline{8}4x10^{-3}) = 2.73$$

b) Addition of 20.0 mL of NaOH neutralizes some of the acetic acid, as shown below:

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$ +	H_2
Initi	9.00 mr	3.60 mm	0	----
Δ	–3.60 m	–3.60 mm	+3.60 mm	----
Fina	5.40 mr	0	3.60 mm	----

$$[HC_2H\ O_2] = \frac{5.40\ mmol}{65.0\ mL} = 0.0831\ M \qquad [C_2H_3O_2^-] = \frac{3.60\ mmol}{65.0\ mL} = 0.0554\ M$$

$$K_a = 1.7 \times 10^{-5} \qquad pK_a = \log K_a = 4.77$$

$$pH = pK_a + \log\frac{[base]}{[acid]} = 4.77 + \log\frac{0.0554}{0.0831} = 4.59$$

c) At equivalence point all the acid is neutralized by the base and the pH of the solution is based on hydrolysis of the salt formed.

Volume of base at equivalence point:

$$45.0\ mL\ \cancel{acid} \times \frac{0.200\ \cancel{mol}}{1\ \cancel{L}} \times \frac{1\ \cancel{mol}\ base}{1\ \cancel{mol\ acid}} \times \frac{1\ \cancel{L}}{0.180\ \cancel{mol}} = 50.0\ mL\ of\ base$$

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$ +	H_2O
Initial	9.00 mmol	9.00 mmol	0	----
Δ	–9.00 mmol	–9.00 mmol	+9.00 mmol	----
Final	0	0	9.00 mmol	----

$$[C_2H_3O_2^-] = \frac{9.00 \text{ mmol}}{95.0 \text{ mL}} = 0.094\underline{7}4 \text{ M}$$

	$C_2H_3O_2^-$ +	H_2O ⇌	$HC_2H_3O_2$ +	OH^-
Initial	0.094$\underline{7}$4	----	0	0
Δ	–x	----	+x	+x
Equil.	0.094$\underline{7}$4–x	----	x	x

$$K_b = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.\underline{8}8 \times 10^{-10}$$

$$K_b = \frac{[C_2H_3O_2^-][OH^-]}{[HC_2H_3O_2]} = \frac{(x)(x)}{(0.094\underline{7}4 - \cancel{x})} = 5.\underline{8}8 \times 10^{-10}$$

$$x = [OH^-] = \sqrt{(0.094\underline{7}4)(5.\underline{8}8 \times 10^{-10})} = 7.\underline{4}7 \times 10^{-6}$$

$$pOH = -\log(7.\underline{4}7 \times 10^{-6}) = 5.13$$

$$pH = 14.00 - 5.13 = 8.87$$