- **Review Homework Chap 15:** pg. 673 #98, 100, 104, 106*, 107*, 108, 109, 112, 117, 120*, 123*, 125, 129, 133, 138, 139. *See review list for hints for questions with **.

98 **pH = 3.86 for a 0.064 M solution of a monoprotic acid. Is this a strong acid?**

Determine the $[H^+]$ for this solution: $[H^+] = 10^{-pH} = 10^{-3.86} = 1.4\times10^{-4}$ M. Since this is << 0.064, <u>the acid must be a weak acid</u>. To further confirm, the % ionization would be:

$$\%\text{ionization} = \frac{[H^+]}{[HA]_0}\times100\% = \frac{1.4\times10^{-4}\text{ M}}{0.064\text{ M}}\times100\% = 0.22\% \ll 100\%$$

100 **HA and HB are both weak acids, but HB is a stronger acid than HA. Will it take more volume of 0.1M NaOH to neutralize 50.0 mL of 0.10M HA or 50.0 mL or 0.10M HB?**

<u>The volume of NaOH required will be the same.</u> Since a strong base (NaOH) is being added to the weak acids, both reactions will be driven to completion (100% to right).

$$HA + NaOH \rightarrow H_2O + NaA$$
$$HB + NaOH \rightarrow H_2O + NaB$$

At the equivalence point, the moles of acid must equal the moles of base. Thus, since the moles of acid is the same for both titrations, the moles of NaOH added must be the same for both titrations. Therefore the volume of base required to react with the same concentration of acid solutions (either both weak, both strong, or one strong and one weak) will be the same.

104 **Use the data in Table 15.3 to calculate the equilibrium constant for the following reaction:**

We can write two equilibria that add up to the equilibrium in the problem.

$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$ $\qquad \boldsymbol{K_a} = \dfrac{[H^+][CH_3COO^-]}{[CH_3COOH]} = 1.8\times10^{-5}$

$H^+(aq) + NO_2^-(aq) \rightleftharpoons HNO_2(aq)$ $\qquad \boldsymbol{K'} = \dfrac{1}{\boldsymbol{K_a}(HNO_2)} = \dfrac{1}{4.4\times10^{-4}} = 2.2\times10^3$

$$\boldsymbol{K'} = \frac{[HNO_2]}{[H^+][NO_2^-]}$$

$CH_3COOH(aq) + NO_2^-(aq) \rightleftharpoons CH_3COO^-(aq) + HNO_2(aq)$ $\qquad \boldsymbol{K} = \dfrac{[CH_3COO^-][HNO_2]}{[CH_3COOH][NO_2^-]} = \boldsymbol{K_a}\times\boldsymbol{K'}$

$$\boldsymbol{K} = \boldsymbol{K_a}\times\boldsymbol{K'} = (1.8\times10^{-5})(2.2\times10^{3}) = \mathbf{4.0\times10^{-2}}$$

106 **Calculate the pH of a 0.20 M ammonium acetate solution.**

An ammonium acetate solution consists of the ions, NH_4^+ and CH_3COO^-. This is a <u>salt</u> solution.

Ammonium is a weak acid. $K_a(NH_4^+) = 5.6\times10^{-10}$

Acetate is a weak base. $K_b(CH_3COO^-) = 5.6\times10^{-10}$

In this specific case the K_a of ammonium ion is the same as the K_b of acetate ion. The two are of exactly (to two significant figures) equal strength. The solution will have **pH 7.00**.

However, be aware— not all salt solutions have a pH of 7!!!

107 Novacaine is a weak base, $K_b = 8.91 \times 10^{-6}$. What is the ratio of the concentration of the base to that of its acid in the blood plasma (pH = 7.40) of a patient?

Write the **base as Nov** and the **acid as $HNov^+$.** Then the equilibrium reaction is:

$$Nov(aq) + H_2O(l) \rightleftharpoons HNov^+(aq) + OH^-(aq) \quad K_b = 8.91 \times 10^{-6}$$

pH = 7.40 so pOH = 14 – 7.40 = 6.60

$[OH^-] = 10^{-6.60} = 2.51 \times 10^{-7}$

$$K_b = \frac{[HNov^+][OH^-]}{[Nov]}, \text{ so } \frac{[Nov]}{[HNov^+]} = \frac{[OH^-]}{\boldsymbol{K_b}} = \frac{2.51\times10^{-7}}{8.91\times10^{-6}} = \mathbf{0.028}$$

108 Which of the following is the stronger base: NF_3 or NH_3? (Hint: F is more electronegative than H.)

The fact that fluorine attracts electrons in a molecule more strongly than hydrogen causes NF_3 to be a poor electron pair donor and a poor base (poor Lewis base). **NH_3** is the stronger base.

109 Which of the following is a stronger base: NH_3 or PH_3? (Hint: The N–H bond is stronger than the P–H bond.)

Because the P–H bond is weaker, there is a greater tendency for PH_4^+ to ionize. Therefore, PH_3 is a weaker base than NH_3.

112 What is the pH of 250.0 mL of an aqueous solution containing 0.616 g of the strong acid trifluoromethane sulfonic acid?

First we must calculate the molarity of the trifluoromethane sulfonic acid. (Molar mass = 150.1 g/mol)

$$Molarity = \frac{0.616\ g \times \dfrac{1\ mol}{150.1\ g}}{0.250\ L} = 0.0164\ M$$

Since trifluoromethane sulfonic acid is a strong acid and is 100% ionized, the $[H^+]$ is 0.0165 *M*.

$$\mathbf{pH} = -\log(0.0164) = \mathbf{1.79}$$

117 Given the equation: $HbH^+ + O_2 \rightleftharpoons HbO_2 + H^+$

(a) What form of hemoglobin is favored in the lungs where oxygen concentration is highest?
From the equilibrium equation, high oxygen concentration puts stress on the left side of the equilibrium and thus shifts the concentrations to the right to compensate. **HbO_2** is favored.

(b) In body tissues, where the cells release CO_2, the blood is more acidic. What form of hemoglobin is favored under this condition?
High acid, H^+ concentration, places stress on the right side of the equation forcing concentrations on the left side to increase, thus releasing oxygen and increasing the concentration of **HbH^+**.

(c) When a person hyperventilates, the concentration of CO_2 in his or her blood decreases. How does this action affect the above equilibrium? Why should a person who is hyperventilating, breath into a paper bag? Removal of CO_2 decreases H^+ (in the form of carbonic acid), thus shifting the reaction to the **right**. More HbO_2 will form. Breathing into a paper bag increases the concentration of CO_2 (re–breathing the exhaled CO_2), thus causing more O_2 to be released as explained above.

120 Calculate the concentrations of all species in a 0.100 M Na_2CO_3 solution.

$0.100\ M\ Na_2CO_3 \rightarrow 0.200\ M\ Na^+ + 0.100\ M\ CO_3^{2-}$

First stage:

	$CO_3^{2-}(aq) + H_2O(l)$	$\rightleftharpoons HCO_3^-(aq)$	$+\ OH^-(aq)$
Initial (M):	0.100	0	0
Change (M):	$-x$	$+x$	$+x$
Equilibrium (M):	$0.100 - x$	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.8 \times 10^{-11}} = 2.1 \times 10^{-4}$$ (Or just look up K_b of CO_3^{2-} on p. 651)

$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]} = \frac{x^2}{0.100 - x} \approx \frac{x^2}{0.100} = 2.1 \times 10^{-4}$$

$x = 4.6 \times 10^{-3}\ M = [HCO_3^-] = [OH^-]$

Second stage:

	$HCO_3^-(aq) + H_2O(l)$	$\rightleftharpoons H_2CO_3(aq)$	$+\ OH^-(aq)$
Initial (M):	4.6×10^{-3}	0	4.6×10^{-3}
Change (M):	$-y$	$+y$	$+y$
Equilibrium (M):	$(4.6 \times 10^{-3}) - y$	y	$(4.6 \times 10^{-3}) + y$

$$K_b = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]} = \frac{y[(4.6 \times 10^{-3}) + y]}{(4.6 \times 10^{-3}) - y} \approx \frac{(y)(4.6 \times 10^{-3})}{(4.6 \times 10^{-3})} = 2.4 \times 10^{-8}$$

$y = 2.4 \times 10^{-8}\ M$

At equilibrium:
$\mathbf{[Na^+] = 0.200}\ \boldsymbol{M}$**;** $\mathbf{[CO_3^{2-}]} = 0.100 - (4.6 \times 10^{-3})\ M = \mathbf{0.0954}\ \boldsymbol{M}$

$\mathbf{[HCO_3^-]} = (4.6 \times 10^{-3})\ M - (2.4 \times 10^{-8})\ M \approx \mathbf{4.6 \times 10^{-3}}\ \boldsymbol{M}$

$\mathbf{[H_2CO_3] = 2.4 \times 10^{-8}}\ \boldsymbol{M}$

$\mathbf{[OH^-]} = (4.6 \times 10^{-3})\ M + (2.4 \times 10^{-8}]\ M \approx \mathbf{4.6 \times 10^{-3}}\ \boldsymbol{M}$

$$\mathbf{[H^+]} = \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-3}} = \mathbf{2.2 \times 10^{-12}}\ \boldsymbol{M}$$

123 How many grams of NaCN would you need to dissolve in enough water to make exactly 250mL of solution with a pH = 10.00?

When the pH=10.00, the pOH= 4.00. Thus, $[OH^-] = 1.0 \times 10^{-4}\ M$. The concentration of HCN must be the same. (Why?) If the concentration of NaCN is x, the table looks like:

	$CN^-(aq) + H_2O(l)$	$\rightleftharpoons HCN(aq)$	$+\ OH^-(aq)$
Initial (M):	x	0	0
Change (M):	-1.0×10^{-4}	$+1.0 \times 10^{-4}$	$+1.0 \times 10^{-4}$
Equilibrium (M):	$(x - 1.0 \times 10^{-4})$	(1.0×10^{-4})	(1.0×10^{-4})

$$K_b = \frac{[HCN][OH^-]}{[CN^-]} = \frac{(1.0 \times 10^{-4})^2}{(x - 1.0 \times 10^{-4})} = 2.0 \times 10^{-5}$$

$x = 6.0 \times 10^{-4}\ M = [CN^-]_0$

$$\textbf{Amount of NaCN} = 250\text{ mL} \times \frac{6.0 \times 10^{-4}\text{ mol NaCN}}{1000\text{ mL}} \times \frac{49.01\text{ g NaCN}}{1\text{ mol NaCN}} = \mathbf{7.4 \times 10^{-3}\ g\ NaCN}$$

125 Calculate the pH of a 1 L solution containing 0.150 mol of CH_3COOH and 0.100 mole of HCl.

A weak acid (CH_3COOH) is being added to a strong acid (HCl).

Thus, the pH is totally determined by the HCl because the contribution of H^+ by the weak acid is negligible.

[HCl] = 0.100 M, so $[H^+] = 0.100\ M$ **pH = 1.000**

129 Describe the hydration of SO_2 as a Lewis acid–base reactions.

Like carbon dioxide, sulfur dioxide behaves as a Lewis acid by accepting a pair of electrons from the Lewis base water. The Lewis acid–base adduct rearranges to form sulfurous acid in a manner exactly analogous to the rearrangement of the carbon dioxide–water adduct to form carbonic acid that is presented on page 665 of the textbook.

133 Both NH_2^- and N^{3-} ions do not exist in aqueous solutions because they are stronger bases than hydroxide.

Write equations showing the reactions of these ions in water.

(a) $NH_2^- + H_2O \rightarrow NH_3 + OH^-$

$N^{3-} + 3H_2O \rightarrow NH_3 + 3OH^-$

(b) Which ion is the stronger base?

N^{3-} is the stronger base since each ion produces 3 OH^- ions.

138 Which of the following does not represent a Lewis acid–base reaction?

a) $H_2O + H^+ \rightleftharpoons H_3O^+$

b) $NH_3 + BF_3 \rightleftharpoons H_3NBF_3$

c) $PF_3 + F_2 \rightleftharpoons PF_5$

d) $Al(OH)_3 + OH^- \rightleftharpoons Al(OH_4)^-$

(c), $PF_3 + F_2 \rightarrow PF_5$, does not represent a Lewis acid–base reaction. In this reaction, the F–F single bond is broken and single bonds are formed between P and each F atom. For a Lewis acid–base reaction, the Lewis acid is an electron–pair acceptor and the Lewis base is an electron–pair donor.

139 True or False? If false explain why the statement is wrong

(a) All Lewis acids are Brønsted acids: False.

A Lewis acid such as CO_2 is not a Brønsted acid. It does not have a hydrogen ion to donate.

(b) The conjugate base of an acid always carries a negative charge. False.

Consider the weak acid, NH_4^+. The conjugate base of this acid is NH_3, which is neutral.

(c) The percent ionization of a base increases with its concentration in solution. False.

The percent ionization of a base decreases with increasing concentration of base in solution.

(d) A solution of barium fluoride is acidic. False.

A solution of barium fluoride is basic. The fluoride ion, F^-, is the conjugate base of the weak acid, HF. It will hydrolyze to produce OH^- ions.