**Ch 18 Aqueous Ionic Equilibrium Name:**

*18.1 The Danger of Antifreeze*

How is pH maintained in your blood? How does acidosis result from antifreeze poisoning?

*18.2 Buffers: Solutions That Resist pH Change*

Weak acids/bases partially ionize to produce their conjugates – why can’t they be buffers on their own?

A classic buffer is acetic acid mixed with sodium acetate. Why not use aluminum acetate? *(last chapter!)*

How does this buffer resist changes in pH when strong base or acid is added?

*CC18.1 ­\_­­\_\_\_ (Explain your choice.)*

On pg790 the ionization equation for acetic acid is shown. This equation can also be written as:

HC2H3O2(aq) ⇋ H+(aq) + C2H3O2−(aq) or using abbreviations as HAc(aq) ⇋ H+(aq) + Ac−(aq)

What happens to this equilibrium when sodium acetate is added to make a buffer solution? What is this known as?

*Example 18.1 Shows the calculation for determining the pH of the HAc/Ac− buffer we will use in lab.*

*Practice 18.1 (How did the pH change since the [HAc] was doubled compared to the Example?)*

Why does the [H+] for a buffer equal its *K*a? (assuming ionization is minimal)

Review how the Henderson-Hasselbalch equation is derived from the *K*a expression. *Example 18.2* shows how you can use an equilibrium approach (ICE) to calculate the pH of a buffer and then the much easier approach using the

H-H equation. *AP does not require this calculation but it does require qualitative understanding of the equation.*

*Practice 18.2 (Use only the H-H method.)*

*CC18.2 \_\_\_\_ (explain why) CC18.3 \_\_\_\_*

*Calculating pH Changes in a Buffer Solution* This section shows how we will complete calculations for the Buffer Lab. The BAA charts help keep track of the stoichiometric changes as small amounts of acid or base are added to the buffer they are similar to ICE charts. In *Example 18.3, my set up for the chart looks like this:*

OH− + HAc ⇋ H2O + Ac− I feel this better shows how the added base (or acid) is

0.010 0.100 ---- 0.100 “absorbed” by the buffer and resultant changes in [ ]

-0.010 -0.010 ---- +0.010 (the stoich!). Once you have the [HAc] and [Ac−] use H-H

0.000 0.090 ---- 0.110 to calculate pH (shown in blue on pg 797).

The last calculation of *Example 18.3* is an interesting comparison to show the effect of a buffer.

*More Practice 18.3 (Use H+ + Ac− ⇋ HAc for your chart and use H-H to calculate pH)*

*CC18.4 \_\_\_\_*

*Figure 18.4* shows the “classic” basic buffer. The weak base \_\_\_\_\_ is mixed with \_\_\_\_\_\_\_ which contain \_\_\_\_+

which is the conjugate \_\_\_\_\_\_\_ of ammonia. The system works in a similar manner as the HAc/Ac− buffer.

*Practice 18.4 Use OH− + NH4+ ⇋ NH3 + H2O More Practice 18.4 NH4+ ⇋ NH3 + H+*

*86.3 Buffer Effectiveness: Buffer Range and Buffer Capacity*

When is a buffer most effective? By how much should the relative concentrations not differ?

What do the charts and calculations at the bottom of pg800 show?

Since the relative [acid] to [base] in a \_\_\_\_\_\_\_\_\_\_ should not differ by more than 10 the effective range of a buffer is

p*K*a ± \_\_\_\_\_. *Review Example 16.5*

Read *Chemistry and Medicine: Buffer Effectiveness in Human Blood* and the paragraph on *Buffer Capacity*

*18.4 Titrations and pH Curves*

Review the terminology of titration on pg803.

Pgs 805-806 go over in detail the changes that occur during the titration of a strong \_\_\_\_\_\_ with a strong \_\_\_\_\_\_.

The most important part is the equivalence point. At this point what is present in the solution and why is the pH =7?

In AP it is important to understand the analysis of the overall pH curve on pg806. The curve at the bottom of the page shows that the titration of a strong base with a strong acid is just the reverse curve.

*Review Example 18.6 CC18.5 \_\_\_\_*

Why does it take the same amount of NaOH to titrate the same volume of a weak acid (same conc) as it did for the strong acid?

Notice how the initial pH of the weak acid solution can be calculated by finding the [H+] using an ICE chart and the *K*a

expression. As you add strong base to the weak acid a stoichiometric amount of the \_\_\_\_\_\_\_\_\_\_ate base is formed.

This creates a \_\_\_\_\_\_\_\_\_\_\_ and the H-H equation can be used to calculate pH *(center of pg775)*. At the \_\_\_\_-

equivalence point *(12.5mL in the table)* the concentration of the formic acid and its conjugate base are \_\_\_\_\_\_\_\_.

Note in the H-H calculation at this point [base]/[acid] = 1 and log 1 = \_\_\_\_ so pH = \_\_\_\_\_.

At the equivalence point the solution contains water and the ion \_\_\_\_\_\_ (conjugate base of HCHO2) which undergoes

hydrolysis according to the equation: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

*(notice the Na+ are not shown because they are spectator ions and do not hydrolyze)*

The calculations on pg810 show that the pH at the equivalence point is determined by the hydrolysis of the conjugate

base. The titration of a weak \_\_\_\_\_\_\_ by a strong base ALWAYS has a \_\_\_\_\_\_\_\_\_ equivalence point. The

calculations on pg811 show that after the eq pt there is an excess of \_\_\_− ions and the pH curve flattens due the

presence of the strong base. Note the curve at the bottom of the page shows the formation of a buffer range (zone)

where the pH of the solution does NOT change appreciably. *Review Example 18.7 CC18.7 \_\_\_\_ 18.8 \_\_\_\_*

The titration of a weak base with a strong acid is basically the opposite graph. Notice *Figure 86.8* also shows a buffer

zone at the beginning of the titration. At the eq pt the conjugate \_\_\_\_\_\_ of the weak base controls the pH (due to

hydrolysis) which is less than \_\_\_\_. *Figure 18.9* shows the titration of a polyprotic acid *(AP does not require*

*quantitative calculations for these acids)*

*CC 18.7 \_\_\_\_ (Explain your choice.)*

*Figure 18.11* shows that if the correct indicator is chosen the \_\_\_\_ pt of the indicator occurs at the titration eq point.

An indicator is a \_\_\_\_\_\_\_ acid that has a different \_\_\_\_\_\_\_\_ than its \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ base. Another way to write

general eqn is: HIn(aq) ⇋ H+ + In− \_\_\_\_\_\_\_\_phthalein is a common indicator. Most indicators are

large organic molecules similar to what is shown in *Figure 16.12*. Note phenolphthalein is \_\_\_\_\_\_\_less in \_\_\_\_\_\_\_ic

solutions and bright \_\_\_\_\_\_\_\_ in basic solutions. *(The equilibrium shifts based on the [base]/[acid] ratio for the*

*indicator. Table 18.1* shows the color changes for several indicators.

*18.5 Solubility Equilibria and the Solubility Product Constant*

This section helps to that compounds we previously considered to be \_\_\_soluble can have \_\_\_\_\_\_\_\_\_\_\_s of solubility.

Why is the solid not included in the *K*sp expression?

*K*sp is NOT \_\_\_\_\_\_\_\_\_\_\_\_\_ solubility (in mol/\_\_) of the compound. The AgCl calculations on pg784 and *Example 18.8* show how molar solubility can be calculated. They use “S” instead of “x” as we did before.

*Practice 18.8 Practice 18.9*

Read *Chemistry in Your Day: Hard Water* What makes water hard? *(soap scum forms due to hard water)*

Why does magnesium hydroxide have a smaller solubility than iron(II) carbonate even though it has a smaller *K*sp?

A \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ ion can reduce the solubility according to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ principle.

*Practice 18.10*

*CC 18.8 \_\_\_ (Explain your choice.)*

*Review Example 16.11 and complete Practice 16.11*

*18.6 Precipitation*

When *Q < K*sp in a solution the “reaction” will proceed to the \_\_\_\_\_\_\_ and solid will continue to \_\_\_\_\_\_\_\_\_\_\_\_ as

long as the solution is \_\_\_saturated (until it reaches equilibrium). When *Q = K*sp in a solution the solution is

\_\_\_\_\_\_\_\_\_\_\_\_ed. When *Q > K*sp the reaction will proceed to the \_\_\_\_\_\_\_\_\_ and solid will \_\_\_\_\_\_\_\_\_\_\_\_\_\_ate.

*(unless the solution is \_\_\_\_\_\_\_\_saturated – Figure 18.14)*

*Review Example 18.12 and complete Practice 18.12 CC18.11 \_\_\_\_*

*Review Example 18.13 (Why would magnesium hydroxide precipitate before calcium hydroxide?*

*18.7 Qualitative Chemical Analysis* This section shows a method which had been used to identify ions in an unknown solution before instruments that use spectroscopy were available.

*18.8 Complex Ion Equilibria* AP will use complex ion equations in questions but they do not require in-depth knowledge of how the ions form. The silver test tube ornaments are made using this type of reaction chemistry. If time allows I demonstrate the amphoteric properties of aluminum hydroxide showing it can be soluble in both high and low pH solutions.

*Exercises* (pgs 803-809) **#**27, 28, 29&37, 31, 33, 35, 37, 39(a,c), 41b, 45, 47, 53, 61-63, 65, 75-77, 81, 85(a,b), 87(a,b), 95(a,c), 101, 102