# Notes#54 Additional Aspects of Aqueous Equilibrium: The Common Ion Effect/AP Chem

I. The Common Ion Effect and Buffers. Things that can SUPPRESS a pH change.

a. What is the Common Ion Effect? The common ion effect is the _____ caused by the addition of a _____ that has an ion in common with ions already dissolved in the solution. The common ion effect _____ pH changes in solution.

b. Common ion effect is just an extension of **LeChatelier's Principle**. How does it work? Let us consider our favorite salad dressing acid, CH_3COOH. Below is the ionization expression for acetic acid.

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$
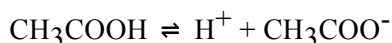
Now, what will happen when we add _____? Will we make as much H^+? _____

What will happen to the [H+]?_____ What will happen to the pH? _____

c. Calculations with Common ions

EX 1: What is the pH of a solution containing 0.30 M HCOOH and 0.52 M HCOOK? (Ka for formic acid = 1.7×10^{-4})

We are going to do an example of such a calculation TWO ways...

WAY #1 - "The Old Skool Way" I.C.E. Tables

WAY #2 - "The New School Way" Henderson-Hasselbalch

Consider the dissociation of a weak acid:

$$HA \Leftrightarrow H^+ + A^-$$

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

$$[H+] = \frac{Ka\,[HA]}{[A^-]}$$

Take –(log) of this entire equation

This is the Henderson-Hasselbach Equation:

$$pH = pKa + log([A^-]/[HA])$$

OR

$$pOH = pKb + log([HA]/[A^-])$$

*** How could you prove that a adding a common ion to your solution really DOES suppress a pH change?

EX 2: You can explore the common ion effect with solutions containing a weak base too. Check out hmwk prob. 16.4. (example in your book on page 667)

II. BUFFERS...just an extension of common ion effect...

a. What are Buffers? A buffer is a solution of (1) a _____ acid/weak base and (2) its salt (conjugate partner) used to _____ a pH change upon the addition of small amounts of either acid or base. DEMO

*** Our bodies are actually a cocktail of ESSENTIAL buffer systems......(insert pg. 669)

b. Why a weak acid/base and its salt? A buffer must contain an ACID to react with any OH^- ions that are added *and* it must contain a BASE to react with any H^+ ions. Let's take a look at an example buffer system: CH_3COONa/CH_3COOH

- If you add some base it will be neutralized by _____. - If you add some acid, it will be neutralized by _____.

c. Do buffer systems have limitations? In other words, the MORE concentrated the buffer components, the MORE H^+ and OH^- that can be "eaten up," the _____ the buffer capacity of the system.

EX 3: Notice that *only* weak acids/base systems can act as buffers. Why is this? Why can't strong acids and their conjugate partners act as buffers???
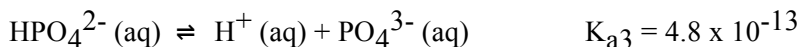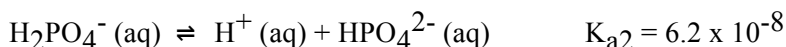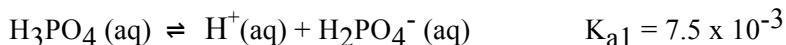
d. Calculations with Buffers EX 4: a) Calculate the pH of the 0.30 M NH_3/0.36 M NH_4Cl buffer system.

b) What is the pH *after* the addition of 20.0 mL of 0.050 M NaOH to 80.0 mL of the buffer solution.

c) What would happen to the pH if 20 mL of 0.050 M NaOH was added to 80.0 mL of pure water? Comment

on the effectiveness of your buffer system.

e. Choosing the Buffer system that is right for YOU. For example, how would you prepare a "Phosphate Buffer" with a pH of about 7.40. Logic for solving problem is based on Henderson Hasselbalch Eq.....

1. Choose a buffer system where the weak acid has a pK_a value close to the pH that you are trying to maintain.

$$H_3PO_4 \text{ (aq)} \rightleftharpoons H^+\text{(aq)} + H_2PO_4^- \text{ (aq)} \qquad K_{a1} = 7.5 \text{ x } 10^{-3}$$

$$H_2PO_4^- \text{ (aq)} \rightleftharpoons H^+ \text{ (aq)} + HPO_4^{2-} \text{ (aq)} \qquad K_{a2} = 6.2 \text{ x } 10^{-8}$$

$$HPO_4^{2-} \text{ (aq)} \rightleftharpoons H^+ \text{ (aq)} + PO_4^{3-} \text{ (aq)} \qquad K_{a3} = 4.8 \text{ x } 10^{-13}$$

2. Determine the ideal ratio of [Conjugate base]/[acid] by plugging in pH and pK_a into Henderson Hasselbalch Eq and solving for ratio.