

In this Chapter you will learn about 3 more types of equilibrium systems:

1. Buffer Solutions
2. Titration Systems
3. Equilibrium of Slightly Soluble (almost insoluble) salts & metal hydroxides

I. Section 17.1 & 17.2

A. Buffer is an equilibrium that protects an existing Hydronium or Hydroxide ion concentration. A buffer solution once formed works against the addition of a Strong acid or a strong base. This is a twofold problem;

1. Setting up the buffer system solution then stressing it with either a strong Acid or base.
2. Then stoichiometric problems can be conducted

B. There are 2 types of "Buffer" systems;

1. An "Acid Buffer" which is made from adding a weak acid (p.685) and an ionic salt containing a metal and the conjugate base ion of the weak Acid. (polyprotic acids can be used also p.692)

Ex: HCN with NaCN then add either NaOH or HCl

Ex: H<sub>2</sub>CO<sub>3</sub> with CsHCO<sub>3</sub> then add either KOH or HI

2. A "Base Buffer" which is made by adding a weak Bronsted base (p. 694) And a salt containing the conjugate acid of the weak base.

Ex: NH<sub>3</sub> with NH<sub>4</sub>Br then add either LiOH or HNO<sub>3</sub>

C. There are two ways to find the pH (or pOH) of a buffer system before adding a stressing strong acid or strong base:

1. If the initial concentration of the weak acid and its common ion salt are equal then the  $[H_3O^+] = K_a$  because their concentrations will cancel out in the equilibrium expression.  $pH = -\log K_a$  and we now call the pH of the buffer  $pK_a$

2. If the initial concentration of the weak Bronsted base and its common ion salt are equal the the  $[OH^{-1}] = K_b$  because their concentrations will cancel out in the equilibrium expression.  $pOH = -\log K_b$  and we now call the pOH of the buffer  $pK_b$ .

3. If the initial concentration of the weak acid and its common ion salt are NOT equal then you will use the Henderson-Hasselbalch equation (p.727).

A. for an acid Buffer system 
$$\text{pH} = \text{pK}_a + \log \frac{[\text{conjugate base}]}{[\text{weak acid}]}$$

\*\*remember  $\text{pK}_a = -\log K_a$

B. for a base Buffer sytem 
$$\text{pOH} = \text{pK}_b + \log \frac{[\text{conjugate acid}]}{[\text{weak base}]}$$

\*\*remember  $\text{pK}_b = -\log K_b$

\*\*See page 727 for example problem acid buffer

\*\*See page 728 for example problem base buffer

Page 728 Buffers only have a certain range at which they are effective. READ

D. Once the Buffer system and its pH or pOH have been determined we can now stress it with a strong acid or base and re-calculate the pH or pOH and see how effective the Buffer system is. LeChaetlier's Principle comes into play her.

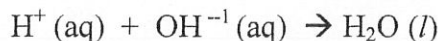
\*\*See sample problem page 730

Try the practice problem pg. 730

II. Section 17.3 *Acid - Base Titrations* (p. 732)

A. In this section you will mix a Strong Acid and a Strong Base. The Reaction is *Neutralization*.

B. In a Strong Acid you eliminate the spectator ion and same in the Strong Base, thus the net ionic equation is:



C. Some key terms: equivalence point is when the moles of acid equals moles of base  
end-point is when an indicator changes color

1. For a Strong Acid -Strong Base Titration, the equivalence point and end point are the same. The pH = 7. If a weak acid or weak base is involved they are not.

D. The pH can be calculated at different stages: \*\*see page 733 for these stages

1. Understand the graph on page 733
2. Do the practice problem on page 734

E. For a weak acid - strong base titration, at the equivalence point and end point are not equal because at the equivalence point (when moles of acid and base are equal) the strong conjugate base of the weak acid interacts with water and forms a Hydroxide ion (hydrolysis) thus the pH will be above 7 and this is when the indicator will change.

1. Finding pH at various stages are listed on pages 735 and 736
2. Look at the sample problem on page 737
3. Do the practice problem on page 737

F. For a weak Bronsted base and a strong acid, at the equivalence point (when moles of acid and base are equal) the conjugate acid of the weak base will interact with water and form hydronium ions, so the end point will be below 7.

1. The same steps can be use to calculate pOH at various stages as used to find pH

Tools of solution: Molarity =  $\frac{\text{moles}}{\text{L}}$

Henderson-Hasselbalch equation

Limiting/Excess problems

III. Sections 17.4 - 17.6 *Slightly Soluble Salts & Metallic Hydroxides* (p. 739)

A. This section deals with salts and metallic hydroxides that are almost insoluble. But the fact that they do dissociate to a very small extent is still very important on a microbiological level.

1. On page 1126 is a list of these substances.
2. As you can see their "K" values are extremely small indicating a very low concentration of ions in solution.
  
3. Remember to omit the solid from the equilibrium expression, so in these expressions there will be no denominator.
  - a) see problem page 740
  - b) see problems page 742

These equilibrium systems can be stressed and Lechatelier's Principle can be applied.