Name:_____
Blk:_____Date:_____

Chemistry 12
ACID BASE PART II Lesson # 19
BUFFERS

A BUFFER is a solution containing appreciable amounts of a weak acid and it's conjugate base !!!

usually 1 M 0.5 M

Example: $CH_3COOH + H_2O$ <-- --> $CH_3COO^- + H_3O^+$

1M 1M

If you have "appreciable" amounts then $[CH_3COOH] \cong [CH_3COO^-]$

$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = 1.8 \cdot 10^{-5}$; $[CH_3COOH] = [CH_3COO^-]$

table

CONCLUSION: $K_a = [H_3O^+]$ $\rightarrow$ $pK_a = pH$

$-\log(K_a) = -\log([H_3O^+])$

IMPT: a solution of CH_3COOH by itself IS NOT A BUFFER!!! You need substantial amounts of both weak acid and conjugate base to have a buffer!

The purpose of a BUFFER is to maintain specific pHs !!!
Therefore, the addition of small quantities of acid or base to a buffer results in a shift in the equilibrium to counter the added substance.

ex. $HCO_3^- + H_2O$ <-- --> $CO_3^{2-} + H_3O^+$

If a base is added to the above buffer system, the $[H_3O^+]$ that is present will react with the added OH^- and the buffer will shift to the products with little effect on the pH.

ex. $HCO_3^- + H_2O(\ell) \rightleftharpoons CO_3^{2-} + H_3O^+ + OH^-$

If an acid is added to this buffer system, the $[H_3O^+]$ will increase and the reaction will shift to the reactant with little effect on the pH.

ex. $HCO_3^- + H_2O(\ell) \rightleftharpoons CO_3^{2-} + H_3O^+$

NOTE: There is a LIMIT to the amount of H_3O^+ or OH^- that can be neutralized by a buffer. This is referred to as the buffer capacity ! If it is exceeded the pH will not be maintained.

There are TWO TYPES OF BUFFERS

1. acidic buffers (pH < 7)
2. basic buffers (pH > 7)

ACIDIC BUFFERS- try to maintain a pH in the ACIDIC REGION 0 to 7. An acidic buffer is made by adding appreciable amounts of a weak acid and it's conjugate base (in the form of a SOLUBLE SALT) into solution.

ex. add 1.0 M HF and 1.0 M NaF → Na^+ + F^- (NaF: 1.0 M; Na^+: 1.0 M; F^-: 1.0 M)

$HF_{1.0M} + H_2O(l) \rightleftharpoons F^-_{1.0M} + H_3O^+$

$K_a = \frac{[F^-][H_3O^+]}{[HF]} = 3.5 \cdot 10^{-4}$

$[H_3O^+] = 3.5 \cdot 10^{-4}$

$-\log(3.5 \cdot 10^{-4})$

$pH = 3.46$

This buffer would maintain a pH of 3.46.

BASIC BUFFERS- try to maintain a pH in the BASIC REGION 7 to 14. A basic buffer is made by adding apprecable amounts of a weak acid and it's conjugate base into solution.

ex. 1.0 M NH_3 and 1.0 M NH_4Cl → NH_4^+ + Cl^- (NH_4Cl: 1.0 M; NH_4^+: 1.0 M; Cl^-: 1.0 M)

NH_3 (1.0 M) $+ H_2O(l) \rightleftharpoons NH_4^+$ (1.0 M) $+ OH^-$

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \left(\frac{1.00 \cdot 10^{-14}}{5.6 \cdot 10^{-10}}\right) = 1.8 \cdot 10^{-5} = [OH^-]$

$pOH = -\log(1.8 \cdot 10^{-5})$

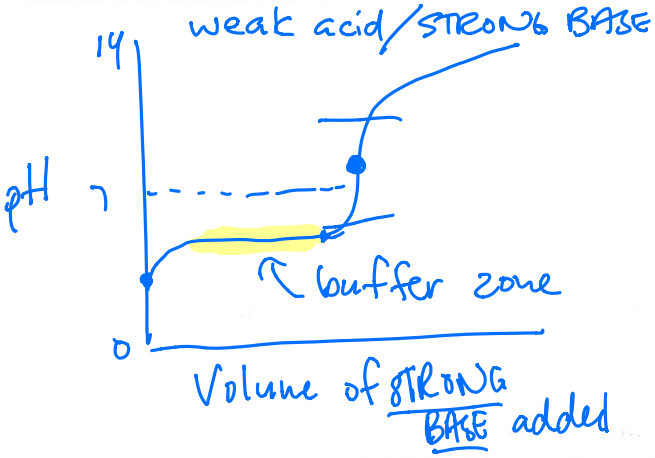
$= 4.74$

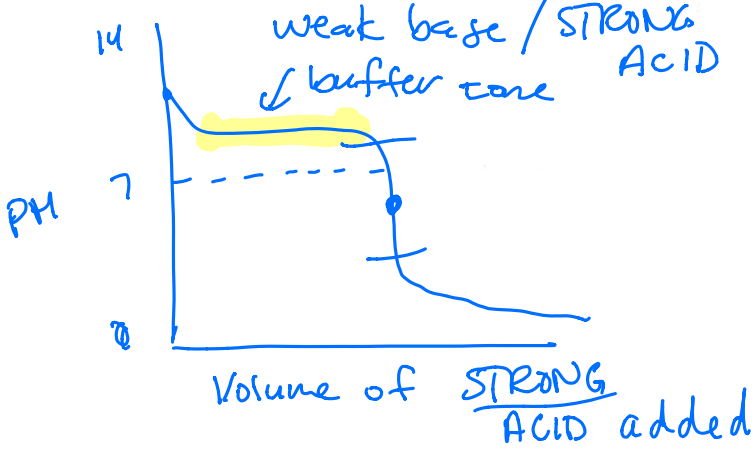
$pH = 14.000 - 4.47$

$= 9.26$

This buffer would maintain a pH of 9.26.

There is a BUFFERING REGION on both the weak acid/ STRONG BASE and weak base/STRONG ACID titration curves. These regions are illustrated below:





SEATWORK/HOMEWORK: Excercises 131-140 pgs 181-182
READ pgs 182-183 and do Exercises 141-143 pg 183
PLO's: Q1-Q6