Name: Key
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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 weak base!!!

Example: $CH_3COOH + H_2O$ <-- --> $CH_3COO^- + H_3O^+$
1M 1M

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

$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$ = if $[CH_3COOH] = [CH_3COO^-]$

CONCLUSION: $K_a = [H_3O^+]$ ∴ pka = pH of buffer!

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 a certain pH value !!! 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. $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+ \downarrow$

If an acid is added to this buffer system, the $\uparrow [H_3O^+]$ ~~that is present will react with the~~ { the new acid added ~~and the buffer~~ will shift to the REACTANTS with little effect on the pH.

ex. $HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+ \uparrow$

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 less 7)
2. BASIC BUFFERS. (pH greater than 7)

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

ex. adding appreciable amounts of HF and NaF (soluble salt)

$$HF + H_2O \rightleftharpoons F^- + H_3O^+$$
(1M) (1M)

$$K_a = \frac{[F^-][H_3O^+]}{[HF]} \Rightarrow 3.5\times10^{-4} = [H_3O^+]$$

$$-\log(3.5\times10^{-4}) = pH$$

This buffer would maintain a pH of 3.46.

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

ex. adding appreciable amounts of NH_3 and NH_4Cl (soluble salt)

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$
(1M) (1M)

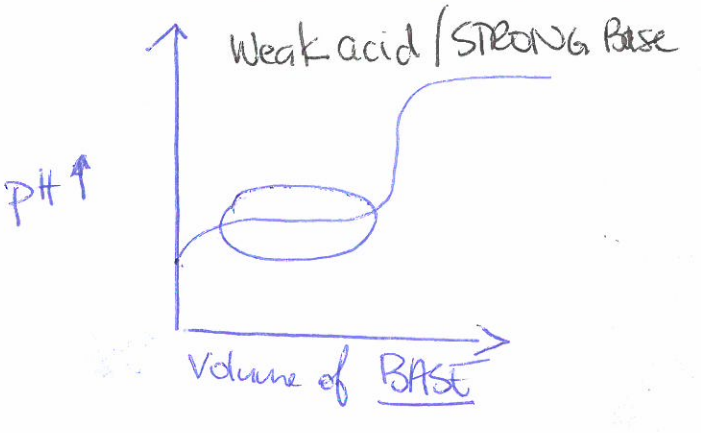
$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \Rightarrow \frac{1.00\times10^{-14}}{5.6\times10^{-10}} = [OH^-] \rightarrow -\log(1.8\times10^{-5}) = pOH$$

$$4.74 = pOH$$

$$14.000 - 4.74 = pH$$

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