

LM 21 - due tom.  
 HW 6 - due Thurs.  
 LM 22 - due Thurs.

2/28/13

Buffer

→ look at worksheets from chemical equilibria to neutralization  
 Quiz 1) The pH of a solution of a soluble salt will be:  
 - any of the above, depends on salt.

2) Calculate pH of following solution 0.15M HF, given:  $K_a = 7.2 \times 10^{-4}$ .

$$pH = -\log \sqrt{(0.15)(7.2 \times 10^{-4})} \quad \text{HF-weak acid.}$$

$$\boxed{1.983288124}$$

3.) Calculate pH of 0.15M NaF.

given  $K_a \text{ HF} = 7.2 \times 10^{-4}$

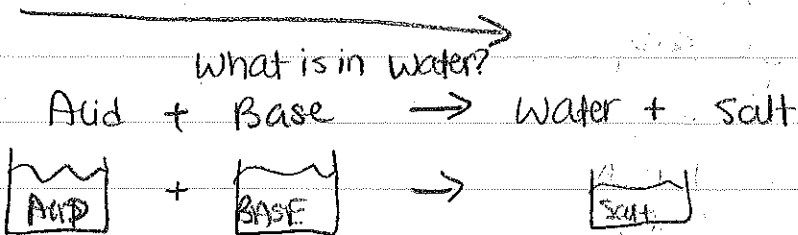
$$K_b = 1.388 \times 10^{-11}$$

$$pOH = -\log \sqrt{(M)(K_b)} = 5.84$$

$$pH = 14 - 5.84 = \boxed{8.15}$$

NaF =  $K_b$ .

Neutralization



weak

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad K_b = \frac{[OH^-][BH^+]}{[B]} \quad K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Strong

$$[acid] = [H_3O^+] \quad [base] = [OH^-] \quad \text{neutral.} \quad K_b = \frac{[OH^-][BH^+]}{[B]}$$

ionization

Neutralization

excess acid + base → water + salt

• If have excess acid, acid will dominate pH.

Acid + excess base → water + salt

• Strong excess base, pH will be dominated by excess strong base.

ionization - behavior of acid, is it going to dissociate completely

Prob. 1) pH of 0.1M aqueous solution of salts,

$\text{NaCH}_3\text{COO}$ ,  $\text{NH}_4\text{Cl}$ ,  $\text{KCl}$

↓

↓

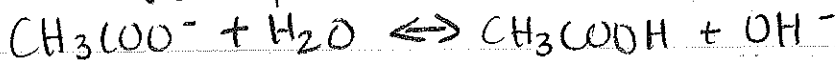
↓ neither a base or an acid

HELP!

look at  $\text{COO}^-$  → can get a H so that's why basic.

$\text{NH}_4^+$  → give up a H to be  $\text{NH}_3$ .

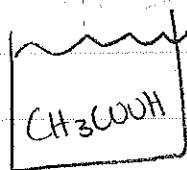
$\text{NH}_4^+$  conjugate acid pair for  $\text{NH}_3$ .



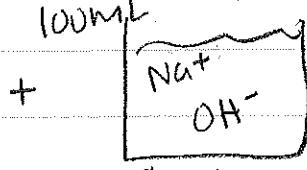
• If a strong acid or strong base won't have a  $K_b$  or  $K_a$ .



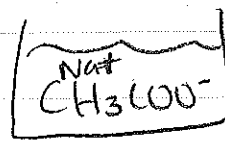
100ml



Major component

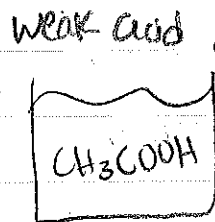


Strong base

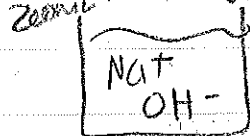


Weak base.

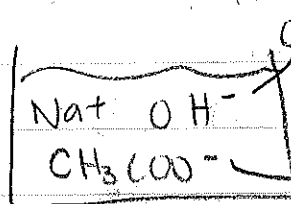
100ml



Weak acid dissociate partially.



→



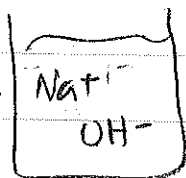
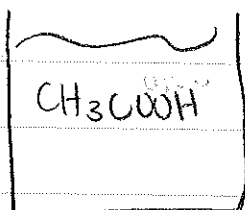
0.01mol

0.01mol

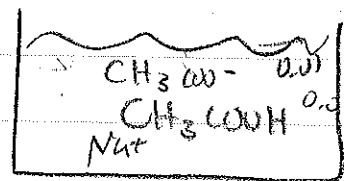
The pH of solution would be 0.11 since dominated.

200 mL

100 mL



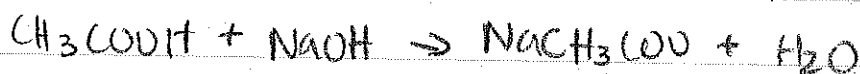
→



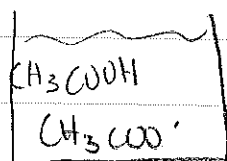
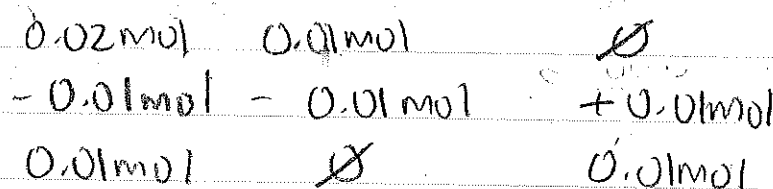
Dominant species → Na<sup>+</sup>, CH<sub>3</sub>COO<sup>-</sup>, CH<sub>3</sub>COOH

How would you determine the pH?

Neutralization.



I  
C  
E



Vol = 300 mL

Ionization:



$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{[x][0.033 + x]}{[0.033 - x]}$$

What happens to ionization in presence of common ion?

Shifts left. X will be VERY SMALL.

$$X = 1.8 \times 10^{-5} = -\log(1.8 \times 10^{-5}) = \boxed{4.74}$$

## Common Ion Effect.

The % of ionization is suppressed in the presence of a common ion.. Le Chatelier's principle.

$$\% \text{ ionization} = \frac{\text{amount ionized}}{\text{amt initial}} \times 100\%$$

? % ionization

0.1M CH<sub>3</sub>COOH

pH = 2.87

$$[\text{H}_3\text{O}^+] = 0.00135 \text{ M}$$

$$\% \text{ ionization} = \frac{0.00135}{0.1 \text{ M}} \times 100\% = 1.35\%$$

mostly CH<sub>3</sub>COOH

? % ionization

0.1M CH<sub>3</sub>COOH solution containing 0.1M NaCH<sub>3</sub>COO.

pH = 4.74

$$\frac{[\text{H}_3\text{O}^+]}{C_a} = \frac{1.8 \times 10^{-5}}{0.1} \times 100\%$$

0.018%

(can ignore x)

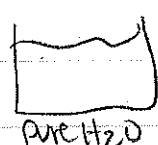
## Demo:

Carbon dioxide dissolved in pure water so its acidic. 4.28.

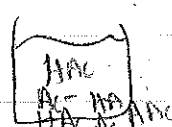
She poured in 3 drops of Sodium OH pH = 10.64

When added NaOH 1:1 acetic acid

Write down neutralization reaction for demo:



hydroxide dominated.



HAc - acid

Ac - anion.



hydroxide don't dominate.

pH doesn't change.

Buffer → if pH is constant.

Buffer - a solution in which the pH resists change when a strong acid is added.

$$pH = -\log(1.8 \times 10^{-5}) = -\log(K_a) = pK_a.$$

$$\underline{pH = pK_a.}$$

What happens if we keep adding NaOH to solution...  
eventually neutralize all acid. Excess NaOH of weak  
base. } exhaust buffer.

• Calculate pH of buffer solution.

4.74.

→ Derive a short cut formula for pH of buffers:

$$\log \left( K_a = \frac{[H_3O]^+ [A^-]}{[HA]} \right)$$

$$-\log K_a = \frac{-\log [H_3O]^+ - \log [A^-]}{-\log [HA]}$$

$$pK_a = pH - \log \frac{[A^-]_0}{[HA]_0} \left. \begin{array}{l} \leftarrow @ \text{ equilibrium.} \\ = \text{assume there are initial} \\ \text{concentration due to common} \\ \text{ion effect.} \end{array} \right\}$$

$$pH = pK_a + \log \frac{[A^-]_0}{[HA]_0} \quad \text{of buffer.}$$

If you have similar concentration of  $A^-$  &  $HA$  then pH is  
close to  $pK_a$ .

$$pOH = pK_b + \log \frac{[BH^+]}{[B]}$$

Pool

$$pH = 3.18 + \log \frac{[2]}{[1]} \quad \begin{array}{l} 100\text{mL} \\ 0.1\text{M HF} \\ 100\text{mL} \\ 0.2\text{M NaF.} \end{array}$$

Slightly more than 3.18.

Calculate ratio of molarities of acetate ions & acetic acid  
needed to buffer a solution pH = 5.25  $pK_a$   $CH_3COOH$  is 4.75