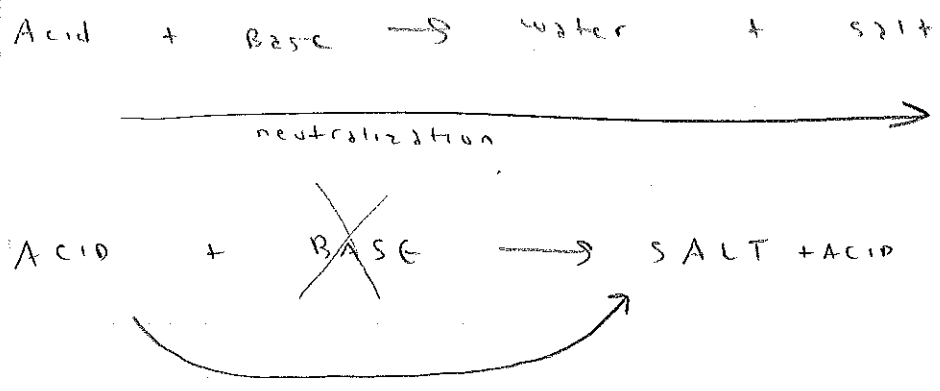


✓ c1) pH of a solution of a soluble salt will be
 - any of the above, depends on the salt

✓ c2) Calc pH of following solution, 0.15 M HF; given
 $K_a = 7.2 \times 10^{-4}$
2

✓ c3) Calc pH of following solution 0.15 M NaF , given
 K_a of HF = $K_a = 7.2 \times 10^{-4}$.
8.1 $\text{pOH} = 5.9$
 $14 - 5.9 = 8.1 = \text{pH}$

ionizing (depends on strength)



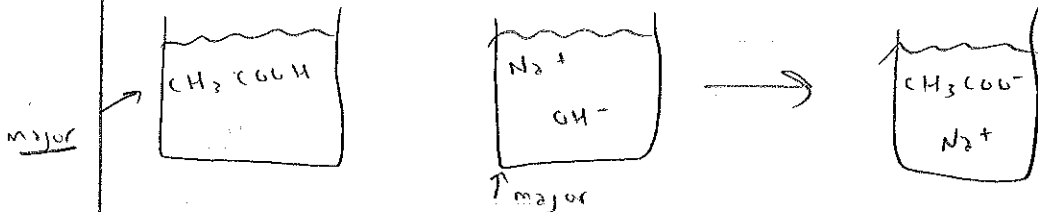
First neutralize, then go to equilibrium

c4) +10 points (pull)
 The pH of a 0.1 M aq solution of salts
 NaCH_3COO , NH_4Cl , KCl will be
Basic, Acidic, Neutral

what are the components?



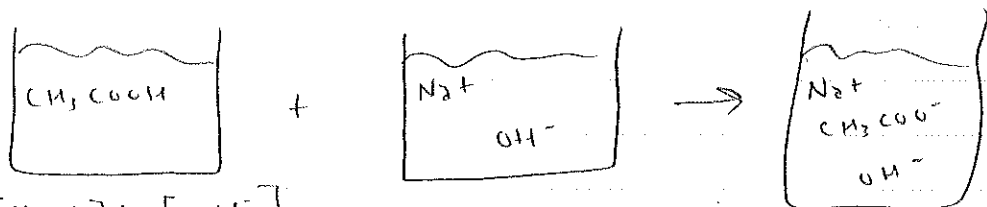
what are major species in solution as a result of mixing: 100 mL 0.1 M CH_3COOH + 100 mL 0.1 M NaOH



what are components?

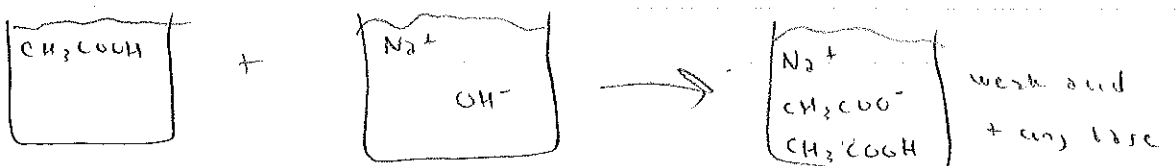


what are major species in solution as a result of mixing: 100 mL 0.1 M CH_3COOH + 200 mL 0.1 M NaOH

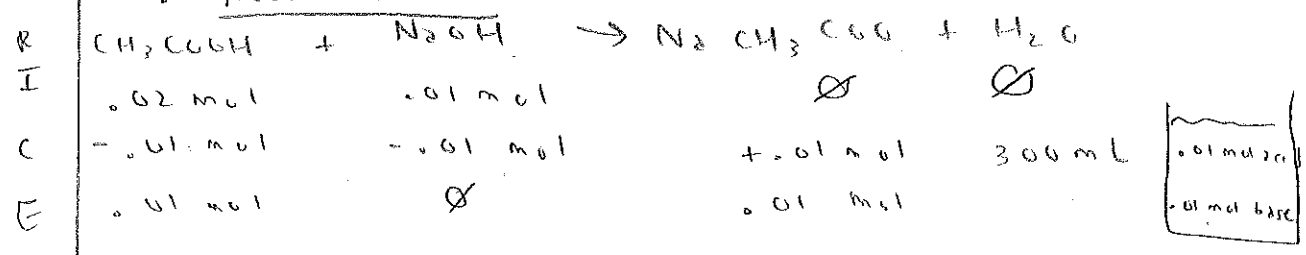


$[\text{CH}_3\text{COOH}] > [\text{OH}^-]$
both small

mixing: 200 mL 0.1 M CH_3COOH + 100 mL 0.1 M NaOH

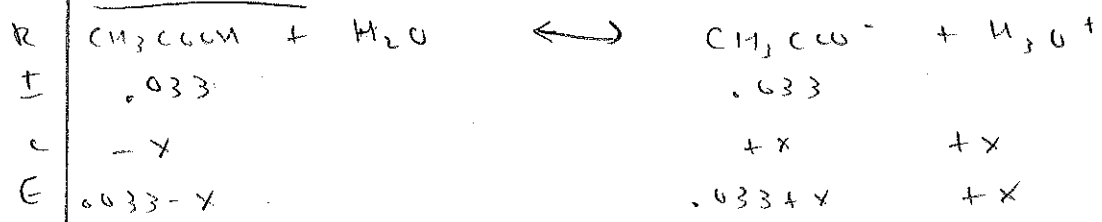


PH? Neutralization



.033 M acid/base

Ionization



$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{(x)(.033+x)}{(.033-x)}$$

x is really small b/c K_a is small + common ion!

$$1.8 \times 10^{-5} = \frac{x(.033)}{(.033)} \leftarrow \text{ignore change in } x$$

$$1.8 \times 10^{-5} = x = [H_3O^+]$$

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

Common ion effect

$$\frac{\text{amt that ionized}}{\text{starting amt}} = \% \text{ ionization} = \frac{[H_3O^+]}{C_{acid}} \times 100\%$$

0.1 M CH₃COOH

pH = 2.87

[H₃O⁺] = .00135 M

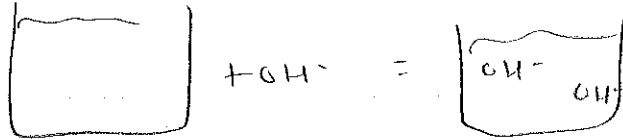
% ion = $\frac{.00135 M}{.1 M} \times 100 = 1.35\%$

.1 M CH₃COOH + .1 M NaCH₃COO

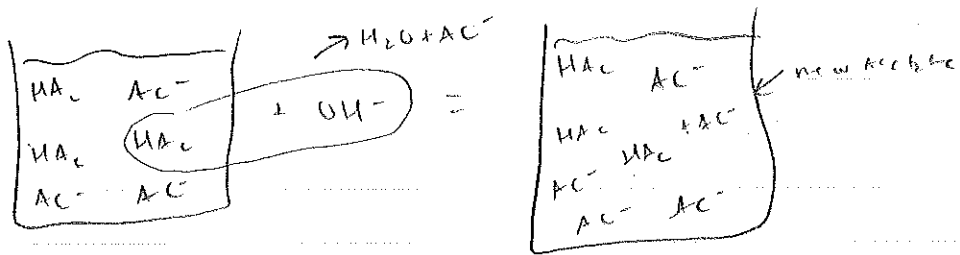
pH = 4.74

$\frac{[H_3O^+]}{C_2} = \frac{1.8 \times 10^{-5}}{.1} = .018\%$

Add NaOH to pure water!

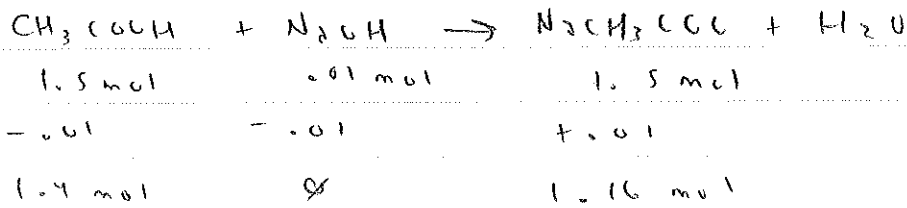


Acetic acid + acetate



pH doesn't change very much \therefore it is called a

BUFFER SOLUTION



Buffers are used to maintain pH of a particular level; resist pH change when strong acid/base is added.

Buffers can be acidic/basic; depends on conjugate acid/base pair

What happens if we keep adding NaOH to solutions

- little more basic each time; eventually neutralize

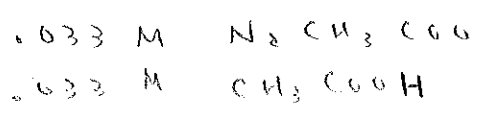
all acid \rightarrow

excess NaOH w/ weak base

pH dominated by strong base

"EXHAUST" buffer

Calc pH of buffer solution



pH = 4.74

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

same

$pK_a = -\log K_a$

$$K_a = [\text{H}_3\text{O}^+] \quad pK_a = \text{pH}$$

Formula? $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$

$$-\log K_a = -\log [\text{H}_3\text{O}^+] - \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$pK_a = \text{pH} - \log \frac{[\text{A}^-]}{[\text{HA}]}$$

Henderson Hasselbalch

$$\rightarrow \text{pH} = pK_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

← initial conc conj base
← initial conc acid

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

(5) + 10 points (pull)

pK_a of HF is 3.18. What is pH of solution of 100 mL of 0.1 M HF and 100 mL of 0.2 M NaF

$$\text{pH} = pK_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

slightly more than 3.18

adding conj base to acid, there was more base