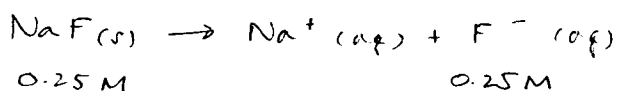


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$K_a = 3.5 \times 10^{-4}$   
 $K_b = 2.86 \times 10^{-11}$

$K_b = \frac{[HF][OH^-]}{[F^-]}$       Let  $x = [OH^-] = [HF]$   
 $[F^-] = 0.25 - x \approx 0.25$  if  $x$  is small  
 $2.86 \times 10^{-11} = \frac{x^2}{0.25}$

$x = 2.67 \times 10^{-6} M = [OH^-]$

$pOH = 5.57$

$pH = 8.43$

2) 5.5% acetic acid      0.100 M NaOH

10.0 mL

$D = 1.05 g/mL$

$5.5\% = \frac{\text{mass of acid}}{\text{mass of sample}} = \frac{x}{10.5g}$

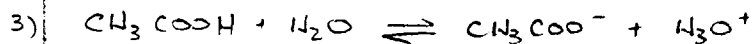
note:

$\text{mass of acid} = 0.5775g / \text{FW}$

$= 0.5775g / 60.0g/mol$

$= 9.625 \times 10^{-3} \text{ mol } CH_3COOH = \text{mol } H^+ = \text{mol } OH^- = \text{mol } NaOH$

$V = \frac{n}{M} = \frac{9.625 \times 10^{-3} \text{ mol } NaOH}{0.100M} = 9.6 \times 10^{-2} L$  or 96 mL



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

$pH = 4.50$

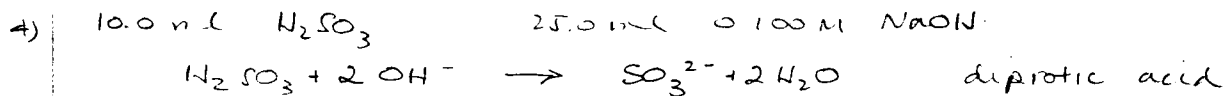
$[H_3O^+] = 3.16 \times 10^{-5} M$

$K_a = 1.8 \times 10^{-5}$

$1.8 \times 10^{-5} = \frac{[CH_3COO^-]}{[CH_3COOH]} \times 3.16 \times 10^{-5}$

$0.5696 = \frac{[CH_3COO^-]}{[CH_3COOH]} \therefore \frac{[CH_3COOH]}{[CH_3COO^-]} = 1.8$

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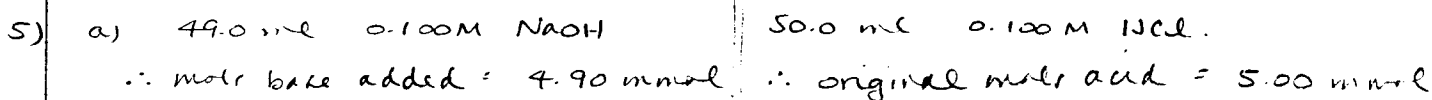


mol NaOH =  $0.100 \text{ M} \times 25.0 \text{ mL} = 2.5 \text{ mmol NaOH} = \text{mol OH}^-$

mol  $\text{OH}^- = \text{mol H}^+ = 2 (\text{mol acid})$

mol acid =  $\frac{\text{mol H}^+}{2} = \frac{2.5 \text{ mmol}}{2} = 1.25 \text{ mmol}$

$[\text{H}_2\text{SO}_3] = \frac{1.25 \text{ mmol}}{10.0 \text{ mL}} = 0.125 \text{ M}$



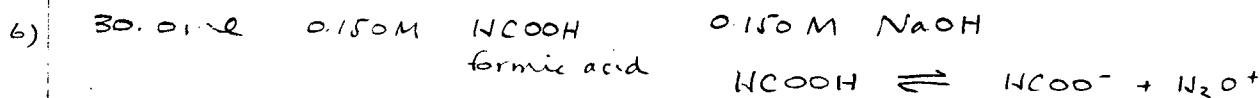
$[\text{H}^+] = \frac{5.00 \text{ mmol acid} - 4.90 \text{ mol base added}}{99.0 \text{ mL}} = 1.01 \times 10^{-3} \text{ M}$

pH = 2.996 (3 sf)



$[\text{OH}^-] = \frac{5.10 \text{ mmol base added} - 5.00 \text{ mmol acid}}{101.0 \text{ mL}} = 9.90 \times 10^{-4} \text{ M}$

pOH = 3.004      pH = 10.996



a) Let  $x = [\text{H}_3\text{O}^+] = [\text{HCOO}^-]$       ( $K_a = 1.8 \times 10^{-4}$ )  
 $[\text{HCOOH}] = 0.150 - x \approx 0.150 \text{ M}$  if  $x$  is small.

$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$

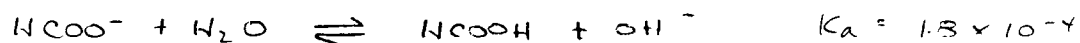
$1.8 \times 10^{-4} = \frac{x^2}{0.150}$        $x = 5.196 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]$

pH = 2.284

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b) continued

b) all  $\text{HCOOH}$  "used up"  $\rightarrow \text{HCOO}^-$  ( $\text{HCOOH} + \text{OH}^- \rightarrow \text{HCOO}^- + \text{H}_2\text{O}$ )



$$K_b = \frac{[\text{HCOOH}][\text{OH}^-]}{[\text{HCOO}^-]}$$

$$K_b = 5.56 \times 10^{-11}$$

$$[\text{salt}] = \frac{4.50 \text{ mmol}}{60.0 \text{ mL}} = 0.0750 \text{ M}$$

$$\text{Let } x = [\text{OH}^-] = [\text{HCOOH}]$$

$$[\text{HCOO}^-] = 0.0750 \text{ M} - x \approx 0.0750 \text{ if } x \text{ is small.}$$

$$5.56 \times 10^{-11} = \frac{x^2}{0.0750}$$

$$x = 2.04 \times 10^{-6} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = 5.690$$

$$\text{pH} = 8.310$$

c) original moles acid  $30.0 \text{ mL} \times 0.150 \text{ M} = 4.50 \text{ mmol}$   
 moles base added  $40.0 \text{ mL} \times 0.150 \text{ M} = 6.00 \text{ mmol}$

$$[\text{OH}^-] = \frac{6.00 \text{ mmol} - 4.50 \text{ mmol}}{70.0 \text{ mL}} = 0.0214 \text{ M}$$

$$\text{pOH} = 1.669$$

$$\text{pH} = 12.331$$