

Part I

Solution.

Titration of formic Acid with NaOH.

Volume of Formic Acid = 100 mL.

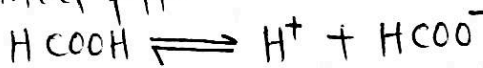
$[HCOOH] = 0.50 M.$

pK_a of $HCOOH = 3.75.$

$K_a = 10^{-3.75}$

$K_a = 1.778 \times 10^{-4}.$

① Initial pH.



I 0.50 0 0

C -0.50α $+0.50\alpha$ $+0.50\alpha$

E $0.50 - 0.50\alpha$ 0.50α 0.50α

Where, α = degree of dissociation
 α = degree of ionization.

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

$$K_a = \frac{(0.50\alpha)(0.50\alpha)}{0.50 - 0.50\alpha} = \frac{(0.50\alpha)^2}{(0.50)(1-\alpha)}$$

$$1.78 \times 10^{-4} = \frac{(0.50\alpha)^2}{(1-\alpha)}$$

α is very small.

$$1.778 \times 10^{-4} = \frac{0.50\alpha^2}{1}$$

$$\alpha^2 = 3.556 \times 10^{-4}$$

$$\alpha = 0.018857$$

now, $[H^+] = 0.50\alpha$

$$[H^+] = (0.50 \times 0.018857)$$

$$[H^+] = 0.0094286 M.$$

Now, $pH = -\log[H^+] = -\log[0.0094286]$

$$pH = 2.0255$$

$$pH = 2.02$$

Initial $pH = 2.02$

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② Percent of ionization

$$= \left(\frac{0.018857}{1} \times 100\% \right)$$

$$= 1.8857\%$$

$$= \underline{\underline{1.886\%}}$$

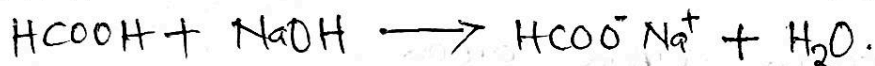
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③. Addition of 10ML, NaOH.

$$[\text{NaOH}] = 1.0\text{M}$$

$$\text{moles of NaOH added} = \left(\frac{1.0\text{M} \times 10\text{ML}}{1000} \right) = 0.01 \text{ moles.}$$

Reaction



$$\text{moles of HCOO}^- \text{ formed} = 0.01 \text{ moles}$$

$$\text{Initial moles of HCOOH} = \text{molarity} \times \text{volume (L).}$$

$$= \left(\frac{0.50\text{M} \times 100}{1000} \right) = 0.05$$

$$\text{Final moles of HCOOH} = (0.05 - 0.01) = 0.04 \text{ moles.}$$

HCOOH/HCOO⁻ makes an acidic buffer solution.

$$\text{pH} = \text{pKa} + \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$\text{pH} = 3.75 + \log \frac{[0.01]}{[0.04]}$$

$$\text{pH} = 3.75 + \log (0.25)$$

$$\text{pH} = 3.75 - 0.602$$

$$\text{pH} = 3.15.$$

④. 25ml NaOH added.

$$\text{moles of NaOH} = \frac{(1.0\text{M} \times 25\text{ml})}{1000} = 0.025 \text{ moles}$$

$$\text{moles of HCOO}^- \text{ formed} = 0.025 \text{ moles.}$$

$$\text{remaining moles of HCOOH} = (0.05 \text{ moles} - 0.025 \text{ moles}) \\ = 0.025 \text{ moles}$$

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

$$\text{pH} = \text{pK}_a + \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$\text{pH} = \text{pK}_a + \log \frac{[0.025]}{[0.025]}$$

$$\text{pH} = \text{pK}_a + 0 = 3.75$$

$$\text{pH} = 3.75$$

⑤ At equivalence point volume required

$$\text{moles of Formic Acid} = \frac{(0.50\text{M} \times 100\text{ml})}{1000} \\ = 0.05 \text{ moles}$$

$$\text{moles of NaOH} = \text{moles of Formic Acid} \\ = 0.05 \text{ moles.}$$

∴ volume of NaOH at equivalence point.

$$= \frac{(0.05 \text{ moles} \times 1000)}{1.0\text{M}}$$

$$= \underline{\underline{50 \text{ ml}}}$$

6. At equivalence point there is HCOO^-

$$\text{Total volume} = (100 + 50) \text{ mL} \\ = 150 \text{ mL}$$



$$[\text{HCOO}^-] = \left(\frac{0.05 \text{ mol} \times 1000}{150} \right) \\ = \underline{\underline{0.333 \text{ M}}}$$

I	0.333	0	0
C	-x	+x	+x
E	0.333-x	x	x

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

$$\frac{K_w}{K_a} = \frac{x \cdot x}{0.333 - x}$$

$$\frac{K_w}{K_a} = \frac{x^2}{0.333}$$

$$x^2 = 0.333 \times \frac{K_w}{K_a}$$

$$x^2 = \left(\frac{0.333 \times 1.0 \times 10^{-14}}{1.778 \times 10^{-4}} \right)$$

$$x^2 = 1.8797 \times 10^{-11}$$

$$x = 4.3298 \times 10^{-6}$$

$$[\text{OH}^-] = x = 4.3298 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log [4.3298 \times 10^{-6}]$$

$$\text{pOH} = 5.36 \text{ at equivalence point.}$$

⑦ pH at equivalence point

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ &= 14 - 5.36 \\ &= \underline{\underline{8.64}} \end{aligned}$$

⑧ After 100ml addition of NaOH

50ml is equivalence point.

So extra 50ml NaOH is added after end point.

$$\begin{aligned} \text{extra moles of NaOH} &= \left(\frac{50 \text{ mL} \times 1.0 \text{ M}}{1000} \right) \\ &= \underline{\underline{0.05 \text{ moles}}} \end{aligned}$$

$$\text{total volume} = 100 + 100 = 200 \text{ mL}$$

$$\begin{aligned} [\text{OH}^-] \text{ final} &= \left(\frac{0.05 \text{ moles} \times 1000}{200} \right) \\ &= \underline{\underline{0.25 \text{ M}}} \end{aligned}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{[0.25]}$$

$$[\text{H}^+] = 4 \times 10^{-14}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$= -\log [4 \times 10^{-14}]$$

$$\text{pH} = 13.40$$

$$\text{pH} = 13.40$$

$$\textcircled{9} \text{ Molarity} = \left(\frac{\text{moles} \times 1000}{\text{volume}} \right)$$

moles = 0.05 moles of HCl.

Volume = 100 mL

$$= \left(\frac{0.05 \text{ moles} \times 1000}{100} \right)$$

$$= \underline{\underline{0.5 \text{ M}}}$$

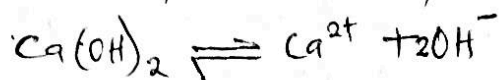
Percentage of dissociation of formic Acid.

$$= \left(\frac{0.5}{1.0} \right) \times 100\%$$

$$= \underline{\underline{50\%}}$$

Part II. Solution.

①. Reaction for solubility of $\text{Ca}(\text{OH})_2$.



②. Molar solubility of $\text{Ca}(\text{OH})_2$.

$$K_{sp}[\text{Ca}^{2+}][\text{OH}^-]^2 = (s)(2s)^2 = 4s^3$$

$$4s^3 = 4.68 \times 10^{-6} \quad s = \sqrt[3]{\frac{4.68 \times 10^{-6}}{4}}$$

$$s = \underline{\underline{1.05 \times 10^{-2} \text{ M}}}$$

③. Concentration of $\text{OH}^- = 2s = 2(1.05 \times 10^{-2} \text{ M})$
 $= \underline{\underline{2.10 \times 10^{-2} \text{ M}}}$

④. $\text{pH} = 14 - \text{pOH}$
 $\text{pOH} = -\log[\text{OH}^-]$

$$\text{pH} = 14 - [-\log[2.10 \times 10^{-2} \text{ M}]]$$

$$= (14 - 1.677)$$

$$= \underline{\underline{12.32}}$$

⑤. a. Decreasing pH \rightarrow Right
decreasing pH means reducing OH^- concentration

b. Add $\text{CaCl}_2 \rightarrow$ Left
(increasing Ca^{2+} conc. would shift the equilibrium to the left.)

c. Decreasing temperature \rightarrow Left.
(reaction is exothermic).

d. Add $\text{NaBr} \rightarrow$ Right
(Ca^{2+} would get consumed).