

# ALMUSTAQBAL UNIVERSITY COLLEGE

## College of Medical and Health Techniques

### Medical Laboratories Techniques Department

Stage : First year students

Subject : Chemistry 1 - Lecture 8

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## Buffer solutions

A buffer solution is a combination of either a weak acid and its conjugate base (supplied by a salt) or a weak base and its conjugate acid (supplied by a salt). Buffer systems resist changes in the pH of a solution due to dilution or the addition of small amounts of strong acid or base. Buffers are used to maintain the pH of solutions at a relatively constant and predetermined level. Usually, buffers have a useful pH range =  $pK \pm 1$ , Buffers play a vital role in maintaining the pH of body fluids.

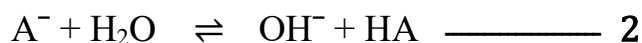
### Calculation of the pH of different types of Buffer solutions

#### 1. Acidic buffers

Consists of weak acid (HA) and its salt ( $A^-$ ). Typical example is (acetic acid – acetate salt (  $CH_3COOH - CH_3COO^-$  )



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



$$K_b = \frac{[OH][HA]}{[A^-]} = \frac{K_w}{K_a}$$

equilibrium ① will decrease  $C_{HA}$  by the amount  $[H_3O^+]$  and equilibrium ② will increase it by the amount  $[OH^-]$ .

$$[HA]_{\text{equil.}} = C_{HA} - [H_3O^+] + [OH^-]$$

Similarly, equilibrium ① will increase  $[A^-]$  by the amount  $[H_3O^+]$

while equilibrium ② will decrease  $[A^-]$  by the amount  $[OH^-]$  then

$$[A^-]_{\text{equil.}} = C_{A^-} + [H_3O^+] - [OH^-]$$

As we have acid then

$$[HA] = C_{HA} - [H_3O^+]$$

$$[A^-] = C_{A^-} + [H_3O^+]$$

And because we have weak acids then

$$[HA] \cong C_{HA}$$

$$[A^-] \cong C_{A^-}$$

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

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

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

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

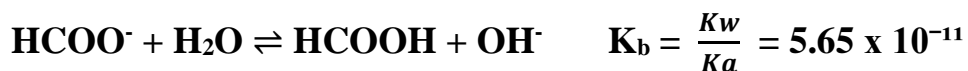
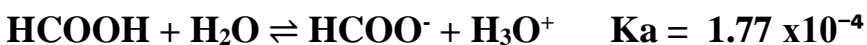
$$\text{pH} = \text{pK}_a + \log \frac{C_A}{C_{\text{HA}}}$$

$$** \quad \text{pH} = \text{pK}_a + \log \frac{C_{\text{salt}}}{C_{\text{acid}}} \quad (\text{Henderson equation})$$

**Example:** what is the pH of a solution that is (0.40 M) in formic acid ( $K_a = 1.77 \times 10^{-4}$ ) and (1.0 M) in sodium formate? compare it with the pH of 0.4 M formic acid solution.

**Solution :**

The pH of the solution will be affected by  $K_a$  of formic acid ( $\text{HCOOH}$ ) and  $K_b$  of formate ion ( $\text{HCOO}^-$ )



Since  $K_a$  of formic acid  $\gg K_b$  for formate the solution will be acidic and  $K_a$  will determine the  $\text{H}_3\text{O}^+$  concentration.

$$\text{pH} = \text{pK}_a + \log \frac{C_{\text{salt}}}{C_{\text{acid}}}$$

$$\text{pK}_a = -\log K_a = -\log (1.77 \times 10^{-4}) = 3.75$$

$$\text{pH} = 3.75 + \log \frac{1.0}{0.4} = 4.14$$

Check if  $\frac{[\text{H}_3\text{O}^+]}{[\text{HCOOH}]} \times 100 \% < 10 \%$  Then approximation is valid

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.14} = 7.2 \times 10^{-5}$$

$$\frac{7.2 \times 10^{-5}}{0.4} \times 100\% = 0.018\% \text{ (The approximation is valid)}$$

B) The pH of the acid-only solution

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \cdot C}$$

$$[\text{H}_3\text{O}^+] = \sqrt{1.77 \times 10^{-4} \times 0.4} = 8.41 \times 10^{-3}$$

$$\text{pH} = -\log(8.41 \times 10^{-3}) = 2.07$$

Practice Exercise :

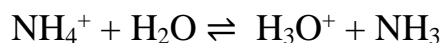
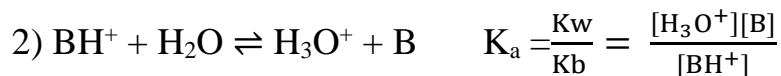
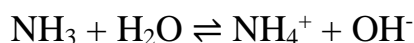
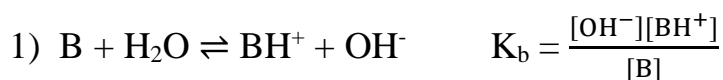
1. What is the pH of a solution containing 0.3 M HCOOH and 0.52 M HCOOK? Compare your result with the pH of a 0.3 M HCOOH solution ( $K_a = 1.77 \times 10^{-4}$ ).

2. Determine the pH of (a) a 0.40 M CH<sub>3</sub>COOH solution ( $K_a = 1.8 \times 10^{-5}$ ),

(b) a solution that is 0.40 M CH<sub>3</sub>COOH and 0.20 M CH<sub>3</sub>COONa.

## B) Basic Buffers

It is composed of a solution of a weak base (B) and its conjugate acid (Salt) BH<sup>+</sup>  
e.g : NH<sub>3</sub>-NH<sub>4</sub>Cl .



[B] will decrease in equilibrium ① by amount [OH<sup>-</sup>] & increase in equilibrium ② by [H<sub>3</sub>O<sup>+</sup>]

$$\text{Then } [B] = C_B - [OH^-] + [H_3O^+]$$

Similarly  $[BH^+]$  will increase in equilibrium ① By  $[OH^-]$  and decrease in equilibrium ② by  $[H_3O^+]$ .

$$\text{Then } [BH^+] = C_{BH^+} + [OH^-] - [H_3O^+]$$

$$[B] \equiv C_B \quad \text{and} \quad [BH^+] \equiv C_{BH^+} \quad (\text{by approximation})$$

$$K_b = \frac{[OH^-][BH^+]}{[B]}$$

$$[OH^-] = K_b \frac{C_B}{C_{BH^+}}$$

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

$$pOH = pK_b + \log \frac{C_{\text{salt}}}{C_{\text{base}}} \quad (\text{Henderson equation})$$

In General:

$$\text{pH} = pK_a + \log \frac{C_{\text{salt}}}{C_{\text{acid}}} \quad (\text{for acidic buffer})$$

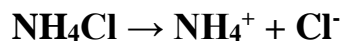
$$\text{pOH} = pK_b + \log \frac{C_{\text{salt}}}{C_{\text{base}}} \quad (\text{for basic buffer})$$

$$pH = 14 - pOH$$

### Example:

Calculate the pH of a solution that is 0.1 M in  $NH_3$  ( $K_b = 1.75 \times 10^{-5}$ ) and 0.15 M in  $NH_4Cl$ .

### Solution:



$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+ \quad K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.75 \times 10^{-5}} = 5.7 \times 10^{-10}$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \quad K_b = 1.75 \times 10^{-5}$$

because  $K_b \gg K_a$  the solution is assumed to be basic

$$\text{pOH} = \text{pK}_b + \log \frac{C_{\text{salt}}}{C_{\text{base}}} \quad (\text{for basic buffer})$$

$$\text{pK}_b = -\log K_b = -\log (1.75 \times 10^{-5}) = 4.75$$

$$\text{pOH} = 4.75 + \log \frac{0.15}{0.1} = 4.93$$

To check the validity of approximation we calculate  $[\text{OH}^-]$

then Check if  $\frac{[\text{OH}^-]}{[\text{Base}]} \times 100 \% < 10 \%$  Then approximation is valid

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-4.93} = 1.17 \times 10^{-5}$$

$$\text{Then } \frac{1.17 \times 10^{-5}}{0.2} \times 100 \% = 5.85 \times 10^{-3} \quad (\text{approximation is valid})$$

$$\text{pH} = 14 - 4.93 = 9.07$$

Exercise:

Determine the pH of (a) a 0.20 M  $\text{NH}_3$  solution, (b) a solution that is 0.20 M in  $\text{NH}_3$  and 0.30 M  $\text{NH}_4\text{Cl}$ .

### Properties of buffer solution:

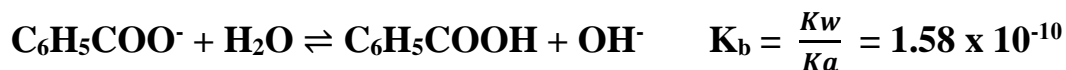
#### ① Effect of dilution:

The pH of the buffer solution remains independent of dilution until the concentration of species it contains is decreased to the point where the approximation mentioned above becomes invalid.

**Example:** calculate the change in pH of a buffer containing (0.4M) Benzoic acid  $\text{C}_6\text{H}_5\text{COOH}$  ( $K_a = 6.3 \times 10^{-5}$ ) and (1M) sodium benzoate  $\text{C}_6\text{H}_5\text{COONa}$  after dilution by a factor of 10 times .

#### Solution:





Since  $K_a$  of benzoic acid  $\gg K_b$  for benzoate the solution will be acidic.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{C}_6\text{H}_5\text{COONa}]}{[\text{C}_6\text{H}_5\text{COOH}]}$$

$$\text{p}K_a = -\log(6.3 \times 10^{-5}) = 4.2$$

**a. Before Dilution**

$$\text{pH} = 4.2 + \log \frac{1}{0.4} = 4.59$$

Check if  $\frac{[\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_5\text{COOH}]} \times 100\% < 10\%$  Then approximation is valid

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.59} = 2.57 \times 10^{-5}$$

$$\frac{2.57 \times 10^{-5}}{0.4} \times 100\% = 0.0064\% \quad (\text{The approximation is valid})$$

**b. After dilution with 10 times.**

$$M_1 V_1 = M_2 V_2$$

for  $[\text{C}_6\text{H}_5\text{COOH}]$

$$0.4 \times 1 = M_2 \times 10$$

$$[\text{C}_6\text{H}_5\text{COOH}] = \frac{0.4}{10} = 0.04\text{M}$$

For  $[\text{C}_6\text{H}_5\text{COO}^-]$

$$1 \times 1 = M_2 \times 10$$

$$[\text{C}_6\text{H}_5\text{COO}^-] = \frac{1}{10} = 0.1\text{M}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{C}_6\text{H}_5\text{COONa}]}{[\text{C}_6\text{H}_5\text{COOH}]} = 4.2 + \log \frac{0.1}{0.04} = 4.59$$

Check if  $\frac{[\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_5\text{COOH}]} \times 100 \% < 10 \%$  Then approximation is valid

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.59} = 2.57 \times 10^{-5}$$

$$\frac{2.57 \times 10^{-5}}{0.04} \times 100\% = 0.0064 \% \quad (\text{The approximation is valid})$$

$\therefore$  NO change in pH occurs after 10 times dilution(  $\Delta\text{PH} = 0$ )

## ② Effect of adding strong acid or base:

Buffer solution resists the pH change after the addition of a small amount of strong acid or base.

### Example:

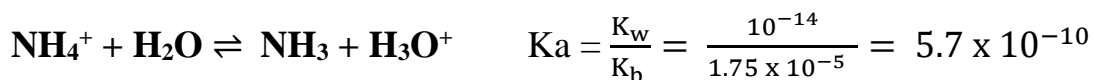
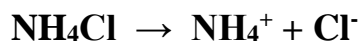
Calculate the pH change that takes place when 100 mL of:

a) 0.05 M NaOH                      b) 0.05 M HCl

is added separately to 400 mL of buffer solution of (0.3M)  $\text{NH}_4\text{Cl}$  and (0.2M)  $\text{NH}_3$  ( $K_b = 1.75 \times 10^{-5}$ ).

### Solution:

#### **A. The original buffer before addition**



because  $K_b \gg K_a$  the solution is assumed to be basic

$$\text{pOH} = \text{p}K_b + \log \frac{C_{\text{salt}}}{C_{\text{base}}} \quad (\text{for basic buffer})$$

$$\text{p}K_b = -\log K_b = -\log (1.75 \times 10^{-5}) = 4.76$$

$$\text{pOH} = \text{p}K_b + \log \frac{C_{\text{NH}_4\text{Cl}}}{C_{\text{NH}_3}} = 4.76 + \log \frac{0.3}{0.2} = 4.93$$



To check the validity of approximation we calculate  $[\text{OH}^-]$  then

Check if  $\frac{[\text{OH}^-]}{[\text{Base}]} \times 100 < 10\%$  Then approximation is valid

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-4.93} = 1.17 \times 10^{-5}$$

$$\text{Then } \frac{1.17 \times 10^{-5}}{0.2} \times 100\% = 5.85 \times 10^{-3} \% \text{ (approximation is valid)}$$

$$\text{pH} = 14 - 4.93 = 9.07$$

## B. after the addition of a strong base or acid

1) addition of NaOH converts part of  $\text{NH}_4^+$  in the buffer to  $\text{NH}_3$



The analytical concentration of  $\text{NH}_3$  and  $\text{NH}_4\text{Cl}$  become :

$$C_{\text{NH}_3} = \frac{\text{original No. of moles of NH}_3 + \text{moles of produced NH}_3}{\text{Total New volume (L)}}$$

$$\text{Or } C_{\text{NH}_3} = \frac{\text{original No. of mmoles of NH}_3 + \text{mmoles of produced NH}_3}{\text{Total New volume (mL)}}$$

No. of moles of produced  $\text{NH}_3$  = No. of moles of reacted NaOH

$$C_{\text{NH}_3} = \frac{M_{\text{NH}_3} V_{\text{NH}_3} + M_{\text{NaOH}} V_{\text{NaOH}}}{V_{\text{NH}_3} + V_{\text{NaOH}}}$$

$$C_{\text{NH}_3} = \frac{400 \times 0.2 + 100 \times 0.05}{[400 + 100]} = \frac{85}{500} = 0.17\text{M}$$

$$C_{\text{NH}_4^+} = \frac{\text{original No. of moles of NH}_4^+ - \text{moles of reacted NH}_4^+}{\text{Total New volume (L)}}$$

$$\text{Or } C_{\text{NH}_4^+} = \frac{\text{original No. of mmoles of NH}_4^+ - \text{mmoles of reacted NH}_4^+}{\text{Total New volume (mL)}}$$

No. of moles of consumed  $\text{NH}_4^+$  = No. of moles of reacted NaOH

$$C_{\text{NH}_4\text{Cl}} = \frac{M_{\text{NH}_4\text{Cl}} V_{\text{NH}_4\text{Cl}} - M_{\text{NaOH}} V_{\text{NaOH}}}{V_{\text{NH}_4\text{Cl}} + V_{\text{NaOH}}}$$

$$C_{\text{NH}_4\text{Cl}} = \frac{400 \times 0.3 - 100 \times 0.05}{[400 + 100]} = \frac{115}{500} = 0.23\text{M}$$

$$\text{pOH} = 4.76 + \log \frac{0.23}{0.17} = 4.89 \quad (\text{Henderson equation})$$

To check the validity of the approximation we calculate

if  $\frac{[\text{OH}^-]}{[\text{Base}]} \times 100 < 10\%$  Then the approximation is valid

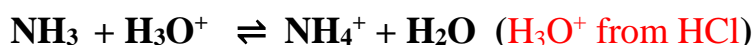
$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-4.89} = 1.29 \times 10^{-5}$$

$$\text{Then } \frac{1.29 \times 10^{-5}}{0.17} \times 100\% = 7.59 \times 10^{-3} \quad (\text{approximation is valid})$$

$$\text{pH} = 14 - 4.89 = 9.11$$

$$\Delta\text{pH} = 9.11 - 9.07 = 0.04$$

2) The addition of HCl converts part of  $\text{NH}_3$  to  $\text{NH}_4\text{Cl}$



$$C_{\text{NH}_3} = \frac{\text{original No. of moles of NH}_3 - \text{moles of reacted NH}_3}{\text{Total New volume (L)}}$$

$$\text{Or } C_{\text{NH}_3} = \frac{\text{original No. of mmoles of NH}_3 - \text{mmoles of reacted NH}_3}{\text{Total New volume (mL)}}$$

No. of moles of consumed  $\text{NH}_3$  = No. of moles of reacted HCl

$$C_{\text{NH}_3} = \frac{M_{\text{NH}_3} V_{\text{NH}_3} - M_{\text{HCl}} V_{\text{HCl}}}{V_{\text{NH}_3} + V_{\text{HCl}}}$$

$$C_{\text{NH}_3} = \frac{400 \times 0.2 - 100 \times 0.05}{[400 + 100]} = 0.150\text{M}$$

$$C_{\text{NH}_4^+} = \frac{\text{original No of moles of NH}_4^+ + \text{moles of produced NH}_4^+}{\text{Total New volume (L)}}$$

$$\text{Or } C_{\text{NH}_4^+} = \frac{\text{original No of mmoles of NH}_4^+ + \text{mmoles of produced NH}_4^+}{\text{Total New volume (mL)}}$$

No. of moles of produced  $\text{NH}_4$  = No. of moles of reacted HCl

$$C_{\text{NH}_4\text{Cl}} = \frac{M_{\text{NH}_4\text{Cl}} V_{\text{NH}_4\text{Cl}} + M_{\text{HCl}} V_{\text{HCl}}}{V_{\text{NH}_4\text{Cl}} + V_{\text{HCl}}}$$

$$C_{NH_4Cl} = \frac{400 \times 0.3 + 100 \times 0.05}{[400+100]} = 0.25M$$

$$pOH = pK_b + \log \frac{C_{NH_4Cl}}{C_{NH_3}} \quad (\text{Henderson equation})$$

$$pOH = 4.76 + \log \frac{0.25}{0.15} = 4.98$$

To check the validity of approximation we calculate  $[OH^-]$  then

Check if  $\frac{[OH^-]}{[Base]} \times 100 < 10\%$  Then approximation is valid

$$[OH^-] = 10^{-pOH} = 10^{-4.98} = 1.05 \times 10^{-5}$$

$$\text{Then } \frac{1.05 \times 10^{-5}}{0.15} \times 100\% = 0.007\% \quad (\text{approximation is valid})$$

$$pH = 14 - 4.98 = 9.02$$

$$\Delta pH = 9.02 - 9.07 = -0.05$$

Addition	$\Delta pH$
100 mL 0.05 M NaOH	0.04
100 mL 0.05 M HCl	- 0.05

### Example

Calculate the pH change that take place when 100 mL of:

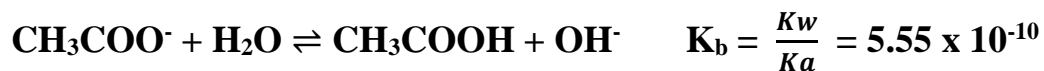
- a) 0.05 M NaOH                      b) 0.05 M HCl

is added separately to 400 mL of buffer solution of (0.1M)  $CH_3COOH$  and (0.2M)  $CH_3COONa$  ( $K_a = 1.8 \times 10^{-5}$ ).

solution:

a. The pH of the buffer before addition:





Since  $K_a$  of acetic acid  $\gg K_b$  for acetate the solution will be acidic and  $K_a$  will determine the  $\text{H}_3\text{O}^+$  concentration.

$$\text{pH} = \text{p}K_a + \log \frac{C_{\text{salt}}}{C_{\text{acid}}}$$

$$\text{p}K_a = -\log K_a = -\log (1.8 \times 10^{-5}) = 4.74$$

$$\text{pH} = 4.74 + \log \frac{0.2}{0.1} = 5.04$$

Check if  $\frac{[\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \times 100 < 10\%$  Then approximation is valid

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-5.06} = 8.7 \times 10^{-6}$$

$$\frac{8.7 \times 10^{-6}}{0.1} \times 100 = 0.009\% \text{ (The approximation is valid)}$$

## **B. addition of 100 mL of 0.05 M NaOH**



From the above reaction  $[\text{CH}_3\text{COOH}]$  is decreased and  $[\text{CH}_3\text{COONa}]$  is increased

$$C_{\text{CH}_3\text{COONa}} = \frac{\text{original No. of mmoles of CH}_3\text{COONa} + \text{mmoles of produced CH}_3\text{COONa}}{\text{Total New volume (mL)}}$$

No. of moles of  $\text{CH}_3\text{COONa}$  produced = No. of moles of  $\text{NaOH}$  reacted

$$C_{\text{CH}_3\text{COONa}} = \frac{\text{original No. of mmoles of CH}_3\text{COONa} + \text{mmoles of NaOH reacted}}{\text{Total New volume (mL)}}$$

$$C_{\text{CH}_3\text{COONa}} = \frac{400 \times 0.2 + 100 \times 0.05 \text{ mmoles}}{400 + 100} = 0.17$$

$$C_{\text{CH}_3\text{COOH}} = \frac{\text{original No. of mmoles of CH}_3\text{COOH} - \text{mmoles of CH}_3\text{COOH consumed}}{\text{Total New volume (mL)}}$$

No. of moles of CH<sub>3</sub>COOH consumed = No. of moles of NaOH reacted

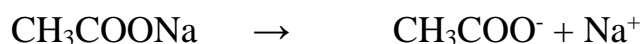
$$C_{\text{CH}_3\text{COOH}} = \frac{\text{original No. of mmoles of CH}_3\text{COOH} - \text{mmoles of NaOH reacted}}{\text{Total New volume (mL)}}$$

$$C_{\text{CH}_3\text{COOH}} = \frac{(400 \times 0.1 - 100 \times 0.05) \text{ mmoles}}{\text{Total New volume (mL)}} = \frac{(35) \text{ mmoles}}{500} = 0.07$$

$$\text{pH} = 4.74 + \log \frac{0.17}{0.07} = 5.12$$

$$\Delta \text{pH} = 5.12 - 5.06 = 0.06$$

**c) Addition of 100 mL of 0.05 HCl to the buffer**



**From the above reaction [CH<sub>3</sub>COONa] is decreased and [CH<sub>3</sub>COOH] is increased**

$$C_{\text{CH}_3\text{COONa}} = \frac{\text{original No. of mmoles of CH}_3\text{COONa} - \text{mmoles of CH}_3\text{COONa consumed}}{\text{Total New volume (mL)}}$$

No. of moles of CH<sub>3</sub>COONa consumed = No. of moles of HCl reacted

$$C_{\text{CH}_3\text{COONa}} = \frac{\text{original No. of mmoles of CH}_3\text{COONa} - \text{mmoles of HCl reacted}}{\text{Total New volume (mL)}}$$

$$C_{\text{CH}_3\text{COONa}} = \frac{400 \times 0.2 - 100 \times 0.05 \text{ mmoles}}{400 + 100} = 0.15$$

$$C_{\text{CH}_3\text{COOH}} = \frac{\text{original No. of mmoles of CH}_3\text{COOH} + \text{mmoles of CH}_3\text{COOH produced}}{\text{Total New volume (mL)}}$$

No. of moles of CH<sub>3</sub>COOH produced = No. of moles of HCl reacted

$$C_{\text{CH}_3\text{COOH}} = \frac{\text{original No. of mmoles of CH}_3\text{COOH} + \text{mmoles of HCl reacted}}{\text{Total New volume (mL)}}$$

$$C_{CH_3COOH} = \frac{(400 \times 0.1 + 100 \times 0.05) \text{ m moles}}{\text{Total New volume (mL)}} = \frac{(45) \text{ m moles}}{500} = 0.09$$

$$\text{pH} = 4.74 + \log \frac{0.15}{0.09} = 4.96$$

$$\Delta \text{pH} = 4.96 - 5.06 = -0.1$$

Addition	$\Delta \text{pH}$
100 mL 0.05 M NaOH	0.06
100 mL 0.05 M HCl	- 0.1

### **Exercise:**

1. Calculate the pH of the buffer solution of 0.40 M HF ( $K_a = 3.2 \times 10^{-4}$ ) and 0.48 M KF. What is the pH after the addition of 15 mL of 0.10 M HCl to 50 mL of the buffer solution?

2. Determine the pH of (a) a 0.20 M  $\text{NH}_3$  ( $K_b = 1.75 \times 10^{-5}$ ) solution,

(b) a solution that is 0.20 M in  $\text{NH}_3$  and 0.30 M in  $\text{NH}_4\text{Cl}$ .

### **Preparation of buffer:**

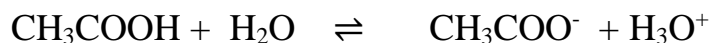
To prepare a buffer, it is to choose the acid with the  $\text{p}K_a$  close to the desired  $\text{pH}$ . Usually, buffers have a useful  $\text{pH}$  range =  $\text{p}K_a \pm 1$ , but the closer it is to the weak acid's  $\text{p}K_a$ , is the better.

### **Example:**

Describe how you might prepare approximately (500 mL) of pH 4.5 buffer solution from 1 M acetic acid  $\text{CH}_3\text{COOH}$  ( $K_a = 1.74 \times 10^{-5}$ ) and sodium acetate  $\text{CH}_3\text{COONa}$  (82.03 g /mole).

### **Solution:**

**For acidic buffer (pH= 4.5)**



$$\text{pH} = \text{pK}_a + \log \frac{C_{\text{salt}}}{C_{\text{acid}}} \quad (\text{Henderson equation for acidic buffers})$$

$$\text{pK}_a = -\log(1.8 \times 10^{-5}) = 4.74$$

$$4.5 = 4.74 + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$4.5 = 4.74 + \log \frac{[\text{CH}_3\text{COO}^-]}{[1]}$$

$$4.5 - 4.74 = \log [\text{CH}_3\text{COO}^-] - \log 1$$

$$\log [\text{CH}_3\text{COO}^-] = -0.24$$

$$[\text{CH}_3\text{COO}^-] = 10^{-0.24} = 0.575 \text{ M}$$

$$\text{Mass of CH}_3\text{COONa needed} = \text{Molarity (M)} \times \text{V (liter)} \times \text{Mwt}$$

$$\text{Mass of CH}_3\text{COONa} = 0.575 \text{ (mol/L)} \times \frac{500}{1000} \text{ L} \times 82.03 \text{ (g/mol)} = 23.58 \text{ g}$$

The required buffer is to be made by dissolution of 23.58 g of CH<sub>3</sub>COONa and completing the volume to 500 mL with 1M CH<sub>3</sub>COOH