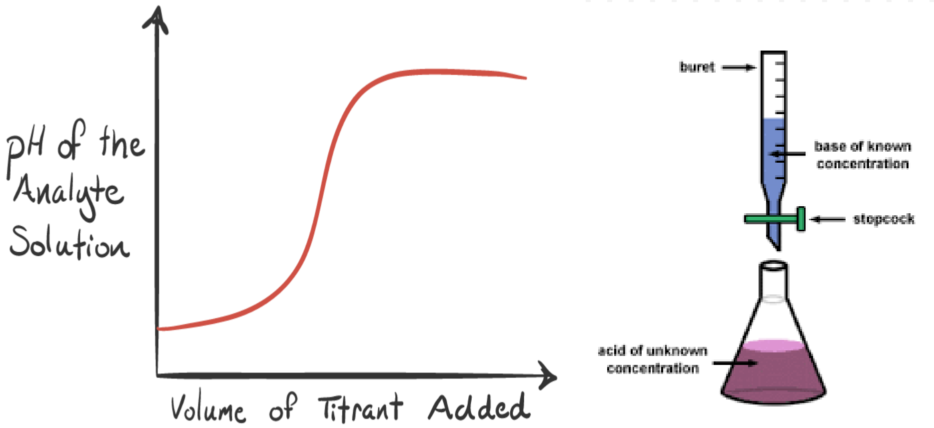
Titration Of A Weak Acid With Strong Base



BCH 312 [Practical]

Weak Acid :

- Weak acids or bases do not dissociate completely, therefore an equilibrium expression with **Ka must be used.**

- **The Ka is a quantitative measure of the strength of an acid in solution.**

since it's value is always very low, Ka is usually expressed as pKa , where:

$$pKa = -\log Ka$$

- As an acid/base get **weaker**, its Ka/Kb gets smaller and pKa/pKb gets **larger.**

- **For example:**

- HCl is a strong acid , it has **1×10^{7} Ka** value and **-7 pKa value**.
- CH_3COOH is a weak acid , it has **1.76×10^{-5} Ka** value and **4.75 pKa value.**

- **Type of weak acid:**

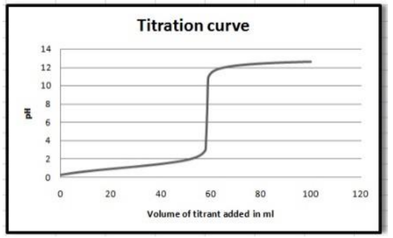
- Monoprotic (contain 1 group 'hydrogen ion'). ➔ Ex: $CH_3COO\underline{H}$
- Diprotic (contain two group). ➔ Ex: $\underline{H_2}SO_4$
- Triprotic (contain three group). ➔ Ex: $\underline{H_3}PO_4$

➔ **each group has own Ka value.**

- **Which dissociation group will dissociate first?**

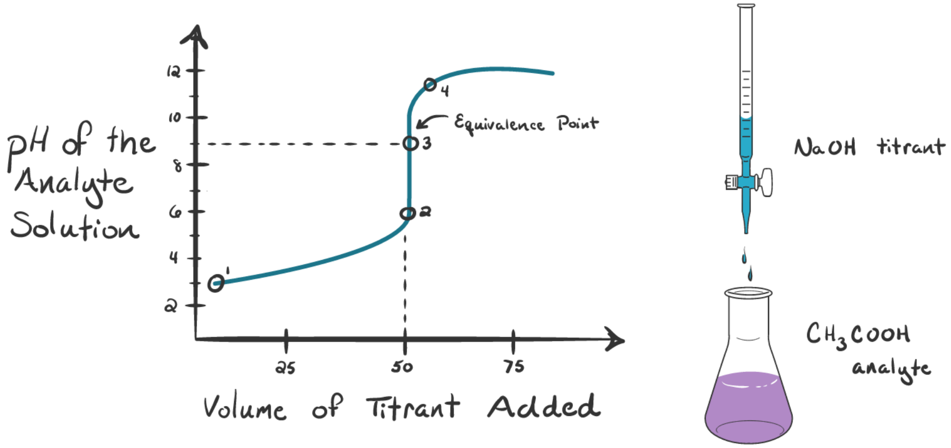
➔ The group that has higher Ka value or i.e that has lower pKa value

- **pKa** values of weak acids can be determined **mathematically** or **practically** by the use of **titration curves.**



Titration Curve:

- Titration Curves are produced by monitoring the pH of a **given volume** of a sample solution after successive **addition of acid or alkali**.
- The curves are usually plots of pH against the volume of titrant added (acid or base).
- There are many uses of titration, one of them is to indicate the pKa value of the weak acid by using the titration curve.
- Each dissociation group represent **one stage** in the titration curve.

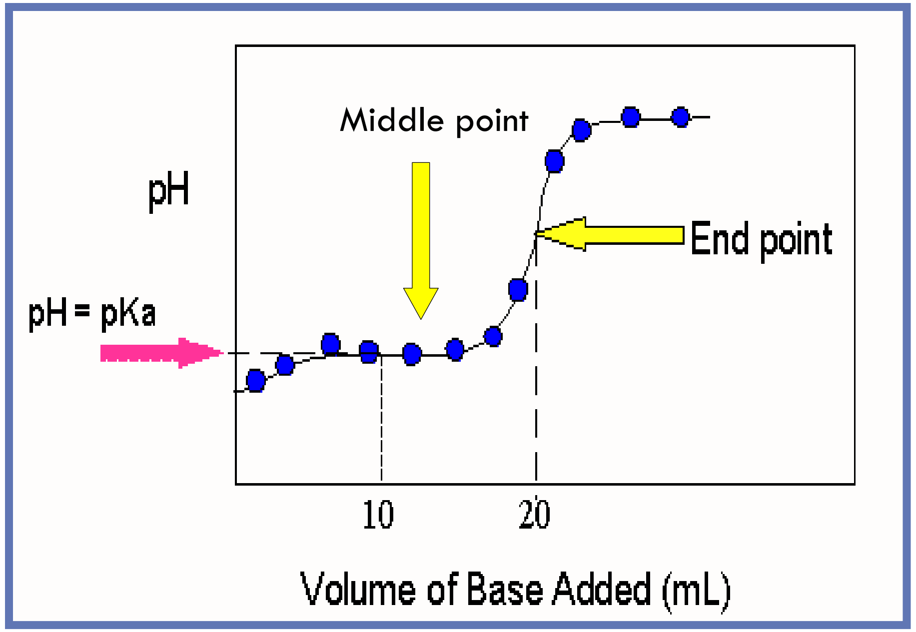


Titration curve of a weak acid with strong base:

- **[1] Before any addition of strong base the (starting point):**

  - ALL the weak acid is in the **full protonation** form [CH_3COOH] (electron donor) .
  - In this point pH of weak acid **<** pKa.
  - We can calculate the pH from:

    pH = (pKa + p[HA]) / 2



[2] When certain amount of strong base added (any point before the middle of titration):

- The weak acid is starting to dissociate

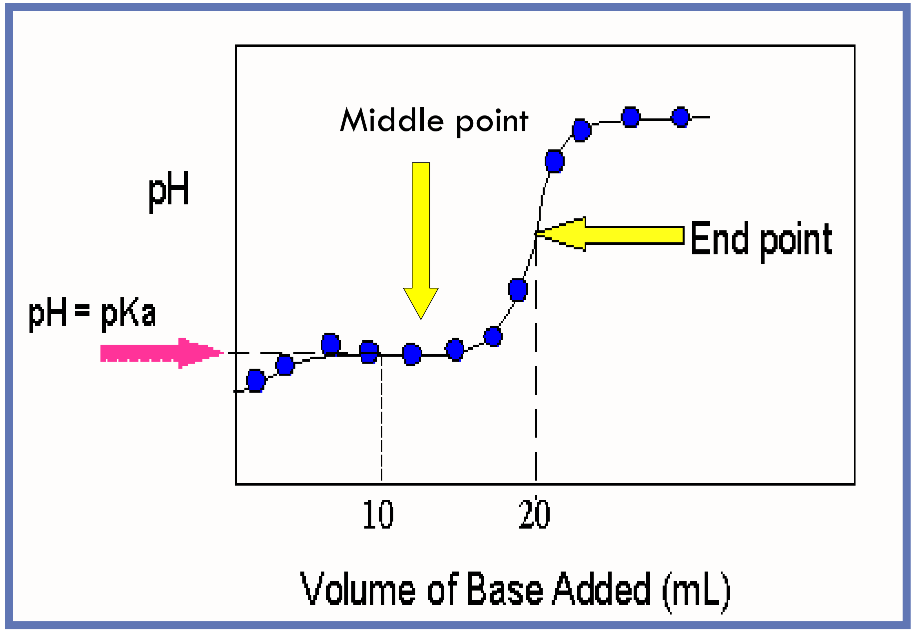
$[CH_3COOH] > [CH_3COO^-]$

- (Donor > Acceptor).

- In this point pH of weak acid **<** pKa.

- We can calculate the pH from:

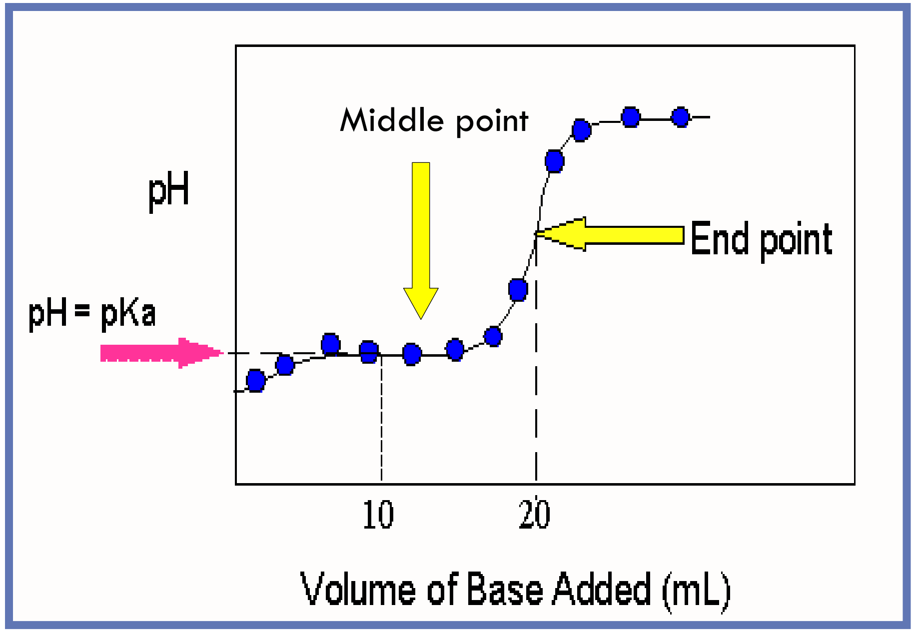
$pH = (pKa + \log [A^-] / [HA])$



[3] At middle of titration:

- $[CH_3COOH]=[CH_3COO^-]$.
- (Donor=Acceptor).
- In this point pH = pKa.
- The component of weak acid work as a **Buffer** (A solution that can resistant the change of pH) .
- Buffer capacity= pKa ± 1
- pKa is defined as the pH value at middle of titration at which they will be [donor]=[acceptor].
- We can calculate the pH from:

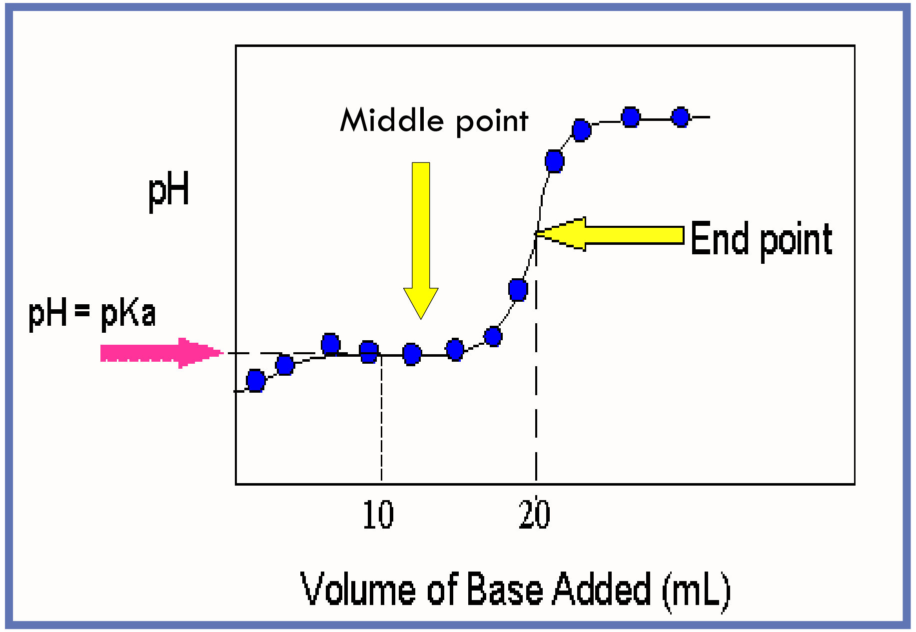
pH = (pKa + log [A-] / [HA])



[4] At any point after mid of titration and before end point:

- $[CH_3COOH] < [CH_3COO^-]$.
- (Donor< Acceptor) .
- In this point pH > pKa.
- We can calculate the pH from:

pH = (pKa + log [A⁻] / [HA])



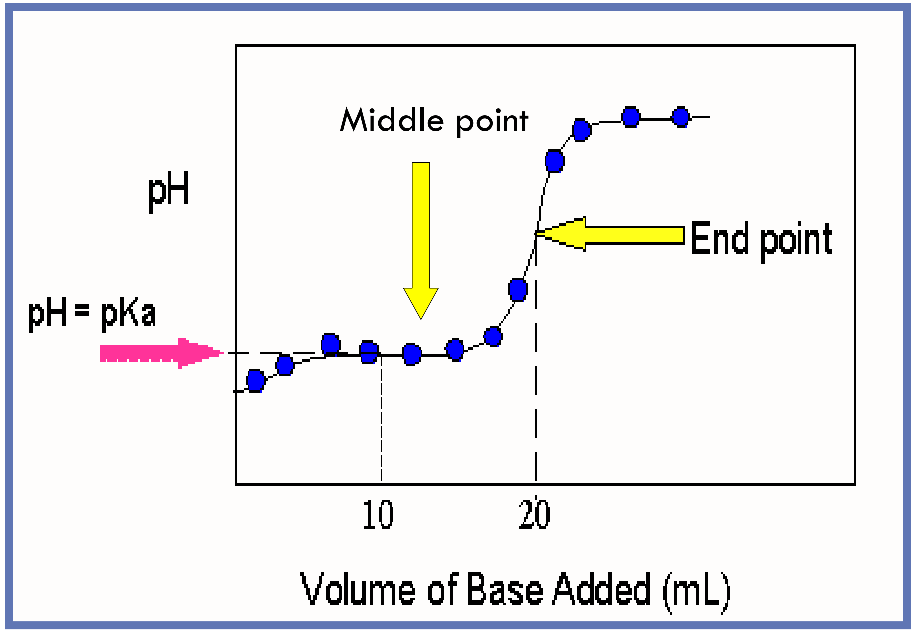
[5] At the end point :

- The weak acid is fully dissociated $[CH_3COO^-]$.
- (electron acceptor).

- In this point pH **>** pKa.

- Approximately, all the solution contains CH_3COO^- , so we first must calculate pOH, then the pH:

pOH = (pKb + p[A⁻]) / 2
pH = pKw – pOH
***pKb = pKw - pKa**



Calculating the pH at different point of the titration curve :

- [1] At start point [Weak acid only]:

$$pH = (pKa + p[HA]) / 2$$

- [2] At any point within the curve [weak acid and conjugated base mix]:

$$pH = (pKa + \log [A^-] / [HA])$$ -Henderson-Hasselbalch equation-

- [3] At the end point [approximately conjugated base only] :

$$pOH = (pKb + p[A^-]) / 2 \rightarrow pH = pKw - pOH$$

- **Henderson-Hasselbalch equation is an equation that is often used to :**

1. To calculate the pH of the Buffer.
2. To preparation of Buffer.
3. To calculated the pH in any point within the titration curve (Except starting and ending point)

Note:

- **If you start titration using 20 ml of the weak acid, In titration curve………..**

➔ The total volume of weak acid is 20 ml , we need 20 ml of strong base to full dissociate the group of weak acid.

➔ We can reach to **middle** titration if we add 10 ml of strong base (half the amount of 20 ml).

- **Bearing in mind that :**

1. the weak acid and the strong base (titrant) should have the <u>same concentration</u>.
2. the weak acid and strong base should have the <u>same protonation and hydroxylation state</u> respectively (ex: monoprotic acid and monohydroxy base).

Example: **Determine the pH value of 500 ml of monoproteic weak acid (0.1M) , titrated with 0.1M KOH (pKa=5), after addition of:**
(1) 100 ml. (2) 250 ml (3) 375 (4) 500 ml of KOH?

[1] pH after addition of 100 ml of KOH?

➔ **SECOND STAGE**
- pH= pKa + log[A-]/[HA]

-HA + KOH ➔ KA + H_2O

-we should calculate the No. of moles of **remaining [HA]** first because it is reflect the pH value at this stage.

-Mole of HA [original] – mole of KOH [added] = mole of HA remaining.
-No. of KOH [A^-] mole = 0.1 X 0.1 L = 0.01 mole
-No. of HA mole originally = 0.1 X 0.5 L =0.05 mole
-No. of HA mole remaining = 0.05 – 0.01 = 0.04 mole

So,
pH = 5 + log [0.01]/[0.04]
pH=4.4 ➔ pH<pKa

[2] pH after addition of 250 ml of KOH?

➔ **MIDDLE STAGE**
- pH= pKa + log[A-]/[HA]

-Mole of HA [original] – mole of KOH [added] = mole of HA remaining.

-No. of KOH [A^-] mole = 0.1 X 0.25 L = 0.025 mole
-No. of HA mole originally = 0.1 X 0.5 L =0.05 mole
-No. of HA mole remaining = 0.05 – 0.025 = 0.025 mole

So,
pH = 5 + log [0.025] / [0.025]
pH=5=pKa➔ (at mid point , The component of weak acid work
as a Buffer , has a buffering capacity 5 ± 1)

[3] pH after addition of 375 ml of KOH?

➔ **FOURTH STAGE**

- pH= pKa + log[A-]/[HA]

-Mole of HA [original] – mole of KOH [added] = mole of HA remaining.

-No. of KOH [A^-] mole = 0.1 X 0.375 L = 0.0375 mole

-No. of HA mole originally = 0.1 X 0.5 L =0.05 mole

-No. of HA mole remaining = 0.05 – 0.0375 = 0.0125 mole

So,

pH = 5 + log [0.0375] / [0.0125]

pH=5.48 ➔ pH>pKa "slightly"

[4] pH after addition of 500 ml of KOH?

➔ **END STAGE (**Note: 500 ml is the same volume of weak acid that mean the all weak acid are as [$CH3COO^-$]).

- pOH = (pKb + p[A^-])/2 ➔ pKb= pKw-pKa

➔pKb=14-5=9

- p[A-]= - log [A^-] ➔ [A^-]=??

No. of a mole KOH= 0.1 X 0.5 **(volume of base added)** = 0.05 mole

-[A-] = 0.05/1 =0.05 M **(total volume = 500+500=1000= 1L)**

So ➔ p[A-]= - log 0.05 = 1.3

-pOH=(9+1.3)/2 = 5.15

-pH=pKw-pOH

pH = 14 – 5.15 = 8.85 ➔ pH>pKa "slightly"

Practical Part

Objectives

- To study titration curves.
- Determine the pKa value of a weak acid.
- Calculate the pH value at a given point.

Method:

- You are provided with 10 ml of a **0.1M** **CH_3COOH** weak acid solution, titrate it with **0.1M NaOH.**
- Add the base drop wise mixing, and recording the pH after each **0.5 ml** NaOH added.
- Stop when you reach a pH=9.

ml of 0.1M NaOH	PH
0	
0.5	
1	
1.5	
….	

1. Record the values in titration table and **plot a Curve** of pH versus ml of NaOH added.
2. **Calculate the pH** of the weak acid HA solution after the addition of 3ml, 5ml, and 10ml of NaOH.
3. **Determine the pKa** value of weak acid.
4. **Compare** your calculated pH values with those obtained from Curve.
5. At what pH-range did the acid show buffering behavior? What are the chemical species at that region, what are their proportions? What is the **buffer capacity range**?

ml of 0.1M NaOH	PH
0	
0.5	
1	
1.5	
….	