

KWAME NKRUMAH UNIVERSITY OF SCIENCE AND TECHNOLOGY

COLLEGE OF ENGINEERING
DEPARTMENT OF CHEMICAL ENGINEERING

TITLE: ACID-BASE EQUILIBRIUM



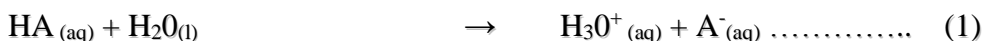
NAME: SENYO DENNIS A. AKANDE
COURSE: BSC. CHEMICAL ENGINEERING
YEAR: FIRST YEAR
EXPERIMENT NO. : P.1.2.1.
I.D. NO: 1314507
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Aims and Objectives:

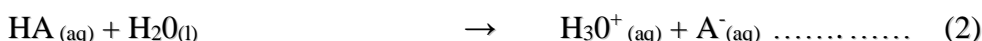
- 1) To be able to construct titration curves for strong and weak acids in their reactions with a strong base and
- 2) Observe the different characteristics of the curves.
- 3) To examine the behaviour of indicators during the titrations and determine the dissociation constant, K_a for an unknown weak organic acid.

INTRODUCTION

According to the Bronsted – Lowry definition, an acid is a species which donates a proton and a base is a species which accepts a proton in aqueous solutions. A strong acid is characterized as a substance which completely dissociates in aqueous solutions as shown in the equation below:



Whereas a weak acid is a substance which only dissociates partially in aqueous solutions as shown in the equation below:



The difference between (1) and (2) lies in the extent of the reaction. The major species in (2) is HA, the undissociated weak acid, while the major species in (1) are $\text{H}_3\text{O}^+_{(\text{aq})}$ and $\text{A}^-_{(\text{aq})}$, the products of the reaction are a conjugate acid and base.

For a strong monoprotic acid, the number of moles of A^- formed is equal to the number of moles of H_3O^+ formed (which is equal to the number of moles of HA dissociated). The same holds for the weak monoprotic acid but these concentrations are small in comparison to the concentration of HA remaining in solution. Equilibrium lies far to the left in equation (2) and the extent of dissociation of the weak acid is described by K_a , the acid dissociation constant.

The acid dissociation constant (K_a) of an acid is:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \quad (3)$$

Taking the negative log of each term, we then have:

$$-\log K_a = -\log[\text{H}_3\text{O}^+] - \log[\text{A}^-] + \log[\text{HA}] \quad (4)$$

The acidity of the solution can be described by a pH value defined as the negative logarithm of the hydroxonium ion concentration.

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \quad (5)$$

The negative log of the acid dissociation constant (K_a) also gives a constant, $\text{p}K_a$. Therefore (4) rearranges to:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{Acid}]} \quad (6)$$

When a strong base is added to:

- 1 A strong acid,
- 2 A weak acid,
- 3 A polyprotic acid,

considerable changes occur in terms of pH.

The equivalence points of the different titration reactions will be determined using the pH meter but a different indicator will be for each one and from observations made from the aftermath of the reaction, it will be determined whether the indicator used was appropriate.

Part A: Titration of HCl (strong acid) with NaOH.

With the aid of a pipette, 25.00 ml of 0.100M HCl is transferred into a clean Erlenmeyer flask and 2 drops of an indicator is added and about 75.00 ml of distilled water. The acid solution is stirred thoroughly and the colour of the indicator is noted. Measure the pH of the acid solution after standardizing the pH meter with buffer solutions of 4.00 and 7.00.

About 5.00ml of NaOH from the burette is added to the acid solution and the solution is allowed to mix and equilibrate and the pH is measured. NaOH is added 5.00ml at a time and the pH is measured after each addition. The pH range over which the indicator changes are noted and as the solution reaches equivalence point, smaller amounts of the NaOH is added to cause a change in pH. The titration is continued until the pH reaches about 12 to 14.

The results of the first titration are shown below:

ACID USED	INDICATOR	VOLUME OF BASE USED	pH	COLOUR CHANGE
HCl	Phenolphthalein	0.00	0.866	None
		5.50	1.017	None
		10.00	1.171	None
		15.00	1.457	Pink
		20.00	2.115	Pink
		22.30	7.00	Pink

		25.00	12.095	Pink

Part B: Titration of $\text{HC}_2\text{H}_3\text{O}_2$ (weak acid) with NaOH .

The titration process is repeated as in Part A but here, a solution of 0.100M $\text{HC}_2\text{H}_3\text{O}_2$ is used, with a different indicator. The results are shown below:

ACID USED	INDICATOR	VOLUME OF BASE USED	pH	COLOUR CHANGE
$\text{HC}_2\text{H}_3\text{O}_2$	Methyl Red	0.00	2.603	Pink
		5.00	3.558	Pink
		10.00	4.052	Pink
		15.00	4.435	Light Pink
		17.50	4.639	Orange
		23.00	5.379	Orange
		26.00	6.202	Orange
		31.00	12.394	Yellow

Part C: Titration of H_3PO_4 (a polyprotic acid) with NaOH .

The titration process is repeated as in Part A but here, a solution of 0.100M H_3PO_4 is used, with a different indicator. The results are shown below:

ACID USED	INDICATOR	VOLUME OF BASE USED	pH	COLOUR CHANGE
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H ₃ PO ₄	Methyl Orange	0.00	1.387	None
		5.00	2.062	None
		10.00	6.562	None
		15.00	7.741	None
		20.00	11.901	None
		23.00	12.279	None

Part D:

Titration of an unknown organic acid with NaOH.

The titration process is repeated as in Part A but here, a dilute solution of the unknown is used, with a different indicator. The results are shown below:

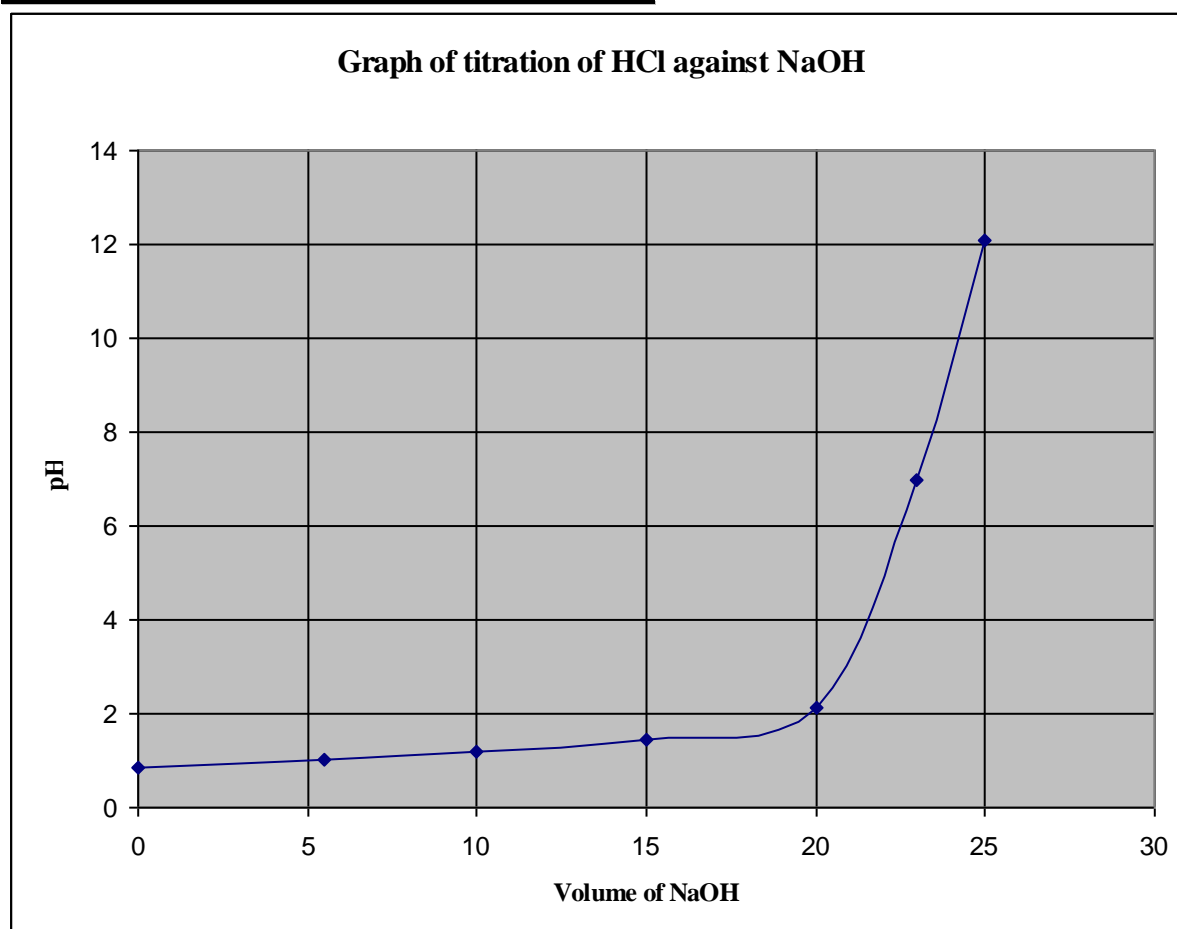
ACID USED	INDICATOR	VOLUME OF BASE USED	pH	COLOUR CHANGE
Unknown Acid	Phenolphthalein	0.00	0.370	None
		5.00	0.650	None
		10.00	1.070	None
		15.00	1.222	None
		20.00	1.604	None
		27.00	4.077	None
		33.00	5.341	None

		40.00	5.937	None
		50.00	6.503	None
		65.00	10.514	Purple
		70.00	13.627	Purple

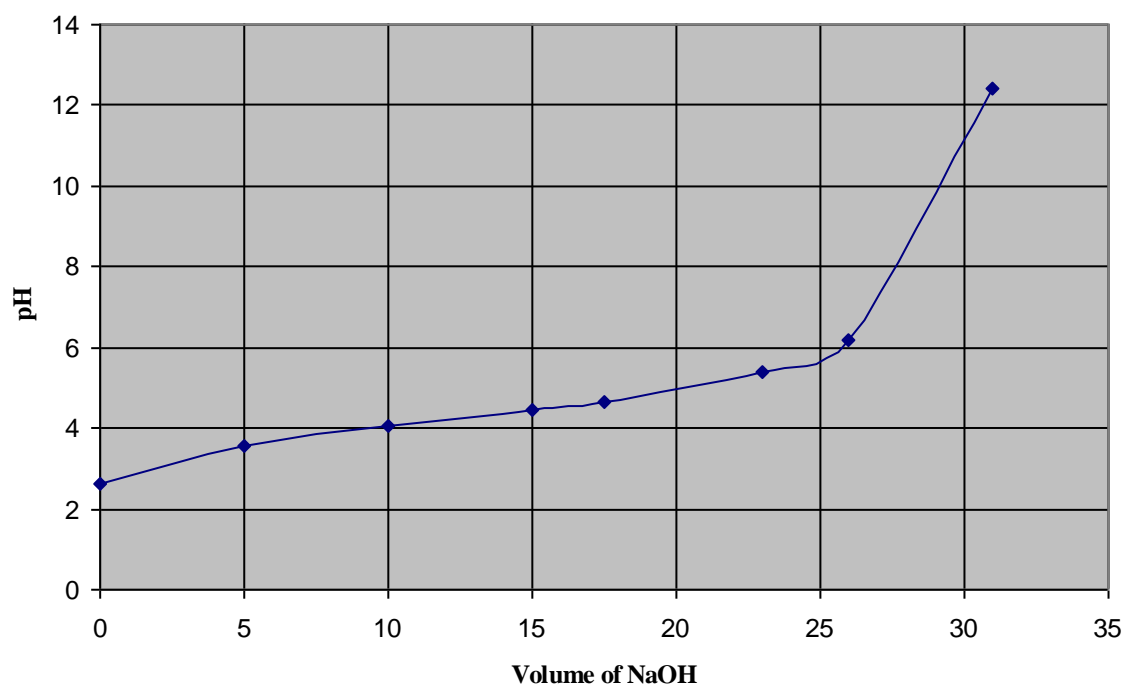
Precautions:

- 1 I made sure I put on my safety glasses and lab coat.
- 2 I rinse the electrode with distilled water before and after each measurement.
- 3 I made sure I do not temper with the electrical cord of the pH meter.

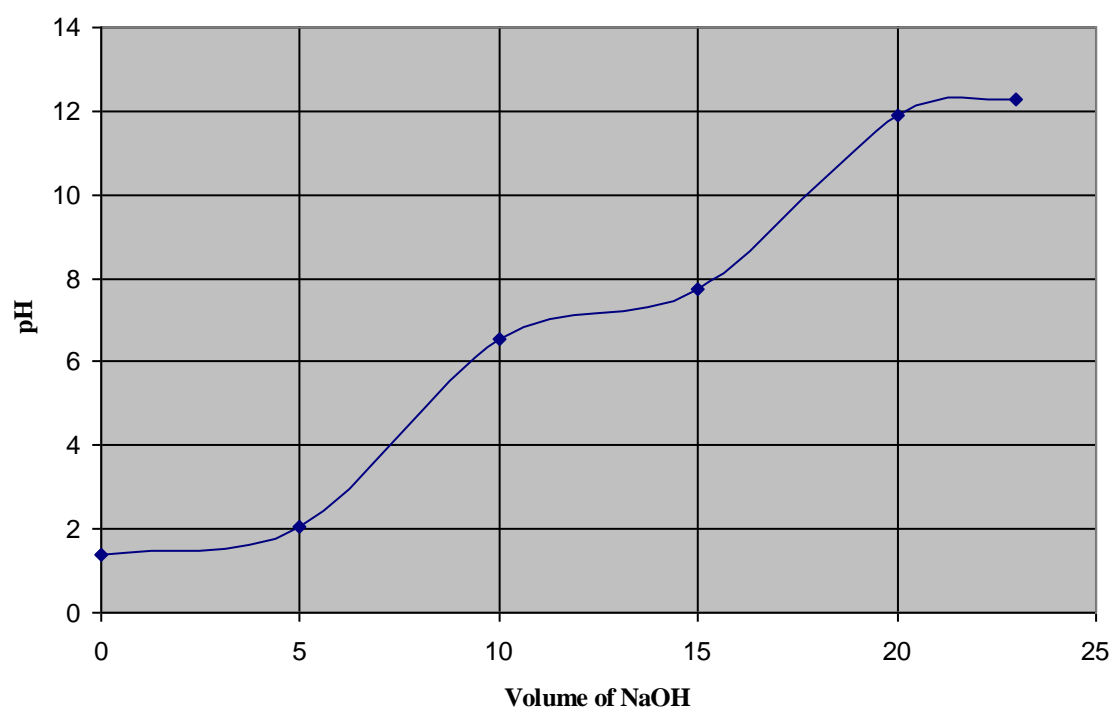
CALCULATIONS/POST – LAB QUESTIONS:

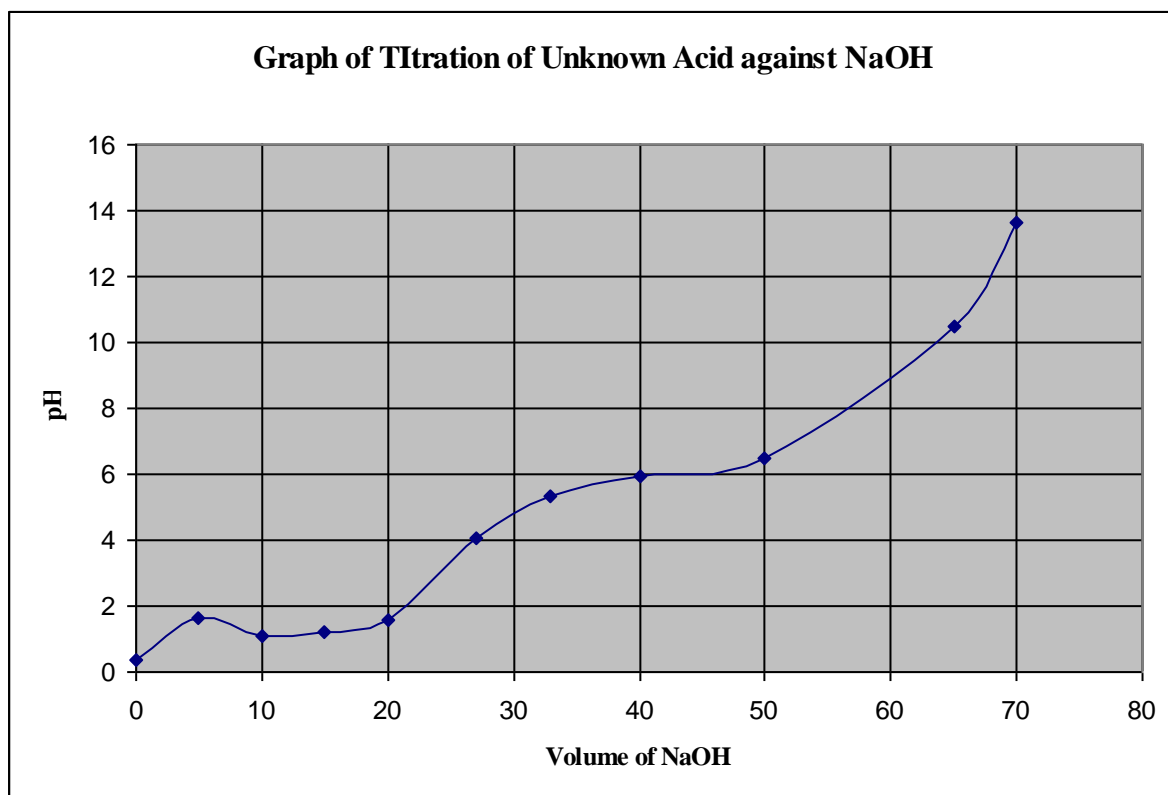


A Graph of Titration of $\text{HC}_2\text{H}_3\text{O}_2$ against NaOH



Graph of Titration of H_3PO_4 against NaOH





2. The bottom half of both curves correspond to each other.
Some differences in both curves include:

- 1 A higher initial pH.
- 2 A gradual rising portion of the curve, called the buffer region appears before the steep rise to the equivalence point.
- 3 The pH at the equivalence point is greater than 7.

3. The first half – way to equivalence point of the Phosphoric acid maps on the graph to a pH of 1.9. This gives a pK_{a1} value of 1.9 and a K_{a1} value of $10^{-1.9}$ which equals 12.6×10^{-3} .

The second half – way to equivalence point of the Phosphoric acid maps on the graph to a pH of 7.4. This gives a pK_{a2} value of 7.4 and a K_{a2} value of $10^{-7.4}$ which equals 3.98×10^{-8} .

4. The first point from the graph is 1.2. Since this equals the pK_{a1} , the K_{a1} is $10^{-1.2}$ which equals 6.3×10^{-2} . This value almost equals the K_{a1} value of 6.5×10^{-2} for Oxalic acid. Therefore the acid may be Oxalic acid.

The second point from the graph equals 6.1. Therefore pK_{a2} equals 6.1 and K_{a2} is $10^{-6.1}$ which equals 7.9×10^{-7} .

5. For the weak acid ($HC_2H_3O_2$) titration, methyl red was used which changes colour (pink to yellow) earlier and slower over a large volume ($\approx 10ml$) of titrant, thereby giving a vague and false indication of the equivalence point.

Phenolphthalein is suitable in this titration because its colour change lies within the range of the basic region.

For the strong acid (HCl) titration, phenolphthalein was used. Even though the range of its colour change (colourless to pink) does not occur at the equivalence point, it occurs on the vertical position of the curve, where a single drop of base causes a large pH change and this shows that the titration is within a fraction of a drop from the equivalence point.

For the polyprotic acid (H_3PO_4) titration, methyl orange was used. From the results of the titration, there was no colour change meaning that the indicator was not suitable for the titration.

For the unknown acid titration, phenolphthalein was used. The colour change occurred at the basic region and this shows the acid is a weak one and the indicator was quite suitable for the titration.

