EXPERIMENT 4
ACID-BASE TITRATION CURVES BY PH VARIATION.

Introduction.
The objective of this experiment is to compare the titration curves of a weak and a strong acid with a strong base. The purpose of titrations, say an alkaline solution with a standard solution of an acid, is to determine the amount of the acid which is exactly chemically equivalent to the amount of the base present. The point at which this is reached is the equivalent point, stoichiometric point or the theoretical end point.

If both the acid and the base are strong electrolytes, the resultant solution will be neutral and will have a pH value of 7. However, if either the acid or the base is a weak electrolyte, the resulting salt solution will be hydrolysed to a certain degree and the solution at the equivalent point will be slightly alkaline or slightly acidic. For example ammonium hydroxide is a weak base and its titration with a strong acid like hydrochloric acid occurs as below:

$$HCL + NH_4OH \longrightarrow NH_4Cl + H_2O \qquad (1)$$

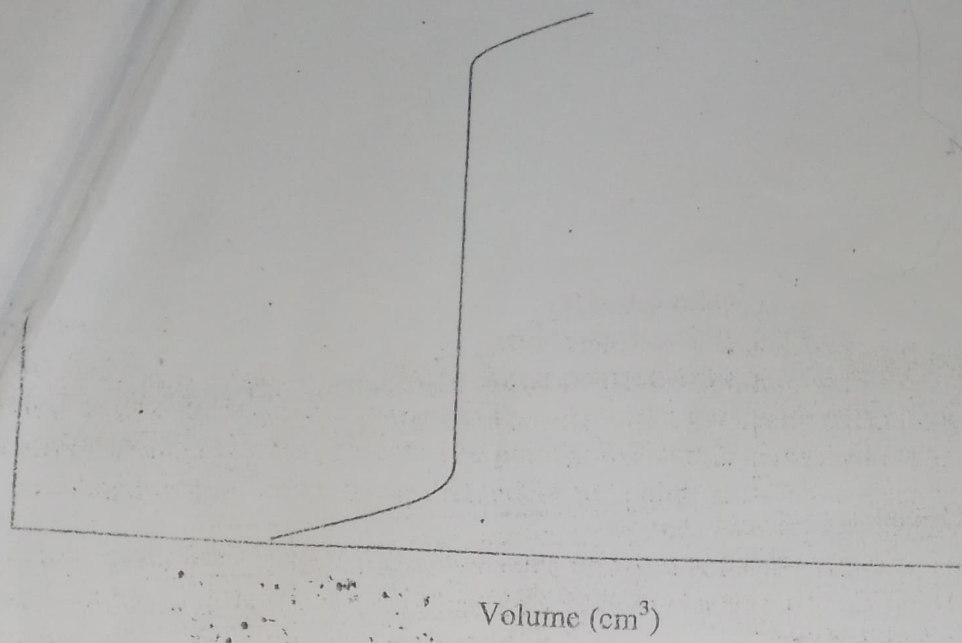
The resulting salt is hydrolysed as below

$$NH_4Cl + H_2O \longrightarrow NH_4OH + H^+Cl^- \qquad (2)$$

And thus the resulting solution is slightly acidic due to the presence of H^+ ions produced in the hydrolysis reaction 2.
The pH of a solution is defined as the negative logarithm of hydrogen ion concentration. i e. pH = -log [H].
Example: For 0.1M HCL, the acid is a strong acid and ionises fully. The [H] = [HCL] = 0.1
Then the pH of this solution is given as;

$$PH = -\log [H]$$
$$= -\log (0.1)$$
$$pH = 1$$

If the pH of a solution can be measured during a titration and then be plotted against the volume of the base added, then the end point can be read off from the graph.
A characteristic titration curve for a strong acid titrated against a strong base would look like below:



In this experiment, the pH will be determined during the titrations of a strong and a weak acid both by a strong base using universal indicator. The result is that it is then possible to extend the colour change of the indicator over a considerable portion of pH range. Such indicators can be used to approximate the pH of solutions resulting from acid-base titration. The colour shown by the universal indicator with a chart one is able to approximate the pH of the solution at the various stages of the titration. (Fully ionize in solution while weak acids and bases).

Ionisation of acids and bases: Strong acids and bases will be fully ionized while weak strength of that acid or base. This is expressed as the equilibrium constant for the ionization of the electrolyte.

Example, for acid, HA, the ionization constant (K_a) is given by;

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

For the ionization reaction HA $\Leftrightarrow H^+_{(aq)} + A^+_{(aq)}$

In practice it is usually convenient to express this value as a pka value.

$$pka = -\log ka.$$

1. Pka values can be used to predict the strength of a weak acid. pka values can be approximated from the titration curves of a weak acid with a strong base as will be seen later in the experiment.

paratus
4 test tubes
Pipette, 50cm^3
Burette, 50cm^3
Teat pipette

Reagents
0.1M hydrochloric acid
0.1M sodium hydroxide
0.1M acetic acid
0.1M sodium hydroxide
Universal indicator, B.D.H.

PROCEDURE

1. Into of the 14 test tubes, pipette 5 cm^3 of 0.1M HCL.
2. To each test tube, add one drop of universal indicator solution.
3. Add different vol... of 0.1M NaOH according to the table.
4. Obtain the pH of each test tube by comparing the colour with a chart (Re... Vogel A.I., A test book of quantitative inorganic analysis, Longman Ltd., London,
5. Repeat using 0.1M acetic acid in place of HCL.

Test tube no.1 5	Volume (cm^3) of 0.1M NaOH Added to 5.0cm^3 of 0.1M	pH 0.1M HCL	pH 0.1M CH_3COOH
1	0.0		
2	1.0		
3	2.0		
4	2.5		
5	3.0		
6	4.0		
7	4.5		
8	4.75		
9	5.0		
10	5.25		
11	6.0		
12	7.0		
13	8.0		
14	10.0		

Questions

1. For each acid, plot pH against volume of base added and determine the volume and the pH at the end point.
2. Determine the pka value for acetic acid (HAC)

$$\text{Since pH} = \text{pka} + \log\frac{[Ac^-]}{[HAc]}$$

And $[Ac^-] = [Hac]$ at half-titration.

Then pH = pka at half titration.

3. Compare your pka value for Hac with the literature value and account for the erro
4. Does the end points of titrations of NaOH with Hac and HCL occur at the same pH? If the difference, explain the difference.