Al-Karkh University of Science
College of science
First year level
General chemistry Labs

Supervisor: Assis. Prof. Dr. Mohammed A. B. Abdul Jabar

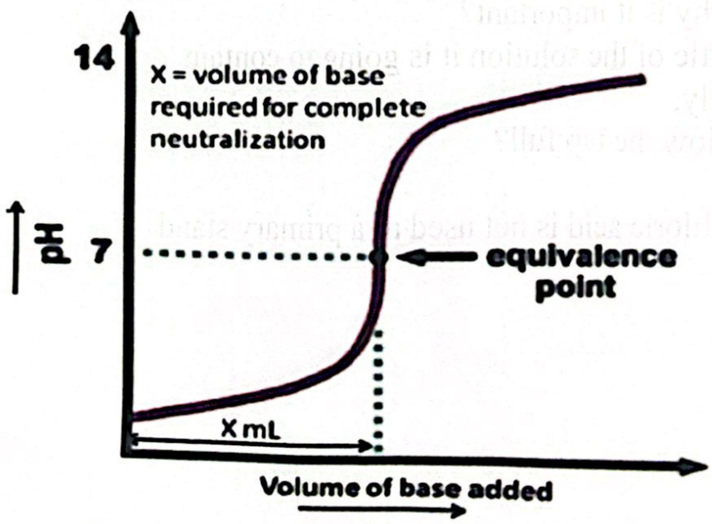
Lab -7-

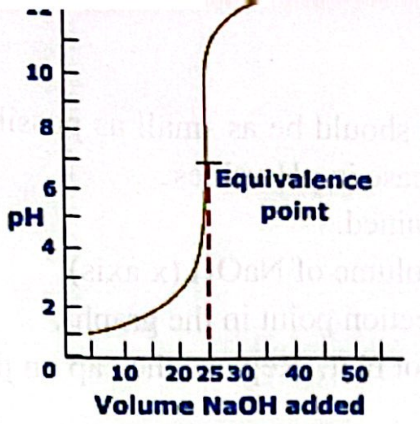
pH - metric Titration

Determination of the strength of a given hydrochloric acid solution against a standard sodium hydroxide solution

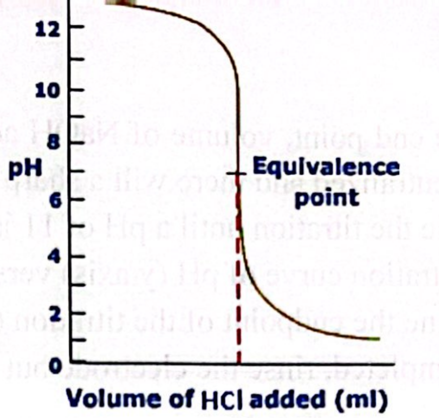
1. Introduction:

- The pH of a solution can be measured accurately with the help of a pH meter.
- The pH values of the solution at different stage of acid–base neutralization are determ and plotted against the volume of alkali added on adding a base to an acid, the pH slowly in the initial stages as the concentration of H^+ ion decreases gradually.
- Nevertheless, at the equivalence point, it increases rapidly as at the equivalent point ion concentration is very small. Then it flattens out after the end point.
- The end point of the titration can be detected where the pH value changes most rap However, the shape of the curve depends upon the ions ability of the acid and the used and on the acidity of base and basicity of the acid.
- In this experiment we will use pH meter by putting the electrode of pH meter in HCl(aq) solution and add definite small volumes of NaOH(aq) from the burette recording the read out of the instrument as function of added volume of NaOH(aq).
- The end point of the titration can be determined by graphic method by using a titra curve – a plot of the pH versus added volume of NaOH(aq). As in the figure below





Titration curve of strong acid (HCl) with a strong base (NaOH)



Titration curve of strong base (NaOH) with strong acid (HCl)

2. Objective:

- The purpose of this experiment is to determine the concentration of an unknown acid solution via the endpoint of an acid-base titration.

3. Apparatus and Chemicals:

HCl (unknown conc.)	250-mL beaker
0.10 M NaOH	10 mL volumetric pipet
Magnetic stirrer	pH meter and electrode
50-mL burette	150 mL beaker
pH buffer solutions (pH 4 and 6.86)	burette clamp
Funnel	wash bottle

4. Procedure:

1. Plug in the pH meter, and allow it to warm up for about 10 minutes.
2. The temperature knob should be set between 20 and 25°C.
3. Place the electrode in the pH 6.86 buffer, turn the knob to pH, and adjust the pH to 6.86 with the standardization knob.
4. Place the instrument on standby. Rinse and blot the electrode.
5. Place the electrode in pH 4.00 buffer, turn the knob to pH, and adjust the pH to 4.00 with the slope knob.
6. Place the beaker on the magnetic stirrer, and add a stirring bar.
7. Measure 10.0 mL of unknown HCl into a 250-mL beaker, and dilute it with ~90 mL of distilled water.
8. Carefully add the NaOH to the buret. You will need to drain some of the NaOH from the buret to fill the tip. Fill the buret so that the meniscus of the NaOH is sitting on the 0.00 mL mark with the tip filled.
9. Put the selector at the expected range (0–7). The reading shown on the scale of pH meter is pH value of the HCl solution.
10. Add NaOH solution drop wise from the burette (maximum 2.0 mL at a time), shake the solution well and note the corresponding pH values.

11. Near the end point, volume of NaOH added should be as small as possible because acid is neutralized and there will a sharp increase in pH values.
12. Continue the titration until a pH of 11 is obtained.
13. Plot a titration curve of pH (y axis) versus volume of NaOH (x axis).
14. Determine the endpoint of the titration (inflection point in the graph).
15. When completed, rinse the electrode but do not blot. Replace the cap on the electrode

5. Calculation:

pH	mL NaOH

➤ Plot a graph between pH and volume of NaOH added and find out the volume of Na required (V2 mL) for complete neutralization of HCl from the graph. Then find out strength of HCl (M1).

$$10 \text{ ml} \times M1 = V2 \text{ (from graph)} \times 0.1 \text{ M}$$

Strength of HCl (M1) = (V2 (from graph) x 0.1/10) (N)