CH1202 EXP-4: *Determination of strength of a solution of a strong acid by conductometric titration.*

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AIM:

- To discuss how the conductance changes with a change in the concentration of ions during titration.
- To perform a conductometric titration of a strong acid with a strong base.
- To determine the equivalence point of the titration by plotting the titration curve using conductance values and the amount of the base added during titration.
- To estimate the strength of the given solution of a strong acid.

PRINCIPLE:

The principle of conductometric titration is based on the fact that during the titration, one of the ions is replaced by the other and invariably these two ions differ in the ionic conductivity with the result that the conductivity of the solution varies during the course of titration. The equivalence point may be located graphically by plotting the change in conductance as a function of the volume of titrant added.

In order to reduce the influence of errors in the conductometric titration to a minimum, the angle between the two branches of the titration curve should be as small as possible

(see Fig). If the angle is very obtuse, a small error in the conductance data can cause a large deviation. The following approximate rules will be found useful.

- The smaller the conductivity of the ion which replaces the reacting ion, the more accurate will be the result. Thus it is preferable to titrate a silver salt with lithium chloride rather than with HCl. Generally, cations should be titrated with lithium salts and anions with acetates as these ions have low conductivity

- The larger the conductivity of the anion of the reagent which reacts with the cation to be determined, or vice versa, the more acute is the angle of the titration curve.

- The titration of slightly ionized salt does not give good results, since the conductivity increases continuously from the commencement. Hence, the salt present in the cell should be virtually completely dissociated; for a similar reason; the added reagent should also be a strong electrolyte.

- Throughout a titration the volume of the solution is always increasing, unless the conductance is corrected for this effect, nonlinear titration curves result. The correction can be accomplished by multiplying the observed conductance either by total volume $(V+V^{'})$ or by the factor $(V+V^{'})/V$, where V is the initial volume of solution and $V^{'}$ is the total volume of the reagent added. The correction presupposes that the conductivity is a linear function of dilution, this is true only to a first approximation.

- In the interest of keeping V small, the reagent for the conductometric titration is ordinarily several times more

concentrated than the solution being titrated (at least 10-20 times). A micro burette may then be used for the volumetric measurement.

The main advantages of conductometric titration are its applicability to very dilute, and colored solutions and to systems that involve relatively incomplete reactions. For example, neither a potentiometric, nor indicator method can be used for the neutralization titration of phenol ($K_a = 10^{-10}$) a conductometric endpoint can be successfully applied.

Acid-base titration, especially at trace levels can be carried out using conductometry. Relative precision is better than 1% at all levels. There are also a few disadvantages to this technique. As you know the conductance is a non-specific property, concentration of other electrolytes can be troublesome.

Titration curve of a strong acid with a strong base, e.g. HCl with NaOH: Before NaOH is added, the conductance is high due to the presence of highly mobile hydrogen ions. When the base is added, the conductance falls due to the replacement of hydrogen ions by the added cation as H^+ ions react with OH^- ions to form un-dissociated water. This decrease in the conductance continues till the equivalence point. At the equivalence point, the solution contains only NaCl. After the equivalence point, the conductance increases due to the large conductivity of OH^- ions. You will get a titration curve that resembles the one shown in (Fig). The point of

intersection of the two lines gives the point of neutralization i.e. equivalence point.
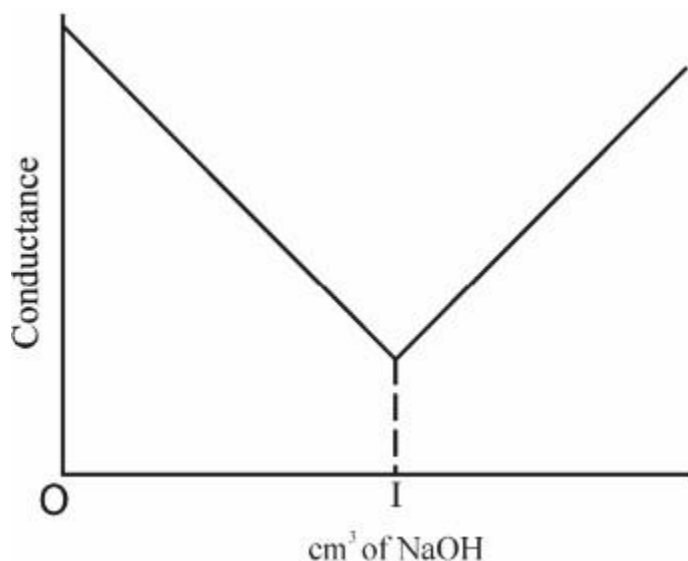

FIGURE: Conductometric titration of a strong acid (HCl) vs. a strong base (NaOH)

REQUIREMENTS:

Apparatus

Conductometer, conductance Cell, burette, pipette, beaker (100 cm^3), standard flask(100 cm^3), glass Rod, burette stand with clamp.

Chemicals

Hydrochloric acid and sodium hydroxide

Solutions Provided

0.1 M Sodium hydroxide solution: which was standardized with standard oxalic acid and~ 0.01 M HCl solution.

PROCEDURE:

1. Take 30 cm^3 of HCI (~ 0.01) solution in a 100 cm^3 beaker with the conductance cell.
2. Take NaOH solution in the burette.
3. Connect the conductometer to the mains and to the conductance cell. Switch on the instrument keeping the meter switch at **'CAL'.**
4. Calibrate the meter keeping the selector knob at **'20 ms'** by rotating the 'sensitivity knob till the meter reads 1.0.
5. Shift the meter switch to **'Read'**. Read the conductance of the solution (Keep the stirrer above the solution). Record this value in Observation Table 1.
6. Make additions of NaOH from the burette as given in the Observation Table1. After each addition, stir the solution well and read the conductance, keeping the stirrer above the solution. Enter all the conductance values in Observation Table.
7. Plot conductance versus volume of NaOH added on a graph sheet.

OBSERVATIONS:

1)Standardization of NaOH against freshly prepared 0.2 N Oxalic acid:

Weight taken = 1.2607 g
Weight to be taken = 1.26 g
For making 0.2 N 100 ml solution.

Volume of NaOH taken in burette (ml)	Volume of oxalic acid aliquot (ml)
10.1	10
10.1	10
10.1	10

We know,

$$\Rightarrow \quad N_{NaOH} \times V_{NaOH} = N_{OXALIC} \times V_{OXALIC}$$

$\Rightarrow M_{NaOH} \times V_{NaOH} \times n - factor_{NaOH} = N_{OXALIC} \times V_{OXALIC}$

$\Rightarrow M_{NaOH} \times 10.1(ml) \times (1) = 0.2(N) \times 10(ml)$

$\Rightarrow M_{NaOH} = 0.1980M$

*2) Titration of the supplied acid with the standardized NaOH solution:*

Volume of the supplied acid solution taken V_{HCL}=25 ml

Volume of NaOH added(ml)	Total Volume of NaOH(ml)	Observed conductance(mS)
0	0.0	3.00
0.5	0.5	3.12
0.5	1.0	2.94

0.6	1.6	2.73
0.6	2.2	2.52
0.5	2.7	2.34
0.5	3.2	2.15
0.5	3.7	1.947
0.5	4.2	1.744
0.5	4.7	1.581
0.5	5.2	1.388
0.2	5.4	1.303
0.2	5.6	1.249
0.2	5.8	1.172
0.2	6.0	1.094
0.2	6.2	1.027
0.2	6.4	0.957
0.2	6.6	0.983
0.2	6.8	1.022
0.2	7.0	1.073
0.2	7.2	1.138
0.2	7.4	1.153
0.2	7.6	1.236
0.2	7.8	1.286

0.1	7.9	1.320
0.1	8.0	1.356
0.1	8.1	1.383
0.2	8.2	1.442
0.2	8.4	1.497
0.2	8.6	1.545
0.2	8.8	1.589
0.3	9.1	1.684
0.3	9.4	1.735
0.3	9.7	1.819
0.3	10.0	1.905
0.3	10.3	1.976
0.3	10.6	2.05
0.5	11.1	2.19
0.5	11.6	2.31
0.5	12.1	2.45
0.5	12.6	2.56
0.5	13.1	2.67
0.5	13.6	2.31

3)Plot of Conductance vs volume of NaOH:
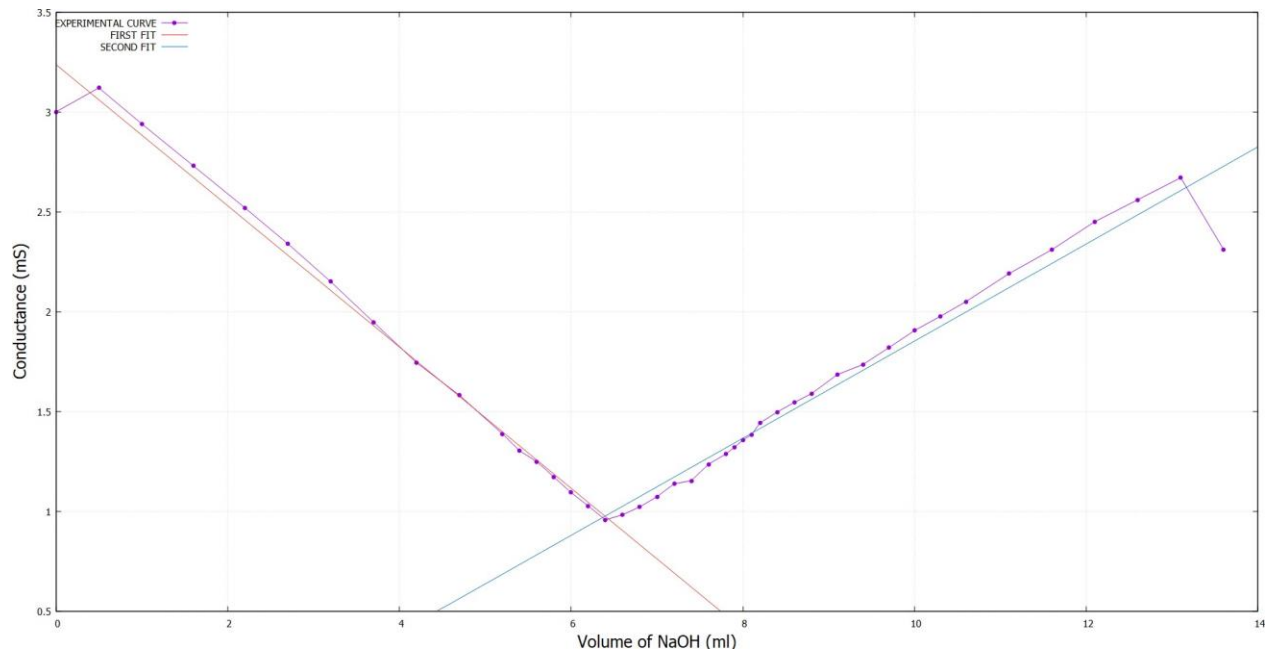
From the plot, the volume of NaOH to meet the endpoint is,

$V_{NaOH} = 6.3958ml$

We know,

$$\Rightarrow N_{NaOH} \times V_{NaOH} = N_{HCl} \times V_{HCl}$$

$$\Rightarrow M_{NaOH} \times V_{NaOH} \times n - factor_{NaOH} = M_{HCl} \times V_{HCl} \times n - factor_{HCl}$$

$$\Rightarrow 0.1980(M) \times 6.3958(ml) \times (1) = M_{HCl} \times 25(ml) \times 1$$

$$\Rightarrow M_{HCl} = 0.0506M$$

CONCLUSION:

The molarity of the HCl is calculated to be 0.0506 M.

PRECAUTIONS:

1. After switching on the instrument (conductometer), it should be allowed to stabilize prior to starting the experiment.

2. The conductance cell must always be dipped either in solution or in distilled water.

3. The platinum electrodes of the conductance cell must be completely immersed in the solution during the measurement of conductance.

4. There should be no air bubbles between the two electrodes.

5. The titrant must be at least ten times more concentrated than the analyte.