

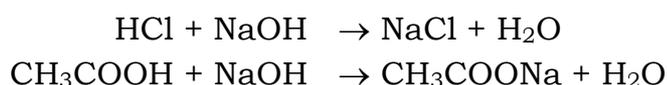
Aim : To determine the strength of the two acids (Weak + strong) in a mixture by conductometric method.

Requirement : Given acid mixture solution, N/10 NaOH solution, Oxalic acid salt.

Glassware and apparatus: 50 mL measuring cylinder, 100 mL volumetric flask, burette (25 mL), beaker (100 mL and 250 mL), magnetic stirrer, conductivity cell and conductivity meter.

Theory in brief:

Conductometric titration are based on the principle that the electrolytic conductance depends upon the number and mobility of ions present in the solution. The titration of mixture of two acids in which one is weak (acetic acid) and other one strong (HCl), by a strong base (NaOH) can be carried out conductometrically efficiently. The reactions take places are as follows:



Since HCl is strong acid, it will dissociate completely and hence due to common ion effect it will also suppress the dissociation of acetic acid. Therefore, HCl will get titrated first followed by acetic acid. Since the number of fast moving H⁺ ions decrease, so conductance decreases first and then increases because of excess of fast moving OH ions. So the end points are characterized by a sharp change in conductance.

Procedure:

1. The given N/10 NaOH was standardized with standard N/10 oxalic acid solution (need to be prepared) and then the standardized NaOH was filled in burette.
2. 10 mL of mixture of two acids was taken in 50 mL measuring cylinder and 40 mL distilled water was added and then transferred to 100 mL beaker.
3. Magnetic bead was put in the beaker containing diluted mixture of two acids and conductivity cell was dipped in it and connected to the conductometer.
4. The solutions were titrated with standardized NaOH and conductance readings were noted down.
5. A graph between conductance (in milli-Siemens) vs volume of alkali added (mL) was plotted to get the end point.

Observation:

- Standardization of given N/10 NaOH with standard oxalic acid solution.

Wt. of oxalic acid taken, w (should be approx.. 0.630 g in 100 mL)

$$\text{So, Normality, } N_o = \frac{WX1000}{\text{Eq.wt}X 100}$$

Table 1: Titration results for standardization of NaOH by oxalic acid solution.

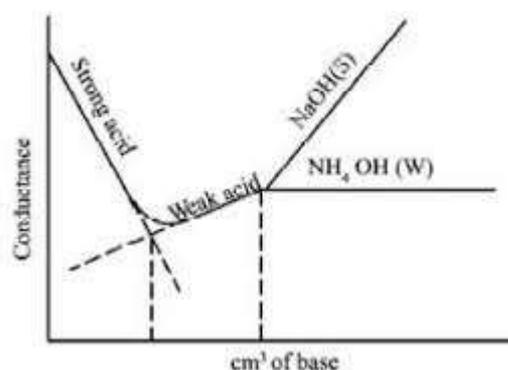
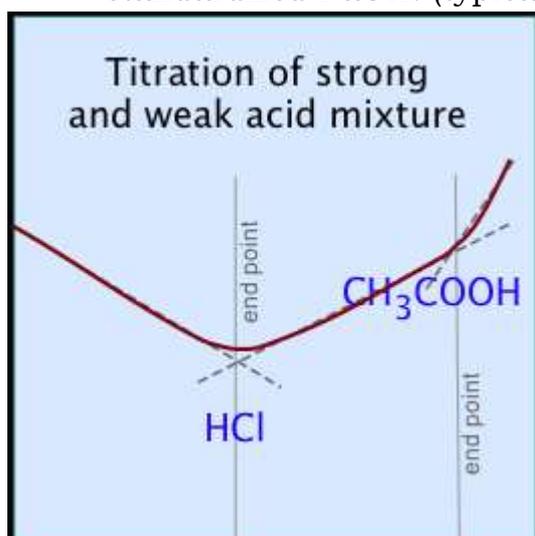
V_o (Oxalic acid) = 10 mL.

Sl. no	Initial burette reading of NaOH	Final burette reading of NaOH	Volume of NaOH used	Concordant readings
1.				

Table 2: Conductometric titration results of mixture of two acids with standardized NaOH solution:

Sl. No.	Volume of alkali added (mL)	Conductance (mSiemens)	Volume of alkali added (mL)	Conductance (mSiemens)
1.			contd.	

Graph 1: Change of conductance of mixture of acids on addition of standardized NaOH. (typical graphs shown below)



Results: The strength of strong acid and weak acids are found to be X.XX N and Y.YY N respectively.

Typical graphs:

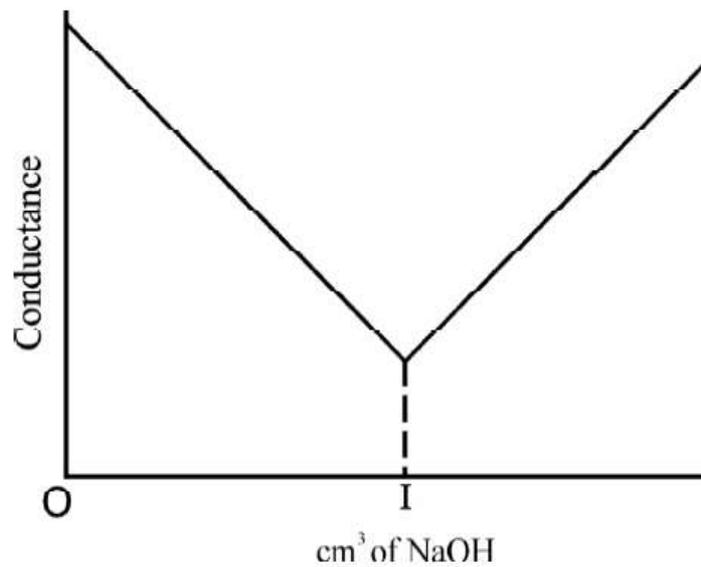


Fig. 6.2: Conductometric titration of a strong acid (HCl) vs. a strong base (NaOH)

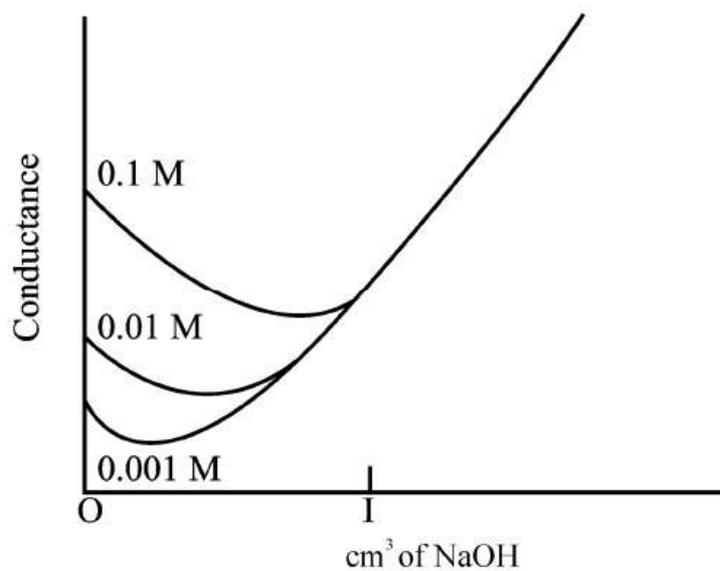


Fig. 6.3: Conductometric titration of a weak acid (acetic acid) vs. a strong base (NaOH)

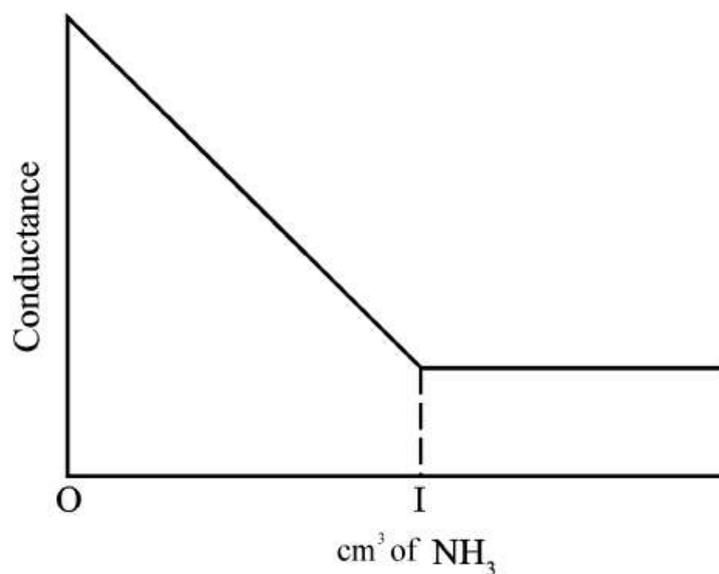


Fig. 6.4: Conductometric titration of a strong acid (H_2SO_4) vs. a weak base (NH_4OH)

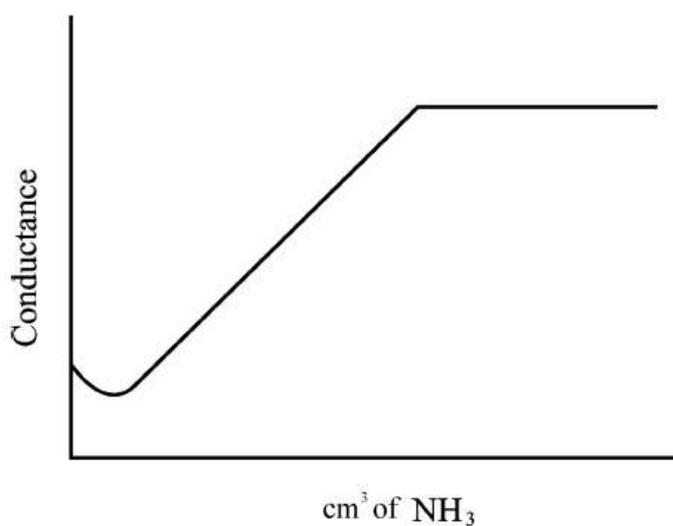


Fig. 6.5: Conductometric titration of a weak acid (acetic acid) vs. a weak base (NH_4OH)

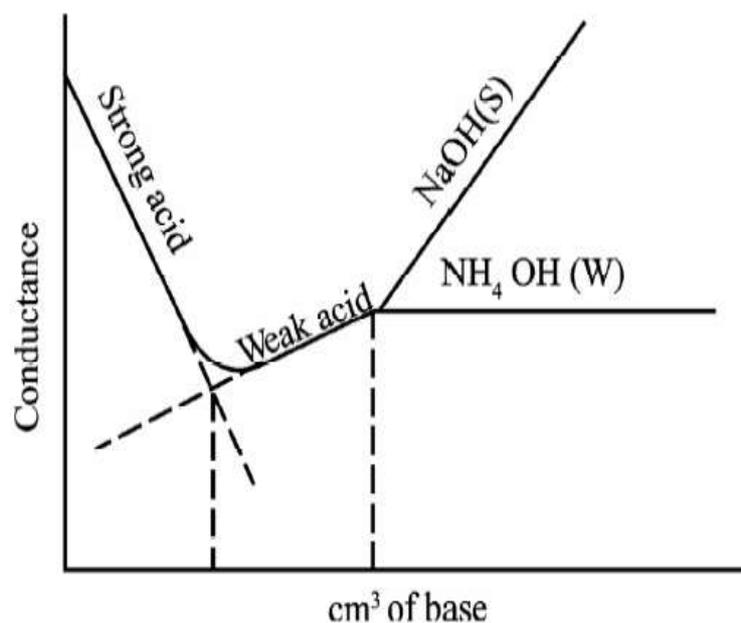


Fig. 6.6 Conductometric titration of a mixture of a strong acid (HCl) and a weak acid (CH₃COOH) vs. a strong base (NaOH) or a weak base (NH₄OH)

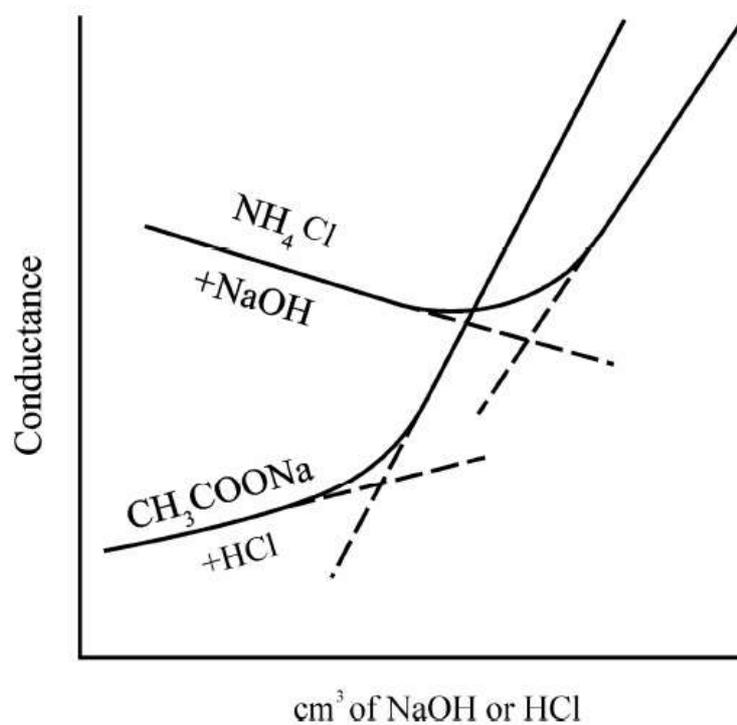


Fig. 6.7: Conductometric titration of a salt of weak acid (sodium acetate) vs. strong acid (HCl); salt of a weak base (NH₄Cl) vs. a strong base (NaOH)

Some important points to remember:

- 1) 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 or KCl rather than with HCl.
- 2) 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 titration curve.
- 3) The titration of a 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 as strong electrolyte.
- 4) Throughout a titration the volume of the solution is always increasing, unless the conductance is corrected for this effect, *non-linear 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.
- 5) 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).
- 6) Advantage: The main advantages to the conductometric titration are its applicability to *very dilute, and coloured solutions* and to *system that involve relative incomplete reactions*. For example, which 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.
- 7) Disadvantage: The conductance is a non-specific property, presence or slight concentration of another electrolyte can be troublesome. For this reason, direct conductance measurement is of little use unless the solution contains only the electrolyte to be determined. Conductometric titrations, in which the species in the solution are converted to non-ionic for by neutralization, precipitation, etc. are of more value.



Typical conductivity meter.