

LIST OF EXPERIMENTS.

1. Determination of the amount of sodium carbonate and sodium hydroxide in a mixture by titration.
2. Determination of strength of an acid by Conductometry.
3. Determination of hardness of water using EDTA
- Complexometry method.
4. Determination of Fe^{2+} ion using $\text{K}_2\text{Cr}_2\text{O}_7$ by Potentiometric titration.
5. Estimation of the amount of chloride content of a water sample.
6. Determination of molecular weight of a polymer by viscosity average method
7. Determination of the strength of a mixture of AcOH & HCl by Conductometry.
8. Determination of strength of an acid using pH meter.

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Titration 2 : Standardisation of HCl.

S.No	Volume of Pipette (ml.)	Burette Initial (ml.)	Burette Final (ml.)	Concordant Value (ml.)	Indicator
					Methyl Orange.

DETERMINATION OF THE AMOUNT OF SODIUM CARBONATE AND SODIUM HYDROXIDE IN A MIXTURE BY TITRATION.

AIM:
To determine the amount of Na_2CO_3 and NaOH in a mixture using hydrochloric acid.

PRINCIPLE :

When a known volume of the mixture is titrated with HCl in presence of phenolphthalein, the acid reacts with all the sodium hydroxide and with only half of the carbonate.

When the titration is continued with methyl orange indicator, the remaining half of CO_3^{2-} ions will be neutralised with HCl at the end point.

PROCEDURE :

Titration 2 : Standardisation of HCl

Pipette out 20 ml of standard Na_2CO_3

solution in a 250 ml clean conical flask. Add 2-3 drops of Methyl Orange indicator. The solution will turn yellow in colour. Titrate the mixture against Hydrochloric acid taken in the burette. The end point is the colour change from yellow to orange.

Titration 3 : Estimation of the mixture ($\text{Na}_2\text{CO}_3 + \text{NaOH}$)

Make up the given solution in a 100 ml standard measuring flask. Pipette out 20 ml of the made up solution in a 250 ml conical flask. Add 2-3 drops of Phenolphthalein indicator. The solution will turn pink in colour. Titrate the mixture

Titration I : Estimation of the mixture $\text{Na}_2\text{CO}_3 + \text{NaOH}$

S.No	Volume of Burette soln. ($\text{Na}_2\text{CO}_3 + \text{NaOH}$)	Burette Reading cm ³		Concordant Value	Indicator.
		Initial H ₂ S ₂ O ₈	Final H ₂ S ₂ O ₈		
		M.O	M.O	Phenolphthalein	

A - Phenolphthalein End point - Volume of acid used upto H₂S₂O₈ end point.

B - Methyl Orange End point - Total volume of acid used till M.O end point.

(B - A) = C - Volume of acid used for CO_3^{2-} neutralisation.

2C - Total volume of acid used for complete neutralisation of CO_3^{2-} ions.

Calculations :

1. Estimation of Na_2CO_3

$$\text{Volume of HCl} (V_1) = 20 \text{ ml}$$

$$\text{Normality of HCl} (N_1) =$$

$$\text{Volume of the mixture} = 20 \text{ ml}$$

$$\text{Normality of the mixture} = ?$$

$$\frac{V_1 N_1}{HCl} = \frac{V_2 N_2}{\text{Mixture (Na}_2\text{CO}_3)}$$

$$N_2 = \frac{V_1 N_1}{V_2} = N.$$

Amount of Na_2CO_3 present in the whole of the given soln. = Normality of eq.wt Na_2CO_3 of Na_2O_3

$$= \frac{N_2 \times 53}{10} = g.$$

against standard Hydrochloric acid in a burette.
The end point is the colour change from pink to colourless.

RESULT:

Amount of NaOH present in 100ml of the given solution = 31L.

Amount of Na_2CO_3 present in 100ml of the given solution = 31L.

Estimation of NaOH:

$$\text{Volume of HCl} \cdot V_1 = A - c \text{ ml.}$$

Normality of HCl N_1 :

Volume of mixture $V_2 = 20 \text{ ml}$

Normality of " $N_2 =$

$$V_1 N_1 = V_2 N_2$$

Mixture (NaOH)

$$N_2 = \frac{V_1 N_1}{V_2} = N.$$

$$\left. \begin{array}{l} \text{Amount of NaOH present in the} \\ \text{whole of the given solution} \end{array} \right\} = \frac{\text{Normality} \times \text{Eq. wt.}}{\text{of NaOH}}$$

$$= \frac{N_2 \times 40}{10} = 9.$$

Simple Procedure:

	Titration I	Titration II
Burette Solution	Hydrochloric acid.	std. Hydrochloric acid.
Pipette solution	20 ml of std. Na_2W_3 normal. (0.05N)	20 ml of made up mixture ($\text{Na}_2\text{W}_3 + \text{NaOH}$)
Indicator	Methyl orange	Phenolphthalein
End Point	Yellow to orange pink to colourless (2 end pt.)	Yellow to orange yellow

Titration 1: Standardisation of EDTA.

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S. No	Volume of Std. Hard water	Burette Reading (ml) Initial Final	Volume of EDTA (V_1)	Indicator EBT

Calculations:

1 ml of std. hard water contains 1 mg of CaCO_3 equivalent hardness.

∴ 20 ml of std. hard water will contain 20 mg of CaCO_3 eq. hardness.

20 ml of std. hard water consumes = V_1 ml of EDTA.

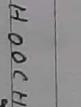
(or) V_1 ml of EDTA = 20 mg of CaCO_3 eq. hardness.

1 ml of EDTA = $\frac{20}{V_1}$ mg of CaCO_3 eq. hardness.

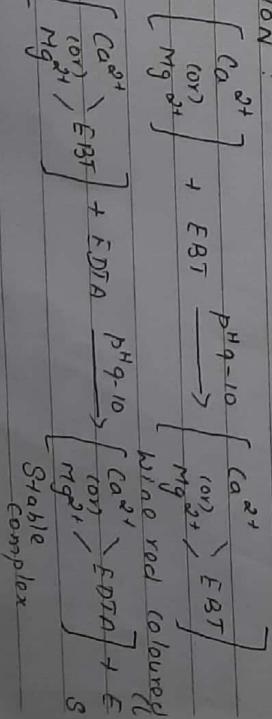
Principle:

Dissodium salt of ethylene diamine tetra acetic acid (EDTA) is used to determine the total hardness of the given hard water. The hardness causing ions (Ca^{2+} & Mg^{2+}) form a wine red coloured unstable complex with Eriochrome Black-T indicator (EBT) in presence of buffer solution. When EDTA is added, the indicator is replaced by EDTA and a stable complex is formed and the EBT indicator is leached out free. Hence, wine red colour changes to steel blue. This is the end point of a titration between EDTA & hard water.

The structural formula of EDTA is as follows:

$$\text{HOOCCH}_2\text{N}-\text{CH}_2-\text{CH}_2-\text{N}-\text{CH}_2\text{COOH}$$


REACTION:



DETERMINATION OF HARDNESS OF A WATER SAMPLE BY EDTA METHOD

AIM:

To estimate the amount of total hardness, permanent hardness and temporary hardness of given sample of water by EDTA method.

Titration 2 : Determination of Total Hardness.

No	Volume of sample hard water	Burette Reading (ml) Initial	Burette Reading (ml) Final	Volume of EDTA (V_2)	Indicator.
					Fig 7.

20 ml of sample hard water consumes = V_2 ml of EDTA.

$$\begin{aligned} &= V_2 \times 1 \text{ ml of EDTA} \\ &= V_2 \times \frac{20}{V_1} \text{ mg of CaCO}_3 \\ &\text{eq. hardness.} \end{aligned}$$

$$\therefore 1000 \text{ ml of sample hard water} \left\{ \text{will consume } V_2 \times \frac{20}{V_1} \times \frac{1000}{20} \text{ mg of CaCO}_3 \right. \\ \left. \text{eq. hardness.} \right\}$$

$$\begin{aligned} \text{Total Hardness} &= 1000 \times \frac{V_2}{V_1} \text{ mg of CaCO}_3 \\ &\text{eq. Hardness.} \end{aligned}$$

Titration 2 : Determination of total hardness.

Pipette out 20 ml of sample hard water into a clean conical flask. Add 5 ml of buffer soln. and 2-3 drops of EBT Indicator. The solution turns wine red in colour. Titrate it against EDTA taken in the burette. The colour change from wine red to steel blue is the end point. Repeat the titration for concordant values. Let the titre value be V_2 ml.

Titration 3 : Determination of permanent hardness
Pipette out 20 ml of boiled water into a clean conical flask. Add 5 ml of buffer &

When the sample water is boiled, bicarbonates of Ca Mg are converted into carbonates and hydroxides, which can be removed by filtration.

The permanent hardness which is not removed by boiling, is once again estimated by titration against EDTA with EBT.

PROCEDURE:

Titration 1 : Standardisation of EDTA.

Pipette out 20 ml of standard hard water in a 25 ml conical flask. Add 5 ml of buffer soln. and 2-3 drops of EBT Indicator. The solution turns wine red in colour. Titre it against EDTA taken in the burette. The

colour change from wine red to steel blue is the end point. Repeat the titration for concordant values. Let the titre value be V_1 ml

Titration 3 : Determination of Permanent Hardness

S.No	Volume of boiled hard water (ml)	Burette Reading (ml)		Volume of EDTA.	Indicator
		Initial	Final		
					EBT.

20 ml of boiled hard water sample } = V_3 ml of EDTA.

consumed

$$= V_3 \times 1 \text{ ml of EDTA}$$

$$= V_3 \times \frac{20}{V_1} \text{ mg of CaCO}_3 \text{ eq. hardness.}$$

$$\therefore 1000 \text{ ml of boiled water sample } = V_3 \times \frac{20}{V_1} \times \frac{1000}{20} \text{ mg of CaCO}_3 \text{ eq. hardness.}$$

$$\therefore \text{Permanent Hardness} = 1000 \times \frac{V_3}{V_1} \text{ mg of CaCO}_3 \text{ eq. hardness.}$$

$$\text{Total Hardness} = \text{Temporary Hardness} + \text{Permanent Hardness.}$$

$$\therefore \text{Temp. Hardness} = \text{Total Hardness} - \text{Permanent Hardness.}$$

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and 2-3 drops of EBT Indicator. The solution turns wine red in colour. Titrate it against EDTA taken in the burette. The colour change from wine red to steel blue is the end point. Repeat the titration for concordant values. Let the titre value be V_3 ml.

	Titration I	Titration II	Titration III
1. Burette Solution	EDTA	EDTA	EDTA
2. Pipette Solution	20 ml of standard hard water.	20 ml of sample hard water	20 ml of boiled hard water.
3. Additional Reagent	5 ml of buffer	5 ml of buffer	5 ml of buffer.
4. Indicator	2-3 drops of EBT	2-3 drops of EBT	2-3 drops of EBT
5. End point	colour change from wine red to steel blue.	Colour change from wine red to steel blue.	colour change from wine red to steel blue.

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RESULT:

1. Total Hardness = ppm.
2. Permanent Hardness: ppm.
3. Temporary Hardness = ppm.

Titration 2: Standardization of AgNO₃.

S.No	Volume of NaCl (ml)	Burette Reading (ml) Initial	Burette Reading (ml) Final	Volume of AgNO ₃ (ml)	Indicator
				K ₂ Cro ₄	

Calculation :

$$\begin{aligned} \text{Volume of AgNO}_3 &= V_1 \text{ ml} \\ \text{Normality of AgNO}_3 &= N_1 \\ \text{Volume of NaCl} &= 20 \text{ ml} \\ \text{Normality of NaCl} &= 0.02 N \end{aligned}$$

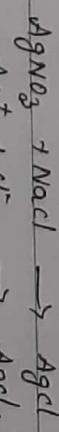
$$\begin{aligned} V_1 N_1 &= \frac{V_2 N_2}{N_1} \\ \text{AgNO}_3 & \\ \therefore N_1 &= \frac{V_2 N_2}{V_1} \\ &= \frac{V_2 N_2}{N} \end{aligned}$$

ESTIMATION OF THE AMOUNT OF CHLORIDE CONTENT IN A WATER SAMPLE

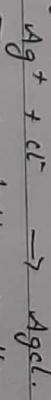
AIM:
To estimate the amount of chloride content in a water sample by Mohr's method.

PRINCIPLE:

T-1 is an example of precipitation reaction. The reaction between chloride and silver nitrate is direct and simple. It proceeds as follows:



The completion of the reaction in this case is observed by employing K₂Cro₄ solution as the indicator. At the end point, the yellow colour changes into reddish brown precipitate due to the reaction,



K₂Cro₄ will not be precipitated as Ag₂Cro₄ until all the chlorides in the solution have been precipitated as AgCl.

PROCEDURE:

Titration 2: Standardization of AgNO₃ solution.

Pipette out 20 ml of standard NaCl solution (N/20) into a clean conical flask. Add 1 ml of 2% K₂Cro₄ indicator and the solution will turn yellow in colour. Titrate it against std. AgNO₃ solution taken in the burette. Shake the solution well in the conical flask during each addition of

Titration 2 : Estimation of chloride.

S.No	Volume of given chloride soln. (ml)	Burette Reading (ml)	Volume of Indicator
	Initial	Final	AgNO ₃
			K ₂ CrO ₄

Volume of AgNO₃ = V₁ ml

Normality of AgNO₃ = N₁

Volume of chloride solution = 20 ml

Normality of chloride solution = N₂

$$\begin{aligned} V_1 N_1 &= V_2 N_2 \\ \text{AgNO}_3 &\text{Chloride} \\ V_1 N_1 &= 20 \times N_2 \\ N_2 &= \frac{V_1 N_1}{V_2} = N. \end{aligned}$$

Amount of chloride present in the whole of the given solution } = Normality × Eq.wt. of chloride

$$= \frac{N_2 \times 35.46}{10}$$

$$= 3.12$$

AgNO₃. Near the end point, the discharge of red colour will be seen in the solution which will not be discharged by further shaking. Repeat the titration for concordant values

Titration 3 : Estimation of chloride.

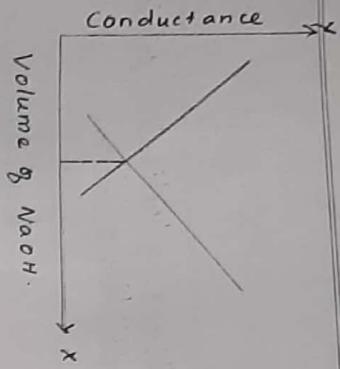
Make up the given chloride solution in a 100 ml standard measuring flask. Pipette out 20 ml of this solution into a clean conical flask.

Add 1 ml of 2% K₂CrO₄ solution to this and titrate it against standardised AgNO₃ taken in the burette. Continue the titration till it produces reddish brown hinge. Repeat the titration for concordant values.

Result: Amount of chloride present in the whole
of the given solution = 91.

	Titration I	Titration II
Burette Solution	AgNO ₃	AgNO ₃
Pipette Solution	20 ml of Std NaCl (0.02 N)	20 ml of made up Chloride solution
Additional Reagent	-	-
Indicator	1ml of 2% K ₂ Cro ₄	1ml of 2% K ₂ Cro ₄
End point	Appearance of Red ppt.	Appearance of Red ppt.

MODEL GRAPH:



TH 1 : Titration between Std NaOH & Unknown HCl (Pilot)

S.No	Volume of NaOH (ml)	conductance.
—	—	—

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DETERMINATION OF STRENGTH OF AN ACID BY
CONDUCTOMETRIC TITRATION

AIM :

To determine the strength of a given HCl solution with Std NaOH by conductometric titration.

PRINCIPLE :

It is based on the measurement of the change of conductance with the help of a conductivity meter. The conductance of the solution depends on the number of ions and their mobility.

During the titration, the base is added to the strong acid and H^+ ions are replaced by slow moving Na^+ ions. Hence the conductance will decrease after the neutralization point, further addition of excess alkali introduces fast moving OH^- ions hence the conductance increases. The end point of the titration is determined graphically by plotting the conductance of HCl against the volume of alkali added. The point of intersection of the straight line gives the end point at the volume axis (x-axis).

PROCEDURE :

make up the given HCl solution in a 100 ml standard measuring flask. Pipette out 10 ml of the made up HCl into a beaker. Dilute the solution with distilled water, so that the conductivity cell can be immersed well in the solution (upto 100 ml). Stir the solution well with the help of a glass rod. Note down the conductance of the solution from the conductivity meter

Titration 2: Titration between Std. NaOH & Unknown HCl
(Fair)

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S.NO	Volume of NaOH (ml)	Conductance

Fill the burette with standard NaOH solution and run down into the beaker in small increments (1 ml) with gentle stirring of the contents of the beaker. After each addition, stir the contents of the beaker and after an equilibrium time of 2-3 minutes note the corresponding conductance value and tabulate it. Continue the titration upto 15 increments, after the conductance reaches a minimum and starts increasing. After the completion of the titration, wash the conductance cell with distilled water and immerse in distilled water.

Plot a graph between conductivity and volume of NaOH added. The intersection of two lines to the volume axis gives the end point.

Volume of NaOH required for neutralization is taken from graph (Fair) on intersection point which is corresponding to the volume axis (x). Then to get accurate results, perform a fair titration by adding NaOH in small amounts (0.2 ml) near and beyond the end point. Calculate the strength of the given strong acid from the given NaOH strength.

Calculations:

$$\text{Volume of HCl} = V_1 \text{ ml (10 ml)}$$

$$\text{Normality of HCl} = N_1$$

$$\text{Volume of NaOH} = V_2 \quad (\text{from the 3rd graph})$$

$$\text{Normality of NaOH} = N_2 \quad (0.1 N)$$

$$\frac{V_1 N_1}{N_1} = \frac{V_2 N_2}{V_1}$$

$$= \frac{V_2 N_2}{V_1} = N$$

Ans. The strength of given HCl solution = N.

SIMPLE PROCEDURE:

Burette solution - Standard NaOH (0.1N)

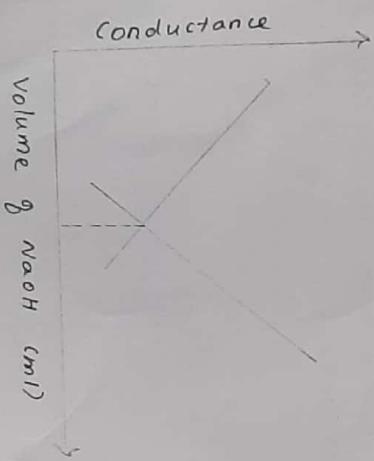
Pipette solution - 10 ml of made up HCl solution with 100 ml of distilled water in a 250 ml beaker.

Conductivity cell is immersed in the beaker solution & conductivity is measured for every 1ml addition of NaOH from the burette. Initially conductance decreases and after some stage, it will start increasing. The point at which conductance starts increasing is the end point. This is the 'pilot titration' which will be done for 1ml increments of NaOH from the burette.

A 'fair titration' is conducted in the same manner before and after 2ml of the end point range (0.2ml increments) to get the accurate value.

From the point of intersection of fair titration graph, the volume of NaOH is found out. From the volume of NaOH, the strength of given HCl solution is calculated.

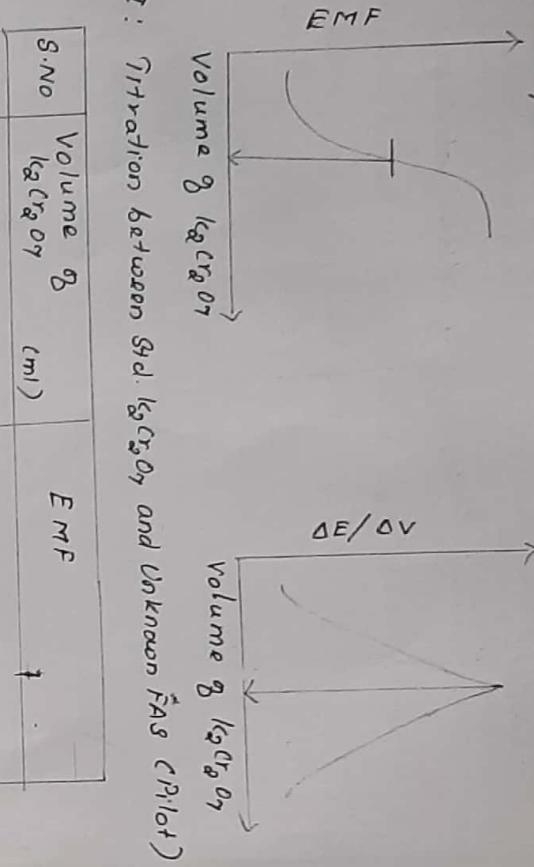
Model Graph:



RESULT:

The strength of given HCl solution = _____

Model Graphs:



Tit I: Titration between std. $K_2Cr_2O_7$ and Unknown Fe^{2+} soln.

ESTIMATION OF THE AMOUNT OF Fe^{2+} IONS BY POTENSIOMETRY

AIM:

To estimate the amount of Fe^{2+} ions by potentiometric titration.

PRINCIPLE:

Measurement of EMF of an electrochemical cell using potentiometer and the change in emf due to a redox reaction is monitored. In this potentiometric titration set up, an indicator electrode (Cu electrode) and reference electrode (Calomel) is coupled to form an electrochemical cell for $Fe^{2+} \rightarrow Fe^{3+}$.

Fe^{2+} is oxidised to Fe^{3+} as $K_2Cr_2O_7$ is progressively added. Pt electrode, which is kept in contact with a mixture of Fe^{2+} & Fe^{3+} ions act as a redox electrode. The reduction potential of this single electrode depends on Fe^{2+} ions in H_2SO_4 medium with the ratio of Fe^{2+}/Fe^{3+} . This ratio varies to a little extent at the beginning and suddenly near the end point. After the end point, the ratio change is very little. It is noted that, the sudden change in the ratio denotes that the equivalence point is reached. This causes a sudden increase in the emf of the cell at equivalence point.

Cell set up: $Hg/HgCl_2(s)$, $KCl(aq)$ || Fe^{2+}/Fe^{3+} ; Pt.

Chemical Reaction: $6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightleftharpoons 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

PROCEDURE:

Make up the given borax 100 ml solution in a 100 ml standard measuring flask. Pipette out 10 ml of the made up solution into a clean 250 ml beaker. Add 10 ml of dil. H_2SO_4 & 100 ml of distilled water to

Titration between Std. $\text{K}_2\text{Cr}_2\text{O}_7$ and Unknown FAS (Fair)

S.No	Volume of $\text{K}_2\text{Cr}_2\text{O}_7$ (ml)	ENF	ΔE	ΔV	$\Delta E / \Delta V$

Fit the pH electrode coupled with calomel electrode in this solution. Incorporate the resulting cell into the potentiometer circuit. Fill the burette with Std. $\text{K}_2\text{Cr}_2\text{O}_7$ solution and add in 1 ml increments into the beaker and measure the cell emf after each addition by proper rinsing. Continue the process till and well before neutralisation point as indicated by an abrupt change in emf. Note the volume of $\text{K}_2\text{Cr}_2\text{O}_7$ required for complete oxidation of Fe^{2+} from the plot of emf vs volume of $\text{K}_2\text{Cr}_2\text{O}_7$ added. Calculate the range at which sudden rise in emf occurs. Perform the similar titration by adding 0.1 ml of std. Fe^{2+} close to the end point. 1 ml on either side of the range tabulate the emf measured corresponding to each addition. Plot 2 graphs by taking

1. ENF vs Volume of $\text{K}_2\text{Cr}_2\text{O}_7$ added
2. $\Delta E / \Delta V$ vs volume of $\text{K}_2\text{Cr}_2\text{O}_7$ added

calculations:

$$\text{Volume of FAS} = V_1 \text{ (10 ml)}$$

Normality of FAS = N_1

Volume of $\text{K}_2\text{Cr}_2\text{O}_7$ = V_2 (from 1st graph)

Normality of $\text{K}_2\text{Cr}_2\text{O}_7$ = 0.1 N.

$$\frac{V_1 N_1}{V_2 N_2} = \frac{V_2 N_2}{\text{K}_2\text{Cr}_2\text{O}_7}$$

$$N_1 = \frac{V_2 N_2}{V_1} = N$$

Amount of Fe^{2+} present in the whole of the given soln. } = Normality of Fe^{2+} $\times \frac{\text{Eq. wt}}{\text{Molar mass}}$

SIMPLE PROCEDURE

Burette soln. - Std. $\text{K}_2\text{Cr}_2\text{O}_7$ (0.1N) Pipette 10ml - 10ml of made up FAS 10%.

Additional soln. - 10ml of dil. H_2SO_4 .

Electrodes : Reference - calomel electrode, Indicator electrode - Fe^{2+}

Eq. wt. of Fe^{2+} - 56.85

10ml of made up 10% of FAS & 10ml of dil. H_2SO_4 to 100ml of distilled water is taken in a 250ml beaker. The two electrodes are immersed in the beaker and a glass rod is put for stirring. $\text{E}_{\text{H.F}}$ is measured for each ml of $\text{K}_2\text{Cr}_2\text{O}_7$ from the burette till 15ml. $\text{E}_{\text{H.F}}$ increases gradually and at the neutralisation point, there is a steep increase in the $\text{E}_{\text{H.F}}$. Fair titration is also conducted in the same manner before and after 8ml of the end points (in portions of 2ml) to get accurate value. From the plot of $\text{E}_{\text{H.F}}$ vs. Volume of $\text{K}_2\text{Cr}_2\text{O}_7$, the end point is obtained. The plot of $\Delta E/\Delta V$ vs. Volume of $\text{K}_2\text{Cr}_2\text{O}_7$ gives the accurate end point. From the volume of $\text{K}_2\text{Cr}_2\text{O}_7$ in the graph, the amount of given Fe^{2+} solution is calculated.

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RESULT:

The amount of Fe^{2+} present in the whole of the given solution is = 91.

Titration between Std. NaOH and mixture of acids.

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S.No	Volume of NaOH added (ml)	Conductance
1	0	Max
2	10	Min
3	20	Max
4	30	Min
5	40	Max
6	50	Min
7	60	Max
8	70	Min
9	80	Max
10	90	Min
11	100	Max

DETERMINATION OF THE STRENGTH OF A MIXTURE OF ACETIC ACID AND HCl BY CONDUCTOMETRY.

Aim:

To estimate the strength of mixture of acetic acid and HCl by present in a given mixture by conductometry.

PRINCIPLE:

The conductivity of the solution is related to the mobility of ions which in turn related to the size of the ions. When a mixture of acids & a strong acid (HCl) and a weak acid (acetic acid) are titrated against a strong base (NaOH), strong acid reacts first followed by a weak acid. When the titration of strong acid and strong base are carried out, there is a decrease in conductivity as highly mobilised hydrogen ions are replaced by sodium ions.

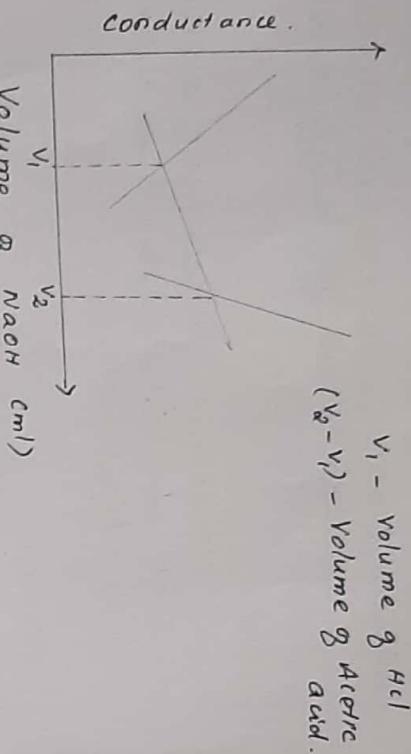


When the whole strong acid is consumed base reacts with weak acid and conductivity increases as unionised weak acid becomes the ionised salt.



After both the acids are consumed, there is a steep increase in conductivity which gives the end point and this increase in conductivity is due to the fast moving OH⁻ ions from the base. From this, amount of base consumed and in turn, the amount of acids present is calculated.

Model Graph:



Calculation:

$$V_1 = \text{Volume of HCl}$$

$$(V_2 - V_1) = \text{Volume of Acetic acid.}$$

$$\text{Normality of HCl } N_1 = \frac{V_1}{20} \text{ M.L.}$$

$$\text{Normality of NaOH } N_2 = \frac{V_2}{20} \text{ M.L.}$$

$$\text{Volume of NaOH } V_2 = V_1 \text{ ml. (1st end point)}$$

$$\text{Normality of NaOH } N_2 = 0.05 \text{ N}$$

$$\frac{V_1 N_1}{N_1} = \frac{V_2 N_2}{N_2}$$

$$N_1 = \frac{V_2 N_2}{V_1} = N$$

$$\text{Volume of CH}_3\text{COOH } V_1 = 20 \text{ ml.}$$

$$\text{Normality of CH}_3\text{COOH } N_1 = \frac{V_1}{20} \text{ M.L.}$$

$$\text{Volume of NaOH } V_2 = (V_2 - V_1) \text{ ml}$$

$$\text{Normality of NaOH } N_2 = 0.05 \text{ N.}$$

$$\frac{V_1 N_1}{N_1} = \frac{V_2 N_2}{N_2}$$

$$N_1 = \frac{V_2 N_2}{V_1} = N$$

RESULT :

1. The strength of HCl present in the whole of the given solution } = N

2. The strength of acetic acid present in the whole of the given solution } = N

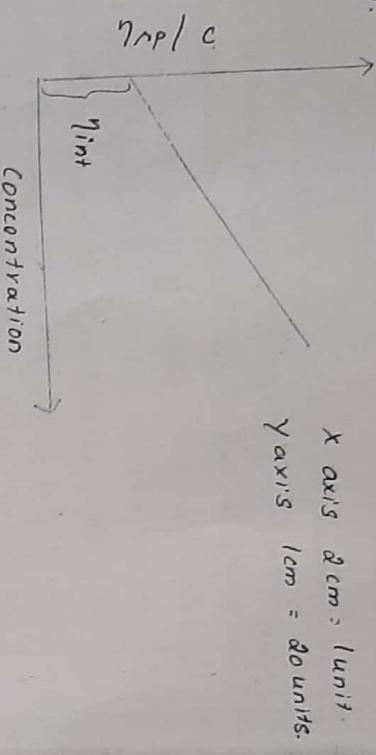
PROCEDURE :

Make up the given mixture of acids to 100 ml using distilled water in a standard measuring flask.

Pipette out 20 ml of this made up soln. in a 250 clean beaker and add 100 ml distilled water to it.

Dip the conductivity cell into a test solution and add the base (NaOH) from the burette at 0.5 ml. interval with uniform stirring. Measure the conductance after each addition of NaOH. After complete neutralisation, determine the amount of acid present in the given mixture based on the volume of base consumed. It is found out by plotting a graph between conductance and volume of base added where the first end point corresponds to strong acid and second end point corresponds to weak acid.

Model Graph:



Aim:

To determine the molecular weight of a polymer by viscosity average method using Ostwald viscometer.

PRINCIPLE:

Sno	Concentration of HO polymer solution	Time of flow (sec)	Relative viscosity $\eta_r = \frac{t_f}{t_0}$	Specific viscosity $\eta_{sp} = \frac{\eta_r - 1}{c}$	Reduced viscosity $\eta_{sp}/c \times 100$
1.	Pure Solvent				
2.	0.1% (2.5 ml)				
3.	0.2% (5 ml)				
4.	0.3% (7.5 ml)				
5.	0.4% (10 ml)				
6.	0.5% (12.5 ml)				

The value of η/η_0 is known as the relative viscosity η_{rel} . In dilute solutions, which are often employed for molecular weight determinations, ρ is not much different from ρ_0 and hence

$$\eta_{rel} = \frac{\eta}{\eta_0} = \frac{t}{t_0}$$

Calculation:

$$\text{Formula} \quad \eta_{int} = k M^\alpha$$

$$M = \left[\frac{\eta_{int}}{k} \right]^{\frac{1}{\alpha}}$$

$$k = 45.3 \times 10^{-3}; \alpha = 0.64.$$

$$M = \left[\frac{\eta_{int} \times 10^3}{45.3} \right]^{\frac{1}{0.64}}$$

=

SIMPLE PROCEDURE:

From 1 v. of the given solution of

PVA (Poly Vinyl Alcohol) five different concentrations (0.1%, 0.2%, 0.3%, 0.4%, & 0.5%) are prepared as follows.

$$1. \quad 0.1\% = \frac{25 \times 0.1}{1} = 2.5 \text{ ml}$$

$$2. \quad 0.2\% = \frac{25 \times 0.2}{1} = 5 \text{ ml}$$

$$3. \quad 0.3\% = \frac{25 \times 0.3}{1} = 7.5 \text{ ml}$$

$$4. \quad 0.4\% = \frac{25 \times 0.4}{1} = 10 \text{ ml}$$

$$5. \quad 0.5\% = \frac{25 \times 0.5}{1} = 12.5 \text{ ml.}$$

12 ml of pure distilled water is pipetted into the viscometer and the flow time is measured. The flow time is the time taken for liquid from upper mark to lower mark. The flow time is measured for all the concentrations as before.

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The specific viscosity η_{sp} is defined as

$$\eta_{sp} = \eta_{sp} - 1$$

A plot of η_{sp} vs c is a straight line for dilute solutions. No intercept

$$\frac{\eta_{sp}}{c} = \eta_{int}$$

of which is known as the intrinsic viscosity.

The Staudinger-Hark-Hawkins equation

$$\eta_{int} = k(M)^\alpha$$

where k is an empirical parameter characteristic of a particular solute-solvent pair and α is also a parameter which can vary from about 0.5 for polycyclic polymers. From the known values of k & α , molecular weight can be determined.

Plot of η_{sp} vs concentration of a polymer solution to find out intrinsic viscosity.

PROCEDURE:

Preparation of various concentrations of polymer in water (solvent)

1% solution of polymer in water will be supplied. We need to prepare at least 5 dilutions viz 0.1%, 0.2%, 0.3%, 0.4% & 0.5% polymer in water before carrying out the experiment.

Dilutions can be done by using volumetric equation $V_1 N_1 = V_2 N_2$.

Set up the Ostwald viscometry and measure the flow time (t₀) of a fixed volume of the pure solvent

Relative viscosity is found out by using $\eta_{rel} = \frac{t_1}{t_0}$.
 Specific viscosity is found out by using $\eta_{sp} = \frac{\eta_{sp}}{c} - 1$.

Reduced viscosity is found out by using $\eta_{red} = \frac{\eta_{sp}}{c} \times 100$.
 The flow time keeping the flow volume the same.

A plot of η_{sp} vs concentration gives a straight line. By extrapolation η_{int} is found out.

Using Mark-Houwink's equation

$$\eta = \left(\frac{\eta_{int}}{K} \right)^{\alpha}$$

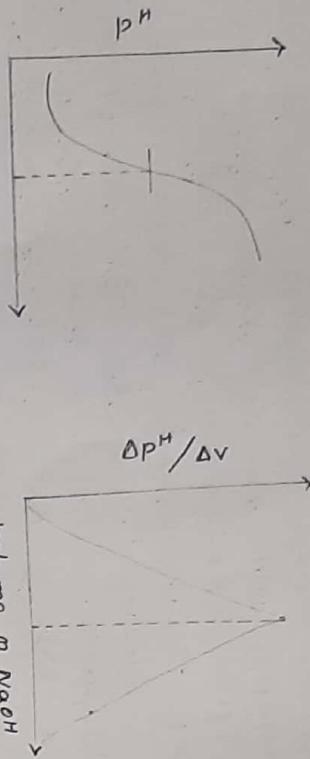
is found out.

$$K = 45.3 \quad \alpha = 0.64$$

RESULT:

The molecular weight of the given polymer is =

Model Graph:



TITRATION: Titration between Std. NaOH & Unknown HCl (Pilot)

B.NO	Volume of NaOH (ml)	pH

AIM:

To find out the strength of given HCl by titrating it against NaOH (0.1N) using pH meter.

PRINCIPLE:

When an alkali is added to an acid solution, the pH of the solution increases slowly, but at vicinity of the end point, the rate of change of pH of the solution is very rapid. From the sharp break in the curve, we can find out the end point, from which the strength of HCl can be calculated.

PROCEDURE:

Make up the given HCl solution to 100 ml using distilled water. Pipette out 10 ml of this made up solution into a clean 250 ml beaker and add 100 ml of distilled water. Dip the pH meter into the solution and note the pH. Fill the burette with Std. NaOH (0.1N) solution and add 1 ml increments into the beaker, and measure the pH after each addition. Continue the process till and well beyond the neutralisation point as indicated by an abrupt change in pH. Calculate the range at which sudden rise in pH occurs. Perform the similar titration by adding 0.2 ml of Std. NaOH close to the end point. (ml on either side of the range). Tabulate the pH measured corresponding to each addition.

Tit 2 : Titration between Std. NaOH & Unknown HCl (Fair)

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Plot 2 graphs by taking

1. p_H vs volume of NaOH added
2. $\Delta p_H / \Delta V$ vs volume of NaOH added.

The exact end point should be calculated from the 2nd graph.

RESULT:

The strength of given HCl solution is = N_1

S.No	Volume of NaOH (cm ³)	p_H	Δp_H	ΔV	$\Delta p_H / \Delta V$
1					
2					
3					
4					
5					
6					
7					
8					
9					
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11					
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100					

Calculations:

$$\text{Volume of HCl} = V_1 \text{ ml (10m)}$$

$$\text{Normality of HCl} = N_1$$

$$\text{Volume of NaOH} = V_2 \quad (\text{from the 2nd graph})$$

$$\text{Normality of NaOH} = (0.1 \text{ N}) N_2.$$

$$V_1 N_1 = V_2 N_2$$

$$N_1 = \frac{V_2 N_2}{V_1} = N.$$

The strength of given HCl solution = N.