DEPARTMENT OF CHEMISTRY
Titration 1: Standardisation of EDTA

Standard hard water vs EDTA

<table>
<thead>
<tr>
<th>S. No</th>
<th>Volume of Std. hard water (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of EDTA ($V_1$) (ml)</th>
<th>Concordant value (ml)</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Initial</td>
<td>Final</td>
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</tr>
<tr>
<td>1.</td>
<td>20</td>
<td>0</td>
<td></td>
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<tr>
<td>2.</td>
<td>20</td>
<td>0</td>
<td></td>
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</tr>
</tbody>
</table>

Calculations

1 ml of standard hard water contains 1 mg of CaCO$_3$
20 ml of standard hard water contains 20 mg of CaCO$_3$
20 ml of standard hard water consumes $V_1$ ml of EDTA
$V_1$ ml of EDTA is equivalent to 20 mg of CaCO$_3$
1 ml of EDTA = $\frac{20}{V_1}$ mg of CaCO$_3$ equivalent.
1. ESTIMATION OF HARDNESS OF WATER BY EDTA METHOD

AIM

To estimate the amount of total hardness, permanent hardness and temporary hardness of a given hard water sample by EDTA method. You are provided with a standard solution of calcium of strength 0.01 M.

CHEMICALS REQUIRED

EDTA, Eriochrome black-T, water sample, ammoniacal buffer solution.

PRINCIPLE

The amount of calcium and magnesium ions present in a given sample of water is estimated by titration with standard EDTA. The hardness is expressed in terms of calcium carbonate equivalence.

Hardness

It is one of the important characteristics of water. Hardness of water is due to the presence of calcium and magnesium salts. Lather is not produced until the cations, Ca\(^{2+}\) and Mg\(^{2+}\) are completely precipitated out in the form of insoluble salts. For example,

\[
2C_{17}H_{35}COONa + CaSO_4 \rightarrow (C_{17}H_{35}COO)_2Ca + Na_2SO_4
\]

sodium stearate (in water) (ppt)

\[
2C_{17}H_{35}COONa + MgCl_2 \rightarrow (C_{17}H_{35}COO)_2Mg + 2NaCl.
\]

sodium stearate (in water) (ppt)

Temporary hardness

It is due to bicarbonates of calcium and magnesium. It can be removed by boiling. The salts are precipitated as carbonates and hydroxides respectively, which can be removed by filtration leaving only permanent hardness producing salts.

Permanent hardness

It is due to the sulphates and chlorides of calcium and magnesium.

Hardness is expressed in terms of CaCO\(_3\) equivalence

The reason for choosing CaCO\(_3\) as the standard for calculating hardness of water is due to its molecular weight being 100, which makes the calculations easier.

- It is the most insoluble salt, thus can be easily precipitated in water treatment processes.
Titration 2: Determination of Total Hardness

Sample hard water vs standard EDTA

<table>
<thead>
<tr>
<th>S. No</th>
<th>Volume of sample hard water (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of EDTA ($V_2$) (ml)</th>
<th>Concordant value (ml)</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Initial</td>
<td>Final</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1.</td>
<td>20</td>
<td>0</td>
<td>0</td>
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<tr>
<td>2.</td>
<td>20</td>
<td>0</td>
<td>0</td>
<td></td>
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</tr>
</tbody>
</table>

Calculations

20 ml of hard water sample
\[= V_2 \times \frac{20}{V_1} \text{ mg of CaCO}_3 \text{ equivalent}\]

1000 ml of hard water sample
\[= V_2 \times \frac{20}{V_1} \times 1000/20 \text{ mg of CaCO}_3 \text{ equivalent}\]

The total hardness of the given hard water sample = \(1000 \times \frac{V_2}{V_1}\) ppm.
Determination of hardness

The hardness of water can be determined by complexometric titration. EDTA is used as a complexing agent. The Ca$^{2+}$ and Mg$^{2+}$ present in water are titrated with EDTA using Eriochrome Black T as indicator.

The Chemical structure of EDTA (Ethylene diamine tetra acetic acid) is represented as below:

\[
\text{HOOCH}_{2}\text{C} \quad \text{N} - \text{CH}_2 - \text{CH}_2 - \text{N} \quad \text{CH}_2\text{COO}^-\text{Na}^+ \\
\text{Na}^+\text{COO}^-\text{H}_2\text{C} \\
\text{STRUCTURE OF EBT}
\]

The indicator Eriochrome Black–T (EBT), forms a wine red colored unstable complex with calcium and magnesium ions in hard water at a pH of 8-10.

\[
\text{Mg}^{2+}/\text{Ca}^{2+} + \text{EBT} \rightarrow [\text{Mg/Ca} - \text{EBT}] \\
\text{unstable complex (wine red)}
\]

As this solution is titrated against EDTA, the Ca$^{2+}$ and Mg$^{2+}$ ions from the indicator complex forms stable metal ion–EDTA complex leaving the indicator free. The free form of indicator is steel blue in colour. The metal–EDTA complex is stable at pH 8-10. This pH range can be maintained by using ammoniacal buffer (NH$_4$Cl + NH$_4$OH).

\[
\text{Mg/Ca} - \text{EBT} + \text{EDTA} \rightarrow [\text{Mg/Ca} - \text{EDTA}] + \text{EBT} \\
\text{Wine red complex} \quad \text{Colourless stable complex} \quad \text{Steel blue}
\]

Thus the amount of EDTA used corresponds to the hardness of water.

PROCEDURE

Titration 1: Standardisation of EDTA.

The burette is washed well with distilled water and then rinsed with the given EDTA solution. It is then filled up to the zero mark without any air bubble. 20 ml of standard hard water is pipetted out into a clean conical flask. 5ml of ammonical buffer (NH$_4$Cl + NH$_4$OH) and 2 drops of EBT indicator is added. The solution turns wine red in colour. It is then titrated against EDTA solution taken in the burette. The change of colour from wine red to
Titration 3: Determination of Permanent Hardness

Boiled water sample vs Standardised EDTA

<table>
<thead>
<tr>
<th>S. No</th>
<th>Volume of boiled hard water Sample (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of EDTA ($V_3$) (ml)</th>
<th>Concordant value (ml)</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Initial</td>
<td>Final</td>
<td></td>
<td>Eriochrome black –T</td>
</tr>
<tr>
<td>1.</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>2.</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Calculations

20 ml of hard water sample

\[
20 \text{ ml of hard water sample } = V_3 \times \frac{20}{V_1} \text{ mg of CaCO}_3 \text{ equivalent}
\]

100 ml of hard water sample

\[
100 \text{ ml of hard water sample } = V_3 \times \frac{20}{V_1} \times \frac{100}{20}
\]

1000 ml of hard water sample

\[
1000 \text{ ml of hard water sample } = V_3 \times \frac{20}{V_1} \times \frac{100}{20} \times 10
\]

\[
= \frac{V_3}{V_1} \times 1000 \text{ ppm}
\]

The permanent hardness of the given sample water = \[
\frac{V_3}{V_1} \times 1000 \text{ ppm}
\]
steel blue is the end point. Repeat the titration for concordant values. Let the titre value be $V_1$ ml.

**Titration 2: Estimation of total hardness of sample water**

20 ml of the given hard water sample is pipetted out into a clean conical flask. 5 ml of ammonical buffer solution and 2 drops of EBT indicator are added. The solution turns wine red in colour. The solution is then titrated against EDTA solution taken in the burette. The change of colour from wine red to steel blue is the end point. The titration is repeated for concordant values. Let the titre value be $V_2$ ml. This gives the total hardness of the sample water.

**Titration 3: Estimation of permanent Hardness**

100 ml of the given sample water is taken in a clean 250 ml beaker and boiled for 10-15 minutes until it is reduced to 1/5th of the original volume. It is then cooled and filtered. The filtrate is collected in a 100 ml standard flask and made up to the mark using distilled water. 20 ml of this made up solution is pipetted out into a clean conical flask and the titration is repeated in the same way as in titration 1. Let the titre value be $V_3$ ml. This gives the permanent hardness of the sample water.

**Estimation of temporary hardness**

Temporary hardness of the water sample is calculated by subtracting permanent hardness from total hardness.

Temporary hardness = Total Hardness – Permanent hardness.

**RESULT**

1. The total hardness of the given water sample = _______ ppm
2. The permanent hardness of the given water sample = _______ ppm
3. The temporary hardness of the given water sample = _______ ppm.
Step I : Standardization of sodium thiosulphate

Titration I (Standard dichromate vs Sodium thiosulphate)

<table>
<thead>
<tr>
<th>Sl.No</th>
<th>Volume of Potassium dichromate (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of Sodium thiosulphate (ml)</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Initial</td>
<td>Final</td>
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<td></td>
</tr>
<tr>
<td>1</td>
<td>20</td>
<td>0</td>
<td></td>
<td>$V_1$</td>
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<td>2</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Volume of Potassium dichromate ($V_1$) = 20 ml
Strength of Potassium dichromate ($N_1$) = 0.0125 N
Volume of Sodium thiosulphate ($V_2$) = --- ml

Strength of Sodium thiosulphate ($N_2$) = --- N

According to volumetric formula

$$V_1N_1 = V_2N_2$$

$$N_2 = \frac{V_1N_1}{V_2}$$

Strength of Sodium thiosulphate ($N_2$) = --- N
2. ESTIMATION OF DISSOLVED OXYGEN OF BOILER FEED WATER

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<thead>
<tr>
<th>Expt. No.</th>
<th>Date</th>
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</thead>
</table>

AIM

To determine the dissolved oxygen in the boiler feed water.

CHEMICALS REQUIRED

Na$_2$S$_2$O$_3$ (N/40), MnSO$_4$ solution, KI, starch, conc. H$_2$SO$_4$

PRINCIPLE

Oxygen dissolves in water to the extent of 7 - 9 mgs/lit at a temperature range of 25 - 35°C. The estimation of dissolved oxygen in water is useful in studying water pollution. Water sample is collected carefully avoiding aeration/deaeration in ground stoppered flask. Initially manganous sulphate and alkali-iodide reagents are added and the reaction occur as follows

\[
\text{Mn}^{2+} + 2\text{OH}^- \rightarrow \text{Mn(OH)}_2 \downarrow \text{ (White)}
\]

\[
\text{Mn(OH)}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{MnO(OH)}_2 \downarrow \text{ (Yellow brown)}
\]

Potassium iodide and the precipitate react with concentrated sulphuric acid liberating iodine and the liberated iodine is titrated against Na$_2$S$_2$O$_3$

\[
\text{MnO(OH)}_2 + 2\text{H}_2\text{SO}_4 \rightarrow \text{Mn(SO}_4)_2 + 3\text{H}_2\text{O}
\]

\[
\text{Mn(SO}_4)_2 + 2\text{KI} \rightarrow \text{MnSO}_4 + \text{K}_2\text{SO}_4 + \text{I}_2
\]

\[
2\text{Na}_2\text{S}_2\text{O}_3 + \text{I}_2 \rightarrow \text{Na}_2\text{S}_4\text{O}_6 + 2\text{NaI}
\]

PROCEDURE

Titration I

**Standardisation of sodium thiosulphate**

The burette is washed and rinsed with sodium thiosulphate solution. Then the burette is filled with the given sodium thiosulphate solution. 20 ml of 0.0125N potassium dichromate solution is pipetted out into a clean conical flask. To this, 5 ml of sulphuric acid and 50 ml of 5% potassium iodide are added. This is titrated against sodium thiosulphate solution. When the solution becomes straw yellow colour, starch indicator is added and then titration is continued.

The end point is disappearance of blue colour and appearance of light green colour. The titration is repeated to get concordant values.
Step II : Estimation of dissolved oxygen

Titration II (Water sample vs Sodium thiosulphate)

<table>
<thead>
<tr>
<th>Sl.No</th>
<th>Volume of Potassium dichromate (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of Sodium thiosulphate (ml) ( V_2 )</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>20</td>
<td>0</td>
<td></td>
<td>Starch</td>
</tr>
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<td>20</td>
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<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Volume of Sodium thiosulphate \( (V_1) \) =

Strength of Sodium thiosulphate \( (N_1) \) =

Volume of water sample \( (V_2) \) = 100 ml

Strength of water sample \( (N_2) \) = —— N

According to volumetric formula

\[
V_1N_1 = V_2N_2
\]

\[
N_2 = \frac{V_1N_1}{V_2}
\]

Amount of dissolved oxygen in one litre of tap water =

Normality \( \times \) Eq.wt.of \( O_2 \) \( \times \) 1000 mg.

\[
= —— N \times 8 \times 1000
\]

\[
= —— mg/lit
\]
Titration II

Estimation of dissolved oxygen

100-150ml of the water sample is taken in the iodine flask, 2ml of manganese sulphate and 2 ml of alkali-iodide are added. The stopper is replaced and the flask is inverted and shaken several times for thorough mixing of the reagents. The flask is left aside for sometime. When half of the precipitate settles down, the stopper is removed and 2 ml of concentrated sulphuric acid is added.

The stopper is replaced and the flask is inverted several times for complete dissolution of the precipitate. 100 ml of the brown coloured solution is pipetted out and titrated against standardized sodium thiosulphate solution. Starch indicator is added when the solution becomes light yellow. The titration is continued until the blue colour disappears. From the titre value, the strength of dissolved oxygen and hence the amount of dissolved oxygen in the water sample is calculated.

RESULT

Amount of dissolved oxygen in water sample = --------- mg/lit
Titration 1: Standardisation of silver nitrate

Std. NaCl vs AgNO₃

<table>
<thead>
<tr>
<th>Sl. No</th>
<th>Volume of std. NaCl (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of AgNO₃ (V₁ ml)</th>
<th>Concordant value (ml)</th>
<th>Indicator</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
<td>K₂CrO₄</td>
</tr>
<tr>
<td>2.</td>
<td>20</td>
<td>0</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Calculation of the Strength of silver nitrate

Volume of std. NaCl \( (V₂) = 20 \text{ ml} \)

Strength of NaCl \( (N₂) = 0.01 \text{ N} \)

Volume of AgNO₃ \( (V₁) = \_\text{ ml} \)

Strength of AgNO₃ \( (N₁) = ? \)

According to volumetric formula

\[
V₁N₁ = V₂N₂ \text{ i.e., } N₁ = V₂ \times \frac{N₂}{V₁}
\]

\[
N₁ = 20 \text{ ml} \times 0.01 \text{ N/V₁}
\]

Strength of AgNO₃ \( (N₁) = \_\text{ N} \)
3. ESTIMATION OF CHLORIDE CONTENT IN WATER BY ARGENTOMETRIC METHOD (MOHR’S METHOD)

Expt. No.  Date

AIM

To estimate the amount of chloride ion present in the water sample by Argentometric method (Mohr’s method). You are provided with standard NaCl solution of strength 0.01 N and a link solution of AgNO₃ (approximately 0.01N).

CHEMICALS REQUIRED

Standard NaCl solution, AgNO₃ solution, Potassium chromate indicator

PRINCIPLE

Natural water contains chloride ions in the form of NaCl, KCl, CaCl₂, MgCl₂. The concentration of chloride ions more than 250 ppm is not desirable for drinking purposes.

This determination is based on precipitation titration. When AgNO₃ solution is added to the water sample, in presence of K₂CrO₄, the chlorides present in it are precipitated first as AgCl.

\[
\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} \downarrow + \text{NaNO}_3
\]

White ppt.

When all the Cl⁻ ions is removed, AgNO₃ added from the burette will react with K₂CrO₄ to give a reddish brown colour due to silver chromate (Ag₂CrO₄). This is the end point.

\[
2 \text{AgNO}_3 + \text{K}_2\text{CrO}_4 \rightarrow \text{Ag}_2\text{CrO}_4 \downarrow + 2\text{KNO}_3
\]

Yellow colour \quad reddish brown

PROCEDURE

Titration 1: Standardisation of AgNO₃

The burette is washed well with distilled water and rinsed with small amount of AgNO₃ solution. The pipette is washed with distilled water and rinsed with small amount of standard NaCl solution. 20ml of this solution is pipetted out into a clean conical flask. 1ml of 2% K₂CrO₄ indicator solution is added and titrated against AgNO₃ solution taken in the burette. The end point is the change of colour from yellow to reddish brown. The titration is repeated for concordant values.
Titration 2 : Estimation of Chloride

Water sample vs Standard AgNO₃

<table>
<thead>
<tr>
<th>Sl. No</th>
<th>Volume of water sample (ml)</th>
<th>Burette reading (ml)</th>
<th>Volume of AgNO₃ (ml)</th>
<th>Concordant value (ml)</th>
<th>Indicator</th>
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<tbody>
<tr>
<td></td>
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<td>Initial Final</td>
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<tr>
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<td></td>
<td>K₂CrO₄</td>
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<tr>
<td>2.</td>
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<td>0</td>
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</table>

Calculation of the normality of water sample (Chloride ion)

Volume of water sample \( (V₁) = 20 \text{ ml} \)
Strength of water sample \( (N₁) = ? \)
Volume of AgNO₃ \( (V₂) = \text{___} \text{ ml} \)
Strength of AgNO₃ \( (N₂) = \text{___} \text{ N} \)

According to volumetric formula
\[
V₁N₁ = V₂N₂
\]
\[
N₁ = \frac{V₂ \times N₂}{20}
\]

Strength of water sample \( (N₁) = \text{___} \text{ N} \)

Calculation of amount of chloride

Amount of chloride present in 1 litre of the given water sample = Eq wt of chloride ion \( \times \) Normality of chloride ion

Amount of chloride ion present in 100 ml = \( 35.46 \times \text{ Normality of chloride ion} \times \frac{100}{1000} \text{gm} \)
Titration 2: Estimation of Chloride ion

The given water sample is made up into 100 ml in a standard flask using distilled water. 20ml of this solution is pipetted out into a clean conical flask and 1ml of 2% $K_2CrO_4$ indicator solution is added. It is then titrated against standard $AgNO_3$ solution taken in the burette. The end point is the change of colour from yellow to reddish brown. The titration is repeated for concordant values.

RESULT

The amount of chloride ion present in 100 ml of the given water sample = ______ gms.
UV - Vis Spectrophotometer - Block diagram
4. DETERMINATION OF IRON IN WATER SAMPLE BY SPECTROPHOTOMETRY (USING THIO CYANATE)

AIM:

To estimate the amount of Fe$^{3+}$ ion or Fe$^{2+}$ ion present in the given aqueous solution by spectrophotometry. You are provided with a standard solution of ferrous ammonium sulphate containing 0.0838g in 1000ml.

MATERIAL REQUIRED:

Spectrophotometer, ferrous ammonium sulphate, nitric acid, potassium or ammonium thiocyanate, glass wares.

PRINCIPLE:

When a light passed through a solution, a portion of the light is absorbed. The portion of absorbed light depends on the concentration of the solution.

According to Lamberts Beer Law, if $I_o$ and $I_t$ are the intensities of the incident and transmitted light.

$$\log \frac{I_o}{I_t} = \varepsilon cl = A$$

$\varepsilon$, $l$ are constants. $A = \text{absorbance}$, $c = \text{concentration}$

Thus by finding out the change in intensity of colour of the light absorbed by a solution, its concentration can be calculated.

In order to get a visible colour, Fe$^{2+}$ ion is oxidized to Fe$^{3+}$ by nitric acid and treated with potassium thiocyanate to get deep red colour.

$$\text{Fe}^{3+} + 6\text{KSCN} \rightarrow \text{(Fe (SCN)$_6$)$_{3+}$ (red coloured thiocyanato complex of iron which absorbs in the visible region of 480 – 500 nm)}$$

PROCEDURE:

A standard solution may be prepared by dissolving 0.0838 g of pure ferrous ammonium sulphate in distilled water with 1 ml of con. HNO$_3$ and making upto 1000 ml (containing 1 ppm of iron / ml). 2,4,6 and 8 ppm iron solutions are prepared by pipetting out 2,4,6 and 8 ml solution in to a measuring cylinder,
Absorbance of Solution at 480 – 500 nm

<table>
<thead>
<tr>
<th>S.No.</th>
<th>Concentration of Solution (mg/l)</th>
<th>Absorbance(nm)</th>
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<tbody>
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<td>1</td>
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<td>2</td>
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<tr>
<td>4</td>
<td></td>
<td></td>
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<tr>
<td>5</td>
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Graph
1ml of 1:1 HNO₃ and 1 ml of 10% thiocyanate and remaining distilled water are added to make the total volume of 10 ml.

The spectrophotometer is adjusted for $\lambda = 480$nm. The transmittance ($I/I_o$) of the blank solution taken in the cell is adjusted for 100% absorbance. Solutions of various known concentrations are taken one after the other and the respective absorbances are measured. Lastly, the unknown solution is taken in the cell after carrying out the same treatment with nitric acid and thiocyanate.

Absorbance is measured. A graph is drawn connecting absorbance with concentration. A straight line is expected. From the graph and the absorbance of the unknown solution its concentration can be found out.

**RESULT:**

Amount of iron present in the given sample of water is = ------------------ppm.
## Titration I: Standardization of sodium thiosulphate

### Standard dichromate vs Sodium thiosulphate

<table>
<thead>
<tr>
<th>Sl.No.</th>
<th>Volume of Potassium dichromate (ml)</th>
<th>Burette readings (ml)</th>
<th>Volume of sodium thiosulphate (ml)</th>
<th>Concordant Value (ml)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Initial</td>
<td>Final</td>
<td></td>
</tr>
</tbody>
</table>

Volume of Potassium dichromate \((V_1)\) = 20 ml
Strength of Potassium dichromate \((N_1)\) = 0.01 N
Volume of Sodium thiosulphate \((V_2)\) = ---- ml
Strength of Sodium thiosulphate \((N_2)\) = ---- N

According to volumetric formula,  \(V_1N_1 = V_2N_2\)

\[N_2 = \frac{V_1N_1}{V_2}\]

Strength of Sodium thiosulphate \((N_2)\) = ---- N
5. ESTIMATION OF AVAILABLE CHLORINE IN BLEACHING POWDER

<table>
<thead>
<tr>
<th>Expt. No</th>
<th>Date</th>
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</table>

AIM:

To determine the Percentage of Available chlorine in the given sample of bleaching powder solution Iodometrically.

CHEMICALS REQUIRED:

N/10 Hypo (Na₂S₂O₃.6 H₂O), bleaching powder sample solution, Solid KI, dilute H₂SO₄.

PRINCIPLE:

The amount of chlorine liberated by the action of dilute acids on bleaching powder is termed as available chlorine and expressed as percentage weight of bleaching powder. The bleaching powder is a mixture of (CaOCl₂.4H₂O)(CaCl₂ Ca(OH).H₂O) and some free Ca(OH)₂. When dilute H₂SO₄ reacts with bleaching powder then free chlorine is liberated. The liberated chlorine reacts with potassium iodide to give free iodine. This liberated free iodine is then titrated against N/10 Hypo solution, using freshly prepared starch solution.

\[
\text{CaOCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + \text{H}_2\text{O} + \text{Cl}_2 \\
2\text{KI} + \text{Cl}_2 \rightarrow 2\text{KCl} + \text{I}_2 \\
2\text{Na}_2\text{S}_2\text{O}_3 + \text{I}_2 \rightarrow \text{Na}_2\text{S}_4\text{O}_6 + 2\text{NaI} \\
\text{Sodium Tetra Thionate} \\
\text{Starch} + \text{I}_2 \rightarrow \text{Starch . Iodide adsorption Complex} \\
\text{(Deep blue colour)}
\]

PROCEDURE:

Titration 1:

Standardization of sodium thiosulphate

The burette is washed and rinsed with sodium thiosulphate solution. Then the burette is filled with the given sodium thiosulphate solution. 20 ml of 0.0125N potassium dichromate solution is pipetted out into a clean conical flask. To this, 5 ml of sulphuric acid and 50 ml of 5% potassium
Titration 2: Estimation of available chlorine

Bleaching powder vs Sodium thiosulphate

<table>
<thead>
<tr>
<th>Sl.No.</th>
<th>Volume of Bleaching Powder sample taken (ml)</th>
<th>Burette readings (ml)</th>
<th>Volume of sodium thiosulphate (ml)</th>
<th>Concordant Value (ml)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Initial</td>
<td>Final</td>
<td></td>
</tr>
</tbody>
</table>

Volume of Sodium thiosulphate ($V_1$) = ml
Strength of Sodium thiosulphate ($N_1$) = $N$
Volume of Bleaching Powder solution ($V_2$) = 20 ml
Strength of Bleaching Powder solution ($N_2$) = ---- $N$

According to volumetric formula, $V_1N_1 = V_2N_2$

$$N_2 = \frac{V_1N_1}{V_2}$$

Strength of Bleaching Powder solution ($N_2$) = ---- $N$

Amount of Chlorine per Litre of solution = Normality x Equivalent weight

$$= N_2 \times 35.5 \text{ g/lit}$$

Percentage of Available Chlorine

$$= \frac{N_2 \times 35.5 \times 500 \times 100}{1000 \times W}$$

$$= \ldots \ldots \ldots \ldots \ldots \ldots \ldots \ldots \%$$
iodide are added. This is titrated against sodium thiosulphate solution. When the solution becomes straw yellow colour, starch indicator is added and then titration is continued. The end point is disappearance of blue colour and appearance of light green colour. The titration is repeated to get concordant values.

**Titration 2:**

**Estimation of available chlorine**

Fill the burette with Hypo solution. Pipette out 20 ml of bleaching power solution into 250 ml conical flask. Add 2 ml of KI and 10 ml of dilute H₂SO₄. Cover the mouth of conical flask with lid and keep it for 2 minutes. Solution becomes brown. The liberated Iodine against Hypo solution till a light yellow colour persists. Add 1 to 2 ml of starch solution. Solution turns to blue colour. Continue adding Hypo solution till blue colour disappears. This the end point. Repeat the experiment until the concordant readings are obtained.

RESULT:

The percentage of available chlorine in the given sample of bleaching powder is …………………….%