

Hardness by Complexometric titration method

Hard water

Depending on the area of the country, tap water contains various amount of ions such as Ca^{2+} , Mg^{2+} etc. The higher the concentration of these ions, the harder is the water. Hard water is due to metal ions (minerals) that are dissolved in the ground water. These minerals include Ca^{2+} , Mg^{2+} etc.

Our hard water is due to rain water which occurs through the vast amount of limestone rock that occurs in our area. In the oxygen, the CO_2 reacts with the limestone to form carbonic acid which is H_2CO_3 .

Hard water is not a health hazard. People regularly take calcium supplement drinking hard water contributes to small amount of Ca^{2+} and Mg^{2+} toward total human dietary needs of Ca^{2+} and Mg^{2+} . Increasing quantity of hard water could be a major contributor of Ca^{2+} and Mg^{2+} to the diet.

Hard water does cause soap scum. Clay pipes and clay boiled soap scum is formed after the water binds with the soap. This causes an insoluble soap that ppt to form the scum.

B.Tech tutorial Notes

Types —

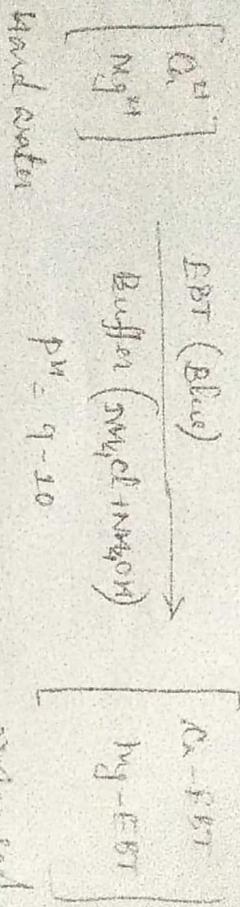
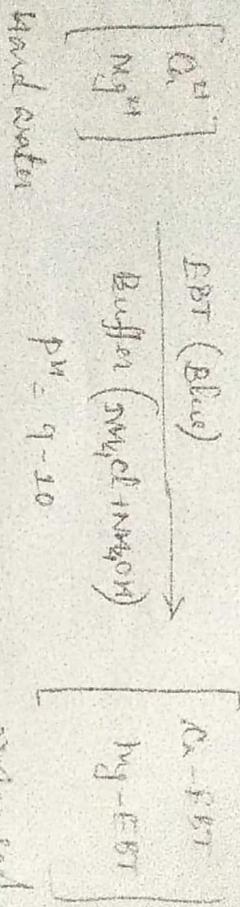
① Temporary

② Permanent

Temporary — due to Mg^{2+} ions being present in water. Can be removed by boiling the water to vaporise the CO_2 .



Permanent — Due to presence of ions Ca^{2+} , Mg^{2+} , Fe^{3+} , SO_4^{2-} . Cannot be eliminated by boiling.



	Soft	0.25% (mfp)
Med. Hard	75 to 150	
Hard	150 to 300	
Very Hard	300 and greater	

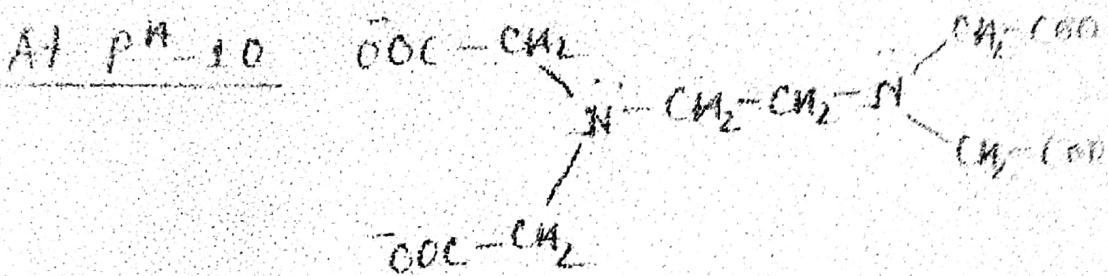
EDTA and EBT are weak acids and their actions are very pH -dependent.

Stains are due to a pH buffer. To hold the pH 's at an appropriate pH for both the EDTA and EBT indicator to work well.

The titration is carried out at a pH of 10, which keeps the EDTA in a half-neutralized form, where it complexes with Mg^{2+} ions very well but does not bind

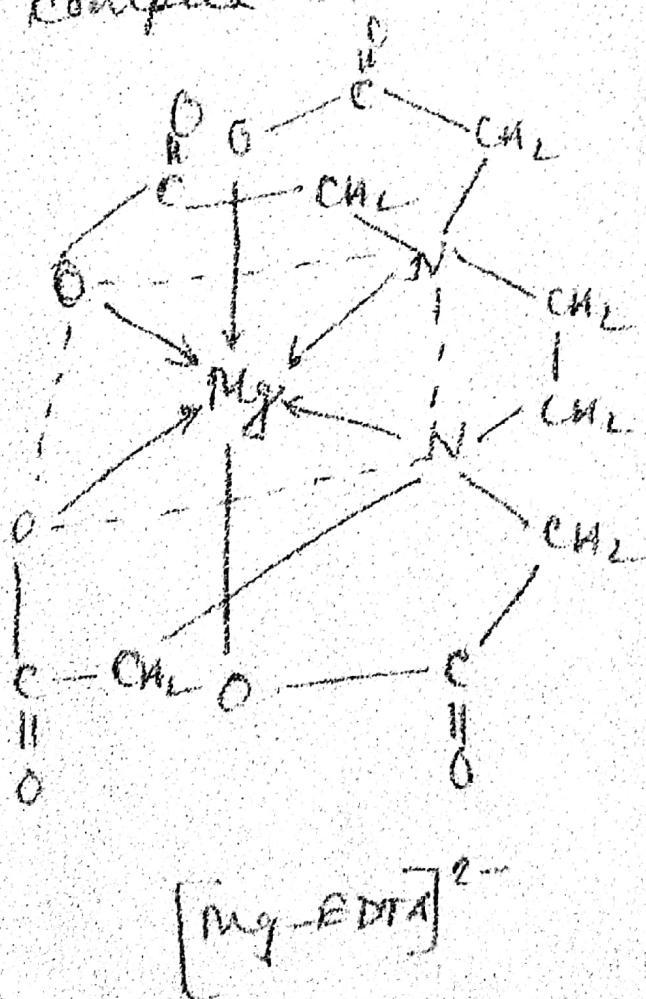
Waters

to react as readily with other cations such as Fe^{3+} that may be present as impurities in the water.



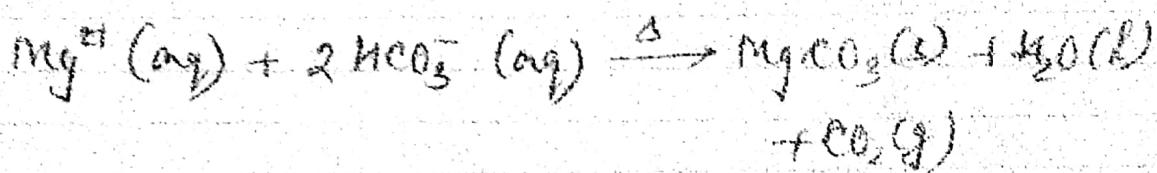
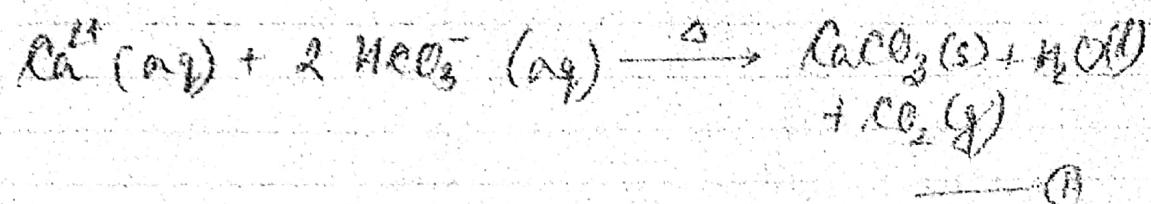
Hexadentate

1:1 complex, known as chelate (chel)



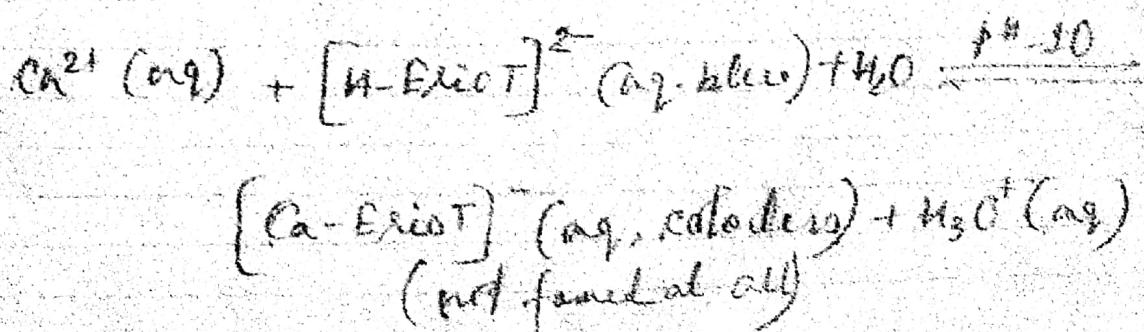
Total hardness of water refers to the total conc' of Ca^{2+} & Mg^{2+} ions in the water.

The temporary or carbonate hardness, refers to the amount of Ca and Mg ions that can be removed as insol. carbonates by boiling the water.



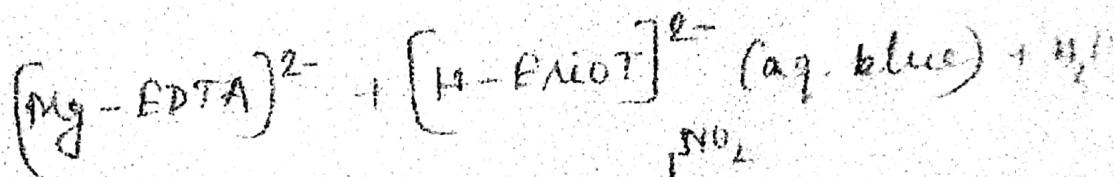
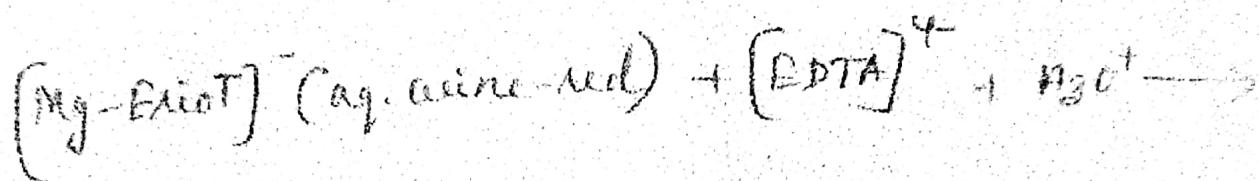
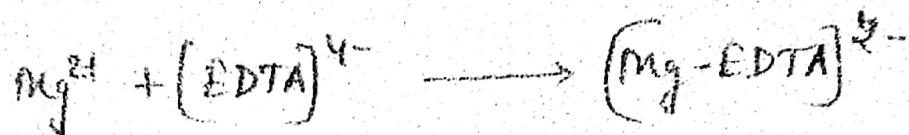
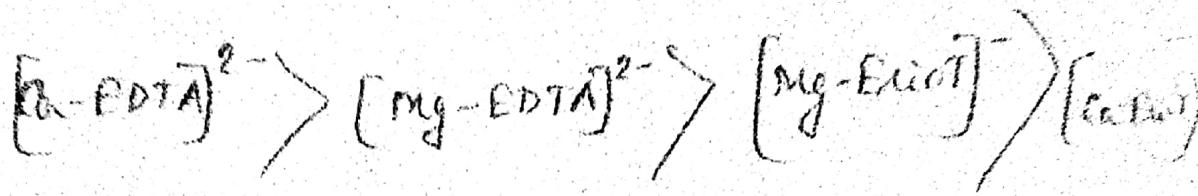
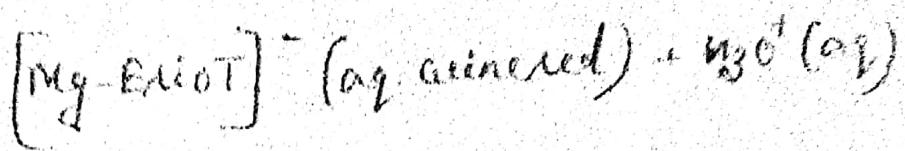
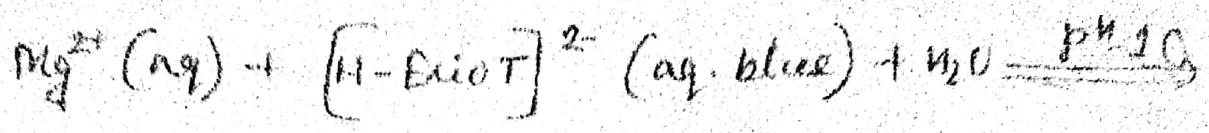
Whenever the conc' of HCO_3^- ion is less than half the total conc' of Ca^{2+} and Mg^{2+} ions, precipitation of these cations is incomplete. Permanent hardness or non-carbonate hardness refers to the amount of Ca^{2+} & Mg^{2+} ions remaining in sol' after removal of temporary hardness.

At pH 10, Eriot becomes $[\text{H}-\text{Eriot}]$

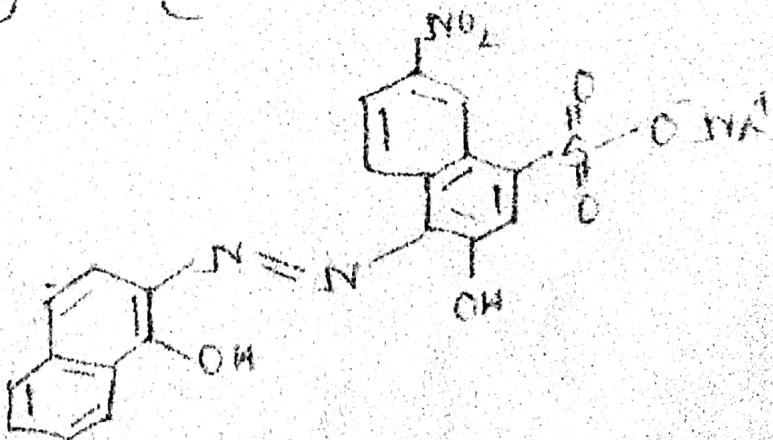


Waters

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Calculation

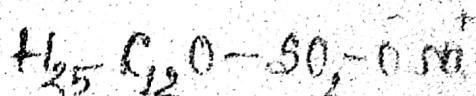


EBT

that it is less likely to stick to itself
and more likely to interact with oil and
grease



sol p-dodecyl benzene
sulfonate



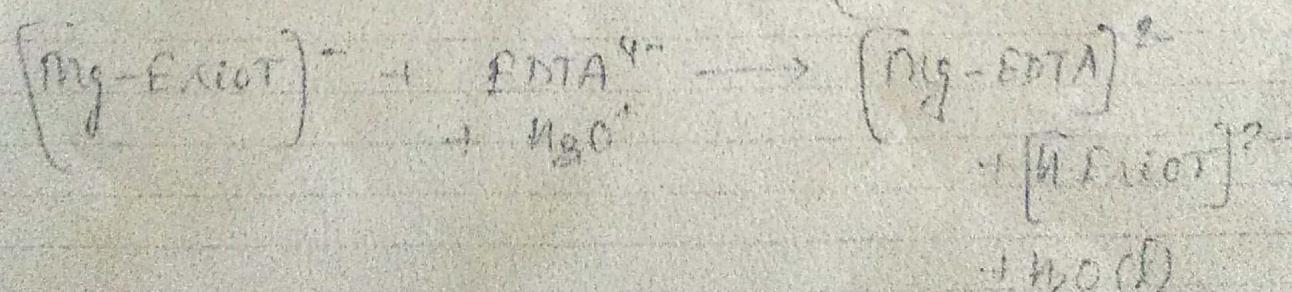
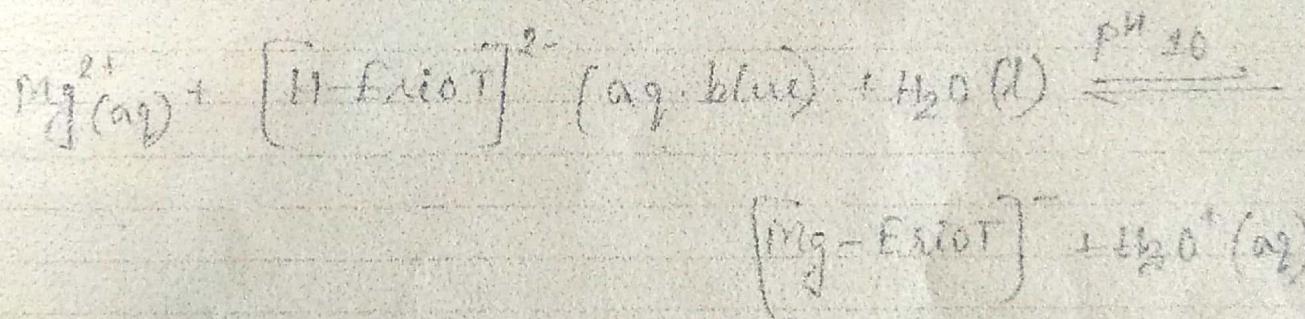
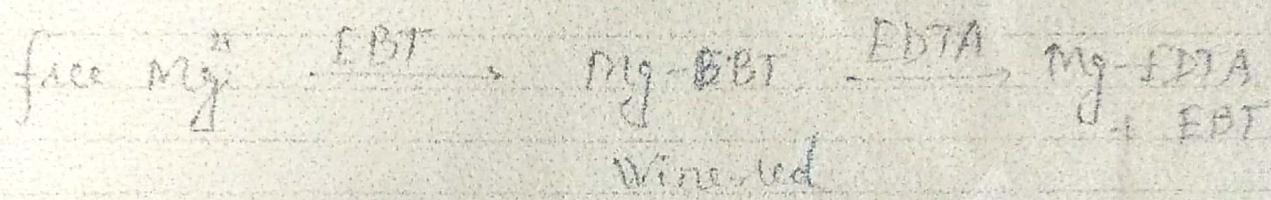
Sodium lauryl
sulfate

1

Expt. 2

Estimation of Magnesium using std. EDTA soln

25 ml Mg-solⁿ + 50 ml water + 2 ml buffer
25 ml + 2-3 drops EBT indicator → titrate with
standard EDTA → end point from wine-red
to blue.

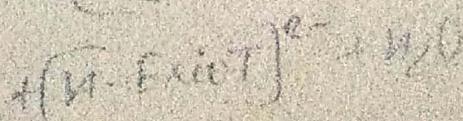
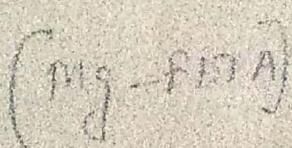
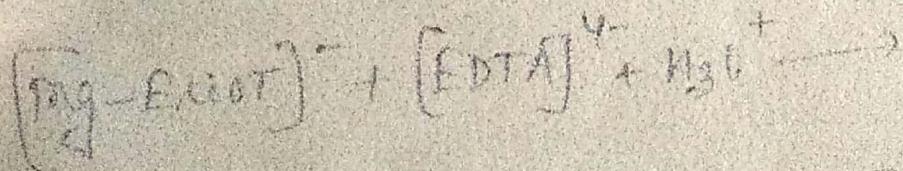
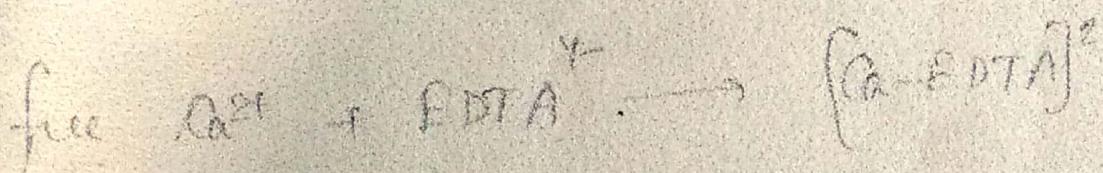
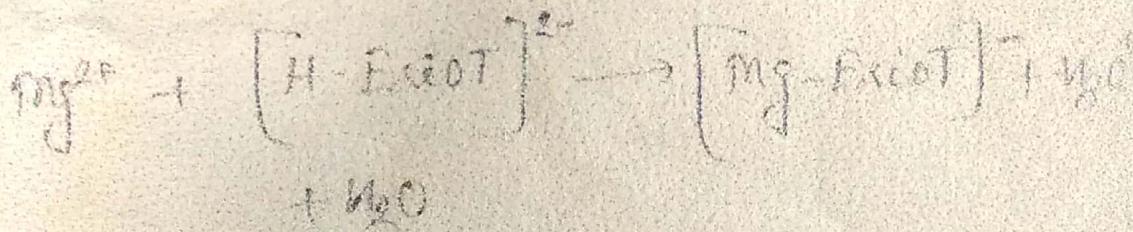
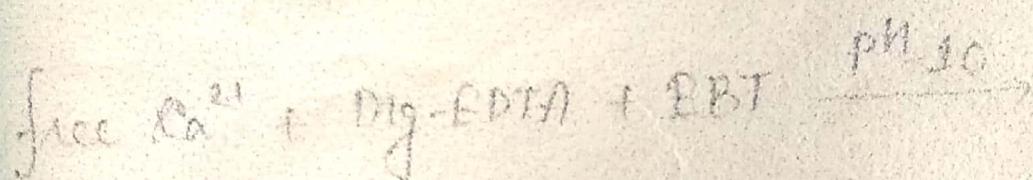


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Estimation of Ca^{2+} using EDTA

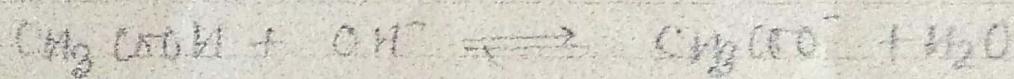
25 ml of Ca^{2+} solution + 25 ml of water
Add 2 ml buffer solution + 1 ml mg-EDTA
+ 3-4 drops of EBT indicator \rightarrow titrate
with standard EDTA soln from yellow
to clear blue.

Inside the reaction:

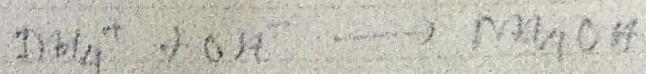


Buffer soln.

A buffer soln. is a mixture of aqueous soln consisting of a mixture of weak acid and its conjugate base. Its pH changes very little when a small or moderate amount of strong acid or base is added to it & thus



Basic buffer



Dissolved Oxygen (determination by Winkler method).

Dissolved oxygen is used as an indicator of the health of a water body, where higher DO conc's are correlated with high productivity and little pollution. The Winkler method is a technique used to measure DO in freshwater systems. The test is performed on-site, as delays between sample collection and testing may result in an alteration in O₂-content.

The Winkler method uses titration to determine DO in H₂O sample. DO analysis can

be used to determine -

The health or cleanliness of a lake or stream.

The amount and type of biomass at a site.

water system can support.

The amount of decomposition occurring in the lake or stream.

The total no. of milliliters of titrant used in steps 6-8 equals the total DO in the sample in mg/l. O₂ saturation is temp dependent - gas is more soluble in cold waters, hence cold waters generally have higher DO conc's. DO also depends on salinity and elevation, or partial pressures.

	(mg/l)
0°C	14.6
5	13.8
10	11.3
15	10.2
20	9.2
25	8.6
100	0

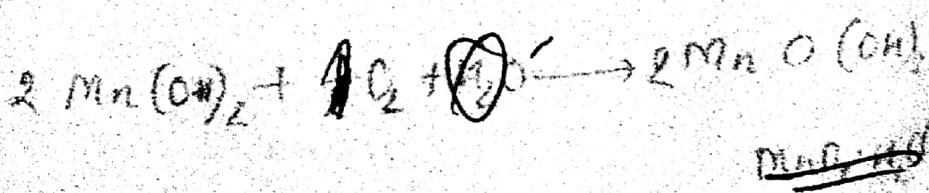
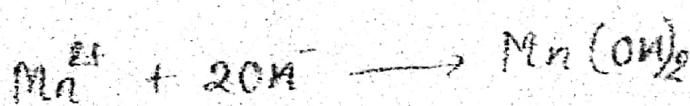
Oxygen in the water sample oxidizes I⁻ to I₂ quantitatively. Oxygen in the water sample oxidizes I⁻ to iodine quantitatively. The amount of iodine generated is then determined by titration with a standard Na₂S₂O₃ (S₂O₃²⁻) solution. The end point is determined by using starch as a visual indicator.

One mole of O₂ reacts with 4 moles of S₂O₃²⁻

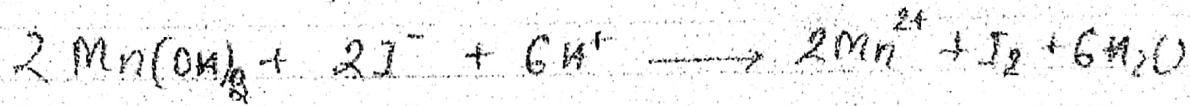
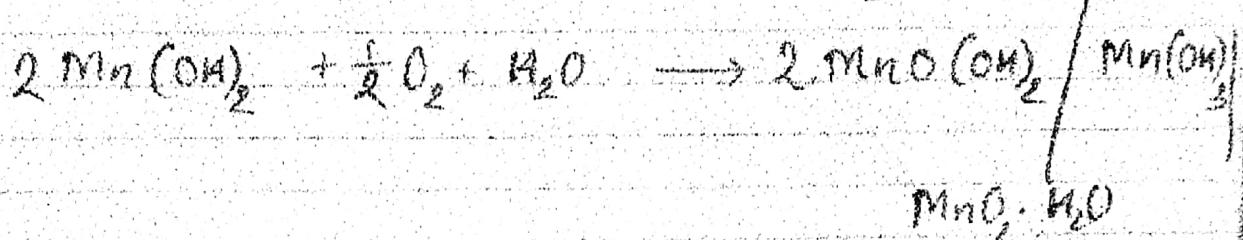
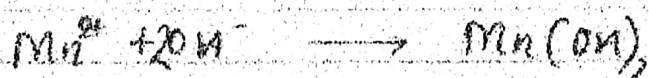
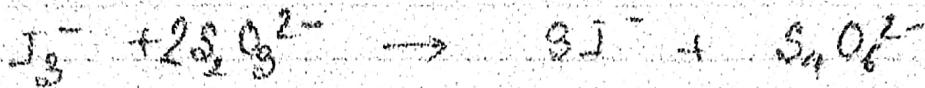
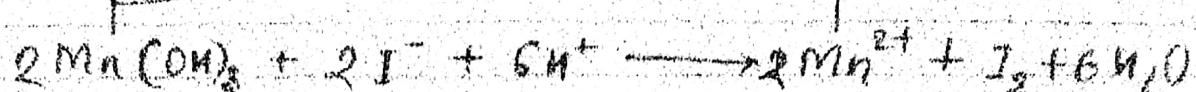
Waters

At the time of sampling, DO is fixed by addition of Mn(II) under basic conditions, resulting in a brown precipitate, Manganic hydroxide $\left[\text{MnO(OH)}_2\right]$. Prior to analysis, the sample is acidified to pH 1.0-2.5. This causes the precipitated hydroxides to dissolve, liberating Mn(III) ions. Mn(III) ions oxidize previously added I⁻ ions to iodine. Iodine forms a complex I_3^- with surplus I⁻ ions. I_2 and I_3^- exists in equil^r, thus I_3^- acts as a reservoir of I_2 .

The I_2 is then titrated with $S_2O_3^{2-}$. Iodine is reduced to I⁻ and thiosulfate is oxidized to tetrathionate.



red



Starch:

Amylose & amylopectin

It is responsible for the form² of a deep blue color in presence of I_2 . The iodine molecule slips inside the amylose coil and release it when $\text{S}_2\text{O}_3^{2-}$ is present near the end point addition.

But if starch is added at the very beginning, the iodine molecules are adsorbed on the surface of the polymer and forms permanent bond.

Waters

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so starch is added near the end point only.

C_2 - is more E.N and forms a weak van-der-Waals bond with water at normal temp. ie why it is soluble in water. The less E.N N_2 - molecules are not soluble in water.

The ext is -

98 ml H_2O + 1 ml MnSO_4 + 1 ml alk. KI

Cover/keep
dark

Brown ppt

} MnO_4^-

pale yellow $\xleftarrow{\text{Titrate with } \text{Na}_2\text{CO}_3}$ Disolve the ppt (brown)

} starch near the end point

blue to a tinge

What is the difference between iodometry and iodimetry?

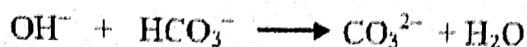
When an analyte that is a reducing agent is titrated directly with a standard Iodine solⁿ, the method is called "iodimetry".

When an analyte that is an oxidizing agent is added to excess iodide to produce iodine, and the iodine produced is determined by titration with Na₂S₂O₃, the method is called "iodometry".

AIM: To Determine the Alkalinity In a Given Water Sample

Theory: Alkalinity is a measure of the capability of water to absorb H⁺ ions without significant change of pH. In other words, alkalinity is a measure of the acid buffering capacity of water. The determination of alkalinity of water is necessary for controlling the corrosion, to calculate the amount of lime and soda needed for water softening; in conditioning the boiler feed water, etc.

Alkalinity of a sample of water is due to the presence of OH⁻ (hydroxide ion), HCO₃⁻ (bicarbonate ion) and CO₃²⁻ (carbonate ion) or the mixture of two ions present in water. The possibility of OH⁻ and HCO₃⁻ ions together is not possible since they combine together to form CO₃²⁻ ions.



The alkalinity due to different ions can be estimated separately by titration against standard acid solution, using selective indicators like phenolphthalein and methyl orange.

- i) $\text{OH}^- + \text{H}^+ \longrightarrow \text{H}_2\text{O}$
- ii) $\text{CO}_3^{2-} + \text{H}^+ \longrightarrow \text{HCO}_3^-$
- iii) $\text{HCO}_3^- + \text{H}^+ \longrightarrow \text{H}_2\text{O} + \text{CO}_2$

The neutralization reaction upto phenolphthalein end point shows the completion of reactions (i) and (ii) (OH⁻ and CO₃²⁻) and (CO₃²⁻ and HCO₃⁻) only. The amount of acid used thus corresponds to complete neutralization of OH⁻ plus half neutralization of CO₃²⁻. The titration of water sample using methyl orange indicator marks the completion of the reactions (i), (ii) and (iii). The amount of acid used after phenolphthalein end point corresponds to one half of normal carbonate and all the bicarbonates. Total amount of acid used represent the total alkalinity due to all ions present in water sample.

Apparatus: Burette, pipette, conical flask, beakers, burette stand and clamp

Chemicals: Dry Na₂CO₃, concentrated 12(N) HCl, phenolphthalein and methyl orange indicator

Procedure:

1. Primary standard solution of Na₂CO₃ (0.1N) is provided.
2. Secondary standard solution of HCl and the water sample are provided.

3. **Standardization of HCl solution by primary standard Na_2CO_3 solution**— Pipette out 10 mL of Na_2CO_3 solution in a conical flask, add 2 drops of methyl orange indicator, fill up the burette with (N/10) HCl solution and titrate till the color of the solution changes from yellow to red.
4. **Analysis of water sample**
- Pipette 20 mL of the sample of water into a 100 mL conical flask and 2 drops of phenolphthalein indicator was added and titrated against (N/10) HCl till the color of the solution changes from pink to colorless. Corresponding burette reading indicates the phenolphthalein end point (V_1).
 - Again pipette out 20 mL of the water sample in a conical flask, add 2 drops of methyl orange indicator. Color of the solution becomes yellow. Continue the titration against the (N/10) HCl solution till the color changes to red. This burette reading corresponds to the methyl orange end point (V_2).

Observations and Calculations

Table 1:- Standardization of HCl solution

Strength of Na_2CO_3 solution (S_1) = 0.1(N)

Entry	Volume of Na_2CO_3 solution (V_1) (mL)	Burette Reading (mL)		Volume of HCl required (mL)	Mean volume of HCl required (V_2) (mL)
		Initial	Final		
1.	10				
2.	10				
3.	10				

Let, the strength of HCl = S_2

As we know, $V_1 S_1 = V_2 S_2$

$$\therefore S_2 = (V_1 S_1 / V_2) \text{ (N)}$$

Table 2:- Analysis of water using phenolphthalein indicator

Strength of HCl solution (S_2) = _____ (N)

Entry	Volume of Water Sample Taken (mL)	Burette Reading (mL)		Volume of HCl required (mL)	Mean volume of HCl required (V_4) (mL)
		Initial	Final		
1.	20				

Estimation of Alkalinity in Water Sample

Expt. No.:

Date:

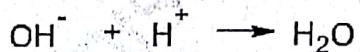
Aim:

To determine the phenolphthalein and methyl orange alkalinity in the given water samples A and B. A standard solution of N sulphuric acid is provided

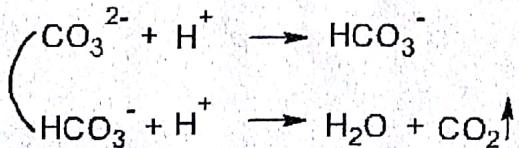
Principle:

Pure water is neutral in nature with pH 7. Due to the presence of dissolved minerals in rain water, the pH increases and becomes alkaline. Alkalinity in water is due to the presence of hydroxide (OH^-), carbonate (CO_3^{2-}) and bicarbonate (HCO_3^-) ions. The various碱碱ities are estimated by titrating the water sample against a standard acid using phenolphthalein and methyl orange indicators successively.

Hydroxide ions is completely neutralized to water using phenolphthalein indicator (single step neutralisation)



Carbonate ions is neutralized to bicarbonate ions using phenolphthalein indicator in the first step. In the second step, these bicarbonate ions is completely neutralized to water and carbon dioxide using methyl orange indicator (double step neutralisation).



Bicarbonate ions is completely neutralized to water and carbon dioxide using methyl orange indicator (single step neutralisation).



Procedure:

The burette is filled with standard sulphuric acid to the zero level, following usual precautions.

20 ml of the given water sample-A is pipetted out into a conical flask. Two drops of phenolphthalein indicator is added and titrated against the standard sulphuric acid. The end point is the disappearance of pink colour. The titre value is noted. A drop of methyl orange indicator is added to the same solution after the phenolphthalein end point and the titration is continued until the solution becomes orange. The total titre value is noted. The titration is repeated to get concordant value and the titre values are noted in the Table 1.

The above procedure is repeated for the water sample-B and the titre values are noted in the Table 2.

Note

$P = M$	OH- only
$P = 0$	HCO_3^-
$P = \frac{1}{2}M$	CO_3^{2-}
$P > \frac{1}{2}M$	both OH^- and CO_3^{2-} are present
$P < \frac{1}{2}M$	both CO_3^{2-} and HCO_3^- are present

Result:

1. The given water sample-A contains

Phenolphthalein alkalinity = ppm

Methyl orange alkalinity = ppm

Since $P \dots \frac{1}{2}M$, and ions are present in the given water sample-A.

2. The given water sample-B contains

Phenolphthalein alkalinity = ppm

Methyl orange alkalinity = ppm

Since $P \dots \frac{1}{2}M$, and ions are present in the given water sample-B.

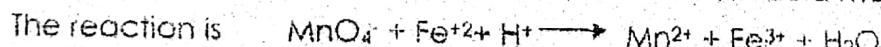
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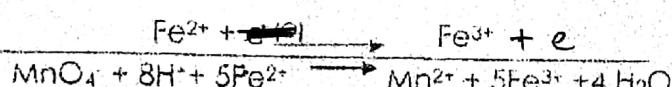
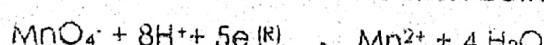
TO DETERMINE THE AMOUNT OF IRON (II) PRESENT IN THE SUPPLIED SOLUTION USING STANDARD KMnO₄ SOLUTION

EXPERIMENT NO. 2

THEORY : FeSO₄ {Fe-(II)} is a reducing agent. To know its strength, it is to be titrated with an Oxidising agent with known strength. In this titration KMnO₄, which is an Oxidising agent, is used to react with FeSO₄ in acid medium.



This is a redox reaction where oxidation and reduction both are taking place.



PROCEDURE

Pipette out 25ml Iron (II) solution in a 500ml conical flask. Add 20ml (one test tube full) Sulphuric Acid and Phosphoric Acid Mixture. Titrate the resulting solution with std. KMO₄ solution until a faint pink colour just appears which is the end point of the titration.

Repeat the titration to get the concordant volume.

TABLE

No of observation	Volm. Of Iron (II) soln. taken (ml)	Burette reading		Volm. Of KMNO ₄ soln. used(ml)	Concordant volume (ml)
		Initial	Final		
					Y

CALCULATION

Given that, Strength of KMnO₄ soln. = 0.1N

A. CALCULATION FOR FeSO₄:

1000 ml 1N KMnO₄ soln. = 151.85 gms of FeSO₄

$$151.85 \times 0.1 \times Y \text{ gm. Of FeSO}_4 \\ Y \text{ ml } 0.1 \text{ N KMnO}_4 \text{ Soln.} = \frac{151.85 \times 0.1 \times Y}{1000}$$

$$\text{Amount of FeSO}_4 \text{ present in gm/L} = \frac{151.85 \times 0.1 \times Y \times 1000}{1000 \times 25}$$

B. CALCULATION FOR IRON:

1ml 0.1N KMNO₄ soln. = 0.005585 gm of Iron

$$0.005585 \times Y \times 0.1 \times 1000 \text{ gm}$$

$$\text{Amount of Iron (II) present in gm/Ltr} = \frac{0.005585 \times Y \times 0.1 \times 1000}{25}$$

Assigned Reading: R.B. Fischer and D.G. Peters, Quantitative Chemical Analysis, Saunders, Philadelphia, 1968, pp.12-18, 51-65, 536-546; this reference is on reserve in the Chemistry Library.

Introduction

Iron is one of the most frequently encountered elements in industrial analyses. Therefore, its accurate determination is of great practical importance. A volumetric determination of iron consists of these steps:

1. dissolution of the sample
2. quantitative conversion of iron (III) to iron (II) with a suitable reducing agent
3. removal of the excess reducing agent
4. addition of special reagents to aid in detection of the endpoint and to ensure that the proper reaction occurs during the subsequent titration
5. titration of iron (II) to iron (III) with a suitable oxidizing agent

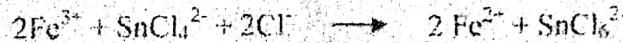
Principles

Dissolution of the Sample

Major components of an iron sample are often iron (II) and iron (III) compounds that are insoluble in water. These can be dissolved in hot concentrated hydrochloric acid. Silica, which is present in most natural iron samples, does not dissolve in hot concentrated hydrochloric acid and may be observed as a white solid material floating on the surface of the solution; silica will not interfere with the iron determination.

Reduction of Iron (III) to Iron (II)

After the sample is dissolved, the iron will exist mainly as iron (III). Titration with a standard oxidizing agent requires that all the iron exist as iron (II). A quantitative reduction of iron (III) to iron (II) can be performed by addition of tin (II) to the hot hydrochloric acid solution of the sample:



8-1

Sandy Anne O'Dowd

Paula Barnes O'Dowd

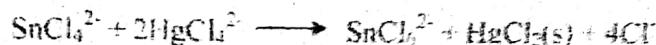
Tin (II) possesses these advantages as a reductant in this system:

1. It reduces iron (III) to iron (II) quantitatively and rapidly
2. It possesses a certain degree of selectivity for iron (III), because it is not so strong a reducing agent that it reduces other species in solution
3. An excess is easily destroyed with $HgCl_2$ so that tin (II) will not interfere with the redox titration

Due to the presence of chloride complexes of iron (III), the sample solution is yellow, whereas iron (II) species are colorless. Therefore, tin (II) solution should be added dropwise to the hot sample solution until the yellow color just disappears. One drop excess should be added to ensure complete reduction of iron (III).

Removal of Excess Reducing Agent

Excess tin (II) must be destroyed as it will interfere with the redox titration. This is accomplished by addition of a $HgCl_2$ solution:



If the amount of tin (II) added is in great excess, the following reaction will also occur:



Elemental mercury is produced in a colloidal form and will react with the standard oxidizing agent, thereby ruining the determination. Mercurous chloride (Hg_2Cl_2) is a white precipitate that will not interfere with the titration. If the precipitate is gray or black, elemental mercury is present, and the sample must be discarded.

Addition of Special Reagents; Titration

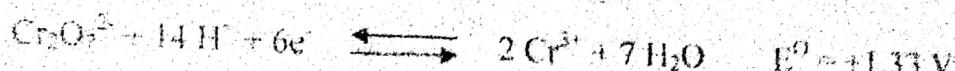
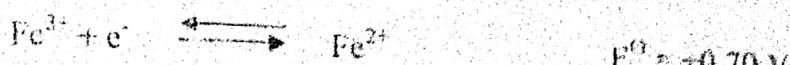
Special reagents are sometimes added before a redox titration to:

1. Act as an indicator
2. Change a half-reaction potential to enhance the sharpness of an endpoint
3. Suppress the formation of certain colored complexes in solution

In an acid medium, the reaction between iron (II) and dichromate ion (the oxidizing agent) is:



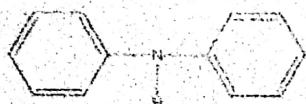
Standard potentials for the half-reactions are:



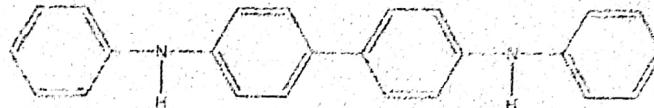
However, in a 1 M hydrochloric acid medium, the formal potential (E^\ominus) for the dichromate-chromic ion couple is +1.00 V. Note that, under the latter conditions, ΔE is only +0.30V. To enhance ΔE , a mixture of phosphoric acid and sulfuric acid is added to the sample prior to the titration. This lowers the potential of the iron (III)-iron(II) couple to +0.61 V. Phosphoric acid forms complexes with iron (III), lowering the concentration of simple Fe^{3+} ions and decreasing the potential of the iron (III)-iron (II) couple. Because the new ΔE is +0.39V, the endpoint coincides more closely with the theoretical equivalence point.

A redox indicator must also be added at this stage. A redox indicator is a compound that changes color over a potential range of interest due to a reduction-oxidation process. Since dichromate cannot serve as its own indicator, the dichromate-chromic ion system requires a redox indicator. Chromic ion is green and the dichromate ion is orange, so the endpoint is masked.

An important family of redox indicators is derived from diphenylamine, a colorless compound which is oxidized by dichromate to diphenylbenzidine



Diphenylamine (colorless)



Diphenylbenzidine (colorless)

And then to diphenylbenzidine violet:



Diphenylbenzidine violet (violet)

Diphenylbenzidine violet has a deep purple color that is easily seen at the endpoint, even in the presence of green chromium (III) ions. Note that the diphenylbenzidine-

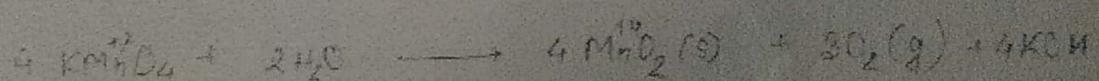
What is KMnO₄?

KMnO₄ (m.w=158.0) is a dark purple or brownish black powder or dark purple or almost black crystals. It is soluble in water and a strong O.A. It decomposes on contact with certain organic substances due to its oxidizing abilities it has disinfectant and deodorizing properties. It is also antiseptic. It is widely used as a standard (volumetric) oxidizing agent, because of its intense colour which serves as an indicator in titrations besides its low cost.

Standardization of KMnO₄ against sodium oxalate (as a primary standard) follows oxid-reduc' reac' in which KMnO₄ is the O.A. whereas Na₂C₂O₄ is the reducing agent



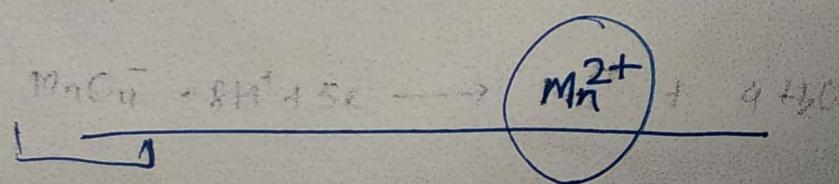
Dq soln of KMnO₄ are not entirely stable because of H₂O oxidation



However, KMnO₄ solns, when properly prepared are reasonably stable because the decomposition reac' is slow. It is catalyzed by light, acids, heat, bases, Mn(II), and MnO₂.

MnO_2 is a contaminant in even the best grade of solid MnO₂. Furthermore, this compound forms in freshly prepared sol's of the reagent as a consequence of the reaction of Mn²⁺ ion with organic matter present in the water used to prepare the sol's.

Sufficient H_2SO_4 is added during standardization because KMnO_4 oxid' ability is better in acidic media and to keep H⁺ ion conc' constant throughout the standardization process.



Temp effect

Keeping the Temp' near to 70°C throughout the standardization process is important because the oxid' of Mn²⁺ is rapid enough at such temp'. Formation of a brown colour during the titration is caused by insufficient acid, applying too high temp or.

Standardized KMnO_4 sol's should be stored in dark.

Acid Effect

1. The Mn^{2+} must be strongly oxidized in order to avoid other side reactions. Hence, H_2SO_4 is normally used. HNO_3 is unsuitable since NO_3^- ion is a moderately strong oxidizing agent, which may interfere with KMnO_4 . If HCl is used, it may oxidize the Cl_2 by permanganate.

If Cl^- ion is present, high results are obtained because some KMnO_4 is used up in oxidizing the HCl to eliminate the interference of chloride ions.



2. The main reason for using acid in the titration is that it transforms oxalate to oxalic acid and this acid so far reduces the KMnO_4 . Any side reaction may hamper the result. So only KMnO_4 H_2SO_4 is used.

Some Mn^{2+} are produced from MnO_2 which reacts with KMnO_4 and forms more MnO_2 . This is known as auto decomposition.



DETERMINE THE MOLARITY AND STRENGTH OF THE GIVEN KMnO₄ SOLUTION.

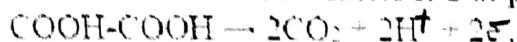
No. of moles of solute
wt. in solution.

AIM - (a) To prepare 100ml of M/20 solution of oxalic acid.

(b) Using this calculate the molarity and strength of the given KMnO₄ solution.

APPARATUS AND CHEMICALS REQUIRED- Oxalic acid, weighing bottle, weight box, volumetric flask, funnel, distilled water, chemical balance, beakers, conical flask, funnel, burette, pipette, clamp stand, tile, dilute H₂SO₄, KMnO₄ solution.

THEORY- (a) Oxalic acid is a dicarboxylic acid having molar mass 126g mol⁻¹. It is a primary standard and has the molecular formula COOH-COOH.2H₂O. Its equivalent mass is 126/2 = 63 as its *n* factor is 2 as per the following reaction:



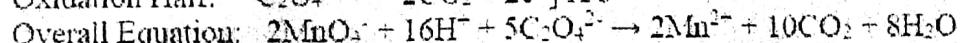
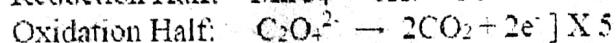
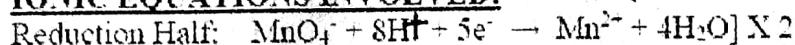
Calculation of amount of oxalic acid to be weighed to prepare 100ml M/20 solution:

$$M = \frac{\text{wt. } X}{\text{Mol. Wt}} \times \frac{1000}{V(\text{ml})}$$

(b) **THEORY-**

1. The reaction between KMnO₄ and oxalic acid is a redox reaction and the titration is therefore called a redox titration.
2. Oxalic acid is the reducing agent and KMnO₄ is the oxidizing agent.
3. KMnO₄ acts as an oxidizing agent in all the mediums; i.e. acidic, basic and neutral medium.
4. KMnO₄ acts as the strongest oxidizing agent in the acidic medium and therefore dil. H₂SO₄ is added to the conical flask before starting the titration.
5. The titration between oxalic acid and KMnO₄ is a slow reaction, therefore heat the oxalic acid solution to about 60°C to increase the rate of the reaction.
- 6.

IONIC EQUATIONS INVOLVED:



INDICATOR- KMnO₄ acts as a self indicator.

END POINT- Colourless to light pink (KMnO₄ in the burette)

PROCEDURE-

1. Fill the burette with KMnO₄ solution.
2. Pipette out 20ml. of oxalic acid solution into the conical flask.
3. Add half a test tube of dil. H₂SO₄ and heat the solution to about 60°C to increase the rate of the reaction.
4. Note down the initial reading of the burette.

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2
O Mn O
O 2 = O

5. Run down the KMnO_4 solution into the conical flask drop wise with shaking.
6. Stop the titration when a permanent pink colour is obtained in the solution.
7. This is the end point. Note down the final burette reading.
8. Repeat the experiment until three concordant values are obtained.

OBSERVATION TABLE:

Volume of Oxalic Acid solution taken = ml

S.I.No	BURETTE READINGS		VOLUME OF KMnO_4 USED (ml)	CONCORDANT VALUE (ml)
	INITIAL(ml)	FINAL(ml)		
1				
2				
3				
4				
5				

CALCULATIONS:

Using formula:

$$N_1 M_1 V_1 = N_2 M_2 V_2$$

Where $N_1 = 5$ (for KMnO_4), $V_1 = \underline{\hspace{2cm}}$, $M_1 = ?$

$N_2 = 2$ (for oxalic acid), $V_2 = 10\text{ml}$, $M_2 =$

$$M_1 = \frac{2 \times M_2 \times V_2}{5 \times V_1}$$

Strength = $M \times \text{Molar Mass of } \text{KMnO}_4$.

$$= \underline{\hspace{2cm}} \times 158$$

$$= \underline{\hspace{2cm}} \text{ g/L}$$

RESULT- The Molarity of KMnO_4 = M

And the strength of KMnO_4 = g/L