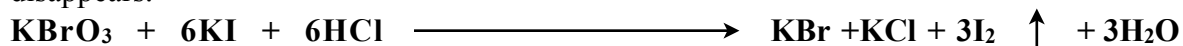


Preparation and Standardisation of 0.1M Sodium thiosulphate :

Sodium thiosulphate - M.Wt.- 248

Preparation of 0.1M Sodium thiosulphate Solution : Dissolve an accurately weighed amount of 25 gm of Sodium thiosulphate and 0.2 gm of sodium carbonate in sufficient amount of water to produce 1000 ml.

Principle : It is standardised by redox titration involving **Iodometric methods** using potassium bromate as oxidising agent. An accurately weighed amount of the primary standard potassium bromate is dissolved in water and the solution is made upto 250 ml with water. To this 50 ml is taken accurately and 20 gm of KI & 30 ml of 2M HCl are added and the liberated Iodine is titrated with sodium thiosulphate solution using starch solution as indicator, until the blue colour disappears.



Procedure :

Procedure for Standardisation of 0.1M Sodium thiosulphate Solution : Dissolve 0.2 gm of KBrO_3 weighed accurately, in sufficient water to produce 250 ml. To 50 ml of this solution, add 2 gm of KI and 3 ml of 2M HCl and titrate with Sodium thiosulphate solution using starch solution as indicator, until the blue colour disappears.

Equivalent Factor : Each ml of 0.1M Sodium thiosulphate \equiv 0.02784 gm of KBrO_3

Calculation :

$$\text{Molarity of Sodium thiosulphate} = \frac{\text{Weight taken}^*(\text{pot. bromate}) \times \text{Req. molarity}}{\text{Titration Volume} \times \text{Equivalent factor}}$$

Note : Weight taken*

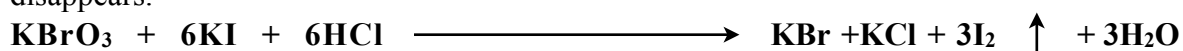
Calculate the amount of potassium bromate that will be present in the volume pipetted out. (i.e., in 50 ml) from the actual amount of weight of sample taken.

Preparation and Standardisation of 0.1N Sodium thiosulphate :

Sodium thiosulphate - M.Wt.- 248

Preparation of 0.1N Sodium thiosulphate Solution : Dissolve an accurately weighed amount of 25 gm of Sodium thiosulphate and 0.2 gm of sodium carbonate in sufficient amount of water to produce 1000 ml.

Principle : It is standardised by redox titration involving **Iodometric methods** using potassium bromate as oxidising agent. An accurately weighed amount of the primary standard potassium bromate is dissolved in water and the solution is made upto 250 ml with water. To this 50 ml is taken accurately and 2 gm of KI & 3 ml of 2M HCl are added and the liberated Iodine is titrated with sodium thiosulphate solution using starch solution as indicator, until the blue colour disappears.



Procedure :

Procedure for Standardisation of 0.1N Sodium thiosulphate Solution : Dissolve 0.2 gm of KBrO_3 weighed accurately, in sufficient water to produce 250 ml. To 50 ml of this solution, add 2 gm of KI and 3 ml of 2M HCl and titrate with Sodium thiosulphate solution using starch solution as indicator, until the blue colour disappears.

Equivalent Factor : Each ml of 0.1N Sodium thiosulphate \equiv 0.02784 gm of KBrO_3

Calculation :

$$\text{Normality of Sodium thiosulphate} = \frac{\text{Weight taken} * (\text{pot. bromate}) \times \text{Req. normality}}{\text{Titration Volume} \times \text{Equivalent factor}}$$

Note : Weight taken*

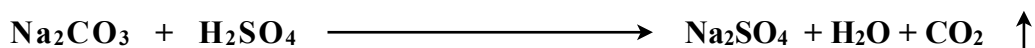
Calculate the amount of potassium bromate that will be present in the volume pipetted out. (i.e., in 50 ml) from the actual amount of weight of sample taken.

Preparation and Standardisation of 1M Sulphuric Acid :

Sulphuric Acid - M.Wt - 98.07

Preparation of 1M Sulphuric Acid Solution : Dissolve an accurately measured volume of 54 ml of Sulphuric Acid in sufficient amount of water to produce 1000 ml.

Principle : It is an example of alkalimetry. When a strong acid is titrated with a strong base, the salt produced in the reaction is not hydrolysed and therefore the pH of the resultant solution at the end-point is not 7.0 . H₂SO₄, a strong acid, is standardised by titrating with a strong base, i.e., Na₂CO₃ (primary standard). The following reaction takes place when Na₂CO₃ is titrated with H₂SO₄.



In this titration, for the detection of the end-point, methyl orange is used as an indicator.

Procedure :

Preparation of 1M Na₂CO₃ Solution : Dissolve an accurately weighed amount of 286.15 gm of Na₂CO₃ in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 1M H₂SO₄: Pipette out exactly 10 ml of 1M Na₂CO₃ solution into a clean conical flask and add 2 or 3 drops of methyl orange indicator. Titrate the contents of the flask with H₂SO₄ until red colour is obtained. Repeat the titration for concordant values. Record the values in the tabular form...

Titration of 1M Na₂CO₃ solution with H₂SO₄ solution

S.No.	Volume of 1M Na ₂ CO ₃ solution (ml)	Burette Reading		Vol. of H ₂ SO ₄ rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Molarity of HCl is calculated using the formula :

$$M_1V_1 = M_2V_2$$

where , V₁ = Volume of 1M Na₂CO₃ solution = 10 ml

M₁ = Molarity of Na₂CO₃ solution = 1M

V₂ = Volume of H₂SO₄ solution rundown (Average Burette Reading)

M₂ = Molarity of H₂SO₄ = ?

Therefore,

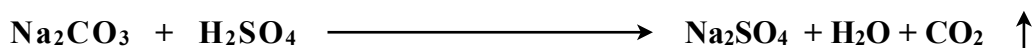
$$M_2 = \frac{M_1V_1}{V_2}$$

Preparation and Standardisation of 0.1M Sulphuric Acid :

Sulphuric Acid - M.Wt - 98.07

Preparation of 0.1M Sulphuric Acid Solution : Dissolve an accurately measured volume of 5.4 ml of Sulphuric Acid in sufficient amount of water to produce 1000 ml.

Principle : It is an example of alkalimetry. When a strong acid is titrated with a strong base, the salt produced in the reaction is not hydrolysed and therefore the pH of the resultant solution at the end-point is not 7.0 . H₂SO₄, a strong acid, is standardised by titrating with a strong base, i.e., Na₂CO₃ (primary standard). The following reaction takes place when Na₂CO₃ is titrated with H₂SO₄.



In this titration, for the detection of the end-point, methyl orange is used as an indicator.

Procedure :

Preparation of 0.1M Na₂CO₃ Solution : Dissolve an accurately weighed amount of 28.62 gm of Na₂CO₃ in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 0.1M H₂SO₄: Pipette out exactly 10 ml of 0.1M Na₂CO₃ solution into a clean conical flask and add 2 or 3 drops of methyl orange indicator. Titrate the contents of the flask with H₂SO₄ until red colour is obtained. Repeat the titration for concordant values. Record the values in the tabular form...

Titration of 0.1M Na₂CO₃ solution with H₂SO₄ solution

S.No.	Volume of 0.1M Na ₂ CO ₃ solution (ml)	Burette Reading		Vol. of H ₂ SO ₄ rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Molarity of HCl is calculated using the formula :

$$M_1V_1 = M_2V_2$$

where , V₁ = Volume of 0.1M Na₂CO₃ solution = 10 ml

M₁ = Molarity of Na₂CO₃ solution = 0.1M

V₂ = Volume of H₂SO₄ solution rundown (Average Burette Reading)

M₂ = Molarity of H₂SO₄ =?

Therefore,

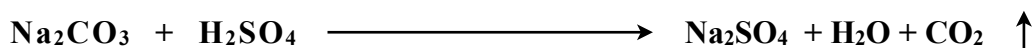
$$M_2 = \frac{M_1V_1}{V_2}$$

Preparation and Standardisation of 1N Sulphuric Acid :

Sulphuric Acid - M.Wt - 98.07

Preparation of 1N Sulphuric Acid Solution : Dissolve an accurately measured volume of 26.6 ml of Sulphuric Acid in sufficient amount of water to produce 1000 ml.

Principle : It is an example of alkalimetry. When a strong acid is titrated with a strong base, the salt produced in the reaction is not hydrolysed and therefore the pH of the resultant solution at the end-point is not 7.0. H_2SO_4 , a strong acid, is standardised by titrating with a strong base, i.e., Na_2CO_3 (primary standard). The following reaction takes place when Na_2CO_3 is titrated with H_2SO_4 .



In this titration, for the detection of the end-point, methyl orange is used as an indicator.

Procedure :

Preparation of 1N Na_2CO_3 Solution : Dissolve an accurately weighed amount of 53 gm of Na_2CO_3 in sufficient amount of water to produce 1000 ml.

Procedure for Standardisation of 1N H_2SO_4 : Pipette out exactly 10 ml of 1N Na_2CO_3 solution into a clean conical flask and add 2 or 3 drops of methyl orange indicator. Titrate the contents of the flask with H_2SO_4 until red colour is obtained. Repeat the titration for concordant values. Record the values in the tabular form...

Titration of 1N Na_2CO_3 solution with H_2SO_4 solution

S.No.	Volume of 1N Na_2CO_3 solution (ml)	Burette Reading		Vol. of H_2SO_4 rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Normality of HCl is calculated using the formula :

$$N_1V_1 = N_2V_2$$

where, V_1 = Volume of 1N Na_2CO_3 solution = 10 ml

N_1 = Normality of Na_2CO_3 solution = 1N

V_2 = Volume of H_2SO_4 solution rundown (Average Burette Reading)

N_2 = Normality of H_2SO_4 = ?

Therefore,

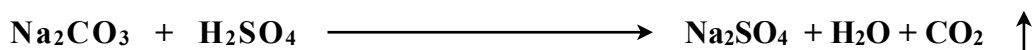
$$N_2 = \frac{N_1V_1}{V_2}$$

Preparation and Standardisation of 0.1N Sulphuric Acid :

Sulphuric Acid - M.Wt - 98.07

Preparation of 0.1N Sulphuric Acid Solution : Dissolve an accurately measured volume of 2.7 ml of Sulphuric Acid in sufficient amount of water to produce 1000 ml.

Principle : It is an example of alkalimetry. When a strong acid is titrated with a strong base, the salt produced in the reaction is not hydrolysed and therefore the pH of the resultant solution at the end-point is not 7.0 . H₂SO₄, a strong acid, is standardised by titrating with a strong base, i.e., Na₂CO₃ (primary standard). The following reaction takes place when Na₂CO₃ is titrated with H₂SO₄.



In this titration, for the detection of the end-point, methyl orange is used as an indicator.

Procedure :

Preparation of 0.1N Na₂CO₃ Solution : Dissolve an accurately weighed amount of 5.3 gm of Na₂CO₃ in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 0.1N H₂SO₄: Pipette out exactly 10 ml of 0.1N Na₂CO₃ solution into a clean conical flask and add 2 or 3 drops of methyl orange indicator. Titrate the contents of the flask with H₂SO₄ until red colour is obtained. Repeat the titration for concordant values. Record the values in the tabular form...

Titration of 0.1N Na₂CO₃ solution with H₂SO₄ solution

S.No.	Volume of 0.1N Na ₂ CO ₃ solution (ml)	Burette Reading		Vol. of H ₂ SO ₄ rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Normality of HCl is calculated using the formula :

$$N_1V_1=N_2V_2$$

where , V₁ = Volume of 0.1N Na₂CO₃ solution = 10 ml

N₁ = Normality of Na₂CO₃ solution = 0.1N

V₂ = Volume of H₂SO₄ solution rundown (Average Burette Reading)

N₂ = Normality of H₂SO₄ =?

Therefore,

$$N_2 = \frac{N_1V_1}{V_2}$$

Preparation and Standardisation of 1M Potassium permanganate :

Potassium permanganate - M.Wt - 158

Preparation of 1M Potassium permanganate Solution : Dissolve an accurately weighed amount of 158 gm of KMnO_4 in 900 ml of water heat on water bath for 1 hour and then cool it, then filter it using a sintered-glass filter and add sufficient amount of water to produce 1000 ml.

Principle : In volumetric analysis many reactions involve the process of oxidation and reduction. An oxidising agent is estimated by titrating with a reducing agent and vice-versa. These titration are called redox titrations. The standardisation of KMnO_4 is an example of redox titration. KMnO_4 is a powerful oxidising agent and in acidic medium, it oxidises oxalic acid to CO_2 .



The temperature is maintained at $60^\circ\text{-}70^\circ\text{C}$ during the titration because, the reaction proceeds slowly at room temperature. No indicator is required as KMnO_4 , is a self-indicator.

Procedure :

Preparation of 1M Oxalic acid Solution : Dissolve an accurately weighed amount of 126.07 gm of Oxalic acid in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 1M KMnO_4 : Pipette out exactly 10 ml of NM Oxalic acid solution into a clean conical flask and add 10 ml of dil. H_2SO_4 . Warm the contents of the flask to 70°C and titrate with KMnO_4 solution. Continue the titration until a faint pink colour is produced. Repeat the titration for concordant values. Enter the values in a tabular form...

Titration of 1M Oxalic acid solution with KMnO_4 solution

S.No.	Volume of 1M Oxalic acid solution (ml)	Burette Reading		Vol. of KMnO_4 rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Molarity of KMnO_4 is calculated using the formula :

$$M_1V_1=M_2V_2$$

where , V_1 = Volume of 1M Oxalic acid solution = 10 ml

M_1 = Molarity of Oxalic acid solution = 1M

V_2 = Volume of KMnO_4 solution rundown (Average Burette Reading)

M_2 = Molarity of KMnO_4 =?

Therefore,

$$M_2 = \frac{M_1V_1}{V_2}$$

Preparation and Standardisation of 0.1M Potassium permanganate :

Potassium permanganate - M.Wt - 158

Preparation of 0.1M Potassium permanganate Solution : Dissolve an accurately weighed amount of 15.8 gm of KMnO_4 in 900 ml of water heat on water bath for 1 hour and then cool it, then filter it using a sintered-glass filter and add sufficient amount of water to produce 1000 ml.

Principle : In volumetric analysis many reactions involve the process of oxidation and reduction. An oxidising agent is estimated by titrating with a reducing agent and vice-versa. These titration are called redox titrations. The standardisation of KMnO_4 is an example of redox titration. KMnO_4 is a powerful oxidising agent and in acidic medium, it oxidises oxalic acid to CO_2 .



The temperature is maintained at $60^\circ\text{-}70^\circ\text{C}$ during the titration because, the reaction proceeds slowly at room temperature. No indicator is required as KMnO_4 , is a self-indicator.

Procedure :

Preparation of 0.1M Oxalic acid Solution : Dissolve an accurately weighed amount of 12.6 gm of Oxalic acid in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 0.1M KMnO_4 : Pipette out exactly 10 ml of 0.1M Oxalic acid solution into a clean conical flask and add 10 ml of dil. H_2SO_4 . Warm the contents of the flask to 70°C and titrate with KMnO_4 solution. Continue the titration until a faint pink colour is produced. Repeat the titration for concordant values. Enter the values in a tabular form...

Titration of 0.1M Oxalic acid solution with KMnO_4 solution

S.No.	Volume of 0.1M Oxalic acid solution (ml)	Burette Reading		Vol. of KMnO_4 rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Molarity of KMnO_4 is calculated using the formula :

$$M_1V_1 = M_2V_2$$

where , V_1 = Volume of 0.1M Oxalic acid solution = 10 ml

M_1 = Molarity of Oxalic acid solution = 0.1M

V_2 = Volume of KMnO_4 solution rundown (Average Burette Reading)

M_2 = Molarity of KMnO_4 = ?

Therefore,

$$M_2 = \frac{M_1V_1}{V_2}$$

Preparation and Standardisation of 1N Potassium permanganate :

Potassium permanganate - M.Wt - 158

Preparation of 1N Potassium permanganate Solution : Dissolve an accurately weighed amount of 31.6 gm of KMnO_4 in 900 ml of water heat on water bath for 1 hour and then cool it, then filter it using a sintered-glass filter and add sufficient amount of water to produce 1000 ml.

Principle : In volumetric analysis many reactions involve the process of oxidation and reduction. An oxidising agent is estimated by titrating with a reducing agent and vice-versa. These titration are called redox titrations. The standardisation of KMnO_4 is an example of redox titration. KMnO_4 is a powerful oxidising agent and in acidic medium, it oxidises oxalic acid to CO_2 .



The temperature is maintained at $60^\circ\text{-}70^\circ\text{C}$ during the titration because, the reaction proceeds slowly at room temperature. No indicator is required as KMnO_4 , is a self-indicator.

Procedure :

Preparation of 1N Oxalic acid Solution : Dissolve an accurately weighed amount of 63.035 gm of Oxalic acid in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 1N KMnO_4 : Pipette out exactly 10 ml of 1N Oxalic acid solution into a clean conical flask and add 10 ml of dil. H_2SO_4 . Warm the contents of the flask to 70°C and titrate with KMnO_4 solution. Continue the titration until a faint pink colour is produced. Repeat the titration for concordant values. Enter the values in a tabular form...

Titration of 1N Oxalic acid solution with KMnO_4 solution

S.No.	Volume of 1N Oxalic acid solution (ml)	Burette Reading		Vol. of KMnO_4 rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Normality of KMnO_4 is calculated using the formula :

$$N_1V_1=N_2V_2$$

where , V_1 = Volume of 1N Oxalic acid solution = 10 ml

N_1 = Normality of Oxalic acid solution = 1N

V_2 = Volume of KMnO_4 solution rundown (Average Burette Reading)

N_2 = Normality of KMnO_4 =?

Therefore,

$$N_2 = \frac{N_1V_1}{V_2}$$

Preparation and Standardisation of 0.1N Potassium permanganate :

Potassium permanganate - M.Wt - 158

Preparation of 0.1N Potassium permanganate Solution : Dissolve an accurately weighed amount of 3.16 gm of KMnO_4 in 900 ml of water heat on water bath for 1 hour and then cool it, then filter it using a sintered-glass filter and add sufficient amount of water to produce 1000 ml.

Principle : In volumetric analysis many reactions involve the process of oxidation and reduction. An oxidising agent is estimated by titrating with a reducing agent and vice-versa. These titration are called redox titrations. The standardisation of KMnO_4 is an example of redox titration. KMnO_4 is a powerful oxidising agent and in acidic medium, it oxidises oxalic acid to CO_2 .



The temperature is maintained at $60^\circ\text{-}70^\circ\text{C}$ during the titration because, the reaction proceeds slowly at room temperature. No indicator is required as KMnO_4 , is a self-indicator.

Procedure :

Preparation of 0.1N Oxalic acid Solution : Dissolve an accurately weighed amount of 6.3 gm of Oxalic acid in sufficient amount of water to produce 1000 ml .

Procedure for Standardisation of 0.1N KMnO_4 : Pipette out exactly 10 ml of 0.1N Oxalic acid solution into a clean conical flask and add 10 ml of dil. H_2SO_4 . Warm the contents of the flask to 70°C and titrate with KMnO_4 solution. Continue the titration until a faint pink colour is produced. Repeat the titration for concordant values. Enter the values in a tabular form...

Titration of 0.1N Oxalic acid solution with KMnO_4 solution

S.No.	Volume of 0.1N Oxalic acid solution (ml)	Burette Reading		Vol. of KMnO_4 rundown (ml)
		Initial	Final	
1	10	0		
2	10			
3	10			

The Normality of KMnO_4 is calculated using the formula :

$$N_1V_1=N_2V_2$$

where , V_1 = Volume of 0.1N Oxalic acid solution = 10 ml

N_1 = Normality of Oxalic acid solution = 0.1N

V_2 = Volume of KMnO_4 solution rundown (Average Burette Reading)

N_2 = Normality of KMnO_4 =?

Therefore,

$$N_2 = \frac{N_1V_1}{V_2}$$

Preparation and Standardisation of 0.1M Ceric ammonium sulphate :

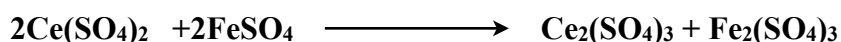
Ceric ammonium sulphate - M.Wt - 632.53

Preparation of 0.1M Ceric ammonium sulphate : Dissolve 63.25 gm of Ceric ammonium sulphate with the aid of gentle heat, in a mixture of 30 ml of sulphuric acid and 500 ml of water. Cool, filter the solution if turbid and dilute to 1000 ml with water.

Principle : Ceric ammonium sulphate in sulphuric acid medium can function as strong oxidising agent and have a high oxidation potential. Sufficient sulphuric acid is to be used to prevent hydrolysis and precipitation of basic salts.

Ceric ammonium sulphate, a salt of suitable solubility for the preparation of the standard solution has the approximate formula $Ce(SO_4)_2 \cdot 2(NH_4)_2SO_4 \cdot 2H_2O$, but the solution has to be standardised against a solution of Arsenic trioxide (As_2O_3). Since Ceric ammonium sulphate is a strong Oxidising agent, we can determine various reducing substances by simple titration.

In presence of reducing agent, it undergoes reduction to the cerous state for example.



Procedure for Standardisation of 0.1M Ceric ammonium sulphate : Weigh accurately about 0.2 gm of **Arsenic trioxide**, previously dried at $105^\circ C$ for 1 hour, and transfer to a 500 ml conical flask. Wash down the inner walls of the flask with 25 ml of a 8% w/v solution of NaOH, Swirl to dissolve, add 100 ml of water and mix.

Add 30 ml of dil. sulphuric acid, 0.15 ml of Osmic acid solution, 0.1 ml of ferroin sulphate solution and slowly titrate with the ceric ammonium sulphate solution until the pink colour is changed to a very pale blue colour.

Equivalent factor : Each ml of 0.1M Ceric ammonium sulphate \equiv 0.04946 gm of As_2O_3

Calculation :

$$\text{Molarity of Ceric ammonium sulphate} = \frac{\text{Weight taken } (Ce(SO_4)_2 \cdot 2(NH_4)_2SO_4 \cdot 2H_2O) \times \text{Req. molarity}}{\text{Titration Volume} \times \text{Equivalent factor}}$$

Preparation and Standardisation of 0.1N Ceric ammonium sulphate :

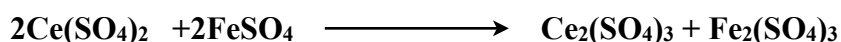
Ceric ammonium sulphate - M.Wt - 632.53

Preparation of 0.1N Ceric ammonium sulphate : Dissolve 63.25 gm of Ceric ammonium sulphate with the aid of gentle heat, in a mixture of 30 ml of sulphuric acid and 500 ml of water. Cool, filter the solution if turbid and dilute to 1000 ml with water.

Principle : Ceric ammonium sulphate in sulphuric acid medium can function as strong oxidising agent and have a high oxidation potential. Sufficient sulphuric acid is to be used to prevent hydrolysis and precipitation of basic salts.

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Procedure for Standardisation of 0.1N Ceric ammonium sulphate : Weigh accurately about 0.2 gm of **Arsenic trioxide**, previously dried at $105^\circ C$ for 1 hour, and transfer to a 500 ml conical flask. Wash down the inner walls of the flask with 25 ml of a 8% w/v solution of NaOH, Swirl to dissolve, add 100 ml of water and mix.

Add 30 ml of dil. sulphuric acid, 0.15 ml of Osmic acid solution, 0.1 ml of ferroin sulphate solution and slowly titrate with the ceric ammonium sulphate solution until the pink colour is changed to a very pale blue colour.

Equivalent factor : Each ml of 0.1N Ceric ammonium sulphate \equiv 0.04946 gm of As_2O_3

Calculation :

Normality of Ceric ammonium sulphate = $\frac{\text{Weight taken } (Ce(SO_4)_2 \cdot 2(NH_4)_2SO_4 \cdot 2H_2O)}{\text{Req. normality}}$

Titration Volume x Equivalent factor

