4.4.2 Practical Exercises on Iodimetry and IodometryTitrations

Experiment 4.4.2.1

You are provided with the following:
FA1 which is 0.06M iodine solution
FA2 which is a solution containing 17.4g of themetal thiosulphate, $XS_2O_3.5H_2O$ in one
You are required to determine the molarity of the metal thiosulphate in FA2 and hence
in $XS_2O_3.5H_2O$.

Theory

Aqueous iodine reacts with thiosulphate ions according to the following equation.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I(aq) + S_4O_6^{2-}(aq)$$

Procedure

i) Using a measuring cylinder, measure and transfer 100cm^3 of FA1 into a 250cm^3 volu
make up to the mark with distilled water. Label the solution FA3.

ii) Pipette 20 or 25cm^3 of FA3 into a clean conical flask and titrate with FA2 from the bu
solution turns pale yellow. Add 1cm^3 of starch solution and continue the titration until
just turns colourless. Repeat the titration until you obtain consistent results. Record your
table below.

Capacity of pipette used..... 25.0 .....cm^3

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA2used(cm^3)			

alues used to calculate average..... cm^3

verage volume of FA2 used..... 17.00 cm^3 cm^3

) Calculate the molarity of iodine in FA3.

1000 cm^3 of FA1 Contain 0.06 mols

100 cm^3 of FA1 Contain $\left(0.06 \times \frac{100}{1000}\right)$ mols = 0.006 moles.

25.0 cm^3 of FA3 Contain 0.006 mols

250 cm^3 of FA3 Contain $\left(0.006 \times \frac{250}{25.0}\right)$ mols = 0.06 mols

250 cm^3 of FA3 Contain 0.06 mols

1000 cm^3 of FA3 Contain $\left(0.06 \times \frac{1000}{250}\right)$ M = 0.24M

b) Determine the:

i) number of moles of thiosulphate ions in FA2 that reacted with the iodine in FA3.

1000 cm^3 of FA3 Contain 0.24 mols

ii) molarity of the thiosulphate ions in FA2 and hence the value of X in $XS_2O_3.5H_2O$.

(S=32, O=16, H=1)

Experiment 4.4.2.2

You are provided with the following:
FA1 which is potassium manganate(VII) solution
FA2 which is sodium thiosulphate solution
FA3 which is 10% potassium iodide solution
FA4 which is 2.0M sulphuric acid
Solid M which is potassium dichromate(VI)
You are required to determine the molar concentration of FA2 and then use it to determine the molar concentration of FA1.

Theory

Dichromate(VI) ions and manganate(VII) ions in an acidic medium, react with iodide ionsaccording the following equations:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq)$$
$$2MnO_4^-(aq) + 16H^+(aq) + 6I^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5I_2(aq)$$

Aqueous iodine reacts with thiosulphate ions according to the following equation.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Procedure I

Weigh accurately 1.4g of solid M into a clean beaker. Measureand transfer about $60cm^3$ of FA4 into the beaker containing solid M and stir well to dissolve. Transfer the solution into a $250cm^3$ volumetric flaskand make up to the mark with distilled water. Label the solution FA5.
Pipette 20 or $25cm^3$ of FA5 into a conical flask. Add $15cm^3$ of FA3 followed by $15cm^3$ of FA4. Titrate the resultant mixture with FA2 from the burette until the solution turns pale yellow. Add $1cm^3$ of Starch solution and continue the titration until the solution turns from dark blue to a pale blue solution. Repeat the titration until you obtain consistent results. Record your results in table I below.

Mass of container + M.....g
Mass of container alone.....g
Mass of solid M.....g
Capacity of pipette used.....25.0.....cm^3

Table I

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA2 used(cm^3)			

Values used to calculate average.....cm^3
Average volume of FA2 used.....22.00cm3.....cm^3

tions

lculate the:

) number of moles of $Cr_2O_7^{2-}$ ions in the 20 or 25cm^3 of FA5 pipetted.
(K=39, Cr=52, O=16).

RFM of $K_2Cr_2O_7$ = (39x2) + (52x2) + (16x7) = 39 + 104 + 112 = 294

294g of FA5 contain 1 mole 0.0476.2

1.4g of FA5 contain $\left(\frac{1 \times 1.4}{294}\right)$ mols = 0.0049 mols

ii) molar concentration of FA2

ocedure II

ette 20 or 25cm^3 of FA1 into a conical flask. Add 15cm^3 of FA3 followed by 15cm^3 of FA4. trate the resultant mixture with FA2 from the burette until the solution turns pale yellow. Add 1cm^3 of arch solution and continue the titration until the end point is reached. Repeat the titration until you tain consistent results. Record your results in table II below.

Pipette = 25.0cm^3

Table II

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA2 used(cm^3)			

alues used to calculate average..... 17.50cm^3 .....cm^3

verage volume of FA2 used.....cm^3

) Determine the molar concentration of FA1.

Experiment 4.4.2.3

You are provided with the following:
HA1 which is a solution containing 1.12g of potassium chromate in $200cm^3$ of solution.
HA2which is a solution made by dissolving 25.0g of a hydrated metal thiosulphate, $MS_2O_3.xH_2O$ in ... litre of solution.
HA3 is 10% potassium iodide solution
HA4 is 2M sulphuric acid
Solid T which is impure potassium iodate, KIO_3.

You are required to determine the:
i) molarity of the hydrated metal thiosulphate in HA2 and the percentage of water of crystallization in the metal thiosulphate, $MS_2O_3.xH_2O$.
ii) percentage purity of the potassium iodate used in the preparation of HA1.

Theory

In acidic medium, CrO_4^{2-} ions are converted to dichromate ions according to the following equation:

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(aq)$$

Still, in acidic medium, IO_3^- and $Cr_2O_7^{2-}$ react with iodide ions according to the equations below:

$$IO_3^-(aq) + 6H^+(aq) + 5I^-(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$$

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq)$$

PART I

Procedure

Pipette 20 or $25cm^3$ of HA1 into a clean conical flask. Add $30cm^3$ of the 2M sulphuric acid followed by $10cm^3$ of the 10% potassium iodide solution and titrate with HA2 from the burette until the solution becomes pale yellow; then add $1cm^3$ of starch indicator and continue the titration until the dark blue solution just turns to a pale bluesolution. Repeat the titration until you obtain consistent results. Record your results in table I below.

Capacity of pipette used……25.0……..cm^3

Table I

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of HA2 used(cm^3)			

Values used to calculate average……21.70cm3……cm³

Average volume of HA2 used……21.70……cm³

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Calculate the number of moles of CrO_4^{2-} ions in the 20 or 25cm^3 of HA1 pipetted.

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Determine the:

i) molarity of the $S_2O_3^{2-}$ ions in HA2.

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ii) value of x in $MS_2O_3.xH_2O$.(M=46, S=32, O=16, H=1)

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iii) percentage by mass of water of crystallisation in the hydrated metal thiosulphate,$MS_2O_3.xH_2O$.

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PART II

Procedure

i) Weigh accurately, 1.0g of solid T into a clean beaker. Add about 150cm^3 of water and stir well to dissolve. Transfer the solution to a 250cm^3 volumetric flask and make up to the mark with distilled water. Label the solution HA5.

ii)Using a measuring cylinder, measure and transfer 25cm^{3}of HA5 into a clean conical flask and add 10cm^3 of the 10% potassium iodide solution followed by 10cm^3 of the 2M sulphuric acid and then titrate the liberated Iodine with HA2 from the burette until the solution is pale yellow; then add 1cm^3 of starch

indicator and continue the titration until the solution just turns colourless.Repeat the titration to obtain consistent results. Record your results in table II below.

Mass of container + T.....g

Mass of container alone.....g

Mass of solid T.....g

Table II pipette = 25.0

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of HA2 used(cm^3)			

Values used to calculate average.....

Average volume of HA2 used.....1.6.80 cm^3.....cm^3

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c) Determine the:

i) number of moles of IO_3^- ions in the $250cm^3$ of HA5.

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ii)mass of pure potassium iodate in the $250cm^3$ of HA5and hence the percentage purity of the potassium iodate sample used in the preparation of HA5.(K=39, I=127, O=16)

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Experiment 4.4.2.4

You are provided with the following:
FA1 which contains 1.8g of potassium dichromate(VI) in 500cm^3 of solution.
FA2 which ishydrogen peroxide solution
FA3 which issodium thiosulphate-5-water
FA4 which is 10% potassium iodide solution
1.5M sulphuric acid
You are required to determine the:
i) concentration of sodium thiosulphate in grams per litre in FA3.
ii)volume strength of the hydrogen peroxide solution in FA2.

Theory

In acidic medium, dichromate(VI) ions and hydrogen peroxide react with iodide ions as shown by the equations below.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq)$$
$$H_2O_2(aq) + 2H^+(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2H_2O(l)$$

Procedure I

Pipette 20 or 25cm^3 of FA1 into a clean conical flask. Add an equal volume of FA4 followed by 30cm^3 of 1.5M sulphuric acid using a measuring cylinder and titrate with FA3 from the burette until the solution turns pale yellow; add 1cm^3 of starch indicator and continue the titration until the blue-black starch-iodine complex turns pale blue.

Capacity of pipette used.....25·0.....cm^3

Table I

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA3 used(cm^3)			

Values used to calculate average.....cm^3

Average volume of FA3 used.....23·00 cm^3.....cm^3

a) Determine the molar concentration of:

i) potassium dichromate(VI) in FA1.

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ii) sodium thiosulphate in FA3.

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Procedure II

Using a measuring cylinder, measure and transfer 100cm^3 of FA3 into a clean beaker. Add 100cm^3 of distilled water, shake well to mix and label the solution FA5.

Pipette 20 or 25cm^3 of FA2 into a clean conical flask. Add an equal volume of FA4 followed by 30cm^3 of 1.5M Sulphuric acid using a measuring cylinder. Leave the mixture to settle for 12 minutes and then titrate with FA5 from the burette as you shake the conical flask and its contents vigorously until the solution turns pale yellow; add 1cm^3 of starch indicator and continue the titration until the blue-black starch-iodine complex turns colourless.

Capacity of pipette used..25.0.....cm^3

Table II

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA5 used(cm^3)			

Values used to calculate average.....31.30.....cm^3

Average volume of FA5 used.....cm^3

b) Calculate the:

i) molar concentration of sodium thiosulphate in FA5.

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ii) molar concentarion of hydrogen peroxide in the FA2 solution.

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iii)volume strength of hydrogen peroxide in the FA2 solution.

(NB: Volume strength is the volume of oxygen gas liberated by 1cm^3 of hydrogen peroxide solution; 1 mole of a gas occupies 24dm^3 at room temperature)

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Experiment 4.4.2.5

You are provided with the following:

FA1 which contains 19.84g of sodium thiosulphate-5-water in one litre of solution.

FA2 which is 10% potassium iodide solution

FA3 which is 2M sulphuric acid

Solution Bwhich is Jik solution [a solution of a bleaching agent that contains hypochlorite/chloric(I) ions]

You are required to determinethe:

i) concentration of sodium thiosulphate in FA1in moldm^{-3}.

ii) percentage by mass of aqueous chlorine in solution B.

Theory

Solutions of bleaching agents such as Jik are prepared by bubbling chlorine gas through a cold dilute solution of sodium hydroxide. Sodium chloride, sodium hypochlorite and water are formed according to the following equation.

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)$$

Addition of a dilute acid to a solution of such a bleaching agent liberates aqueous chlorine according to the equation below.

$$Cl^-(aq) + OCl^-(aq) + 2H^+(aq) \rightarrow Cl_2(aq) + H_2O(l)$$

Since chlorine is more reactive than iodine, the aqueous chlorine has the ability to displace iodide ions from the salt potassium iodide, forming aqueous iodine as shown below.

$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^- + I_2(aq)$$

The aqueous iodine formed can thus be titrated against standard sodium thiosulphate solution in accordance with the equation below.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Procedure

Using a suitable measuring cylinder, measure 20cm^3 of solution B into a 250cm^3 volumetric flask and make up to the mark with distilled water. Label the resultant solution FA4.

Using a measuring cylinder, measure and transfer 25cm^3 of FA4 into a clean conical flask. Using a measuring cylinder, add about 10cm^3 of FA3 followed by 10cm^3 of FA2. Titrate the liberated iodine with FA1 from the burette until the solution turns pale yellow; add 1cm^3 of starch indicator and continue the titration until the blue-black starch-iodine complex just turns colourless. Repeat the titration until you obtain consistent results. Record your results in the table below.

Pipette – 25·0

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA1 used(cm^3)			

Values used to calculate average.....

Average volume of FA1 used..... 16.00 cm^3

a) Determine the concentration of sodium thiosulphate in FA1 in moldm^{-3}.

(Na=23, S=32, O=16, H=1)

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b) Calculate the:

i) number of moles of aqueous iodine liberated by 25cm^3 of FA4.

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massof aqueous chlorine in the $20cm^3$ of solution B.(Cl=35.5)

i) concentration of aqueous chlorine in solution B in gdm^{-3}.

v) percentage by mass of aqueous chlorine in solution B. (Density of solution B=$1gcm^{-3}$)

Experiment 4.4.2.6

You are provided with the following:
GA1 which is 10% potassium iodide solution.
GA2 which contains 8.68gof sodium thiosulphate in $500cm^3$ of solution.
2M sulphuric acid
Solid M which is bleaching powder ($CaOCl_2$).

You are required to determine the:
i) molarity of sodium thiosulphate in GA1.
ii) percentage by mass of available chlorine in the bleaching powder sample.

Theory

The bleaching owder can be dissolved in a specific volume of water and the mixture stirred to make a dilute solution of the calcium hypochlorite.

On addition of a dilute acid to thecalcium hypochlorite solution, aqueous chlorine is liberated from the calcium hypochlorite, $CaOCl_2$according to the equation below.

$$CaOCl_2(aq) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + Cl_2(aq) + H_2O(l)$$

Since chlorine is more reactive than iodine, the aqueous chlorine has the ability adisplace iodide ions from its salt, forming aqueous iodine as shown below.

$$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^- + I_2(aq)$$

The aqueous iodine formed can then be titrated against standard sodium thiosulphate solution in according to the equation below.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Procedure

Weigh accurately, 1.1g of solid M into a beaker. Add $100cm^3$ of water and transfer the resultant solution into a $250cm^3$ volumetric flask. Label the solution GA3.

Pipette 25.0 or $20.0cm^3$ of GA3 into a conical flask. Using a measuring cylinder, add $10cm^3$ of 2M sulphuric acid followed by $25cm^3$ of GA1and titrate the liberated iodine with GA2 from the burette until the solution turns pale yellow. Add1cm^3 of starch indicator and continue the titration until the blue-black starch-iodine complex just turns colourless. Repeat the titration until you obtain consistent results. Record your results in the table below.

Capacity of pipette used.....25.0.....cm^3

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of GA2 used(cm^3)			

Values used to calculate average.....cm^3

Average volume of GA2 used.....24.80cm3.....cm^3

) Calculate the molarity of sodium thiosulphate in GA2.
(Na=23, S=32, O=16, H=1)

b) Calculate the:
i) number of moles of aqueous iodine liberated by 25.0 or 20.0cm^3 of GA3.

ii) mass ofchlorine in the 250cm^3 of GA3.(Cl=35.5)

iii) percentage by mass of available chlorine in solid M.

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Experiment 4.4.2.7

You are provided with the following:

FA1 which is sodium thiosulphate solution.

FA2 which is a solution containing 20.5g of a hydrated copper(II) salt, $CuY.nH_2O$ per litre.

FA3 which is 10% potassium iodide solution

FA4 which is 1M sulphuric acid

Solid U which is potassium iodate

You are required to determine the:

i) concentration of sodium thiosulphate in FA1 in mol dm^{-3}.

ii) value of n in $CuY.nH_2O$ and hence the percentage by mass, of water of crystallisation in the hydrated copper(II) salt.

Theory

In acidic medium, iodate ions react with iodide ions according to the following equation.

$$IO_3^-(aq) + 5I^-(aq) + 6H^+(aq) \rightarrow H_2O(l) + 3I_2(aq)$$

Copper(II) ions react with iodide ions to form both copper(I) iodide and aqueous Iodine as shown by the equation below.

$$2Cu^{2+}(aq) + 4I^-(aq) \rightarrow Cu_2I_2(s) + I_2(aq)$$

The liberated Iodine can then be titrated with sodium thiosulphate as shown by the equation below.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Procedure I

(i) Weigh accurately, 1.0g of solid U into a clean beaker. Using a measuring cylinder, add about $100cm^3$ of water and stir well to dissolve. Transfer into a $250cm^3$ volumetric flask and make up to the mark with distilled water. Label the solution FA5.

Mass of container + U.....g

Mass of container alone.....g

Mass of solid U.....g

(ii) Pipette 20 or $25cm^3$ of FA5 into a clean conical flask. Using a measuring cylinder, add about $10cm^3$ of FA4 followed by $10cm^3$ of FA3.

(iii) Titrate the liberated Iodine with FA1 from the burette until the solution turns pale yellow; add $1cm^3$ of starch indicator and continue the titration until the blue-black starch-iodine complex just turns colourless.

(iv) Repeat the titration until you obtain consistent results. Record your results in the table below.

Capacity of pipette used...25.0.....cm^3

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA1 used(cm^3)			

Values used to calculate average.....31.30cm^3.....

Average volume of FA1 used.....cm^3

.....cm^3

a) Determine the concentration of potassium iodate in FA5 in $moldm^{-3}$. *(K=39, I=127, O=16)*

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b) Calculate the concentration of the sodium thiosulphate in FA1in mol dm^{-3}

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Procedure II

(i) Pipette 20 or 25cm^3 of FA2 into a clean conical flask. Using a measuring cylinder, add about 10cm^3 of FA3.

(ii) Titrate the liberated Iodine with FA1 from the burette until the solution turns pale yellow; add 1cm^3 of starch indicator and continue the titration until the blue-black starch-iodine complex just turns colourless.

(iii)Repeat the titration until you obtain consistent results. Record your results in the table below.

Capacity of pipette used.....cm^3

Pipette – 25.00

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of FA1 used(cm^3)			

Values used to calculate average.....25.00.....

Average volume of FA1 used.....

c) Calculate the:

i) number of moles of copper(II)ions in FA2 which reacted with the iodide ions in FA3.

ii) concentration of copper(II) ions in FA2 in mol dm^{-3}.

d) Determine the value of n in $CuY.nH_2O$ and hence the percentage by mass, of water of crystallization in the hydrated copper(II) salt. (Cu=64, Y=96, H=1, O=16)

Experiment 4.4.2.8

GA1 which is a solution that contains 5.58g of hydrated sodium thiosulphate, $Na_2S_2O_3.5H_2O$ in $250cm^3$ of solution.
GA2 which is potassium manganate(VII) solution.
10% potassium iodide solution.
2M sulphuric acid.
Solid Z which is impure iron(II) oxalate, $Fe(CO_2)_2$.
You are required to determine the:
(i) concentration of the potassium manganate(VII) solution in mol dm^{-3}.
(ii) percentage purity of theiron(II) oxalate sample.

Theory

Acidified manganate(VII) ions react with thiosulphate ions, iron(II) ions and oxalate ions according to the following equations.

$$2MnO_4^-(aq) + 16H^+(aq) + 10S_2O_3^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5S_4O_6^{2-}(aq)$$

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$$

$$2MnO_4^-(aq) + 16H^+(aq) + 5C_2O_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$$

Procedure A

Pipette $25cm^3$ (or $20cm^3$) of GA2 into a conical flask. Add $15cm^3$ of 2.0M sulphuric acid followed by $15cm^3$ of 10% potassium iodide solution and titrate the solution with GA1 from the burette until the solution becomes pale yellow. Add $1cm^3$ of starch indicator and continue the titration until the end point is reached. Repeat the titration until you obtain consistent results. Record your results in table I below.

Capacity of pipette used.....25.00.....cm^3

Table I

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of GA1 used(cm^3)			

Titre values used to calculate verage .....cm^3

Average volume of GA1 used.....25.00.....cm^3

Procedure B

Weigh accurately1.5g of Z into a clean beaker. Add $100cm^3$ of 2.0M sulphuric acid and stir well to dissolve. Transfer the resultant solution into a $250cm^3$ volumetric flask and make up to the mark by addition of more distilled water. Label the solution GA3.

Pipette 25cm^3 (or 20cm^3) of GA3 into a conical flask and heat the solution to 70^{0}C. Titrate the hot solution with GA2 from the burette until the end point is reached. Repeat the titration until you obtain consistent results. Record your results in table II below.

Mass of beaker + Z.....g
Mass of beaker.....g
Mass of Z.....g
Capacity of pipette used.....25.0.....cm^3

Table II

Final burette reading (cm^3)			
Initial burette reading (cm^3)			
Volume of GA2 used(cm^3)			

Titre values used to calculate verage .....27.80 cm^3.....cm^3

Average volume of GA2 used .....cm^3

(a) Calculate the:

(i) molarity of potassium manganate(VII) in GA2. (Na=39, S=32, O=16)

(ii) moles of manganate(VII) ions that reacted with 25cm^3 (or 20cm^3) of GA3.

(iii) moles of iron(II) ions in $25cm^3$ (or $20cm^3$) of GA3.

(b) Determine the:

(i) mass of iron(II) oxalate, FeC_2O_4 in $250cm^3$ of GA3. (Fe = 56, C=12, O=16)

(ii) percentage purity of the iron(II) oxalate sample.