Name: Chik Chun Pong Date of experiment: 9/9/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 1

Title:

1) To prepare a standard solution of the Sulphamic Acid
2) To determine the concentration of a solution of Sodium Hydroxide by titrating against the Sulphamic Acid Solution

Apparatus:

Pipette, burette, conical flask, beaker, volumetric flask, glass rod, wash bottle, funnel, plastic bottle, electronic balance, stand and clamp

Materials:

Sulphamic acid, Sodium hydroxide, phenolphthalein

Theory:

1) As the mass of the Sulphamic acid was known, a standard solution could be prepared.
2) As the concentration of the Sulphamic acid was known, the concentration of the Sodium Hydroxide solution could be obtained by titrating the solution against the Sulphamic acid solution.

Procedure:

1) About 2.41g to 2.45g of Sulphamic acid was weighed in a weighing bottle by the electronic balance.
2) The solid was transferred totally to a beaker and was dissolved in sufficient amount of water.
3) The solution was poured into a 250 cm^3 graduated flask. Water was added to the mark. The flask was inverted several times to mix it well.
4) The burette was filled with Sodium Hydroxide solution provided.
5) 25 cm^3 of the Sulphamic acid solution was transferred into a conical flask by a pipette. 1 to 2 drops of phenolphthalein solution were added.
6) The solution was titrated with Sodium Hydroxide solution until a pink colour is obtained.
7) Steps 4 to step 7 were repeated once.

Results:

Mass of Sulphamic acid used = 2.45 g

	Trial 1	Trial 2	Trial 3
Initial Volume(cm^3)	1.00cm^3	24.00cm^3	4.30cm^3
Final Volume (cm^3)	24.00cm^3	46.50cm^3	27.05cm^3
Volume of NaOH used (cm^3)	23.00cm^3	22.50cm^3	22.75cm^3
Mean Volume (cm^3)	(23.00+22.50+22.75) / 3 =22.75 cm^3		

Calculations:

a. Molar mass of Sulphamic acid =97.02g

No. of moles of Sulphamic acid =2.45 / 97.02
=0.0253 (mol)

Molarity of the Sulphamic acid Sol'n =0.0253 / 0.250
=0.101 (M)

b. No.of moles of 25cm^3 Sulphamic acid =0.0253 / 10
=0.00253 (mol)

No.of moles of Sulphamic acid:No.of moles of sodium hydroxide =1:1

No.of moles of sodium hydroxide = 0.00253 (mol)

Molarity of sodium hydroxide solution = 0.00253 / 0.02258
= 0.112 (M)

Conclusion:

The concentration of the standard sulphamic acid solution is 0.101M. The concentration of the sodium hydroxide is 0.112M .

Answers to Questions:

Primary Standard is a solution made by dissolving a definite amount of a substance in water.

A good primary standard should be obtainable in a high degree of purity. It should be stable and have a high molar mass. It should not be decomposed during storage. It should be a strong electrolyte.

Discussion on the sources of error and improvements:

1) Due to our inexperienced and confusion in determining the colour of the indicator, the error in judgement of end-point may arise. This may be minimized by using a pH meter instead of the indicator.

2) Errors in weighing the materials may also occurred ,a more powerful (with 4 or more decimal places) electronic balance may be used to narrow down the range of error.

3) Solution is not homogeneous; this may be solved by shaking the solution very well.

Name: Chik Chun Pong Date of experiment: 23/9/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 2

Title:

1) To prepare a standard solution of the sodium carbonate
2) To standardize a hydrochloric acid solution
3) To determine the concentrations of a solution of sodium hydroxide and sodium carbonate in a mixtures of solutions by using
 a) phenolphthalein and b)methyl orange as indicators

Apparatus:

Pipette, pipette filter, burette, conical flask, volumetric flask, glass rod, wash bottle, funnel, plastic bottle, electronic balance, stand and clamp, white tile

Materials:

Sodium carbonate solution, hydrochloric acid solution, a mixture of sodium carbonate solution and sodium hydroxide solution, methyl orange, phenolphthalein

Theory:

1) As the concentration of sodium carbonate was known, the concentration of hydrochloric acid solution could be obtained by titrating the solution against the sodium carbonate solution.
2) As the concentration of hydrochloric acid solution was known, the concentrations of sodium carbonate and sodium hydroxide in the mixture could be obtained by the double indicator method.

Procedure:

Part 1:

250 cm^3 of 0.05 M sodium carbonate solution was prepared.

Part 2:

Titrate 25 cm^3 of the sodium carbonate solution with the provided hydrochloric acid solution using methyl orange as indicator.
Repeat twice.

Part 3:

1. Pipette 25cm^3 of the provided mixture of solutions into a conical flask. 1 to 2 drops of phenolphthalein were added. Titrate against hydrochloric acid solution until the colour changes from red to

colourless. The volume of hydrochloric acid used (V_1) was recorded.

2. To the resulting solution add 1 to 2 drops of methyl orange. Continue the titration until the colour changes from yellow to orange. The volume of hydrochloric acid used (V_2) was recorded.
3. repeat twice.

Results:

Part II titrations:

	Titration		
	1	2	3
Initial volume (cm^3)	4.40	3.10	25.60
Final volume (cm^3)	26.90	25.60	47.95
Volume of *HCl* used (cm^3)	22.50	22.50	22.35
Mean volume of the three titrations (cm^3)	22.45		

Part III titrations:

	Titration		
	1	2	3
Initial volume (cm^3)	2.80	24.70	4.00
Intermediate volume (cm^3)	18.90	40.80	20.10
Final volume (cm^3)	24.70	46.60	25.90
Volume of *HCl* used (V_1) (cm^3)	16.10	16.10	16.10
Volume of *HCl* used (V_2) (cm^3)	5.80	5.80	5.80
Mean volume of the 3 titrations (V_1) (cm^3)	16.10		
Mean volume of the 3 titrations (V_2) (cm^3)	5.80		

Calculations:

a. the molarity of hydrochloric acid solution and
b. the molarities of sodium hydroxide and sodium carbonate in the mixture.

Equations: $OH^- + H^+ \rightarrow H_2O$

$CO_3^{2-} + H^+ \rightarrow HCO_3^-$

$HCO_3^- + H^+ \rightarrow H_2O + CO_2$

Chemistry Experiment Laboratory Report (2)

(a) $2HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

22.45 cm^3 25cm^3
? M 0.05M

Number of moles of Na_2CO_3

$= 0.05 \times \dfrac{25}{1000}$

= 0.00125 mol.

From the above equation,
Number of moles of Na_2CO_3 : Number of moles of HCl = 1 : 2

Number of moles of HCl
= 0.00125 × 2
= 0.0025 mol.

Molarity of hydrochloric acid
= 0.0025 / (22.45/1000)
= $\underline{\underline{0.1114M}}$

(b) Volume V_2 is obtained by the following equation:

$HCO_{3\ (aq)}^{-} + H^{+}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

5.80 cm^3
0.1114 M

Number of moles of HCl
= 0.1114 × (5.80 / 1000)
= 0.0006461 mol.

From the equation above,
Number of moles of H^+ : Number of moles of HCO_3^- = 1 : 1

Number of moles of HCO_3^- = 0.0006461 mol

Since HCO_3^- ions came from CO_3^{2-} ions, their numbers of moles are the same.

Number of moles of CO_3^{2-} = 0.0006461 mol

Molarity of sodium carbonate solution

$= 0.0006461 \div \dfrac{25}{1000}$

= $\underline{\underline{0.02584 M}}$

From the equation
$2H^+_{(aq)} + CO_{3\ (aq)}^{2-} \rightarrow H_2O_{(l)} + CO_{2(g)}$
Number of moles of HCl

= 2 × 0.0006885
= 0.001292 mol.

In 16.10cm^3 of *HCl*, we have 0.1114 × (16.10 / 1000) = 0.001794 mol.

Since 0.001292 mol. out of 0.001794 mol. of *HCl* was reacted with CO_3^{2-} ions, there are 0.000502 mol. of *HCl* reacts with *NaOH* in the mixture.

From the equation:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Number of moles of *NaOH* = Number of moles of *HCl* = 0.000502 mol.
Molarity of sodium hydroxide solution

$= 0.000502 \div \dfrac{25}{1000}$

= $\underline{\underline{0.02008 \text{ M}}}$

Conclusion:

The molarity of hydrochloric acid was found to be 0.1114 M.

In the mixture,
the molarity of sodium hydroxide solution was found to be 0.02008 M &
the molarity of sodium carbonate solution was found to be 0.02584 M.

Discussion on the sources of error and improvements:

1) Inaccuracy in judging the end-point and of the instruments used.
   It could be improved by using a pH meter and taking an average of 3 close titration results.
2) Solution is not homogenous.
   This can be minimized by shaking the solution very well.
3) Impurities of materials especially on those solutions.
   Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may arise errors. These can be minimized by the use of tightly fitting stoppers , coloured glass bottles, and soda-lime tubes to absorb carbon dioxide.

-End-

Name: Chik Chun Pong
Class: 6S2
Class number: 8

Date of experiment: 30/9/03

Chemistry Experiment Laboratory Report 3

Title:

To study electrolysis quantitatively

Apparatus:

Ammeter, rheostat, copper electrodes, connecting wires, sand paper, electronic balance, 250 cm^3 beaker, electrode holder

Materials:

1.0 M copper (II) sulphate solution , propanone (or ethanol)

Theory:

The relationship between the mass of a metal deposited during electroplating and the quantity of electricity passed can be studied by measuring the changes in mass of the copper electrodes during the electrolysis of copper (II) sulphate solution for a certain time using a constant current.

Procedure:

1) Clean carefully the copper foil cathode and anode by rubbing with sand paper and then with paper tissue.
2) The copper cathode and anode were weighed respectively.
3) A 250 cm^3 beaker was filled with 1.0 M copper (II) sulphate solution to three quarters and the electrodes were inserted. Use an electrode holder to hold them in place.
4) The electrodes were connected for electrolysis as shown in the above diagram. Start taking time when the circuit was just closed. The rheostat was adjusted to give a current of 0.20 A.
5) The current was allowed to pass through the solution for 30 minutes. A current of 0.2 A was kept by adjusting the rheostat when necessary.
6) The circuit was disconnected after 30 minutes and the copper electrodes were rinsed with water. A strong jet of water from the tap was required for the anode, and a firm wipe with a paper tissue was followed in order to remove the film on the surface.
7) Each electrode was rinsed twice with propanone or ethanol. They were waved in the air for 1-2 minutes to vaporize most of propanone or ethanol, and were dried by warming high above a Bunsen flame.

8) Each electrode was reweighed separately.

Results:

Weights of the copper electrodes:

	ANODE	CATHODE
Before experiment (g)	59.96	60.79
After experiment (g)	59.83	60.92

Calculations:

1. Gain in mass of copper in cathode:
= 60.92 – 60.79
= 0.13 g
Loss in mass of copper in anode:
= 59.96 – 59.83
= 0.13 g
Number of moles of 0.13 g of copper:

$$= \frac{0.13}{63.5}$$

= 0.002047 mol.
From equation:

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow 2Cu_{(s)}$$

Number of moles of electrons required:
= 0.002047 mol.
Number of coulombs required to form 1 mole of copper:

$$= \frac{0.2A \times 60 \times 30\,\text{sec.}}{0.002047\,mol}$$

= $\underline{\underline{175867.12\text{ C}}}$

2. Since

$$\textit{Number of moles of product formed} = \frac{It}{nF},$$

where *I* is the current in Ampere, *t* is the time in seconds, *n* is the number of moles of electrons required to from 1 mole of product during electrolysis , and *F* is the Faraday constant in C mol^{-1}.

By putting $I = 0.2$, $t = 30 \times 60 = 1800$, $n = 2$, number of moles = 0.002047 into the above equation, we have:

$$0.002047 = \frac{0.2 \times 1800}{2F}$$

$$F = \underline{\underline{87933.56\text{ C mol}^{-1}}}$$

Conclusion:

From the experiment, the quantity of electricity required to deposit one mole of copper metal on the cathode was found to be 175867.12 C and the value of one Faraday was found to be 87933.56 C mol^{-1}.

Discussion on the sources of error and improvements:

Discussion:

From the experiment, the calculated value of one Faraday, 87933.56 C mol^{-1}, is 8566.44 C mol^{-1} less than the literature value, 96500.00 C mol^{-1}.

Different in the changes in mass of the cathode and anode respectively after electrolysis, are due to:

- the dirt on the anode was not fully washed away ;
- some of the copper deposited on the cathode may be washed away by the tap water;
- there may be some dirt that are not totally removed and makes the copper atoms cannot deposit on, leading the copper atoms remained in the solution ;
- some of the electricity is used to discharge the hydrogen ions and hydroxide ions in the copper (II) sulphate solution.

Sources of errors:

Errors may arise in the experiment including:

- Errors in weighting the electrodes: This can be improved by using an electronic balance of higher precision.
- Solution is not homogenous: This can be minimized by shaking the solution very well.
- Impurities of materials: Moisture and dusts can be minimized by covering the beaker of copper (II) sulphate with materials like cling wrap.
- Current is not steady: This can be improved by using batteries or power supply units individually, but not sharing electricity on a series circuit.
- Measuring of ampere is not accurate: This can be improved by using an ammeter of higher accuracy or an electronic ammeter.
- Time taking is not accurate: This can be improved by using a stopwatch or a time switch which will disconnect the circuit automatically when the time is up.

Precautions & Important points:

- Since the propanone or ethanol is highly flammable, we should keep them away from naked flames.
- We should disconnect the circuit immediately when the time is up.
- We should also take care when we use a Bunsen flame to vaporize the propanone or ethanol on the electrodes.

-End-

Chemistry Experiment Laboratory Report (4)

Name: Chik Chun Pong Date of experiment: 7/10/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 4(2nd edition)

Title:

To determine the percentage by mass of aspirin in Aspirin Tablets.

Apparatus:

Washing bottle , weighting bottle , electronic balance , beaker , glass rod , volumetric flask , plastic funnel , pipette , pipette filter , conical flask , burette , clamp and stand , white tile , wire gauze , tripod , heat proof mat , Bunsen burner.

Materials:

2 aspirin tablets, 0.05 M sodium carbonate solution, 0.501 M sodium hydroxide solution, methyl orange indicator, phenol red indicator.

Theory:

Aspirin is 2-ethanoyloxybenzenecarboxylic acid (acetyl-salicylic acid) which can be readily hydrolyzed by sodium hydroxide into the sodium salts of two weak acids, ethanoic acid and benzenecarboxylic acid. In this experiment, the excess amount of sodium hydroxide used will later be found by titration with standard sulphuric acid.

$$(CH_3COO)C_6H_4(COOH) + 2NaOH \rightarrow (OH)C_6H_4(COONa) + CH_3COONa$$

Procedure:

Part I

The sulphuric acid provided was standardized by titrating against standard solution of sodium carbonate. Methyl orange was used as indicator.

Part II

1. Two aspirin tablets were weighted.
2. The tablets were dissolved in 50 cm^3 (exactly) of 0.5 M sodium hydroxide and 25 cm^3 (approximately) of water.
3. The mixture was heated gently for 10 to 15 minutes. After heating, the mixture was cooled thoroughly.
4. The mixture was transferred to a 100 cm^3 volumetric flask. Water was added to the graduation mark.

5. 25 cm^3 of the solution was titrated with the standard sulphuric acid using phenol red indicator until the colour changes from red to orange. The titration was repeated once.

Results:

Part I

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	0.00	19.80	28.60
Final volume (cm^3)	19.80	39.80	48.30
Volume of sulphuric acid used (cm^3)	19.80	20.00	19.70
Mean volume (cm^3)	[(19.80+20.00+19.70)/3]=19.83		

Part II

The two aspirin tablets totally weight: 0.80 g

Volume of sodium hydroxide used: 50.00 cm^3

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	2.60	3.20	3.95
Final volume (cm^3)	37.20	37.90	38.50
Volume of sulphuric acid used (cm^3)	34.60	34.70	34.55
Mean volume (cm^3)	[(34.60+34.70+34.55)/3]=34.62		

Calculations:

Equation from part I:

$$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + CO_2$$

0.05 M ? M

25.0 cm^3 19.83 cm^3

Number of moles of Na_2CO_3

$= 0.05 \times \frac{25.0}{1000}$

= 0.00125 mol

From the equation,

Number of moles of Na_2CO_3 : Number of moles of H_2SO_4 = 1 : 1

Number of moles of H_2SO_4 = 0.00125 mol

Molarity of H_2SO_4

$= 0.00125 \div \frac{19.83}{1000}$

= 0.06304 M

Chemistry Experiment Laboratory Report (4)

In part II, the following two reactions occurred:

{1} In the 100 cm^3 solution used in the titration:

$$(CH_3COO)C_6H_4(COOH) + 2NaOH \rightarrow (OH)C_6H_4(COONa) + CH_3COONa$$

{2} In the titration;

$$2\,NaOH \quad + \; H_2SO_4 \; \rightarrow \; Na_2SO_4 \; + \; H_2O$$

not reacted 0.06304 M
with Aspirin 34.62 cm^3

From {2}, number of moles of H_2SO_4
$= 0.06304 \times \frac{34.62}{1000}$
= 0.0021824 mol

From {2},
Number of moles of H_2SO_4 : Number of moles of $NaOH$ = 1 : 2
Number of moles of $NaOH$ in {2} $= 2 \times 0.0021824$
= 0.004365 mol

Number of moles of $NaOH$ not reacted with the aspirin in the 100 cm^3 solution
$=0.004365 \times 4$
=0.0174596

Total number of moles of $NaOH$
$= 0.501 \times \frac{50.0}{1000}$
= 0.02505 mol.

Number of moles of $NaOH$ reacted with aspirin
=0.02505-0.0174596
=0.0075904 mol.

From {1},
Number of moles of $NaOH$: Number of moles of $(CH_3COO)C_6H_4(COOH)$ = 2 : 1
Number of moles of $(CH_3COO)C_6H_4(COOH)$
$=0.0075904 \div 2$
=0.0037952 mol.

Mass of $(CH_3COO)C_6H_4(COOH)$ in two tablets
$=0.0037952 \times 180$
=0.683136 g

Percentage by mass of $(CH_3COO)C_6H_4(COOH)$ in 2 tablets
$= 0.683136 \div 0.80$
= 85.39%

From the package of the aspirin tablets, it claims that there are 300 mg of aspirin per tablet, i.e. 0.3 g per tablet.
From the experiment results, it shows that there are 0.683136 g of aspirin 2 tablets,
Thus, it is 0.083136 g greater than that of the manufacturer's specification.

Discussion on the sources of error and improvements:

Conclusion:

The percentage by mass of aspirin in the aspirin tablet was found to be 85.39%.
The mass of aspirin was found to be 0.083136 g greater than the manufacturer's specification.

Sources of errors:

Errors may arise in the experiment including:

1) Errors in weighting: It can be improved by using an electronic balance of higher precision.
2) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may arise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.
3) Solution is not homogenous: This can be minimized by shaking the solution very well.
4) Inaccuracy in judging the end-point and of the instruments used: It can be improved by using a pH meter and taking average of 3 close titrations.

Precautions:

1) We should wear safety goggles when heating with Bunsen flame.

Chemistry Experiment Laboratory Report (5)

Name: Chik Chun Pong
Date of experiment: 14/10/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 5

Title:

To determine the percentage purity of oxalic acid in a sample by titrating with potassium permanganate solution.

Apparatus:

Washing bottle, weighting bottle, electronic balance, beaker, glass rod, volumetric flask, plastic funnel, pipette, pipette filter, conical flask, burette, clamp and stand, white tile, measuring cylinder, wire gauze, tripod, heat proof mat, Bunsen burner.

Materials:

Ammonium iron (II) sulphate hexahydrate crystals, dilute sulphuric acid, potassium permanganate solution, oxalic acid.

Theory:

1) As the concentration of ammonium iron (II) sulphate solution was known, the concentration of potassium permanganate solution could be obtained by titrating the solution against the ammonium iron (II) sulphate solution.
2) As the concentration of potassium permanganate solution was known, the concentration of oxalic acid solution could be obtained by titrating the solution against the potassium permanganate solution, and hence, its purity.

Procedure:

Part I Standardization of potassium permanganate.

1. 9.8 g of ammonium iron (II) sulphate hexahydrate $(NH_4)_2SO_4 \cdot FeSO_4 \cdot 6H_2O$ (with formula mass = 392) was weighted.
2. The solid was dissolved in 25 cm^3 dilute sulphuric acid.
3. The solution was transferred to a 250 cm^3 volumetric flask. Water was added up to the graduation mark. The mixture was mixed well.
4. 25 cm^3 of the solution was transferred to a conical flask using a pipette. 25 cm^3 of dilute sulphuric acid was added to the conical flask.

5. The solution was titrated with potassium permanganate solution until a light purple colour appears.
6. Steps 4 and 5 were repeated.

Part II Determination of the concentration of oxalic acid solution.

1. 25 cm^3 of the provided oxalic acid solution was transferred to a conical flask using a pipette. 25 cm^3 of dilute sulphuric acid was added to the conical flask.
2. The mixture was heated to about 60℃.
3. The hot solution was titrated with potassium permanganate solution until a light purple colour was obtained.
4. The steps were repeated.

Results:

Part I

Weight of ammonium iron (II) sulphate hexahydrate crystals: 9.82 g

Titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	2.10	26.65	1.70
Final volume (cm^3)	26.65	50.00	26.15
Volume of $KMnO_4$ used (cm^3)	24.55	24.35	24.45
Mean volume (cm^3)	24.45		

Part II

Titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	0.00	24.10	18.00
Final volume (cm^3)	24.10	48.40	42.50
Volume of $KMnO_4$ used (cm^3)	24.10	24.30	24.50
Mean volume (cm^3)	24.30		

Calculation:

In part I, the following reaction occurred:

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

Since one mole of ammonium iron (II) sulphate hexahydrate crystals contains one mole of iron (II) ions,

Molarity of the ammonium iron (II) sulphate solution

$$= \frac{9.82}{392} \div \frac{250}{1000}$$

= 0.1002 M

Number of moles of Fe^{2+} in 25 cm^3 solution

$= 0.1002 \times \frac{25}{1000}$

= 0.0025051 mol

From the equation above,

Number of moles of Fe^{2+} : Number of moles of MnO_4^- = 5 : 1

Number of moles of MnO_4^- = $0.0025025 \div 5$

= 0.0005010 mol

Molarity of $KMnO_4$

$= 0.0005005 \div \frac{24.45}{1000}$

= 0.02049 M

In part II, the following reaction occurred:

$$5H_2C_2O_4 + 2MnO_4^- + 6H^+ \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$$

Number of moles of MnO_4^-

$= 0.02024 \times \frac{24.30}{1000}$

= 0.00049794 mol

From the equation,

Number of moles of $H_2C_2O_4$ $= \frac{5}{2} \times$ Number of moles of MnO_4^-

$= \frac{5}{2} \times 0.00049794$

= 0.00124487 mol

Mass of $H_2C_2O_4$ = $0.00124487 \times [2 \times 1+12 \times 2+16 \times 4+2 \times (1 \times 2+16)]$

= 0.15685 g

Concentration of $H_2C_2O_4$

$= 0.15685 \div \frac{25}{1000}$

= 6.27 g dm^{-3}

The given concentration of oxalic acid was 6.5 g dm^{-3}, which means there will be 6.5 g of hydrated oxalic acid in 1 dm^3 of water.

However, the calculation shows that the concentration is 6.27 g dm^{-3}, which means there will be 6.27 g of hydrated oxalic acid in 1 dm^3 of water.

Hence, the purity of hydrated oxalic acid is given by:

$\frac{6.27}{6.5} \times 100\%$

= 96.5%

Conclusion:

The concentration of oxalic acid in the solution was found to be 6.27 g dm^{-3}.
The purity of hydrated oxalic acid in the sample is 96.5 %.

Answers to the questions:

1) Iron (II) sulphate crystals were not used in this experiment because they were rendered impure by efflorescence and by atmospheric oxidation to form a brown basic sulphate. However, ammonium iron (II) sulphate crystals are free from these advantages.
2) Since potassium permanganate solution was so oxidative that it may be reduced by organic matter from the atmosphere and so rendered impure. Moreover, organic matter organic matter present in the water in which the salt was dissolved may reduce it.
3) An indicator was not used in this experiment because, as soon as potassium permanganate is in excess, the solution becomes pink or light purple and therefore potassium permanganate act as its own indicator.
4) Adding of dilute sulphuric acid in each experiment was used to avoid the formation of manganese (IV) oxide when the solution was lack of hydrogen ions.

Sources of errors:

Errors may arise in the experiment including:

1) Errors in weighting: It can be improved by using an electronic balance of higher precision.
2) Solution is not homogenous: This can be minimized by shaking the solution very well.
3) Inaccuracy in judging the end-point and of the instruments used: It can be improved by using a pH meter and taking average of 3 or more close titrations.
4) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.

Chemistry Experiment Laboratory Report (6)

Name: Chik Chun Pong Date of experiment: 21/10/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 6 (2nd EDITION)

Title:

(i) A study of titrations using iodine solutions.
(ii) Determination of the amount of the active ingredient in bleaching solution.

Apparatus:

Washing bottle, weighting bottle, electronic balance, beaker, glass rod, 100 cm^3 and 250 cm^3 volumetric flasks, plastic funnel, pipette, pipette filter, conical flask, burette, clamp and stand, white tile, measuring cylinder.

Materials:

Potassium iodate (V) solid, dilute sulphuric acid (1.0 M) , 1M potassium iodide, 1.0 M potassium iodide solution, sodium thiosulphate solution, starch solution as indicator, bleaching solution, distilled water.

Theory:

1) As the mass of potassium iodate was known, a standard solution could be prepared.
2) An iodine solution is prepared from reacting potassium iodate with potassium iodide which is known in their masses. Since the concentration of iodine solution was known, the concentration of sodium thiosulphate solution could be obtained by titrating the solution against the iodine solution.
3) Sodium chlorate (I) reacts with iodide ions in acidic medium to give iodine. By titrating the solution with standard sodium thiosulphate, the number of moles of iodine and thus the concentration of sodium chlorate in the solution can be found.

Procedure:

Part I To prepare a standard solution of potassium iodate (V) (M/60)

1. 0.89 g of potassium iodate (V) was weighed and dissolved in suitable amount of water.
2. 10 $cm^3$of dilute sulphuric acid and 10 g potassium iodide (which is in excess) were added to the solution.
3. The mixture was transferred to a 250 cm^3 volumetric flask. Water was added to the mark. The mixture was mixed

thoroughly.

Part II To standardize an unknown solution of sodium thiosulphate

1. 25 cm^3 of the solution prepared in part I was transferred to a conical flask using a pipette.
2. Sodium thiosulphate was added to the conical flask through a burette until the colour changed to light yellow.
3. A few drops of starch solution were added to the mixture. The titration was continued until the dark blue colour of starch was discharged.
4. The steps in part II were repeated.

Part III To determine the amount of sodium chlorate (I) in bleaching solution

1. 5 cm^3 of the bleaching solution was pipetted into a $100cm^3$ volumetric flask using a pipette filler. The solution was made up to $100cm^3$.
2. 25 cm^3 of this solution was pipetted to a conical flask. 10 cm^3 of 1 M potassium iodide solution and 10 cm^3 1 M sulphuric acid were also added to it.
3. The liberating iodine was titrated against the sodium thiosulphate solution using starch solution as indicator.
4. The steps in part III were repeated.

Results:

Part I

Weight of potassium iodate (V) solid used: 0.89 g

Part II titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	1.15	25.40	24.65
Final volume (cm^3)	25.40	49.60	48.95
Volume of $Na_2S_2O_3$ used (cm^3)	24.25	24.20	24.30
Mean volume (cm^3)	24.25		

Part III titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	23.00	23.40	14.20
Final volume (cm^3)	44.30	45.45	36.10
Volume of $Na_2S_2O_3$ used (cm^3)	21.30	22.05	21.90
Mean volume (cm^3)	21.75		

Calculation:

In part I,

$$IO_{3\,(aq)}^{-} + 5I^{-}_{(aq)} + 6H^{+}_{(aq)} \rightarrow 3I_{2(aq)} + 3H_2O_{(l)}$$

0.89 g   in excess

Number of moles of IO_3^-

$$= \frac{0.89}{39 + 127 + 3\times 16}$$

= 0.004159 mol

Number of moles of I_2 = 3× number of moles of IO_3^-

= 3× 0.004159

= 0.01248 mol

Molarity of I_2 in the solution

$$= 0.01248 \div \frac{250}{1000}$$

= 0.04991 M

In part II,

$$I_{2(aq)} + 2S_2O_{3\,(aq)}^{2-} \rightarrow S_4O_{6\,(aq)}^{2-} + 2I^{-}_{(aq)}$$

0.04992M   ? M

25 cm^3 24.25 cm^3

Number of moles of I_2

$$= 0.04992 \times \frac{25}{1000}$$

= 0.001248

Number of moles of $S_2O_3^{2-}$ = 2× number of moles of I_2

= 2× 0.001248

= 0.002495

Molarity of $S_2O_3^{2-} = 0.002496 \div \frac{24.25}{1000}$

= 0.1029 M

In part III,

[1]: $ClO^{-}_{(aq)} + 2I^{-}_{(aq)} + 2H^{+}_{(aq)} \rightarrow I_{2(aq)} + Cl^{-}_{(aq)} + H_2O_{(l)}$

? mol   10 cm^3 10 cm^3 ? mol

1 M   1 M

[2]: $I_{2(aq)} + 2S_2O_{3\,(aq)}^{2-} \rightarrow S_4O_{6\,(aq)}^{2-} + 2I^{-}_{(aq)}$

? mol   0.1029M

21.75 cm^3

Number of moles of $S_2O_3^{2-} = 0.1029 \times \frac{21.75}{1000}$

= 0.002238 mol

From equation {2},

Number of moles of I_2 = Number of moles of $S_2O_3^{2-} \div 2$

= $0.002238 \div 2$

= 0.001119 mol

In equation {1}, number of moles of I_2 = 0.001119 mol.

$\therefore$ Number of moles of ClO^- = 0.001119 mol

Number of moles of ClO^- in 5 cm^3 of bleaching solution

= 0.001119×4

= 0.004476 mol

Mass of *NaClO* in 5 cm^3 bleaching solution

= $0.004476 \times (23 + 35.5 + 16)$

= 0.3335 g

Amount of *NaClO* in the bleaching solution

= $0.3335 \div \frac{5}{1000}$

= $\underline{\underline{66.69 \text{ g dm}^{-3}}}$

Conclusion:

The concentration of sodium thiosulphate was found to be 0.1029 M.
The amount of sodium chlorate (I), *NaClO*, in the bleaching solution was found to be 66.69 g dm^{-3}.

Answers to the questions:

1) Starch is made from a helical polymer of α-glucose. Iodine molecules are trapped inside the helix and form a complex of deep blue colour. Starch was added just before the endpoint, because some iodine will remain bound to the indicator and cannot react with the thiosulphate ions if the indicator was added too early.
The iodine may fuse into the starch complex and make it difficult to be decolourized.

2) KI was used in excess because it provides sufficient I^- ions for the IO^{3-} ions to displace and converts into iodine molecules.
If a student uses much less KI than the specified quantity, quite a large amount of iodate (V) ions will remain in the solution and cannot react with thiosulphate ions in the titration. So, the amount of $Na_2S_2O_2$ needed to titrate would be less than normal.

3) Dilute H_2SO_4 provide H^+ ions for the CIO^- ions to turn I^-(aq) into I_2(aq).

If bleaching solution is added to the acid before the addition of KI, chlorate (I) ions will react with hydrogen ions to form chlorine gas, and hence, affecting the concentration of chlorate (I) ions in the diluted bleaching solution need to be tested.

Sources of errors:

Errors may arise in the experiment including:

1) Errors in judging the end point: It can be improved by taking the average of results of more titrations.
2) Errors in weighting: It can be improved by using an electronic balance of higher precision.
3) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.
4) Solution is not homogenous: This can be minimized by shaking the solution very well.
5) Some of the CIO^- ions in the bleaching solution became Cl_2 gas and escape into the atmosphere, it can be improved by storing the bleaching solution in a covered container before the experiment.

Precautions:

- ✓ Add potassium iodide solution before adding dilute sulphuric acid to the diluted bleaching solution in part III.
- ✓ We should be careful when adding water to the graduation mark in the volumetric flask.

Chemistry Experiment Laboratory Report (7)

Name: Chik Chun Pong Date of experiment: 28/10/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 7

Title:

To study the heat change in the conversion of solid sodium hydroxide to sodium chloride solution.

Apparatus:

Polythene bottle, weighing bottle, two 50 cm^3 measuring cylinders, thermometer, electronic balance.

Materials:

Solid sodium hydroxide, 4.0 M hydrochloric acid.

Theory:

Heat change of the formation of sodium chloride is studied using Hess's Law.

Procedure:

Route 1

1. A polythene bottle, which is the calorimeter in this experiment, was weighted
2. 4.00±0.02 g of sodium hydroxide flakes was weighed in a weighing bottle.
3. 50 cm^3 of water was placed in a measuring cylinder and 50 cm^3 of 4.0 M hydrochloric acid in another measuring cylinder.
4. 25 cm^3 of water was placed in the calorimeter and its temperature was recorded.
5. The solid sodium hydroxide was tipped into the water in three or four stages. It was stirred with the thermometer. The temperature was recorded as soon as the entire solid had dissolved. 25 cm^3 of the hydrochloric acid was added immediately. The mixture was stirred and its temperature was recorded.

Route 2

1. 4.00±0.02 g of sodium hydroxide was weighted.
2. The remaining 25 cm^3 of distilled water was placed into the calorimeter and its temperature was recorded.
3. The temperature of hydrochloric acid was recorded.
4. The remaining 25 cm^3 of the acid was added to the water. It was stirred with the thermometer and the temperature was

recorded.

5. The solid sodium hydroxide was added immediately in three or four quick stages. The mixture was stirred and its temperature was recorded as soon as the entire solid had dissolved.

Equations for route I:

$$NaOH_{(s)} + aq \xrightarrow{\mathrm{H_1}} NaOH_{(aq,4M)}$$

$$NaOH_{(aq,4M)} + HCl_{(aq,4M)} \xrightarrow{\mathrm{H_2}} H_2O_{(l)} + NaCl_{(aq,2M)}$$

Equations for route II:

$$HCl_{(aq,4M)} + aq \xrightarrow{\mathrm{H_3}} HCl_{(aq,2M)}$$

$$HCl_{(aq,2M)} + NaOH_{(s)} \xrightarrow{\mathrm{H_4}} H_2O_{(l)} + NaCl_{(aq,2M)}$$

Results:

For route I:

Weight of solid sodium hydroxide used: g

Temperature of 25 cm^3 water in the calorimeter = ℃ = K

Temperature of the acid = ℃ = K

Temperature of all the solid NaOH dissolved = ℃ = K

Temperature of the mixture of HCl and NaOH = ℃ = K

For route II:

Weight of solid sodium hydroxide used: g

Temperature of 25 cm^3 water in the calorimeter = ℃ = K

Temperature of the original acid = ℃ = K

Temperature of diluted hydrochloric acid = ℃ = K

Temperature of the mixture of HCl and NaOH = ℃ = K

Chemistry Experiment Laboratory Report (7)

Calculation:

Assume the density of the solution is the same as that of water,
i.e. 1 g cm^{-3}.

For route I,
Concerning H_1,
Heat change =

=

Concerning H_2,
Mass of the solution = (+)×1 = g = kg
Heat given out =
=
=

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(aq)}$$

Number of moles of HCl =

=

Number of moles of NaOH =

=

∴ Number of moles of water formed =

Heat given out per mole of water formed =

=

The enthalpy change of neutralization is kJ mol^{-1}.

For route II,
Concerning H_3,
Heat change =

=

Concerning H_4,
Mass of the solution = (+)×1 = g = kg
Heat given out =
=
=
Number of moles of HCl =

=

Number of moles of NaOH =

=

∴ Number of moles of water formed =

Heat given out per mole of water formed =

=

The enthalpy change of neutralization is kJ mol^{-1}.

$$\begin{tikzcd}
& NaOH_{(s)} + HCl_{(aq,4M)} \arrow[dl, "H_1 + aq"'] \arrow[dr, "+ aq\ H_3"] & \\
NaOH_{(aq,4M)} + HCl_{(aq,4M)} \arrow[dr, "H_2"'] & & HCl_{(aq,2M)} + NaOH_{(s)} \arrow[dl, "H_4"] \\
& NaCl_{(aq,2M)} + H_2O_{(l)} &
\end{tikzcd}$$

Conclusion:

The heat change per mole of sodium chloride formed in route I was found to be

and the heat change per mole of sodium chloride formed in route II was found to be

Sources of errors:

Errors may arise in the experiment including:

1) The specific heat capacity of the polythene bottle was ignored, which should be determined and take into calculation.
2) Heat energy may be loss due to evaporation, convection and conduction. It can be improved by replacing the bottle with a vacuum calorimeter.
3) The thermometer is not precise enough, which can be resolved by using a Beckmann thermometer for higher accuracy in temperature reading.
4) Actually, the specific heat capacity of solution is different from that of water. We should find the specific heat capacity of the solution for accurate calculation.

Chemistry Experiment Laboratory Report (9)

Name: Chik Chun Pong
Date of experiment: 19/11/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 9

Title:

Enthalpy of formation of calcium carbonate

Objective:

To determine the enthalpy of formation of calcium carbonate

Apparatus:

Polystyrene cup, weighing bottles, measuring cylinders, thermometer.

Materials:

Powdered calcium carbonate, 2 M hydrochloric acid, calcium metal.

Theory:

The enthalpies of reaction between calcium and hydrochloric acid, calcium carbonate and hydrochloric acid are determined experimentally. By means of an energy cycle the enthalpy change of formation of calcium carbonate can then be calculated.

Procedure:

Part A Reaction of calcium carbonate with dilute hydrochloric acid

1. About 2 g of dry powdered calcium carbonate were weighed out accurately in a clean dry polystyrene cup.
2. About 50 cm^3 of the hydrochloric acid is poured into a measuring cylinder. Its temperature was recorded every half-minute. At exactly the third minute, the acid was poured on the calcium carbonate in the polystyrene cup.
3. The mixture was stirred gently with the thermometer and the temperature was recorded every half-minute until the seventh minute.
4. The results were tabulated. A graph of temperature of the liquid against time was plotted.

Part B Reaction of calcium with dilute hydrochloric acid

1. About 0.5 g of calcium metal was weighed out accurately.
2. 50 cm^3 of hydrochloric acid was poured into a polystyrene cup with the help of measuring cylinder.

3. The metal was added to the acid at exactly the third minute. The temperature of the liquid was continued to be taken every half-minute until the seventh minute. The liquid was stirred throughout the experiment.
4. The results were tabulated. A graph of temperature of the liquid against time was plotted.

<u>Results:</u>

Part A

Time(min.)	0	0.5	1.0	1.5	2.0	2.5
Temp (℃)	24.0	24.0	24.0	24.5	24.5	24.5

Time(min.)	3.0	3.5	4.0	4.5	5.0	5.5	6.0	6.5	7.0
Temp (℃)	React	39.5	39.0	38.5	38.0	37.5	37.0	36.5	36.0

Part B

Time(min.)	0	0.5	1.0	1.5	2.0	2.5
Temp (℃)	24.5	24.5	24.5	24.5	24.5	24.5

Time(min.)	3.0	3.5	4.0	4.5	5.0	5.5	6.0	6.5	7.0
Temp (℃)	React	53.0	51.0	50.0	49.0	48.0	47.0	46.0	45.0

<u>Answers to the questions:</u>

1) Standard enthalpy change of formation is the enthalpy change of the reaction when one mole of the compound in its standard state is formed from its constituent elements under standard conditions.

2) $Ca_{(s)} + C_{(s)(graphite)} + \frac{3}{2}O_{2(g)} \rightarrow CaCO_{3(s)}$

3) The standard conditions are defined as:
   - (i) elements or compounds in their normal physical state
   - (ii) a pressure of 1 atm (101325 Nm^{-2})
   - (iii) a temperature of 25℃ (298 K)
   - (iv) a concentration of 1 mol dm^{-3} for solutions.

4) $CO_3^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

5) It is because the formation of products is not only determined by the concentration of the acid; it is also determined by the mass of

calcium or calcium carbonate. Also the hydrochloric acid is in excess in this experiment.

6) $Ca_{(s)} + 2H^{+}{}_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + H_{2(g)}$

7)

$$\begin{tikzcd}
Ca_{(s)} + C_{(s)(graphite)} + \frac{3}{2}O_{2(g)} + 2HCl_{(aq)} \arrow[r, "\Delta H_f^{\theta}[CaCO_{3(s)}]"'] \arrow[d, "\Delta H_1"'] & CaCO_{3(s)} + 2HCl_{(aq)} \arrow[dddl, "\Delta H_2"] \\
CaCl_{2(aq)} + H_{2(g)} + C_{(s)(graphite)} + \frac{3}{2}O_{2(g)} \arrow[d, "\Delta H_f^{\theta}[CO_{2(g)}]"'] & \\
CaCl_{2(s)} + CO_{2(g)} + \frac{1}{2}O_{2(g)} + H_{2(g)} \arrow[d, "\Delta H_f^{\theta}[H_2O_{(g)}]"'] & \\
CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)} &
\end{tikzcd}$$

8) Given $\Delta H_f^{\theta}[CO_{2(g)}]$ = -393 kJ mol^{-1} and

$\Delta H_f^{\theta}[H_2O_{(g)}]$ = -286 kJ mol^{-1}

In △H_1,

Heat given out = mc△T

= (50 g) (4.2 J g^{-1} ℃$^{-1}$) (32.5 ℃)

= 6825 J

= 6.825 kJ

Heat given out per mole of $CaCl_2$

= 6.825 kJ ÷ $\frac{0.5}{40}$ mol

= 546.00 kJ mol^{-1}

∴ Enthalpy change in △H_1 is -546.00 kJ mol^{-1}.

In △H_2,

Heat given out = mc△T

= (50 g) (4.2 J g^{-1} ℃$^{-1}$) (16.5 ℃)

= 3465J

= 3.465 kJ

Heat given out per mole of $CaCl_{2\ (s)}$

$= 3.465 \text{ kJ} \div \frac{2}{(40+12+16\times 3)}$ mol

$= 173.25 \text{ kJ mol}^{-1}$

∴ Enthalpy change in $\triangle H_2$ is $-173.25 \text{ kJ mol}^{-1}$.

From Hess's Law,

$$\Delta H_f^\theta[CaCO_{3(s)}] = \Delta H_f^\theta[CO_{2(g)}] + \Delta H_f^\theta[H_2O_{(g)}] + \Delta H_1 - \Delta H_2$$

$$= (-286) + (-393) + (-546.00) – (-173.25)$$

$$= -1051.75 \text{kJ mol}^{-1}$$

∴ Enthalpy change of formation of $CaCO_{3(s)}$ is $-1051.75 \text{kJ mol}^{-1}$.

9) Hess's Law.

10) We assume there is no heat loss to the environment, and all the calcium metal or calcium carbonate would react with the hydrochloric acid.

<u>Conclusion:</u>

From the experiment, the enthalpy change of formation of solid calcium carbonate was found to be -1051.75 kilojoules per mole.

<u>Sources of errors and improvements:</u>

Errors may arise in the experiment including:

1) Errors in weighting: It can be improved by using an electronic balance of higher precision.
2) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.
3) Solution is not homogenous: This can be minimized by shaking the solution very well.
4) Heat energy may be loss due to evaporation, convection and conduction. It can be improved by replacing the bottle with a vacuum calorimeter.
5) The thermometer is not precise enough, which can be resolved by using a Beckmann thermometer for higher accuracy in temperature reading.
6) The specific heat capacity of the polythene bottle was ignored, which should be determined and take into calculation.

Actually, the specific heat capacity of solution is different from that of water. We should find the specific heat capacity of the solution for accurate calculation.

-End-

Chemistry Experiment Laboratory Report (10)

Name: Chik Chun Pong Date of experiment: 25/11/03
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 10

Title:

To determine the mass of ascorbic acid in commercial vitamin C tablets.

Apparatus:

Electronic balance, weighing bottles, 250cm^3 volumetric flask, conical flask, measuring cylinders, 25 cm^3 pipette, burette, stand and clamp.

Materials:

Potassium iodate (V), distilled water, 1M potassium iodide, 0.5M sulphuric acid, thiosulphate solution, starch solution, vitamin C tablet.

Theory:

Vitamin C is ascorbic acid which is rapidly and quantitatively oxidized by iodine in acid solution according to the following equation:

The low solubility of iodine makes the direct titration rather difficult. The difficulty may be avoided by the in situ generation of a known excess quantity of iodine by reaction between iodine and iodate. The excess iodine is then back titrated with standard sodium thiosulphate solution.

Procedure:

Part I To standardize the sodium thiosulphate solution provided.

1. 0.6-0.7 g of potassium iodate(V) was weighed out accurately.
2. The solid was dissolved in distilled water and the solution was made up to 250cm^3 in a volumetric flask.
3. 25cm^3 of the solution was pipetted into a conical flask.5 cm^3 of 1M potassium iodide solution was added to it followed by 10 cm^3 of 0.5M sulphuric acid. The solution was titrated immediately with sodium thiosulphate solution using starch solution as indicator. The titration was repeated once.
4. The remaining solution of potassium iodate(V) was

transferred into a beaker and kept for next step.

Part II To analyze the ascorbic acid content

1. The vitamin C tablet provided was dissolved in about 150 cm^3 of 0.5M sulphuric acid. The resulting solution was transferred to a clean volumetric flask and make up to 250 cm^3 using distilled water.
2. 25cm^3 of the vitamin C solution was pipetted into a conical flask and 5cm^3 of 1M potassium iodide solution was added to it.
3. 25cm^3 of the potassium iodate(V) solution was pipetted into the above mixture. The excess iodine was immediately titrated with the standard sodium thiosulphate solution. The titration was repeated once.
4. The mass of ascorbic acid in the tablet was calculated.

Results:

Weight of potassium iodate (V) solid used: 0.68 g

Part I titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	4.50	8.80	9.20
Final volume (cm^3)	35.70	39.60	40.00
Volume of $Na_2S_2O_3$ used (cm^3)	31.20	30.80	30.80
Mean volume (cm^3)	30.93		

Part II titration results:

	Trial 1	Trial 2	Trial 3
Initial volume (cm^3)	9.50	22.50	34.80
Final volume (cm^3)	22.00	34.80	47.50
Volume of $Na_2S_2O_3$ used (cm^3)	12.50	12.30	12.70
Mean volume (cm^3)	12.50		

Calculation:

From part I,

- Molarity of potassium iodate(V) solution
  =[0.68/(39.1+126.9+16x3)]/0.25dm^3
  =0.01271M

From the equation:$IO_3^-(aq) + 5I^-(aq) + 6H^+ (aq) \rightarrow 3I_2 (aq) + 3 H_20(l)$

- Number of moles of I_2 : Numbers of moles of IO_3^- = 1:1
- So number of moles of I_2 = 0.01271M x 0.025 dm^3
  = 0.0003 mol.

From the equation: I_2 (aq) + 2 $S_2O_3^{2-}$ (aq) → $S_4O_6^{2-}$ (aq) + 2I- (aq)

- Number of moles of I_2 : Number of moles of $S_2O_3^{2-}$ (aq) = 1:2
- Number of moles of $S_2O_3^{2-}$ (aq) ions added = 0.000317 mol x 2
  = 0.000635 mol
- Concentration of sodium thiosulphate solution
  =0.000635/ 0.03093 dm^3
  =0.0205M

From part II,

- Numbers of moles of excess iodine : number of moles of $S_2O_3^{2-}$ (aq) ions =1:2
- Numbers of moles of excess iodine = 0.01250 dm^3 x 0.0205M/2
  =0.0001284 mol
- Numbers of moles of iodine initially : number of moles of IO_3^- (aq) ions =3:1
- Numbers of moles of iodine initially = 0.01271M x 0.025 dm^3 x 3
  = 0.00095325 mol
- Thus, number of moles of iodine used = 0.000953 – 0.000128 mol
  = 0.000825 mol
- Number of moles of ascorbic acid : number of moles of iodine = 1:1
- Number of moles of ascorbic acid = 0.000825 mol
- Thus, number of moles of ascorbic acid present in each tablet
  = 0.000825 mol x 10
  = 0.00825 mol
- Mass of ascorbic acid in the tablet = 0.00825 mol x (12x6+6x16+8)
  = 1.45g

Conclusion:

From the experiment, the mass of ascorbic acid in each tablet is 1.45 g.

Sources of errors and improvements:

Errors may arise in the experiment including:

1) Errors in judging the end point: It can be improved by taking the average of results of more titrations or use pH meter instead of indicator in titration.
2) Errors in weighting: It can be improved by using an electronic balance of higher precision.
3) Impurities of materials: Moisture can be minimized by putting them

into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.

4) Solution is not homogenous: This can be minimized by shaking the solution very well.

Precautions:

- ✓ The ascorbic acid is constantly oxidized by air throughout the experiment, so we should perform the experiment quickly to minimize the number of moles of ascorbic acid being oxidized by air.

-End-

Chemistry Experiment Laboratory Report (12)

Name: Chik Chun Pong
Date of experiment: 17/02/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 12

Title:

To study the rate of the reaction between sodium thiosulphate and hydrochloric acid.

Apparatus:

600 cm^3 beaker, conical flask, white tile, mercury thermometer, 3 measuring cylinders.

Materials:

2 M hydrochloric acid, sodium thiosulphate solution.

Theory:

As $Average\ \ rate\ \ \alpha\ \ \frac{1}{Time\ \ taken\ \ for\ the\ \ corss\ \ to\ \ disappear}$

By finding the time taken for the cross to disappear, we can compare the average rate of reactions we want to study.

Procedure:

Part I Change in concentration

1. 20 cm^3 of water was added to 30 cm^3 of sodium thiosulphate solution (40 g dm^{-3}) in a conical flask. The mixture was mixed well.
2. A cross was drawn with a pencil on a piece of paper.

3. 5 cm^3 of hydrochloric acid (2 M) was poured to the above solution and the stop-watch was started at the same time.
4. The cross was looked down through the flask. As soon as the reaction had produced enough precipitate to obscure the cross, the watch was stop and the time was noted.
5. The procedure was repeated for other volumes of sodium thiosulphate, keeping the total volume constant by adding

suitable volume of water. A graph of $\frac{1}{time}$ against the volume of sodium thiosulphate was plotted.

Part II Change in temperature

1. 10 cm^3 of sodium thiosulphate solution and 40 cm^3 of water were added to a conical flask. The flask was placed in a thermo-bath at 30^{o}c.
2. 5 cm^3 of hydrochloric acid was placed in a test-tube and was immersed in the same thermo-bath.
3. The experiment in part I was carried out using the above reactants. The time taken was noted.
4. The experiment was repeated using the same reactants but at different temperatures (40^{o}c, 50^{o}c, 60^{o}c). A graph of $\frac{1}{time}$ against the temperature was plotted.

Results:

Part I

Volume (cm^3)	1	2	3	4	5
$Na_2S_2O_3$	10	20	30	40	50
H_2O	40	30	20	10	0
HCl	5	5	5	5	5
Time (s)	158.73	59.67	39.95	27.87	20.29
$\frac{1}{time}$ (s^{-1})	0.0063	0.0167	0.0250	0.0359	0.0493

Part II

	1	2	3	4	5
Temperature (^{o}c)	22.5 (R.T.)	30	40	50	60
Time (s)	125.22	62.45	40.38	29.89	17.97
$\frac{1}{time}$ (s^{-1})	0.0080	0.0160	0.0248	0.0335	0.0556

* R.T. = Room Temperature

Discussion and Conclusion:

From the graphs, the average reaction rate was found to be directly proportional to the concentration of reactants and temperature.

In part I, the (1/time) showed a positive linear relationship with the volume of sodium thiosulphate. The higher the concentration of the sodium thiosulphate, the higher the chance of reactants colliding with each other and react, the shorter the time needed for the reactant to carry out. A linear relationship of the line means the reaction is a first order reaction and it passes through the origin.

In part II, the (1/time) showed a positive linear relationship with the temperature. The higher the temperature, the more the kinetic energy the molecules and improve their chance of collision, thus speeding up the reaction.

Answers to the questions:

1) We have to provide the same surface area for the hydrochloric acid to act on and allow the precipitate formed in the reaction to diffuse around the same volume.
2) The concentration of sodium thiosulphate.
3) If there is an increase in concentration of thiosulphate, the reaction rate would also increase.

Sources of errors and improvements:

Errors may arise in the experiment including:

1) Errors in judging the end point by our eyes: It can be improved by using a calorimeter to provide a more reliable result.
2) Errors in the temperature of the solutions: It can be improved by putting the conical flask and the test tube in a thermostatic water bath machine for a longer period of time to have a more stable temperature of the reactants.
3) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized

by the use of colored glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.

4) Solution is not homogenous: This can be minimized by shaking the solution very well.
5) The thermometer is not precise enough, which can be resolved by using a Beckmann thermometer for higher accuracy in temperature reading.
6) Measurement of volume not accurate enough: It can be improved by using pipettes to transfer liquids instead of using measuring cylinders.

Precautions:

- ✓ Shake the reaction mixture with same degree since it can speed up the rate of reaction.

-End-

Name: Chik Chun Pong
Date of experiment: 24/02/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 13

Title:

To investigate the order of the reaction of iodine with acetone.

Apparatus:

100 cm^3 conical flask, 150 cm^3 conical flask, white tile, clamp and stand, 10 ml pipette, pipette filler, 10 ml measuring cylinder, burette, plastic funnel, stop watch.

Materials:

1 M acetone, 0.02 M iodine solution (in KI), sodium thiosulphate solution, 1 M sulphuric acid, 0.5 M sodium hydrogen carbonate solution, starch indicator.

Theory:

The volume of sodium thiosulphate used to titrate against iodine in the mixture at different time intervals was plotted and the slope was found. By varying the concentration of acetone, the slopes against concentration can be plotted on a graph and the order of reaction can be found.

Procedure:

1. 10.0 cm^3 of 1 M acetone, 15.0 cm^3 of water and 25.0 cm^3 of 1 M sulphuric acid were mixed in a flask.
2. 50.0 cm^3 of 0.02 M iodine solution (in KI) was added to the above mixture and at once start the clock. It was stirred to mix well.
3. After about 6 minutes, 10 cm^3 of the reaction mixture was transferred into a conical flask containing 10 cm^3 of 0.5 M sodium hydrogen carbonate solution using a pipette. The time of addition was noted.
4. The contents of the flask was mixed thoroughly and titrated with sodium thiosulphate solution using starch solution as indicator.
5. The same procedure was repeated after every 6 minutes. 5 or 6 data was taken.
6. A graph of the volume of sodium thiosulphate needed to react with

remaining iodine against the time at which the 10.0 cm^3 sample of the reaction mixture was added to the sodium hydrogen carbonate was plotted. The slope of the graph was shared with other students doing the same experiment using different concentrations of acetone.

Results:

The relation between the time of addition of reaction mixture to conical flask of Na_2CO_3 and the volume of $Na_2S_2O_3$ used:

Time when reaction mixture was added to Na_2CO_3 (min)	6’’40	13’’05	19’’05	25’’30	32’’10	38’’10	42”28	48’’28
Final burette reading (cm^3)	19.50	19.25	16.50	15.80	23.30	34.90	11.70	21.80
Initial burette reading (cm^3)	1.70	2.25	0.00	0.00	9.65	22.30	0.30	11.80
Volume of $Na_2S_2O_3$ used (cm^3)	17.80	17.00	16.50	15.80	13.65	12.60	11.40	10.00

The slopes of graphs of other students using different concentrations of acetone:

Student Pair	1	2	3	4	5	6	7
Flask A (cm^3) Vol. of I_2	50.0	50.0	50.0	50.0	50.0	50.0	50.0
Flask B (cm^3) Vol. of H_2SO_4	25.0	25.0	25.0	25.0	25.0	25.0	25.0
Vol. of acetone (cm^3)	25.0	22.5	20.0	15.0	12.5	10.0	6.25
Vol. of water (cm^3)	0.0	2.5	5.0	10.0	12.5	15.0	18.75
Mean of slopes of graphs	-0.230	-0.260	-0.190	-0.130	-0.117	-0.095	-0.067

Answers to the Questions:

1) The y-intercept of graph = 19.4

   $\therefore$ 19.4 cm^3 of $Na_2S_2O_3$ can react with 50.0 cm^3 of 0.02 M of iodine.

   According to this equation:

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

   Number of moles of $I_2 = \frac{50.0}{1000} \times 0.02$ = 0.001 mol

   Number of moles of $Na_2S_2O_3$ = 0.002 mol

   Molarity of $Na_2S_2O_3$ = $0.002 \div \frac{19.4}{1000}$ = $\underline{\underline{0.103\text{ M}}}$

2) Sodium hydrogen carbonate was used to quench the sample by reacting with the dilute sulphuric acid, the catalyst, inside it. It also prevents the sulphuric acid from reacting with sodium thiosulphate to form side products.
3) The concentration of iodine is decreasing gradually throughout the experiment.
4) The slope of graph = $\frac{12.5-19.4}{38-0}$ = -0.18
5) The rate of reaction does not vary with different concentrations of iodine.
6) As the rate of reaction does not vary with different concentrations of iodine, the order of reaction with respect to iodine is zero.
7) As iodine is of zeroth order, it does not it does not involved in the rate determining step.
8) From the graph, the higher the concentration of propanone, the higher the rate of the disappearance of iodine.

   As there is a directly proportional relationship between the concentration of propanone and the rate of reaction, the order of reaction with respect to propanone is one.
9) The slope of graph = $\frac{(-0.23)-(-0.075)}{25-7.5}$ = -0.00886

   Since the slope of graph is the relationship between the disappearance of iodine and the concentration of propanone, the rate has a negative sign.

∴ The rate constant is $-(-0.00886s^{-1}) = 0.00886s^{-1}$.

∴ The rate equation is:

$$\text{Rate} = 0.00886[I_2]$$

Discussion and Conclusion:

The order of reaction of iodine and acetone in this reaction were found to be zeroth and first respectively.

Sources of errors and improvements:

1) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.
2) Solution mixture is not homogenous: This can be minimized by shaking the solution very well.
3) Inaccuracy in judging the end-point by using our eyes: It can be improved by taking average of 3 close titration results.
4) Errors in measuring the volumes of liquids: This can be improved by using pipettes to transfer the liquids.

Precautions:

1) Starch indicator should be only added just before the end-point avoiding the combination of starch and iodine molecules.
2) The temperature and the pressure of the surroundings should be kept constant avoiding the increase or decrease in reaction rate.

-End-

Name: Chik Chun Pong Date of experiment: 02/03/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 14

Title:

To determine the order of reaction of the oxidation of an acidified solution of an Iodine by Hydrogen Peroxide (‘Iodine Clock’ Experiment).

Apparatus:

Two 100 cm^3 beakers, 50 ml measuring cylinder, 10 ml measuring cylinder, stop watch, a magnetic box with a/c supply for stirring purpose.

Materials:

0.2 M potassium iodide solution, 0.5 M sulphuric acid, 0.1M sodium thiosulphate solution, 0.2 M hydrogen peroxide solution, starch indicator.

Theory:

An acidified solution of an iodide is oxidized by hydrogen peroxide according to the following equation:

$$H_2O_2\ (aq) + 2I^-(aq) + 2H^+\ (aq) \rightarrow I_2\ (aq) + 2H_2O\ (aq)$$

The rate law is given by

$$\text{rate} = k[H_2O_2]^x\ [I^-]^y\ [H^+]^z$$

where x, y and z are constants. To determine the order of this reaction the initial rate method is applied. The experiment is often known as the ‘Iodine Clock’ experiment.

Procedure:

1. 30 cm^3 of 0.2M hydrogen peroxide solution, 30 cm^3 of 0.5M sulphuric acid and 10 cm^3 of 1% starch solution were mixed in beaker 1. The stirrer was set in motion.
2. 30 cm^3 of 0.2M potassium iodide solution, 10 cm^3 of 0.1M sodium thiosulphate solution and 30 cm^3 of water were mixed in beaker 2. It was stirred to mix well.
3. The contents in beaker 2 were poured into beaker 1 in one portion and at the same time the stop-watch was started. The time for the appearance of a blue color was recorded.

4. The same experiment using different sets of concentrations of reactants were repeated according to the following table:

Results:

Volume in cm^3								
	0.2M KI	0.5M H_2SO_4	0.2M H_2O_2	H_2O	0.1M $Na_2S_2O_3$	1% Starch	Time/ s	$Time^{-1}$ $/s^{-1}$
A	30	30	30	30	10	10	73	0.0137
B	20	30	30	40	10	10	98	0.0102
C	10	30	30	50	10	10	172	0.00581
D	30	20	30	40	10	10	105	0.00952
E	30	10	30	50	10	10	117	0.00855
F	30	30	20	40	10	10	102	0.00980
G	30	30	10	50	10	10	232	0.00431

Calculation:

$$\because \text{ Rate} = k\,[H_2O_2]^x\,[I_-]^y\,[H^+]^z$$

$$\therefore 0.0137 = k\,[30]^x\,[30]^y\,[30]^z/\text{vol.} \text{ --- (1)}$$

$$0.0102 = k\,[30]^x\,[20]^y\,[30]^z/\text{vol.} \text{ --- (2)}$$

$$0.0058 = k\,[30]^x\,[10]^y\,[30]^z/\text{vol.} \text{ --- (3)}$$

$$0.0105 = k\,[30]^x\,[30]^y\,[20]^z/\text{vol.} \text{ --- (4)}$$

$$0.0085 = k\,[30]^x\,[30]^y\,[10]^z/\text{vol.} \text{ --- (5)}$$

$$0.0098 = k\,[20]^x\,[30]^y\,[30]^z/\text{vol.} \text{ --- (6)}$$

$$0.0043 = k\,[10]^x\,[30]^y\,[30]^z/\text{vol.} \text{ --- (7)}$$

$$(6)\div(7),\quad \frac{0.0098}{0.00431} = \left(\frac{20}{10}\right)^x$$

$$x=1.185$$

$$(2)\div(3),\quad \frac{0.0102}{0.00581} = \left(\frac{20}{10}\right)^y$$

$$y=0.812$$

$$(1)\div(5),\quad \frac{0.0137}{0.00855} = \left(\frac{30}{10}\right)^z$$

$$z=0.429$$

Discussion and Conclusion:

The order of reaction of oxidation of potassium iodide by hydrogen peroxide in an acidic medium x,y,z were found to be 1.185,0.812,0429 with respect to $[H^+]$,$[I_-]$ and $[H_2O_2]$ respectively.

Sources of errors and improvements:

1) Impurities of materials: Moisture can be minimized by putting them into a desiccator. The action of light, carbon dioxide, dust particles or oxygen on standard solution may raise errors. This can be minimized by the use of coloured glass bottles, tightly fitting stoppers, soda-lime tubes to absorb carbon dioxide.
2) Solution mixture is not homogenous: This can be minimized by shaking the solution very well.
3) Inaccuracy in judging the end-point by using our eyes: It can be improved by using a calorimeter to provide a more reliable result.
4) Errors in measuring the volumes of liquids: This can be improved by using pipettes to transfer the liquids.
5) Evaporation of solutions causing an increase in molarities, it can be improved by covering the container with a lid to prevent evaporation.
6) Changing of temperature and pressure causing inaccuracy in judgement of rate. This can be minimized by carrying the experiment at a thermo bath with constant temperature.

Precautions:

1) Starch indicator should be only added just before the end-point avoiding the combination of starch and iodine molecules.
2) The temperature and the pressure of the surroundings should be kept constant avoiding the increase or decrease in reaction rate.
3) Start the stop watch when reactants are just in contact.
4) Pour the contents in beaker 2 into beaker 1 in one portion.
5) Transfer the reactants back and fro between beaker 1 and beaker 2 so as to prevent reactants being left in the beaker.

-End-

Chemistry Experiment Laboratory Report (15)

Name: Chik Chun Pong Date of experiment: 09/03/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 15

Title:

To determine the activation energy of the reaction between Br^- and BrO_3^- in acidic solution.

Apparatus:

Boiling tubes (x2), beaker, Bunsen burner, wire gauze, tripod, heat proof mat, thermometer, burette, stop watch.

Materials:

0.01M phenol solution, 0.083M KBr/0.017M $KBrO_3$ solution, 0.5M sulphuric acid, methyl red indicator.

Theory:

The reaction can be represented by
$5Br^-_{(aq)} + BrO_3^-{}_{(aq)} + 6H^+_{(aq)} \rightarrow 3Br_{2(aq)} + 3H_2O_{(l)}$

The progress of the reaction may be followed by adding a fixed amount of phenol together with some methyl red indicator. The bromine produced during the reaction reacts very quickly with phenol. Once all the phenol is consumed, any further bromine bleaches the indicator immediately. So the time for the reaction to proceed to a given point may be determined.

Procedure:

1. 10 cm^3 of 0.01M phenol solution, 10 cm^3 0.083M KBr/0.017M $KBrO_3$ solution and 10drops of methyl red indicator was placed into the first boiling tube.
2. Then 5 cm^3 of 0.5M sulphuric acid was placed into the second tube.
3. The both tubes were placed in a large beaker of water which was maintained between 49-51℃ by careful warming.
4. The first tube was clamped suitably and a thermometer was placed in this solution.

5. The contents of the tubes were allowed to reach the temperature of the water bath and waited until the temperature reached constant.
6. The entire 5 cm^3 of sulphuric acid was poured into the clamped boiling tube and was whirled gently.
7. The clamped should be remained in water throughout the experiment.
8. The time was recorded from the moment of mixing to that of complete disappearance of the red colour.
9. The temperature of the content of the clamped tube at the end of the experiment was recorded also.
10. The experiment was repeated at about 45, 40, 35, 25,10 and 5℃.

Results and calculation:

Set	Temperature/ºC	Temperature/ºK	Time/s	$\ln \frac{1}{t}$	$\frac{1}{T}$ $(\times 10^{-3})$
1	50.5	323.5	27	-3.296	3.09
2	46.5	319.5	37	-3.611	3.13
3	51.5	314.5	54	-3.989	3.18
4	37.0	310.0	85	-4.443	3.23
5	32.5	305.5	120	-4.844	3.27

$\ln \frac{1}{t}$ against $\frac{1}{T}$ was plotted (See graph 1).

Answer to the questions:

1. The first boiling tube contains $Br^-_{(aq)}$ and $BrO_3^-{}_{(aq)}$ while the second boiling tube contains $H^+_{(aq)}$.Both $Br^-_{(aq)}$ and $BrO_3^-{}_{(aq)}$ requires the reducing action from $H^+_{(aq)}$ for the reaction to take place according to the equation.
2. The product of the equation, $Br_{2(aq)}$, reacts with phenol rapidly. However when phenol is used up, the excess Br_2 would bleach the methyl red into colourless. So, the methyl red is used to determine the end point of the reaction between phenol and Br_2 , not the usual use of determinating pH.
3. Reaction time is so short that there is great inaccuracy in time measurement.

4. We only want to take a reference point of concentration to compare the rate of reaction in different set. As $rate = \frac{change\_in\_concentratrion}{time}$, the progress of the reaction is not important.
5. K is the rate constant of the reaction and ΔE is the activation energy.
6. Since k $\propto$ rate and $\frac{1}{t}$ $\propto$ rate, so k $\propto \frac{1}{t}$, i.e. the higher the k, the higher the $\frac{1}{t}$ and vice versa. $\frac{1}{t}$ can't replace k as k $\propto \frac{1}{t}$ doesn't mean k=$\frac{1}{t}$.

7. An equation relating 'k' and '1/T' can be derived as,

$$\ln k - \ln A = -\frac{\Delta E}{RT}$$

As k $\propto \frac{1}{t}$, $\ln c(\frac{1}{t}) = -\frac{\Delta E}{RT} + \ln A$, where c is a constant.

$$\rightarrow \ln \frac{1}{t} = \left(-\frac{\Delta E}{RT}\right)\left(\frac{1}{T}\right) + (\ln A - \ln c) \text{ --- } (*)$$

$$\rightarrow \frac{\ln \frac{1}{t}}{\frac{1}{T}} = -\frac{\Delta E}{R}$$, by observing the format y=mx+c and $\frac{y}{x}$=m

8. The equation of the graph plotted is (*) with slope equals to $-\frac{\Delta E}{R}$.

Slope of the graph : $\frac{-3.296-(-4.844)}{0.00309-0.00327} = -8600$

$$-8600 = -\frac{\Delta E}{R}$$

$$\Delta E = 71.5 \text{ kJ mol}^{-1}$$

9. The sign of ΔE is positive.
As activation energy of a reaction indicates the amount of energy required by the reactants to start reacting and every reaction requires such energy supplied to start, the thermodynamic sign of activation energy must be positive. This signifies that energy must be supplied in starting every reaction.
10. The greater the temperature, the higher the reaction rate. This is because the increase in temperature increases the proportion of the fast moving election in the molecules of reactants. Thus more

particles have energy greater than the activation energy which gives rise to the rate of reaction. Besides that, increase in temperature increase the most probable speed of the particles of the reactants. This increases the frequency of collision between particles and thus increases the rate of reaction.

11. For a reaction with low activation energy, a smaller amount of energy is required to start the reaction, in compare to a reaction with high activation energy. So a reaction with lower activation energy can get the enough amount of energy to start the reaction faster and earlier than a reaction with higher activation energy. Thus it can be concluded that a reaction with low activation energy proceed faster.

Conclusion:

The activation energy of the reaction between Bromide ion and Bromate (V) ion in acid solution is 71.5kJ mol^{-1} .

Sources of errors and improvements:

1. There was an inaccuracy of the measuring instrument (stop watch and measuring cylinder). This can be minimized by using apparatus of higher accuracy.
2. The determination of decoluorization was inaccurate since the colour change is a gradual process; judgement of end point was difficult. Reaction time delay was also involved when naked eye was used to determine colour change. This can be improved by using a calorimeter which would stop the timer when permittivity drops below a pre-set value.

Precautions:

1. During sucking back of methyl orange, the dropper should not be put deeply inside the methyl orange solution, as the precipitate is red in colour and will not change colour during reaction.
2. Do not allow the boiling tube to hit the wall of the beaker while swirling or otherwise the boiling tube will break.
3. Keep the degree of stirring the same in every set to minimize its effect on the rate of reaction.
4. Pour the mixture back and fro in the first and second boiling tube to allow all reactants in each boiling tube to dissolve in the mixture.

-End-

Chemistry Experiment Laboratory Report (16)

Name: Chik Chun Pong Date of experiment: 16/03/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 16

Title:

To determine the equilibrium constant for esterification.

Apparatus:

Round-bottomed flask, pipette, pipette filter, conical flask, burette, white tile, stand and clamp, funnel

Materials:

Glacial ethanoic acid, propan-1-ol, deionized water, phenolphthalein , sodium hydroxide solution, concentrated sulphuric acid.

Theory:

Equilibrium in a reaction is a state where concentrations of the reactants and products remain more or less constant. In this experiment, the equilibrium constant for esterification was studied. By titrating the mixture with NaOH at the beginning of the experiment, the volume of NaOH needed to neutralize the conc. H_2SO_4 was found. When equilibrium is reached, the mixture was titrated with NaOH. Thus, the concentration of CH_3COOH, $CH_3CH_2CH_2OH$, $CH_3COOCH_2CH_2CH_3$ and H_2O could be calculated.

As equilibrium constant = $\frac{[product]eqm}{[reac\tan t]eqm}$, the value Kc could be found.

$$CH_3COOH + CH_3CH_2CH_2OH \Leftrightarrow CH_3COOCH_2CH_2CH_3 + H_2O$$

Procedure:

1. 0.25 mole (equals to 14.3 cm^3) of glacial ethanoic acid (density = 1.05 g cm^{-3}) and 0.25 mole (equals to 18.8 cm^3) of propan-1-ol (density = 0.8 g cm^{-3}) were put into a round-bottomed flask. It was mixed thoroughly.
2. 1.0 cm^3 of the mixture was transferred by pipette to a 150 cm^3

conical flask containing about 25cm^3 deionized water and 2 drops of phenolphthalein indicator. It was titrated to end point with 0.50 M sodium hydroxide solution. The titre (V_1 cm^3)was recorded.

3. 8 drops of concentrated sulphuric(VI) acid was added to the remainder of the acid-alcohol solutions while continuously swirling the flask. Another 1.0 cm^3 sample was titrated immediately. The titre (V_2 cm^3) was recorded. The volume to be subtracted from subsequent titrations to correct for the amount of sulphuric (VI) acid present was represented by the difference between V_1 & V_2.
4. The flask was stoppered and left for 2 or 3 days with occasional swirling.
5. 1.0 cm^3 sample from the flask was removed for titration with the 0.5M sodium hydroxide as before. The titre needed was recorded and corrected for the sulphuric acid. (V_3 cm^3)

Results:

	Initial reading	Final reading	Volume of NaOH needed (cm^3)
V_1	17.10	32.20	15.10
V_2	7.20	22.45	15.25
V_3	22.20	27.40	5.20

Calculation:

- Volume of NaOH required to neutralize the conc.H_2SO_4=V_2-V_1= 0.15
- At equilibrium, volume of NaOH needed to neutralize the ethanoic acid = V3 – 0.15 = 5.05
- No. of moles of NaOH needed to neutralize ethanoic acid = 5.05/1000×0.5 = 0.002525 mol.

The equation of the reaction between ethanoic acid and sodium hydroxide:

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

- No. of moles of CH_3COOH in the pipetted 1 cm^3 = 0.002525 mol.
- Concentration of CH_3COOH = 0.002525×1000= 2.525M

- As the concentration of CH_3COOH ∶ concentration of $CH_3CH_2CH_2OH$ =1 ∶ 1
- Concentration of $CH_3CH_2CH_2OH$ = 2.525 M

- Concentration of CH_3COOH at the beginning (after procedure 1) is:
  $= 15.1\ cm^3 \times 10^{-3} \times 0.5M / 10^{-3}\ cm^3$
  $= 7.55$ M
- Concentration of $CH_3COOCH_2CH_2CH_3$ is:
  $= 7.55$ M - 2.525 M
  = <u>5.025M</u>
- Concentration of water is:
  = <u>5.025M</u>

As $K_C=([CH_3COOCH_2CH_2CH_3]\times[H_2O])/([CH_3COOH]\times [CH_3CH_2CH_2OH])$

$K_C = 5.025 \times 5.025 / (2.525 \times 2.525) = \underline{3.96}$

Conclusion and Discussion:

The concentration of ethanoic acid remained at equilibrium was found to be 2.525M; the concentration of propan-1-ol was 2.525M as well while the concentrations of water and propyl ethanoate were found to be 5.025M. The equilibrium constant for esterification of ethanoic acid and propan-1-ol was found to be 3.96.

In this experiment, concentrated sulphuric acid was used in the esterification. Its usage aims at bringing a clearer picture to the esterification reaction.

Firstly, it acts as a catalyst to speed up the reaction. Also, with its dehydrating properties, it acts as a dehydrating agent too.

Furthermore, it can increase the yield of the ester. It is used also to supply proton to the alkanoic acid during the process of protonation.

Sources of errors and improvements:

- Impurities of materials in the solutions affecting the pH value. This can be minimized by rinsing the containers thoroughly to remove impurities can covering the beakers to prevent impurities in air from reaching.
- Inaccuracy in judging the end-point and of the instruments. This can be improved by using pH meters, which are more accurate in measuring pH value, can be introduced to find out the end-point.
- Evaporation of the solutions causing the increase in molarities. This can be solved by covering the beakers with a lid to prevent

evaporation.

- Increase in volume caused by adding concentrated sulphuric acid. It can be eliminated by measuring the volume of mixture after titration.

Precautions:

- The burette should be rinsed with distilled water to remove impurities and then the sodium hydroxide solution so as to minimize the dilution effect.
- A white tile should always be put under the conical flask when doing titration so that any colour change can be easily observed.
- The mixture must be titrated immediately after adding concentrated sulphuric acid so as to prevent the reaction to go forward which will affect the volume of CH_3COOH & NaOH needed for titration.

Assumption:

1. It is assumed that the volume of 8 drops of concentrated sulphuric(VI) acid is very small which could not greatly affected the volume of the original mixture.

-End-

Chemistry Experiment Laboratory Report (17)

Name: Chik Chun Pong Date of experiment: 23/03/04
Class: 6S2
Class number: 8

Chemistry Experiment Laboratory Report 17

Title:

Investigation of the Oxidation Numbers of Vanadium

Apparatus:

Bunsen burner, test tubes, beakers, test tube holder, test tube rack

Materials:

Solutions of vanadium (V), iron(II), iron(III), iodide, bromide, iodine, solid zinc dust, solid tin pieces, SO_2 in water, KCNS, $Na_2S_2O_3$, starch , 1,1,1-trichloronethane.

Theory:

Every equilibrium system has its own electrode potential for redox reaction. Based on the standard hydrogen electrode, the E.C.S. was set up to slow the relative potentials of all common equilibrium systems, allowing prediction for feasibility of redox reaction. If the subtraction of the potentials gives a positive value, it indicates the reaction is capable of going from left to right and that the reaction is feasible. In this experiment, results of the reactions were predicted first and then the experiments were carried out to verify the predictions. The reactions were observed by the change in colours where the colours of V (II) ions is violet, V (III) pale green, V (IV) blue and V (V) yellow.

Procedure:

1. For the following 5 trials, the results were predicted and experiments were carried out, (i)V(V)+Fe(II), (ii)V(V)+I(-I), (iii)V(V)+S(IV), (iv)V(V)+Sn(0), (v)V(V)+Zn(0).
2. The best reactions which produce V(IV), V(III) and V(II) ions were chosen.
3. The following reactions were predicted and carried out by mixing (i) to (v),(vi) V(IV)+V(II), (vii) V(V)+V(III), (viii) V(IV)+V(III), (ix) V(IV)+Fe(III), (x) V(III)+Fe(III), (xi) V(IV)+I(-I), (xii) V(V)+Br(-I), (xiii) V(III)+I(0), (xiv) V(IV)+I(0).
4. The predictions were verified.

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Ion (hydrated)	**VO_3^- / VO_2^+**	**VO^{2+}**	**V^{3+}**	**V^{2+}**
Colour	yellow	blue	green	violet
Oxidation State	+5	+4	+3	+2

Results:

Reaction (i):

1. VO_2^+ (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) +VO^{2+} (aq) +1.0V
2. Fe^{3+} (aq) + e^- ⇔ Fe^{2+} (aq) +0.77V

V1-V2= +0.23V

Prediction: As V1-V2>0, VO_2^+ should be able to oxidize Fe^{2+} and VO^{2+} to form Fe^{3+}. Experimental: A blue colour was observed, indicating was formed and the reaction proceeded.

Reaction (ii):

1. VO_2^+ (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) +VO^{2+} (aq) +1.0V
2. I_2(s) + $2e^-$ ⇔ $2I^-$ (aq) +0.54V

V1-V2= +0.46V

Prediction: VO_2^+ should be able to oxidize I^- and VO^{2+} and I_2 should be formed. Experimental: A yellow colour was first observed, showing the presence of iodine. After adding $Na_2S_2O_3$ to remove iodine, a blue colour (V (IV) ions) was observed, indicating the reaction has proceeded.

Reaction (iii):

1. VO_2^+ (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) +VO^{2+} (aq) +1.0V
2. $S_2O_6^{2-}$ (aq) + $4H^+$ (aq) +2e- ⇔ $2H_2O$ (l) +$2SO_2$ (aq) +0.49V
3. VO^{2+} (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) + V^{3+} (aq) +0.39V
4. SO_4^{2-}(aq) + $4H^+$ (aq) +2e- ⇔ $2H_2O$ (l) + SO_2 (aq) +0.08V

Prediction: V1-V2= +0.51V, therefore, VO_2^+ should be able to oxidize SO_2 to form $S_2O_6^{2-}$ and VO^{2+} .V3-V4= +0.31V. Therefore, VO^{2+} should also be able to oxidize
SO_2 and V^{3+} and SO_4^{2-} should be form.

Experimental: A blue colour was first observed, showing the first reaction has proceeded. However, no green V (III) was observed, showing the second reaction has not taken place. This might due to the kinetics of S (IV) which is only capable of converting in S (V) but not S (VI).

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Reaction (iv):

1. $VO_2^+(aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O(l) + VO^{2+}(aq)$ +1.0V
2. $Sn^{2+} (aq) + 2e^- \Leftrightarrow Sn(s)$ -0.14V
3. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V

Predictions: V1-V2=+1.14V. Thus, VO_2^+ should be able to oxidize Sn, to form VO^{2+} and Sn^{2+} .V3-V2= +0.53V. So, VO^{2+} should be able to oxidize Sn and V^{3+} should be formed.

Experimental: A blue green colour (V (III) ions) was observed, indicating both reactions have proceeded. The mixture was heated to speed up the reaction and filtered to obtain a clear solution.

Reaction (v):

1. $VO_2^+ (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + VO^{2+}(aq)$ +1.0V
2. $Zn^{2+} (aq) + 2e^- \Leftrightarrow Zn(s)$ -0.76V
3. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V
4. $V^{3+} (aq) + e^- \Leftrightarrow V^{2+} (aq)$ -0.30V

Prediction: V1-V2= +1.76V. So, VO_2^+ should be able to oxidize Zn, forming VO_2^+ and Zn^{2+} .V3-V2= +1.15V. So, VO^{2+} should be able to oxidize Zn and V^{3+} should be formed. V4-V2= +0.46V. So, V^{3+} should be able to oxidize Zn and V^{2+} should be formed.

Experimental: A violet colour (V (II) ions) was observed, indicating all three reactions have proceeded. From the above reactions, the best way to obtain a V (IV) solution would be reaction (iii). The one producing V (III) would be reaction (iv) and that of producing V (II) would be reaction (V). The above reactions were used to verify the following reactions.

Reaction (vi):

1. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V
2. $V^{3+} (aq) + e^- \Leftrightarrow V^{2+} (aq)$ -0.30V

Prediction: V1-V2= +0.69V. So, VO^{2+} should be able to oxidize V^{2+} and V^{3+} should be formed.

Experimental: A pale green colour (V (III) ions) was observed, showing the reaction has taken place.

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Reaction (vii):

1. $VO_2^+(aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O(l) + VO^{2+}(aq)$ +1.0V
2. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V

Prediction: V1-V2 = 0.61V. So, VO_2^+ should be able to oxidize V^{3+} and VO^{2+} should be formed.

Experimental: A blue colour (V (IV) ions) was observed, showing the reaction has taken place.

Reaction (viii):

1. $V^{3+} (aq) + e^- \Leftrightarrow V^{2+} (aq)$ -0.30V
2. $VO_2^+(aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O(l) + VO^{2+}(aq)$ +1.0V
3. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V

Prediction: V1-V2=-1.3V. So, the reaction is not feasible and no change was expected.

Experimental: The mixture remained unchanged, showing no reaction has taken place.

Reaction (ix):

1. $Fe^{3+} (aq) + e^- \Leftrightarrow Fe^{2+} (aq)$ +0.77V
2. $VO_2^+(aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O(l) + VO^{2+}(aq)$ +1.0V

Prediction: V1-V2= -0.23V. So, the reaction is not feasible and no change was expected.

Experimental: The mixture remained unchanged, showing no reaction has taken place.

Reaction (x):

1. $Fe^{3+} (aq) + e^- \Leftrightarrow Fe^{2+} (aq)$ +0.77V
2. $VO^{2+} (aq) + 2H^+ (aq) + e^- \Leftrightarrow H_2O (l) + V^{3+} (aq)$ +0.39V

Prediction: V1-V2= +0.38V. So, Fe^{3+} should be able to oxidize V^{3+} and VO^{2+} and Fe^{2+} should be formed.

Experimental: No significant colour change was observed. The solution measured greenish yellow. Upon addition of KCNS, which stains Fe^{3+} red, the solution remained unchanged, showing the Fe^{3+} ions has reacted and the green colour was caused by the presence of Fe^{2+} ions. The reaction has taken place.

Reaction (xi):

1. VO^{2+} (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) + V^{3+} (aq) +0.39V
2. I_2(s) + $2e^-$ ⇔ $2I^-$ (aq) +0.54V

Prediction: V1-V2 = -0.15V. So, the reaction is not feasible and no change was expected.

Experimental: The mixture remained blue, showing no reaction has taken place. Upon addition of 1,1,1-trichloromethane , no distinct purple layer was observed, showing no iodine was formed due to infeasible reaction.

Reaction (xii):

1. VO_2^+ (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) +VO^{2+} (aq) +1.0V
2. Br_2 (aq) + 2e- ⇔ $2Br^-$(aq) +1.07V

Prediction: V1-V2 = -0.07V. So, the reaction is not feasible and no change was expected.

Experimental: The mixture remained yellow, showing no reaction has taken place. Upon addition of 1,1,1-trichloromethane , no distinct orange layer was observed, showing no bromine was formed due to infeasible reaction.

Reaction (xiii):

1. I_2(s) + $2e^-$ ⇔ $2I^-$ (aq) +0.54V
2. VO^{2+} (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) + V^{3+} (aq) +0.39V

Prediction: V1-V2= +0.15V. Thus, I_2 should be able to oxidize V^{3+} and VO^{2+} and I^- should be formed.

Experimental: A blue colour (V (IV) ions) was observed, showing that V^{3+} has been oxidized to form VO^{2+} .

Reaction (xiv):

1. I_2(s) + $2e^-$ ⇔ $2I^-$ (aq) +0.54V
2. VO_2^+ (aq) + $2H^+$ (aq) + e^- ⇔ H_2O (l) +VO^{2+} (aq) +1.0V

Prediction: V1-V2= -0.46V. So, the reaction is not feasible and no colour change was expected.

Experimental: A green colour, a mix between the blue of V (IV) and the yellow of iodine, remained. Upon addition of $Na_2S_2O_3$ to remove the iodine, a pure blue colour was observed, showing that I_2 was not reacted and no V(V) ions has been formed and so the reaction has not taken place.

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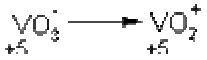
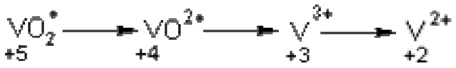
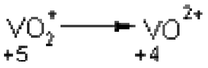
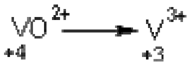
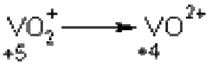
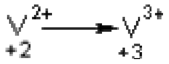
Some chemical reactions of Vanadium

Test	Observations	Summary of reaction
NH_4VO_3 + acid	white solid dissolved to a yellow solution	 $\underset{+5}{VO_3^-} \longrightarrow \underset{+5}{VO_2^+}$
V(V) + Zn	solution changes from yellow via green to blue, then to green and finally deep blue, gas evolves with a rotten-egg odour	 $\underset{+5}{VO_2^+} \longrightarrow \underset{+4}{VO^{2+}} \longrightarrow \underset{+3}{V^{3+}} \longrightarrow \underset{+2}{V^{2+}}$
V(II) + manganate(VII)	solution changes from violet to green, then to blue and finally yellow	 $\underset{+2}{V^{2+}} \longrightarrow \underset{+3}{V^{3+}} \longrightarrow \underset{+4}{VO^{2+}} \longrightarrow \underset{+5}{VO_2^+}$
1.> V(V) + sulphite 2.> Add V(II)	solution changes from yellow via green to blue	 $\underset{+5}{VO_2^+} \longrightarrow \underset{+4}{VO^{2+}}$
	blue solution turns green	 $\underset{+4}{VO^{2+}} \longrightarrow \underset{+3}{V^{3+}}$
V(V) + iodide + thiosulphate	solution changes from yellow via green to blue	 $\underset{+5}{VO_2^+} \longrightarrow \underset{+4}{VO^{2+}}$
V(II) + V(IV)	solution changes from deep blue to green	 $\underset{+2}{V^{2+}} \longrightarrow \underset{+3}{V^{3+}}$

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Conclusion and Discussion:

Most experiment results coincide with the predicted results based on the electro-chemical series. The feasibility of reactions based on the potential difference in the electro-chemical series was thus verified.

Sources of errors and improvements:

- Impurities left in the test tubes reacted with the reaction mixture. Rinsing the test tubes thoroughly should be needed to remove any impurities.
- The V (III) solution obtained by reacting with tin might not be pure as impurities were present on tin surface. Pure tin filling should be used to minimize the effect of impurities in tin.
- It was assumed that all reactions take place in the standard condition, at 25^{o}c and 1 atm. Otherwise, the electrode potentials from the ECS would not apply to the calculations of potential difference. Thermo-bath of temperature 25^{o}c should be set up to let chemicals remain at 25^{o}c before reaction.

Precautions:

1. Reactions involving iodide ions took a longer time to progress. The mixture should be allowed to stand for a while before determining the actual feasibility.
2. Anti-bumping granules should be added into the test tube of mixture before heating to prevent spurting out of the mixture when heating.

-End-