Guide to Chemistry Practicals

Questions and Answers to Selected
NECTA Practicals
1990 - 2006

Matthew C. Reid
Karatu Secondary School

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Introduction

The purpose of this booklet is to guide A-level chemistry students through the questions that are likely to appear on NECTA Paper 3, the practical paper. This booklet is *not* a replacement for the actual practical; it is very important that students are able to perform the practical on their own, and have lots of practice with experimental procedures and data collection. Instead, this booklet is meant to help students with the calculations and questions that come after the data has been collected.

How to use this booklet

The first part of this booklet contains 15 practical questions, all of them from past NECTA exams. The questions are written exactly as they are written on the NECTA exams. As you read the procedure, imagine that you are performing the experiment yourself. Ask yourself: What colour change should I see? What compound is in the pipette? Which is in the burette? What is the reaction that is taking place now?

After a question, experimental data is provided. It is filled into the results tables in the same way that a student would write the data if he or she were doing the experiment.

Now the student should use the provided data to perform any necessary calculations, draw graphs if necessary, and answer the questions. Sample calculations, graphs, and answers are provided at the end of the booklet. Students should attempt the calculations on their own before looking at the answers, however.

Acknowledgements

I am indebted to my students at Karatu Secondary School, especially the 2005-2007 CBA, CBG, and PCB students. They performed many of these practicals, helped to collect much of the data used in this booklet, and provided insight into which questions and calculations will be most difficult for Tanzanian students, and how best to answer these questions in a sensible way.

Notes on Units and Formulas

A. Volume

The units of volume commonly used in this manual are cm3 and dm3. The conversion between cm^3 and dm^3 is 1 dm^3 = 1000 dm^3. Note that a cm^3 are the same as millilitres (ml) and dm^3 are the same as litres (L). 1 cm^3 = 1 ml and 1 dm^3 = 1 L.

B. Concentration

Concentration is a measure of amount per volume. The amount can be grams, kilograms, number of moles, or any other unit of amount. Volume can be cm^3, dm^3, or any other volume. The two types of concentration used commonly are mol/dm^3 and g/dm^3.

(i) mol/dm^3 is also called Molarity or molar concentration. The symbol for molarity is M. Molarity is used in the $M_A V_A$ formula.

(ii) g/dm^3 is called gram concentration.

To convert from molar concentration to gram concentration, you multiply the molar concentration by the molar mass of the molecule.

To convert from gram concentration to molar concentration, you divide the gram concentration by the molar mass of the molecule.

C. The Mole Concept

A mole (mol) is a unit of amount. 1 mol of something is equal to 6.022×10^{23} of that thing. For example, 1 mol of H_2 molecules is equal to 6.022×10^{23} H_2 molecules. 1 mol of goats is 6.022×10^{23} goats.

There are two important formulas for determining the number of moles present of a compound. The first:

$$n = \frac{m \, (in \; grams)}{Molar \, Mass, \, M_r \, (in \; grams \; per \; mol)}$$

You see that when you cancel the units, you are left with moles.

The second formula:

$$n = MV$$

where M = molarity (in mol/dm^3) and V = volume (in dm^3). Notice that when you cancel the units, you are left with moles.

D. Stoichiometry

Stoichiometry is the study of the relationships between amounts (number of moles or masses) of reactants and products. It is fundamental to all of chemistry, and it is essential that you understand the stroichiometry and are able to apply it to numbers of moles and masses in chemical reactions.

Quantitative reactions are reactions that continue until one of the reactants is finished. Reactions with a strong acid or strong base are quantitative, as are most of the reactions in redox titrations. If the reaction proceeds in only one direction (and is not reversible), it is a quantitative reaction. Usually the NECTA exam will state if a reaction is quantitative. If the question states that something "is in excess," this means that there is a very large amount of it, so much that it will not be finished in the reaction.

The mol ratio between two compounds is the ratio in which those two compounds react. For example, in the reaction:

$NH_3 + O_2 \rightarrow NO + H_2O$
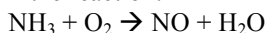
One mol of NH_3 reacts with one mol of O_2 to form one mol of NO and one mol of H_2O. If 8.5 g of NH_3 react with 32 of O_2, how many grams of NO will be formed? When comparing the relationships between two compounds in a reaction, you must look at the mole relation! The mass relation does not tell us very much. Converting to moles, we see that there are 0.5 moles of NH_3 and 1 mole of O_2.

Which of the reactants will be finished first, or will both be finished at the end of the reaction? NH_3 will be finished first, since there are only 0.5 mols of it present and NH_3 and O_2 react in a 1:1 mole ratio.

How many moles of NO will be formed? NH_3 and NO have a 1:1 mol ratio, so 0.5 mols of NH_3 will produce 0.5 mols of NO. 0.5 moles of NO has a mass of 15 grams.

So you can see that in order to determine the relationship between reactants and products, we must first compare the number of moles! Once we have determined the mole relation, we can convert from number of moles to grams.

Another example: In the reaction:

$I_2 + 2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2\ I^-$
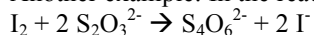
1 mol of I_2 reacts with 2 mol of $S_2O_3^{2-}$ to produce 1 mol of $S_4O_6^{2-}$ and 2 mol of I^-. This gives us no information about the actual number of moles of these compounds, only about the ratio in which they react. It means that 1 mol of I_2 requires 2 mol of $S_2O_3^{2-}$ to react completely. Or: 0.25 mol of I_2 requires 0.5 mol of $S_2O_3^{2-}$ to react completely. The ratio must me 1:2. But the actual number of moles can be 0.25 mol I_2 and 0.5 mol $S_2O_3^{2-}$, or 2 mol I_2 and 4 mol $S_2O_3^{2-}$, or 10 mol I_2 and 20 mol $S_2O_3^{2-}$.

To find the actual number of moles, you can use cross-multiplication. You could be asked: 0.0125 mol of $S_2O_3^{2-}$ requires how many moles of I_2 to react completely? You know that the mole ratio between $S_2O_3^{2-}$ and I_2 is 2:1. Thus:

$$\frac{0.0125\ mol\ S_2O_3^{2-}}{2\ mol\ S_2O_3^{2-}} = \frac{x\ mol\ I_2}{1\ mol\ I_2}$$

We would find that x = 0.00625 mol I_2 would be required to completely react with 0.0125 mol 2 $S_2O_3^{2-}$.

Note that the relationship between the two compounds is between number of moles, *not* mass. For example, it is true that 1 mol of I_2 reacts with 2 mol of $S_2O_3^{2-}$, but this *does not mean* that 1 gram of I_2 reacts with 2 grams of $S_2O_3^{2-}$. You must think of these relationships in terms of numbers of moles!

E. The Dilution Formula

The dilution formula allows you to determine the molarity of a solution after it has been diluted. The formula is:

$M_1V_1 = M_2V_2$

Where M_1 and V_1 are the molarity and volume before a dilution, and M_2 and V_2 are the molarity and volume after a dilution.

Consider this example from the 2006 NECTA: "Hydrochloric acid solution made by diluting 750 cm^3 of 0.25 M HCl to 937.5 cm^3 with distilled water." What is the molarity of this solution after dilution?

The volume before dilution (V_1) was 730 cm^3, and the molarity before dilution (M_1) was 0.25 M. The volume after dilution (V_2) was 937.5 cm^3. To find the molarity after dilution, M_2, we simply substitute the known values into the dilution equation and solve for M_2. We find that $M_2 = 0.2$ M.

Whenever you are given both a molarity and volume before a change (M_1 and V_1) and are asked to calculate a volume or molarity after a change (V_2 or M_2), the dilution formula is a good way to find the answer.

F. M_AV_A

The complete form of M_AV_A is shown below:

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$$

M_A is the molarity of compound A, V_A is the volume of compound A, M_B is the molarity of compound B, and V_B is the volume of compound B. n_A and n_B are the mole ratio between compounds A and B, found from the balanced chemical equation. This equation is extremely important in chemistry practicals. It is used in acid-base titrations, redox titrations, and is sometimes needed in kinetics and solvent extraction questions. You can use it to solve for any part of the formula: M_A, V_B, anything. If you can find values for five of the terms, you will be able to calculate the sixth.

Exercises:

1. In the reaction $I_2 + 2\ S_2O_3^{2-} \rightarrow 2\ I^- + S_4O_6^{2-}$, 12.7 g of I_2 would require how many grams of sodium thiosulphate for complete reaction?

2. You mix 250 cm^3 of a 0.1 M HCl solution with 500 cm^3 of a 0.08 M NaOH solution. Will HCl or NaOH be in excess? How many moles of the compound will remain? What will be the molarity of the solution?

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3. $CuSO_4$ reacts with KI to produce I_2 in the following reaction:
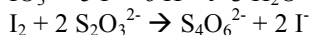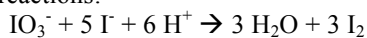
$$2 \ Cu^{2+} + 4 \ I^- \rightarrow Cu_2I_2 + I_2$$
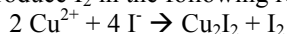
15.9 g of $CuSO_4$ react with 4.15 g of KI. Which reactant will finish first? What amount of I_2 will be formed (in moles and grams)?

4. Consider the two reactions:

$$IO_3^- + 5 \ I^- + 6 \ H^+ \rightarrow 3 \ H_2O + 3 \ I_2$$
$$I_2 + 2 \ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2 \ I^-$$

21.4 g of potassium iodate (KIO_3) would react with an excess of iodide ions to produce what mass of I_2? This I_2 produced in the first reaction would require how many moles of sodium thiosulphate $(Na_2S_2O_3^{2-})$ to be used up? Hint: use the mole ratio between IO_3^- and $S_2O_3^{2-}$ to answer this last question.
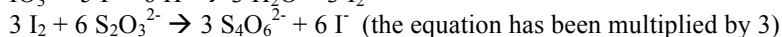
Answers:

1. 15.8 grams
2. NaOH will be in excess. 0.015 moles of NaOH will remain after reaction with HCl, producing a solution with a molarity of 0.02 M NaOH
3. The KI will finish first. The reaction will produce 0.00625 moles if I_2, or 1.59 grams of I_2.
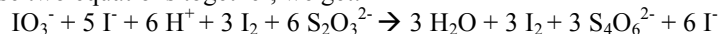4. 21.4 grams of potassium iodate is 0.1 moles of potassium iodate. This will form 0.3 moles of I_2.

0.1 moles of potassium iodate will produce enough iodine to react with how many moles of thiosulphate ions? We can find the mole ratio between iodate ions and thiosulphate ions and then use cross-multiplication.

How can we find the mole ratio between IO_3^- and $S_2O_3^{2-}$? They are in two different equations so they are difficult to compare? Well, we see that there is 1 mol of IO_3^- and 2 mol of $S_2O_3^{2-}$, so the mole ratio must be 1:2. *NO!!!! This is not correct!*

We must add the two equations together in such a way that I_2 cancels. I_2 allows you to translate between these two equations. How can we do this? We see that in the first equation there are 3 I_2 on the product side, and in the second equation there is one I_2 on the reactant side. We must multiply the second equation by 3, and then add the two together.

$$IO_3^- + 5 \ I^- + 6 \ H^+ \rightarrow 3 \ H_2O + 3 \ I_2$$
$$3 \ I_2 + 6 \ S_2O_3^{2-} \rightarrow 3 \ S_4O_6^{2-} + 6 \ I^- \text{ (the equation has been multiplied by 3)}$$

Adding these two equations together, we get:

$$IO_3^- + 5 \ I^- + 6 \ H^+ + 3 \ I_2 + 6 \ S_2O_3^{2-} \rightarrow 3 \ H_2O + 3 \ I_2 + 3 \ S_4O_6^{2-} + 6 \ I^-$$

Cancelling the compounds that appear on both sides:

$$IO_3^- + 6 \ H^+ + 6 \ S_2O_3^{2-} \rightarrow 3 \ H_2O + 3 \ S_4O_6^{2-} + 1 \ I^-$$
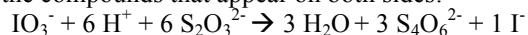
We see that the correct mole ratio between IO_3^- and $S_2O_3^{2-}$ is 1:6.

Using cross-multiplication, you would find that 0.1 moles of potassium iodate would produce enough iodine to react with 0.6 moles of sodium thiosulphate.

Questions

Question 1 – NECTA 1999, Chemistry 3A

Acid-Base Titration

You are provided with
A1: A mixture containing Sodium hydroxide, NaOH and Sodium Carbonate, Na_2CO_3
A2: 1M Hydrochloric Acid, HCl
A3: Phenolphthalein Indicator
A4: Methyl Orange

PROCEDURE

• Pipette 20 cm^3 (or 25 cm^3) of A1 into a clean conical flask and titrate with A2 using A3 until the colour changes from red to colourless. Record the first end point.
• Put A4 to the colourless mixture obtained in the previous procedure, and then continue titrating until the colour changes from yellow to just red. Record the second end point.

RESULTS:
The volume of the pipette used was _____ cm^3

Burette Readings

Titration	Pilot	1	2	3
Second end point (V_2 cm^3)				
First end point (V_1 cm^3)				
Initial volume (V_o cm^3)				
Volume used (V_2-V_o)				
Volume used (V_2-V_1)				

NB: There was an error in the exam question. The fifth row on the exam said: "Volume used (V_2-V_o)." It should say "Volume used (V_1-V_o)

Summary:
_____ cm3 of Solution A1 required _____ cm3 of A2 in the presence of A3 for complete neutralization.
_____ cm3 of Solution A1 required _____ cm3 of A2 in the presence of A4 for complete neutralization.

(a) Write the ionic equations for the reactions taking place in this experiment.

(b) Calculate:
    (i)   The volume of HCl required for complete neutralization of Na_2CO_3
    (ii)   The volume of HCl required for complete neutralization of NaOH
    (iii)   The molarity of Na_2CO_3
    (iv)   The mass of NaOH and that of Na_2CO_3 in the mixture

Experimental Data

The volume of the pipette used was <u>25 </u>cm^3

Titration	Pilot	1	2	3
Second end point (V_2 cm^3)	30.3	30.1	29.7	29.8
first end point (V_1 cm^3)	17.6	17.7	17.4	17.3
initial volume (V_o cm^3)	0	0	0	0
volume used (V_1-V_o)				
Volume used (V_2-V_1)				

NB: The error in the exam question has been fixed, so that the fifth row now says: "Volume used (V_1-V_o), which is what it should say.

Summary:
_____ cm^3 of Solution A1 required _____ cm3 of A2 in the presence of A3 for complete neutralization.
_____ cm^3 of Solution A1 required _____ cm3 of A2 in the presence of A4 for complete neutralization.

Now you, the student, need to fill out the remainder of the table, the summary and answer the questions.

Question 2 -- NECTA 1995, Chemistry Paper 3

Acid-Base Titration

You are provided with the following solutions:
  PP is a solution containing hydrochloric acid and acetic acid
  OQ is a 0.1M sodium hydroxide solution
  MO is Methyl Orange Indicator
  POP is Phenylpthalein Indicator
PROCEDURE:
(a) Pipette 20cm^3 of 25 cm^3 of solution PP into a 250 cm^3 titration flask.
(b) Add 2 or 3 drops of MO
(c) Titrate the solution against solution OQ until a colour change is observed
(d) Record the first titre value
(e) Add 2 or 3 drops of POP
(f) Continue to titrate until a second colour change is observed
(g) Record the second titre value
(h) Repeat your titrations (steps (a) - (g) four or five times and record your results as shown below.

RESULTS:
  The volume of the pipette used was ………….cm^3
  The volume of the burette used was …………cm^3
BURETTE READINGS:

Titration number	PILOT	1	2	3
1. Initial Reading				
2. Final Reading using MO				
3. Final Reading using POP				
4. First titre value (2-1) using MO				
5. Second titre value (3-2) using POP				

(a) The colour change for the titration using MO was from ………… to ………….. and for the titration when POP was used was from ……….. to ……………
(b) (i) MO is a good indicator for titration of strong base against ………… acid
    (ii) POP is a good indicator for titration of strong base against ………….acid
(c) (I) The first titre value signifies the titration of ………… against sodium hydroxide
    (ii) The second titre value signifies the titration of ………. against sodium hydroxide
(d) Give the balanced chemical equations for the titration of sodium hydroxide against
        (i) acetic acid
        (ii) hydrochloric acid
(e) Calculate the concentration of the acid solution PP in mol dm^{-3} when POP was added and when MO was added.
(f) Calculate the amount of hydrochloric acid in 1 dm^3 of solution PP in grams
(g) Calculate the amount of acetic acid in 1 dm3 of solution PP in grams.

Experimental Data

The volume of the pipette used was 20 cm^3
The volume of the burette used was 50 cm^3

BURETTE READINGS:

Titration number	PILOT	1	2	3
1. Initial Reading	0.0	0.0	0.0	0.0
2. Final Reading using MO	10.3	10.1	9.8	9.8
3. Final Reading using POP	26.4	26.2	25.7	25.9
4. First titre value (2-1) using MO				
5. Second titre value (3-2) using POP				

SUMMARY:
………cm^3 of solution PP required …….cm^3 of solution OQ when MO was used and ………cm^3 of solution OQ when POP was used.

The information provided is enough to complete the table, fill in the summary, and answer the questions.

Question 3 -- NECTA 2006, Chemistry Paper 3A

Acid-Base Titration

You are provided with:

AA1: Solution mixture of Na$_2$CO$_3$•xH$_2$O and NaHCO$_3$ made by dissolving 6.4125 of hydrated sodium carbonate contaminated with 1.05 g of pure sodium hydrogen carbonate in distilled water and then making up the volume of the solution to 250 cm^3.

AA2: Hydrochloric acid solution made by diluting 750 cm^3 of 0.25 M HCl to 937.5 cm^3 with distilled water.

AA3: Methyl orange indicator

Theory:

When a mixture of sodium carbonate and sodium bicarbonate solution is titrated with dilute hydrochloric acid in presence of methyl orange indicator, they react quantitatively as shown in equations I and II below:

$Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$ …….. I
$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ ………… II
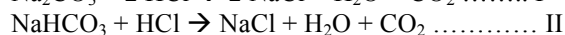
Procedure:

(i) Pipette 20 or 25 cm^3 of Solution AA1 into a titration flask.
(ii) Titrate the pipetted portion of AA1 against AA2 from a burette using two to three drops of methyl orange AA2 as indicator. Note the volume of AA2 used.
(iii) Repeat procedures (i) and (ii) above to obtain three more readings and tabulate your results as shown below.

(a) **RESULTS:**

Burette Readings:

Titration number	Pilot	1	2	3
Final reading (cm^3)				
Initial Reading (cm^3)				
Titre Volume (cm^3)				

(b) _____ cm^3 of AA1 required _____ cm^3 of AA2 for complete reaction
(c) Calculate the molarity of AA2
(d) Calculate the molarity of AA1 with respect to:
(i) $NaHCO_3$ (ii) $Na_2CO_3 \cdot xH_2O$

Experimental Results

Burette Readings:

Titration number	Pilot	1	2	3
Final reading (cm^3)	28.9	28.8	28.8	28.6
Initial Reading (cm^3)	0.00	0.00	0.00	0.00
Titre Volume (cm^3)				

(b) 25 cm^3 of AA1 required _____ cm^3 of AA2 for complete reaction

This information is enough for you to fill in the rest of the results table and answer the questions.

Question 4 -- NECTA 1994, Chemistry Paper 3C

Acid-Base Titration

You are provided with:
    TT: a solution of an ammonium salt whose general formula is $(NH_4)x$ prepared by dissolving
        6.60 g of the salt in 1000 cm^3 of water.
    SS: 0.15M sodium hydroxide solution
    ZZ: 0.025M sulphuric acid.

Procedure:

Pipette 20 cm^3 (or 25cm^3) of solution TT into a beaker and then pipette an equal volume of solution SS into the same beaker. Boil the mixture gently until the steam being evolved is neutral to moist red litmus paper. Cool the mixture in a water bath, then titrate the excess SS in the mixture against solution ZZ using phenolpthalein (POP) as your indicator.

Repeat this procedure to obtain two or three more readings and record your results as shown below.

RESULTS:

The volume of the pipette used was ………… cm^3
The volume of the burette used was ………… cm^3

BURETTE READINGS:

Titration Number	Pilot	1	2	3
Final reading (cm^3)				
Initial reading (cm^3)				
Titre (cm^3)				

SUMMARY:

The solution mixture required …….. cm^3 of solution ZZ for complete neutralization of excess SS.

Calculate:

(a) the volume of excess SS in the mixture
(b) The volume of SS used to liberate ammonia from a single portion of solution TT
(c) The mass of ammonia expelled from the mixture
(d) The total mass of nitrogen in the expelled ammonia
(e) The percentage by weight of ammonia in the salt

Experimental Results

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

Titration Number	Pilot	1	2	3
Final reading (cm^3)	33.8	33.6	33.7	33.7
Initial reading (cm^3)	0.0	0.0	0.0	0.0
Titre (cm^3)				

The solution mixture required …….. cm^3 of solution ZZ for complete neutralization of excess SS.

The information provided above is enough for you to fill in the remainder of the data table, complete the summary, and answer the questions.

Question 5 – NECTA 1993, Chemistry Paper 3

Redox Titration

You are provided with:
KK: a solution of hydrogen peroxide obtained by reacting 7.2g of impure barium peroxide with dilute sulphuric acid and making the solution to 1 dm^3, using distilled water, after filtering.
GG: A solution of potassium permanganate containing 3.0g of KMnO$_4$ per dm^3
HH: A 2M sulphuric acid solution.

A quantitative reaction occurs between barium peroxide and dilute sulphuric acid according to the equation

$$BaO_{2(s)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + H_2O_{2(aq)}$$
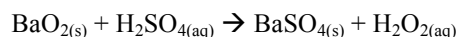
PROCEDURE:
Pipette 20 cm^3 (or 25 cm^3) of solution KK into a conical flask, add to it about 25 cm^3 of 2M HH and titrate this mixture with solution GG until a permanent pink colour is observed.
Repeat this procedure for at least three times. Record your titration results in tabular form as shown below.

RESULTS:
The volume of the pipette used was …………….. cm^3
The volume of the burette used was …………….. cm^3

BURETTE READING:

Titration Number	Pilot	1	2	3
Final reading (cm^3)				
Initial reading (cm^3)				
Titre (cm^3)				

SUMMARY:
…………….. cm^3 of solution KK needed ………….. cm^3 of solution GG for complete reaction.

(a) Write an ionic equation for the reaction between hydrogen peroxide and potassium permanganate in excess of dilute sulphuric acid.
(b) Calculate the concentration of hydrogen peroxide in
    (i) mole/dm^3 (ii) g/dm^3
(c) What was the percentage purity of barium peroxide used?

Experimental Results

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

Titration Number	Pilot	1	2	3
Final reading (cm^3)	16.2	31.8	15.9	31.7
Initial reading (cm^3)	0.0	16.2	0.0	15.9
Titre (cm^3)	16.2	15.6	15.9	15.8

SUMMARY:
…………….. cm^3 of solution KK needed ………….. cm^3 of solution GG for complete reaction

The information provided is enough for you to fill out the remaining boxes in the results table. After filling these boxes, you can complete the summary and answer the questions.

Question 6 – NECTA 1997, Chemistry Paper 3A

Redox Titration

You are provided with:
FA1: A solution made by dissolving 3.25 g of impure Potassium dichromate ($K_2Cr_2O_7$) in 500cm^3 of distilled water.
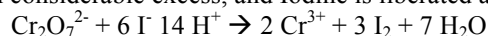FA2: A solution made by dissolving 12.40g of sodium thiosulphate pentahydrate ($Na_2S_2O_3 \cdot 5H_2O$) in 500 cm^3 of distilled water.
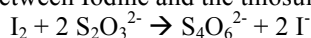FA3: 10% Potassium Iodide Solution

PROCEDURE:
1. Pipette 20.0cm^3 (or 25.0cm^3) of FA1 into a clean conical flask. Add to it 20 cm^3 of bench dilute H_2SO_4 and 20 cm^3 of Potassium Iodide solution.
   The Potassium Iodide is in considerable excess, and Iodine is liberated according to the equation
   $$Cr_2O_7^{2-} + 6\ I^- \ 14\ H^+ \rightarrow 2\ Cr^{3+} + 3\ I_2 + 7\ H_2O$$
2. Titrate the liberated Iodine with FA2, using starch as an indicator.
   Quantitative reaction takes place between Iodine and the thiosulphate ion.
   $$I_2 + 2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2\ I^-$$
   The end point is when the blue colour of the starch-Iodine complex is <u>just</u> replaced by the <u>GREEN</u> colour of Cr^{3+} ions.

RESULTS
The volume of the pipette used was ……………… cm^3

Burette Readings:

Titration	Pilot	1	2	3
Final volume (m^3)				
Initial volume (cm^3)				
Volume used (cm^3)				

Note: There is an error in the question. The units for final volume should be cm^3, not m^3.

Summary:
………………cm^3 of FA1 liberated sufficient Iodine to oxidize ……………cm^3 of FA2.
Calculate:
(a) The molarity of the potassium dichromate, $K_2Cr_2O_7$
(b) The concentration of the potassium dichromate in g/dm^3
(c) The percentage purity of the potassium dichromate

Experimental Results

The volume of the pipette used was 25 cm^3

Titration	Pilot	1	2	3
Final volume (cm^3)	19.4	19.1	19.4	19.2
Initial volume (cm^3)	0.0	0.0	0.0	0.0
Volume used (cm^3)				

Note: the table has been corrected so that the units for final volume are cm^3, not m^3.

Summary:
………………cm^3 of FA1 liberated sufficient Iodine to oxidize ……………cm^3 of FA2

The information provided in this table is enough for you to fill out the summary and then answer the questions.

Question 7 – NECTA 2003, Chemistry 3A

Redox Titration

You are provided with:

A1: a solution of copper(II) sulphate ($CuSO_4$), obtained by reacting 9 g of impure copper (II) oxide (CuO) with dilute sulphuric acid and making the solution to 1 dm^3 using distilled water after filtering.

A2: a solution made by dissolving 12.40g of $Na_2S_2O_3 \cdot 5H_2O$ in 500 cm^3 of distilled water.
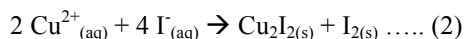
A3: A solution of 10% KI

A4: A starch solution.

Theory:

A quantitative reaction occurs between copper (II) oxide and sulphuric acid according to the equation given below:
$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O \ ..... \ (1)$$
The copper (II) sulphate obtained reacts quantitatively with potassium iodide to liberate iodine as follows:
$$2\ Cu^{2+}_{(aq)} + 4\ I^-_{(aq)} \rightarrow Cu_2I_{2(s)} + I_{2(s)} \ ..... \ (2)$$
The liberated iodine can be titrated against sodium thiosulphate and the reaction taking place during the titration can be represented as follows:
$$2\ S_2O_3^{2-}_{(aq)} + I_{2(s)} \rightarrow S_4O_6^{2-}_{(aq)} + 2I^-_{(aq)} \ ...... \ (3)$$
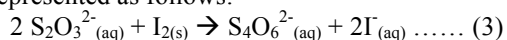
Procedure:
(a) Pipette 20.00cm^3 (or 25.00cm^3) of solution A1 into a conical flask.
(b) Add about 20 cm^3 of A3 solution from a measuring cylinder and shake the mixture.
(c) Titrate the mixture with solution A2 from the burette until the solution changes to light brownish in colour.
(d) Add about 2 cm^3 of solution A4.
(e) Continue with the titration until the pale blue colour just disappears and a milky colour appears.
Record your results in tabular form as shown below: (Table 1)

Results:
The volume of the pipette used was …………………………
The volume of the burette used was …………………………

Table 1: Burette Readings

Titration number	Pilot	1	2	3
Final reading (cm^3)				
Initial reading (cm^3)				
Titre volume used (cm^3)				

Summary: …………… cm^3 of solution A1 liberated iodine which required ………………. cm^3 of A2 for complete reaction.

Questions:
(a) Calculate the concentration of A2 in moles per litre
(b) Calculate the molarity of A1
(c) Calculate the concentration of A1 in g/dm^3
(d) Find the percentage purity of copper (II) oxide (CuO)

Experimental Results

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

Titration number	Pilot	1	2	3
Final reading (cm^3)	23.0	44.4	21.3	42.4
Initial reading (cm^3)	0.0	23.0	0.0	21.3
Titre volume used (cm^3)				

This information is enough for you to fill out the remaining cells in the table and to answer the questions.

Question 8 --NECTA 2005, Chemistry 3A

Redox Titration

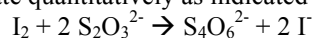You are provided with the following solutions:

D1: is a solution containing 13 g of an impure iodine in 1 dm^3 of aqueous solution
D2: is a solution containing 0.025 moles of $Na_2S_2O_3 \cdot 5H_2O$ in 250 cm^3 of aqueous solution
- Starch Indicator

Theory:
Iodine reacts with sodium thiosulphate quantitatively as indicated in the reaction below:
$$I_2 + 2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2\ I^-$$
Procedure:
(a) Pipette 25 cm^3 (or 20cm^3) of solution D1 into a conical flask.
(b) Put D2 into the burette and note the initial reading
(c) Titrate solution D1 against solution D2 from the burette until the solution is pale yellow. Add starch indicator and continue to titrate until the blur colour is discharged.
(d) Note the volume of D2 used.
(e) Repeat procedures (a) to (d) to get three more readings
(f) Record your results in the table as shown below.

Results:
The volume of the pipette used was _____ cm^3.
Burette readings:

Titration Number	Pilot	1	2	3
Final reading (cm^3)				
Initial Reading (cm^3)				
Volume used (cm^3)				

Summary: _____ cm^3 of solution D$_1$ required _____ cm^3 of solution D$_2$ for complete reaction.

Questions:
(a): Calculate the concentration of D2 in moles per dm^3.
(b) Calculate:
      (i) The molarity of pure iodine
      (ii) The concentration of pure iodine in g/dm^3
      (iii) The percentage purity of iodine
(c) Iodine is only slightly soluble in water. Which chemical substance do you think was added into water to increase the solubility of iodine? Explain

Experimental Results

Results:
The volume of the pipette used was 25 cm^3.
Burette readings:

Titration Number	Pilot	1	2	3
Final reading (cm^3)	21.1	20.7	41.2	20.4
Initial Reading (cm^3)	0.00	0.00	20.7	0.00
Volume used (cm^3)				

Summary: _____ cm^3 of solution D1 required _____ cm^3 of solution D2 for complete reaction.
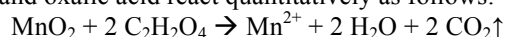
Question 9 -- Necta 1990, Chemistry Paper 3

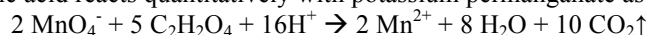*Redox Titration*

You are provided with the following solutions:
EF is a 250 cm^3 solution containing 1 g of an impure sample of Manganese dioxide (MnO_2) and 2.2g of oxalic acid crystals ($C_2H_2O_4 \cdot 2H_2O$).
GH is a standard solution of 0.02 M potassium permanganate, and dilute sulphuric acid.

Manganese dioxide and oxalic acid react quantitatively as follows:
$$MnO_2 + 2\ C_2H_2O_4 \rightarrow Mn^{2+} + 2\ H_2O + 2\ CO_2\uparrow$$
And the excess oxalic acid reacts quantitatively with potassium permanganate as follows:
$$2\ MnO_4^- + 5\ C_2H_2O_4 + 16H^+ \rightarrow 2\ Mn^{2+} + 8\ H_2O + 10\ CO_2\uparrow$$
(**Note**: The first equation is not balanced. If you encounter a mistake like this on the NECTA exam, make a note but then continue to use the given equation. You will note lose points by using the equation that NECTA provides. If balancing the equation is part of the question, however, be sure to balance the equation.)

PROCEDURE:
    Pipette 20 cm^3 (or 25cm^3) of EF into a conical flask, add about 20 cm^3 of dilute sulphuric acid and heat the contents of the flask to about 80ºC. Titrate the hot EF with GH until the purple colour disappears. (**Note**: there is a mistake in the question here. It should read: Titrate the hot EF with GH until the purple colour *no longer* disappears). Repeat this procedure two to three times and record your titration results as indicated in the table below.

RESULTS:
    The volume of the pipette used was ……… cm^3
    The volume of the burette used was ……….cm^3

BURETTE READINGS:

Titration Number	Pilot	1	2	3
Final Reading				
Initial Reading				
Titre (Volume used)				

SUMMARY:
    For complete reaction, ……….. cm^3 of EF required ………… cm^3 of GH.
Calculate:
    (a) the molarity of the oxalic acid
    (b) the concentration of the oxalic acid in g/dm^3
    (c) the percentage purity of the Manganese dioxide in the sample.

Experimental Results

The volume of the pipette used was 20 cm^3
The volume of the burette used was 50 cm^3

Titration Number	Pilot	1	2	3
Final Reading	11.4	11.3	22.7	33.7
Initial Reading	0.0	0.0	11.3	22.7
Titre (Volume used)				

For complete reaction, ……….. cm^3 of EF required ………… cm^3 of GH

These results are enough for you to fill in the remainder of the table, complete the summary, and answer the questions.

Question 10 – NECTA 1990, Paper 3

Reaction Rate – Activation Energy

You are provided with the following:
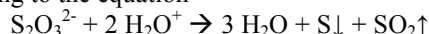
      HI is a solution of 0.05M of sodium thiosulphate
      JK is a solution of 0.1M Nitric Acid
      A stopwatch or stopclock
      A thermometer (0ºC - 100ºC)

A white precipitate of amorphous sulphur can be obtained by the action of a dilute acid on sodium thiosulphate ($Na_2S_2O_3$) according to the equation

$$S_2O_3^{2-} + 2\ H_2O^+ \rightarrow 3\ H_2O + S\downarrow + SO_2\uparrow$$

(Note: There is an error in the question. On the reactant side, H_2O^+ should be H_3O^+)

The precipitated sulphur causes the solution to become opaque. From this phenomenon, you can assess the rate of sulphur precipitation by measuring the time taken for the solution to become totally opaqued by the sulphur

PROCEDURE:
(a) Using a blue/black marker pen, draw a clear letter M on a white piece of paper and place a small beaker/flask provided on top of the letter M such that the letter M is visible through the solution
(b) Pour about 200cm^3 of tap water into a 250cm^3 or 300cm^3 beaker. (Use this as your water bath)
(c) Measure out exactly 10.0 cm^3 of HI and 10.0 cm^3 of JK into separate boiling test tubes.
(d) Put the two boiling test tubes containing HI and JK into the water bath (in (b)) above and warm the contents to about 50ºC.
(e) Immediately pour the hot solutions HI and JK (in (d)) above in the small beaker/flask in (a) above and simultaneously start the stopwatch/stopclock and record the temperature of the reaction mixture.
(f) Using a glass rod, stir the reaction mixture (in (e)) above and record the time taken, in seconds, for the letter M to disappear completely.
(g) Repeat the whole procedure ((c) to (f)) using temperatures 60ºC, 70ºC and 80ºC and record your results as shown below.

RESULTS:

Temperature T/ ºC	Temperature T/K	Time of Reaction (t/sec)	1/T (K^{-1})	Log$_{10}$1/t /sec^{-1}
50ºC				
60ºC				
70ºC				
80ºC				

(i) Plot a graph of Log$_{10}$ 1/t /sec^{-1} against 1/T K^{-1}
(ii) Determine the slope of the graph obtained in (i) above.

(iii) Using the equation $k = Ae^{-E/RT}$ which gives the relationship describing the dependence of the rate constant on temperature determine the value of A and E of the given equation using data from (ii) above.

Experimental Results

Temperature T/ °C	Temperature T/K	Time of Reaction (t/sec)	1/T (K⁻¹)	$Log_{10}1/t$ /sec-1
50°C		346.12		
60°C		184.72		
70°C		92.23		
80°C		46.38		

The information provided is enough to fill in the remainder of the table, plot the graph, and answer the questions.

Question 11 – NECTA 2000, Chemistry 3A

Reaction Rate

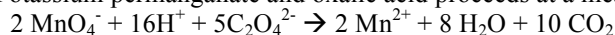You are provided with the following

AA: 0.02M $KMnO_4$
AB: 0.05M $H_2C_2O_4$ made up in 0.5M H_2SO_4
AC: Thermometer
AD: Stop Watch

THEORY:

The reaction between Potassium permanganate and oxalic acid proceeds at a measurable rate as
$$2\ MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2\ Mn^{2+} + 8\ H_2O + 10\ CO_2$$

PROCEDURE
• Prepare a water bath by putting 200 cm³ of distilled water in a 250cm³ conical flask and leave the water boiling
• Put AA and AB in different burettes
• Measure out 10 cm3 of each and put in a different boiling test-tube. Immerse the boiling tubes in the water bath. Boil the water bath up to 50°C. At this temperature pour AB into AA and immediately start AD. Note the time taken for the disappearance of the purple colour. Repeat the experiment for other temperatures 60, 70, 80, and 90° C.

Table of results

Temperature T, °C	Time, t (sec)	T (K)	1/T (K⁻¹)	Log t (sec)
50				
60				
70				
80				
90				

Questions

(a) Write the ionic half-reaction equation for both (i) MnO_4^- and (ii) $C_2O_4^{2-}$ ions.
(b) Plot the graph of log t (sec) against 1/T (K⁻¹)

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(c) Calculate the activation energy given that the slope of the graph is equal to 0.052 Ea mol KJ^{-1}
(**Note:** The units provided are incorrect. It should say: 0.052 Ea mol J^{-1}).

Experimental Data

Temperature T, ºC	Time, t (sec)	T (K)	1/T (K^{-1})	Log t (sec)
50	37.1			
60	14.2			
70	6			
80	3			
90	1.8			

This data is enough to fill the rest of the table. Fill in the final three columns, plot the graph, and answer the questions.

Question 12 – NECTA 1992, Chemistry 3

Reaction Rate

You are provided with:
KK: 0.02M Iodine dissolved in 0.2M KI
LL: 0.25M Acetone solution
MM: 0.50M Aqueous sodium hydrogen carbonate
NN: 001M Aqueous sodium thiosulphate
OO: 1.00M sulphuric acid
DD: Starch Solution

The rate of acid catalyzed iodination of acetone in aqueous solution depends on the concentration of both acetone and iodine.

$$I_2(aq) + CH_3 C CH_{3(aq)} \rightarrow ICH_2 C CH_3 + HI_{(aq)} \quad (1)$$

The dependence of this reaction on the concentration of iodine alone can be studied if a large excess of acetone is used such that its concentration (acetone's) is virtually unchanged throughout the reaction. But this reaction proceeds at a very slow rate whereas the reaction between sodium thiosulphate and iodine proceeds at a very fast rate.

$$2 S_2O_3^{2-} + 2 I_{2(aq)} \rightarrow 2I^-_{(aq)} + S_4O_6^{2-} \quad (2)$$
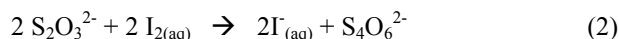
The study of the iodination of acetone may be carried out by titrating portions of the reaction mixture (reaction 1) with sodium thiosulphate using strach indicator.

PROCEDURE:
(i)   Use a measuring cylinder to put 80 cm^3 of KK into the first conical flask, 40 cm^3 of LL and 40 cm^3 of OO into the second conical flask and 15 cm^3 of MM into the third conical flask. Put NN into the burette.
        Then pout the contents of the first conical flask into a second conical flask. Immediately start the stopwatch/clock and shake well the reaction mixture for at least one minute.
(ii)   After about five minutes, pipette 10 cm3 of the reaction mixture into the third conical flask. Note the time when you do this and shake well the flask. When effervescenvce has ceased, titrate the residual iodine in the flask with solution NN using DD solution as indicator.
NB: The starch indicator is only used towards the end point of the titration.
(iii)   Repeat procedure (ii) at intervals of five minutes and tabulate your results as shown below:

Time (Minutes) T	Vol. Of NN used (cm^3)	$Conc^n$ of Iodine (M)	1/$Conc^n$ of Iodine (M)
5			
10			
15			
20			
25			
30			
35			

(**Note**: The units for the final column should be M^{-1}, not M)

(a) Use your results to fill the three blank columns in the table
(b) On the same page of graph paper provided, plot
    (i)   T against M
    (ii)   T against M^{-1}
(c) Determine the rate of reaction (1) with respect to iodine. Give reason(s)
(d) What is the purpose of MM in this experiment?

Experimental Data

The data is shown below in Table

Time (Minutes) T	Vol. Of NN used	$Conc^n$ of Iodine (M)	1/Concn of Iodine (1/M)
5	12.2		
10	11.1		
15	10.3		
20	9		
25	7.9		
30	6.8		

The data provided is enough to fill the last two columns. Use this data to plot the graphs and answer the questions.

Question 13 – NECTA 1999, Chemistry 3A

Reaction Rate

You are provided with:
G1: 0.2M Sodium thiosulphate, $Na_2S_2O_3$
G2: 0.1M Hydrocholoric acid, HCl
G3: Distilled Water
G4: stop watch
G5: Piece of white paper marked "+"

PROCEDURE
• On G5 put a reaction concical flask on top of the mark + in such a way the mark is clearly seen through the bottom of the conical flask.
• Using 10 ml measuring cylinder (or burette) measure out 2 mls of G1 and 8 mls of G3 and then put them in a conical flask on top of the mark G5
• Using another 10 mls measuring cylinder (or burette) measure out 10 mls of G2 and at the convenient time pour G2 in the conical flask containing G1 and G3 and immediately start a stop watch. Record the time taken for the mark to disappear due to formation of the precipitate.

• Repeat the experiment for different sets of concentrations as shown in a table below.

Experiment	Sodium Thiosulphate (cm^3)	Water (cm^3)	0.1M HCl (cm^3)	Time (sec)
A	2	8	10	
B	4	6	10	
C	6	4	10	
D	8	2	10	
E	10	0	10	

RESULTS:

Sodium Thiosulphate (cm^3)	Time (t) Sec	1/Time (1/t) = Rate Sec^{-1}
2		
4		
6		
8		
10		

From the experiment
  (a) Given that the volumes of the individual solutions are directly proportional to their concentration.
     (i)   Plot a graph of 1/t (vertical axis) against the volumes of sodium thiosulphate (note: origin is a point in the graph)
     (ii)   From the graph determine the order of reaction with respect to sodium thiosulphate
  (b) What caused the precipitate to occur in the reaction?
  (c) Write the ionic equation between the acid and sodium thiosulphate

Experimental Data

Experiment	Sodium Thiosulphate (cm^3)	Water (cm^3)	0.1M HCl (cm^3)	Time (sec)
A	2	8	10	699
B	4	6	10	206
C	6	4	10	124
D	8	2	10	86
E	10	0	10	57

Sodium Thiosulphate (cm^3)	Time (t) Sec	1/Time (1/t) = Rate Sec^{-1}
2		
4		
6		
8		
10		

Use the data provided in the first table to fill in the second table.   Then plot the graph and answer the questions.

Question 14 -- NECTA 1994, Chemistry Paper 3A

Reaction Rate

You are provided with the following:

AA: Distilled water
BB: Potassium Iodide Solution
CC: 2M Sulphuric acid
DD: Starch Solution
EE: 0.025M sodium thiosulphate solution
FF: 0.025 M Hydrogen peroxide

You are required to determine the order, n, with respect to hydrogen peroxide, of the reaction

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

whose general rate equation can be expressed as

$$rate = k[H_2O_2]^n[I^-]^p[H^+]^q$$

NB: If the iodine and hydrogen ions are present in excess, the order of the reaction, n, with respect to hydrogen peroxide can be determined.

The cause of the reaction can be followed by noting the time taken by the reaction to produce an amount of iodine just enough to react with a known volume of sodium thiosulphate of specified concentration.

The reaction between iodine and sodium thiosulphate can be represented by the following equation:

$$I_2 + 2 S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$$
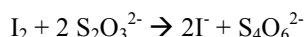
Proceed as follows:

Using a measuring cylinder, pour 70 cm3 of AA into a 250 cm3 conical flask and then add to it the following:

10 cm^3 of BB measured by using a pipette
5 cm^3 of CC measured by using a measuring cylinder
1 cm^3 of DD measured from a burette
2.5 cm^3 of EE measured from a burette

Shake the resulting mixture thoroughly and then pipette into it 10 cm^3 of FF. Stop the stopwatch/clock when you are about half-way through the addition of FF. Shake the mixture well to ensure complete mixing of the contents. Record the time when the dark-blue starch-iodine colour appears.

Without stopping the stopwatch/clock, add a further 2.5 cm^3 portion of EE from the burette, shake well and record the time when the dark-blue starch-iodine colour reappears.

Repeat this procedure by adding a further 2.5 cm^3 portions of EE for a total of six additions Record your results in a tabular form as shown below:

Total volume of EE added (cm^3)	Time, t, at which dark-blur colour appears	Total calculated volume of solution FF which has reacted (x cm^3)	(a-x), where a = initial volume of FF in the mixture (cm^3)	log (a-x)
2.50				
5.00				
7.50				
10.00				
12.50				
15.00				

(a) Plot a graph of (a-x) against time, t.
(b) (i) Plot a graph of log (a-x) against time, t
(ii) Determine the slope of this graph
(c) Using your two graphs (in (a) and (b)) above, what is the order of the reaction with respect to hydrogen peroxide? Explain why.

Experimental Data

Total volume of EE added (cm^3)	Time, t, at which dark-blur colour appears (sec)	Total calculated volume of solution FF which has reacted (x cm^3)	(a-x), where a = initial volume of FF in the mixture (cm^3)	log (a-x)
2.50	148			
5.00	309			
7.50	509			
10.00	768			
12.50	1168			
15.00	1747			

This information is sufficient to fill in the remaining sections of the table, plot the graphs, and answer the questions.

Question 15 -- NECTA 2006, Paper 3A

Partition Coefficient

You are given the following:
W_1: 0.1M sodium hydroxide
W_2: Solution containing an unknown concentration of succinic acid
W_3: Phenolphthalein indicator
W_4: Diethyl ether

Theory:

At a constant temperature, succinic acid dissolves in both water and diethyl ether, while maintaining a constant ratio of concentrations in the solvents under consideration.

PROCEDURE 1:

(i) Pipette 25 cm^3 (or 20 cm^3) of W_2 into a clean conical flask. Add to it 2-3 drops of W_3.
(ii) Put W_1 into the burette
(iii) Titrate very carefully solution W_2 against W_1 till there is a colour change. Record the volume of W_1 used as well as the room temperature.

(a) Results

(i) The volume of the pipette used was _____ cm^3.
(ii) The volume of W_1 used is _____ cm^3.
(iii) The room temperature is _____ ºC.

PROCEDURE 2:

(i) Measure 100 cm^3 of W_4 using a measuring cylinder and place it into a separating funnel.
(ii) Add 100 cm^3 of W_2 by means of a measuring cylinder into the funnel in (i) above. Shake well and allow the system to stand for a few minutes.
(iii) Run off the aqueous layer. Using a measuring cylinder, put 25 cm^3 (or 20 cm^3) of the aqueous layer into a clean conical flask. Titrate very carefully this aliquot against W_1 using W_3 as indicator.

Record the volume of W_1 used.

(b) Results:

(i) Volume of the aqueous layer taken was _____ cm^3.

(ii) Volume of W_1 used was _____ cm^3.

(c) Write down a balanced chemical equation representing the reaction taking place in procedure 1.(iii).
(d) Calculate the
        (i) initial concentration of W_2 in water.
        (ii) final concentration of W_2 in the aqueous layer.
(e) Deduce the concentration of W_2 in the organic layer.
(f) Calculate the partition coefficient of W_2 between water and diethyl ether W_4.

Experimental Results

(a) Results

(i) The volume of the pipette used was 20 cm^3.
(ii) The volume of W_1 used is 20.9 cm^3.
(iii) The room temperature is 27 ºC

(b) Results:

(i) Volume of the aqueous layer taken was 20 cm^3.
(ii) Volume of W_1 used was 16 cm^3

This information is enough for you to answer the questions.

Answers

Answers to Question 1 -- NECTA 1999

Here is the complete set of results:

The volume of the pipette used was __25__ cm^3

Titration	Pilot	1	2	3
Second end point (V_2 cm^3)	30.3	30.1	29.7	29.8
first end point (V_1 cm^3)	17.6	17.7	17.4	17.3
initial volume (V_o cm^3)	0	0	0	0
volume used (V_1-V_o)	17.6	17.7	17.4	17.3
Volume used (V_2-V_1)	12.7	12.4	12.3	12.5

The average V_1-V_o is 17.47 cm^3. This is the volume needed to reach the first colour change, which was the change of A3 (POP) from pink to colourless. In the first line of the summary, we write:

__25__ cm^3 of Solution A1 required __17.47__ cm^3 of A2 in the presence of A3 for complete neutralization.

The average V_2-V_1 is 12.4 cm^3. This is the volume needed to reach the second colour change, which was the change of A4 (MO) from yellow/orange to red. In the second line of the summary, we write:

__25__ cm^3 of Solution A1 required __12.4__ cm^3 of A2 in the presence of A4 for complete neutralization.

Question (a):

There are three neutralization reactions taking place in this experiment. The ionic equations for these reactions are:

$$H^+ + OH^- \rightarrow H_2O$$
$$H^+ \ CO_3^{2-} \rightarrow HCO_3^-$$
$$H^+ + HCO_3^- \rightarrow CO_2 + H_2O$$
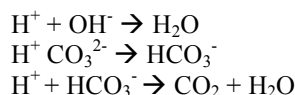
It will be more helpful for us to write the full equation, however.
   i) $NaOH + HCl \rightarrow H_2O + NaCl$
   ii) $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$
   iii) $NaHCO3 + HCl \rightarrow NaCl + H_2O + CO_2$

Equations (ii) and (iii) together represent the **complete neutralisation** of sodium carbonate, Na_2CO_3. To show the complete neutralisation of sodium carbonate, add equations (ii) and (iii) to get the overall reaction:

$$Na_2CO_3 + 2 \ HCl \rightarrow 2 \ NaCl + H_2O + CO_2$$
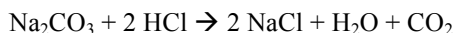
One mole of sodium carbonate reacts with two moles of HCl: one mole of sodium carbonate requires two moles of HCl for complete neutralization.

Question (b) (i):
We must determine how much HCl is needed to neutralise NaOH and how much HCl is needed to neutralise Na_2CO_3. In order to do this, we must write out our three neutralisation reactions along with which indicator shows us the completion of each reaction.
In general, a titration of:
   (i) a strong base and a weak acid requires POP as indicator
   (ii) a strong acid and a weak base requires MO as indicator
   (iii) a strong acid and a strong base can use either POP or MO as indicator.

Therefore:
MO is a good indicator for titration of a weak base against strong acid.
POP is a good indicator for titration of a strong acid against a weak base or a strong base.

In reaction (iii), $NaHCO_3$, a weak base, reacts with HCl, a strong acid. MO must be used for this reaction. In reaction (i), NaOH, a strong base, and HCl, a strong acid, react. This can be indicated by MO or POP. In this pracitcal, Na_2CO_3 will be considered a strong base, so it also can be indicated by POP or MO. The three reactions, and the indicator which shows the completion of each reaction, is shown below:

Indicator to exhibit completion of reaction

i) $NaOH + HCl \rightarrow H_2O + NaCl$ POP
ii) $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$ POP
iii) $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ MO

Because double indicator titrations involving a mixture of Na_2CO_3 and NaOH are so common, it is important to remember the choice of indicators shown above.

Looking back to the summary, we see that 17.47 cm^3 of HCl were required to reach the first colour change, which was POP. This means that 17.47 cm^3 of HCl were needed to complete reaction (i) and reaction (ii) together, because both these reactions are indicated by POP.

We also see that 12.4 cm^3 of HCl were required to reach the second colour change, which was MO. This means that 12.4 cm^3 of HCl were required to complete reaction (iii), because this reaction was indicated by MO.

But we have a problem. We know that 17.47 cm^3 of HCl are needed to complete reactions (i) and (ii) together, but how much of that volume is used in equation (i) by itself, and how much is used in equation (ii) by itself? We need to determine this in order to find the volume of HCl required to neutralise NaOH by itself.

The way to solve the problem is to realise that *equation (ii) and equation (iii) will each use the same volume of HCl*. Why? The $NaHCO_3$ reactant in equation (iii) is formed from the Na_2CO_3 in equation (ii) in a 1:1 ratio. The amount of Na_2CO_3 is therefore the same as the amount of $NaHCO_3$. Because the amounts of the two bases are the same, an equal amount of acid will be needed to neutralise them.

We know that 12.4 cm^3 of HCl were needed to complete reaction (iii). This means that 12.4 cm^3 of HCl were also needed to complete reaction (ii), since reactions (ii) and (iii) use the same volume of HCl for neutralisation. The volume of HCl used in reaction (i) must be 17.47 cm^3 - 12.4 cm^3 = 5.07 cm^3.

Volume of HCl needed to complete reaction

i) $NaOH + HCl \rightarrow H_2O + NaCl$ 5.07 cm^3
ii) $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$ 12.4 cm^3
iii) $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ 12.4 cm^3

Now we can finally answer questions (i) and (ii). The complete neutralisation of Na_2CO_3 is given by reaction (ii) together with reaction (iii). So the volume of HCl needed for complete neutralisation of Na_2CO_3 is 12.4 cm^3 + 12.4 cm^3 = **24.8 cm^3**.
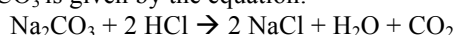
Question (b) (ii):
The complete neutralisation of NaOH is given by equation (i) by itself. The volume of HCl required to complete reaction (i) is 5.07 cm^3, so the volume of HCl needed to completely neutralise NaOH is **5.07 cm^3**.

Question (b) (iii):
To find the molarity of Na_2CO_3, we must use out $M_A V_A$ formula:

$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

The neutralisation of Na_2CO_3 is given by the equation:
$$Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$$
There is a 2:1 mole ratio between HCl and Na_2CO_3, so $n_A = 2$ and $n_B = 1$. $M_A = 1$ M (given by the question), $V_A = 24.8$ cm^3 (the volume of HCl needed for complete neutralisation of Na_2CO_3,

reaction (ii) + (reaction (iii)), and $V_B = 25$ cm^3 (the volume of the pipette). Solving the equation for our unknown, M_B, we get:

$$M_B = \frac{M_A V_A n_B}{n_A V_B}$$

Substituting in the known values, we find $M_B = 0.496$ M, approximately $M_B = 0.5$ M. This tells us that the molarity of Na_2CO_3 is **0.5 M**.

Using a similar method, we can also calculate the molarity of NaOH. There are two differences. For V_A we use 5.07 cm^3, because this was the volume of HCl used to neutralise NaOH by itself. Also, $n_A = 1$ and $n_B = 1$, because in the neutralisation reaction of NaOH there is a 1:1 mole ratio between acid and base. The other values are the same: $M_A = 1$ M (given by the question), and $V_B = 25$ cm^3 (the volume of the pipette). Substituting this value into the equation:

$$M_B = \frac{M_A V_A n_B}{n_A V_B}$$

We find that $M_B = 0.2028$, or approximately **$M_B = 0.2$**.

Question (b) (iv):

To find the masses of Na_2CO_3 and NaOH in the mixture, we first must convert from molarity (mol/dm^3) into gram concentration (g/dm^3). To make this conversion, multiply the molarity (units: mol/dm^3) times the molar mass (units: g/mol). The answer will be in units of g/dm^3.

For Na_2CO_3: 0.5 mol/dm^3 x 106 g/mol = 53 g/dm^3

For NaOH: 0.2 mol/dm^3 x 40 g/mol = 8 g/dm^3

Now we multiply the gram concentration times the volume of the solution in dm^3. The question does not specify the volume of the solution, however, so we will assume that the question refers to a solution with volume 250 cm^3, or 0.25 dm^3. To find the masses:

For Na_2CO_3: 53 g/dm^3 x 0.25 dm^3 = **13.25 g Na_2CO_3**

For NaOH: 8 g/dm^3 x 0.25 dm^3 = **2 g NaOH**.

Answers to Question 2 -- NECTA 1995

The complete set of data is given below:

The volume of the pipette used was 20 cm^3

The volume of the burette used was 50 cm^3

BURETTE READINGS:

Titration number	PILOT	1	2	3
1. Initial Reading	0.0	0.0	0.0	0.0
2. Final Reading using MO	10.3	10.1	9.8	9.8
3. Final Reading using POP	26.4	26.2	25.7	25.9
4. First titre value (2-1) using MO	10.3	10.1	9.8	9.8
5. Second titre value (3-2) using POP	16.1	16.1	15.9	16.1

SUMMARY:

20 cm^3 of solution PP required **9.9 cm^3** of solution OQ when MO was used and **16.03 cm^3** of solution OQ when POP was used.

Note: To find the average titre values, do not use the pilot. Take the average of the titre values of Titration 1, 2, and 3.

Question (a):

PP is a mixture of two acids, so in the presence of PP, Methyl Orange will be red in the presence of acid. When we titrate PP with NaOH, the solution becomes more basic, so methyl orange becomes yellow. So the colour change is from red to yellow.

POP is colourless in the presence of an acid, but pink in the presence of a base. Initially POP will be colourless, because PP is an acidic solution. As we titrate with NaOH, the solution becomes more basic. At the end point, NaOH neutralises the last of the CH_3COOH, so the solution becomes basic, and the POP will becomes pink. So the colour change is from colourless to pink.

The colour change for the titration using MO was from red to yellow and for the titration when POP was used was from colourless to pink.

Question (b):
In general, a titration of:

(i) a strong base and a weak acid requires POP as indicator
(ii) a strong acid and a weak base requires MO as indicator
(iii) a strong acid and a strong base can use either POP or MO as indicator.

Therefore:
MO is a good indicator for titration of a strong base against strong acid.
POP is a good indicator for titration of a strong base against a weak acid or a strong acid.

Question (c):
The first titre value shows the volume of NaOH needed to reach the end point of MO. MO indicates a strong base against a strong acid. In this question, the strong acid is HCl. Therefore, the first titre value shows the volume of NaOH required to neutralise HCl:

The first titre value signifies the titration of HCl against sodium hydroxide.

The second titre value shows the volume of NaOH needed to reach the end point of POP. POP indicates a strong base against a weak acid or a strong acid. The strong acid, HCl, has already been neutralised, so only the weak acid CH_3COOH remains. In this question, the weak acid is CH_3COOH. Therefore, the second titre value shows the volume of NaOH required to neutralise CH_3COOH.

The second titre value signifies the titration of CH_3COOH against sodium hydroxide.

Question (d):

$$(i)\ NaOH + CH_3COOH \rightarrow H_2O + NaCH_3COO$$

$$(ii)\ NaOH + HCl \rightarrow H_2O + NaCl$$
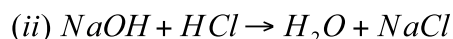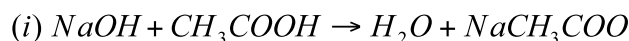
Question (e):
Before continuing, we must be clear about which indicator is indicating which reaction. MO indicates the neutralisation of HCl by NaOH (strong acid-strong base), while POP indicates the neutralisation of CH_3COOH by NaOH (weak acid-strong base).

When POP was added, only acetic acid (CH_3COOH) was present; the HCl had already been finished. To find the concentration of the acetic acid, we use M_AV_A:

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$$

The acid (A) is acetic acid and the base (B) is NaOH. M_A, the molarity of acetic acid, is unknown:

$$M_A = \frac{M_BV_Bn_A}{n_BV_A}$$

M_B, the molarity of the base, is 0.1 M. $V_B = 16.03$ cm^3 (from the volume of OQ required to change the colour of POP). $V_A = 20$ cm^3 (the volume of the pipette. Do not worry that acetic acid is only a portion of this total 20 cm^3. You need to use the *total* volume of the pipette.), n_A=1, and $n_B = 1$. Substituting these values into the equation, we find **$M_A = 0.08$ mol/dm^3 of CH_3COOH**. This is the concentration (in mol/dm^3) of the acid solution when POP was added.

When MO was added, however, both HCl and CH_3COOH were present in the solution PP. We already know the molarity of acetic acid: 0.08 mol/dm^3. How do we find the molarity of the HCl? We know from the summary that 9.9 cm^3 of solution OQ (NaOH) were required to reach the end point of MO, and that MO indicates the neutralisation of HCl. Therefore, 9.9 cm^3 of NaOH were required to neutralise the HCl.

We again use M_AV_A, with M_A as our unknown. In this case, M_A refers to the molarity of HCl:

$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

M_B = 0.1 M, V_B = 9.9 cm^3 (the volume of NaOH required to reach the end point of MO), V_A = 20 cm^3 (volume of the pipette), n_A = 1 and n_B = 1. Substituting the values into the equation, we find M_A = 0.0495 M, or approximately M_A = 0.05 M. This means the **concentration of HCl was 0.05 mol/dm^3**.

When MO was added, both acetic acid and HCl were present. So the concentration of the acid solution when MO was added is the concentration of HCl together with the concentration of acetic acid. We add 0.05M HCl with 0.08M CH$_3$COOH to get a total acid concentration of 0.13 M.

Question (f):
To find the amount in grams of HCl, first we must find the gram concentration. Multiply the molarity times the molar mass. The molar mass of HCl is 36.5 g/mol.

0.05 mol/dm^3 x 36.5 g/mol = 1.825 g/dm^3

This means that in 1 dm^3 of solution, there are **1.825 grams of HCl.**

Question (g):
To find the amount in grams of acetic acid, first find the gram concentration:

0.08 mol/dm^3 x 60g/mol = 4.8 g/dm^3.

This means that in 1 dm3 of solution, there are **4.8 grams of acetic acid**.

Answers to Question 3 -- NECTA 2006

Here are the complete data tables:

Burette Readings:

Titration number	Pilot	1	2	3
Final reading (cm^3)	28.9	28.8	28.8	28.6
Initial Reading (cm^3)	0.00	0.00	0.00	0.00
Titre Volume (cm^3)	28.9	28.8	28.8	28.6

(b) 25 cm^3 of AA$_1$ required 28.73 cm^3 of AA$_2$ for complete reaction

Question (c):
To calculate the molarity of the acid solution AA$_2$, we use the dilution formula, which says:

$$M_1 V_1 = M_2 V_2$$

The left-hand side of the equation refers to a solution before a change, while the right-hand side refers to the solution after a change. We see from the information provided that initially the HCl solution has a molarity of 0.25 M and a volume of 750 cm^3. Thus, M_1 = 0.25 and V_1 = 750 cm^3. This initial solution is diluted to a volume of 937.5 cm3. What is the molarity of this new solution, M_2?

$$M_2 = \frac{M_1 V_1}{V_2} = \frac{0.25\,M \ x\ 750\,cm^3}{937.5\ cm^3} = 0.2\,M$$

Thus, the molarity of AA2 is 0.2 M.

Question (d):
First we will calculate the molarity of AA$_1$ with respect to NaHCO$_3$. Answering this question does not require us to use experimental data: everything we need to answer this question is provided. We are told that solution AA$_1$ is "contaminated with 1.05 g of pure sodium hydrogen carbonate" and that the volume of the solution is 250 cm^3. This is enough to calculate the molarity. Remember, molarity is a measure of concentration in units of mol per dm^3. How many moles are there in 1.05 g of NaHCO$_3$? The molar mass of NaHCO$_3$ is 84 g/mol.   We use the equation below:
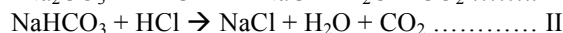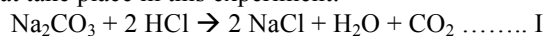
$$n = \frac{m}{M_r} = \frac{1.05g}{84g\,/\,mol} = 0.0125\,mol$$

The volume of the solution is 250 cm^3, or 0.25 dm^3. To find the molarity, we divide the number of moles by the volume:

$$\frac{0.0125\,mol}{0.25\,dm^3} = 0.05\,\frac{mol}{dm^3}$$

The molarity of solution AA_1 with respect to $NaHCO_3$ is thus 0.05 M.

The molarity of AA_1 with respect to $Na_2CO_3 \cdot xH_2O$ is a little more difficult. Consider the two neutralization reactions that take place in this experiment:

$$Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2 \ \text{........ I}$$
$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2 \ \text{............ II}$$

We know the total amount of HCl that is needed to complete both reactions: 28.73 cm^3 (the volume of HCl added from the burette). To find the molarity of Na_2CO_3, we must know how much HCl is used in equation I *only*. But how can we determine what portion of the total HCl was used in equation I and what portion was used in equation II?

We know the molarity of $NaHCO_3$, so we can use $M_A V_A$ to determine the volume of HCl used in equation II. We want to make V_A the subject, since we want to find the volume of the acid HCl used in equation II.

$$V_A = \frac{M_B V_B n_A}{M_A n_B}$$

$M_B = 0.05M$ (the molarity of the base $NaHCO_3$, calculated in Question (d) part (i)). $V_B = 25\ cm^3$ (the volume of the pipette with which solution AA_1, the mixture of bases, was added to the conical flask), and $M_A = 0.2M$ (calculated in Question (c)). For the reaction between HCl and $NaHCO_3$, $n_A = 1$ and $n_B = 1$.

$$V_A = \frac{0.05\ M\ x\ 25 cm^3\ x\ 1}{0.2\ M\ x\ 1} = 6.25\,cm^3$$

6.25 cm^3 of the HCl is used in the neutralization of $NaHCO_3$, equation II. How much HCl was used in the neutralization of Na_2CO_3, equation I? We take the total volume of HCl, 28.73 cm^3, minus the volume used in equation II, 6.25 cm^3. 28.73 cm^3 - 6.25 cm^3 = 22.48 cm^3. The volume of HCl used to neutralize Na_2CO_3 (equation I) is thus 22.58 cm^3.

We can now use $M_A V_A$ again, with M_B as the subject, to find the molarity of Na_2CO_3.

$$M_B = \frac{M_A V_A n_B}{V_B n_A}$$

$M_A = 0.2M$, $V_A = 22.58\ cm^3$ (the volume of HCl used in equation I), and $V_B = 25\ cm^3$ (the volume of the pipette). For equation I, $n_B = 1$ and $n_A = 2$.

$$M_B = \frac{0.2\ M\ x\ 22.58\,cm^3\,x 1}{25\ cm^3\,x 2} = 0.09\,M$$

Thus, the molarity of the $Na_2CO_3 \cdot xH_2O$ was 0.09M. The molarity of AA_1 with respect to $Na_2CO_3 \cdot xH_2O$ was **0.09M**. Do not worry about the hydrated waters; they have no effect in this calculation.


Answers to Question 4 -- NECTA 1994

The complete results are shown below:

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

Titration Number	Pilot	1	2	3
Final reading (cm^3)	33.8	33.6	33.7	33.7
Initial reading (cm^3)	0.0	0.0	0.0	0.0
Titre (cm^3)	33.8	33.6	33.7	33.7

The solution mixture required 33.67 cm^3 of solution ZZ for complete neutralization of excess SS.
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     Before starting, it is important to understand what is happening in this practical. Solution TT is a weakly-acidic solution, with ammonium ion (NH_4^+) acting as a weak acid. When we add Solution SS (NaOH), the NaOH neutralizes the NH_4^+:

$$NaOH + NH_4^+ \rightarrow H_2O + NH_3 + Na^+$$
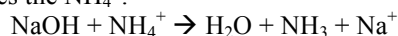
But the NaOH is added in excess, so some amount of it remains after the neutralization of NH_4^+ is complete. This excess NaOH is then neutralized by H_2SO_4, which is added from the burette:

$$2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O$$
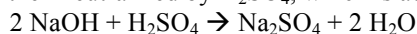
To summarize: 25 cm^3 of NaOH is mixed with NH_4^+. A portion of this NaOH is used in the reaction with NH_4^+, while another portion is in excess. This portion in excess is later neutralized in the titration with H_2SO_4.

Question (a):
     To find the volume of excess SS (NaOH), we can use $M_A V_A$, with V_B as the subject, since the volume of excess SS is the volume of excess NaOH, the base. We use the data collected from the titration, since the purpose of the titration was to neutralize excess NaOH (base) with H_2SO_4 (acid).

$$V_B = \frac{M_A V_A n_B}{n_A M_B}$$

M_A = 0.025 M (the molarity of the sulphuric acid), V_A = 33.67 cm^3 (volume of acid added from the burette), and M_B = 0.15 M (the molarity of the NaOH). From the neutralization reaction between NaOH and H_2SO_4, n_A = 1 and n_B = 2. Substituting these values into the equation, we find V_B = 11.22 cm^3. **So the volume of NaOH (solution SS) in excess is 11.22 cm^3**.

Question (b):
     The total volume of NaOH added was 25cm^3. 11.22 cm^3 of this total was in excess. This means that 25 cm^3 - 11.22 cm^3 = 13.78 cm^3 of NaOH was *not* in excess. This 13.78 cm^3 of NaOH was used up in the reaction with solution TT to produce, or liberate, ammonia. Thus, **13.78 cm^3 of NaOH** was used to liberate ammonia from a single portion of solution TT.

Question (c):
     To find the mass of ammonia expelled from the mixture, first we need to find the number of moles of ammonia expelled from the mixture. To find the number of moles of ammonia, consider the quantitative reaction between NaOH and the ammonium salt:

$$NaOH + NH_4^+ \rightarrow H_2O + NH_3 + Na^+$$

This is a quantitative reaction, which means that the reaction will go to completion: all of the NaOH will react to form products. NaOH and NH_3 are present in a 1:1 mole ratio; this means that x moles of NaOH will produce x moles of NH_3. The number of moles of NaOH will be the same as the number of moles of NH_3. So how many moles of NaOH were used in the reaction? We know the volume of NaOH used in the reaction with NaOH: 13.78 cm^3. How many moles is this? We can use the relationship

$$n = M \times V$$

Where n = number of moles, M = molarity, and V = volume in dm^3. The molarity of the NaOH is 0.15 mol/dm^3, while the volume is 13.78 cm^3. Convert this volume to dm^3 to get 0.01378 dm^3. The number of moles of NaOH is given by

$$n = 0.15\ \frac{mol}{dm^3} \times 0.01378\ dm^3 = 0.0021\ mol$$

If 0.0021 mol of NaOH reacted with NH_4^+, it means that 0.0021 mol of NH_3 were formed. Convert the number of moles of ammonia into the mass of ammonia using the formula:

$$n = \frac{m}{M_r}$$

Rearranging this equation, we find m = n x M_r. The number of moles of ammonia, n, is 0.0021 mol. The molar mass of ammonia, M_r, is 17 g/mol.

$$m = 0.0021 mol \ x \ 17 \frac{g}{mol} = 0.0357 g$$

Thus, 0.0021 moles of ammonia are produced. 0.0021 moles of ammonia are equivalent to 0.0357 g of ammonia. Thus, **0.0357 g of NH$_3$** are produced in the reaction between NaOH and NH$_4^+$ and are expelled from the mixture as gas.

Question (d):
This question asks us to find the total mass of nitrogen in the expelled ammonia. It refers to elemental nitrogen (N) and not nitrogen gas, since no nitrogen gas is produced in this reaction.

In one molecule of NH$_3$ there is one nitrogen atom. This means that in one mole of NH$_3$ there is one mole of nitrogen; similarly, in 0.0021 moles of NH$_3$, there are 0.0021 moles of nitrogen. To find the mass of this nitrogen, use the relationship:

$$m = n \ x \ M_r$$

The number of moles of nitrogen is 0.0021 mol, while the molar mass of nitrogen is 14 g/mol.

$$m = 0.0021 \ mol \ x \ 14 \frac{g}{mol} = 0.0294 \ g$$

The mass of the nitrogen is 0.0294 g. The total mass of nitrogen in the expelled ammonia is **0.0294 grams**.

Question (e):
To find the percentage by weight of ammonia in the salt, first consider the formula for percentage by weight:

$$Weight \ Percentage \ of \ Ammonia = \frac{Weight \ of \ Ammonia}{Total \ Weight} x 100\%$$

In this question, weight has the same meaning as mass. What is the mass of ammonia in the salt? In the procedure, we used 25 cm^3 of the solution of the ammonium salt, NH$_4^+$. In this solution, we determined that the total mass of ammonia expelled from the mixture was 0.0357 g. What mass of ammonia would be expelled from the complete mixture, which had a volume of 1000 cm^3? We can use cross-multiplication. If 25 cm^3 of solution TT produces 0.0357 g, how many grams would 1000 cm^3 of solution TT produce?

$$\frac{25 \ cm^3}{1000 \ cm^3} = \frac{0.0357 \ g}{x}$$

x = 1.428 g of ammonia in the total mixture of 1000 cm^3. The total mass of the salt, given in the beginning of the question, is 6.60 g. Substituting these values into the equation for weight percentage:

$$\frac{1.428 \ g}{6.6 \ g} x 100\% = 21.6\%$$

The weight percentage of ammonia in the salt was **21.6%.**

Answers to Question 5 -- NECTA 1993

The complete results are shown below:

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

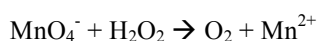| Titration Number	Pilot	1	2	3
Final reading (cm^3)	16.2	31.8	15.9	31.7
Initial reading (cm^3)	0.0	16.2	0.0	15.9
Titre (cm^3)	16.2	15.6	15.9	15.8

SUMMARY:
25 cm^3 of solution KK needed 15.77 cm^3 of solution GG for complete reaction

Question (a):

The reaction between $KMnO_4$ and H_2O_2 is a redox reaction. The permanganate ion, MnO_4^-, is reduced to Mn^{2+}, while the peroxide is oxidized to $O_{2(g)}$. This reduction of permanganate causes the purple colour of $KMnO_4$ to disappear when the permanganate is added to the peroxide. Mn^{7+} (the oxidation state of manganese in MnO_4^-) is purple, while Mn^{2+} is colourless. So, as long as the MnO_4^- reacts with and is reduced by peroxide, the purple colour will disappear. When the purple colour remains, it means that the peroxide has been finished, and thus can no longer reduce the Mn^{7+}.

The unbalanced reaction is:
$$MnO_4^- + H_2O_2 \rightarrow O_2 + Mn^{2+}$$
Following the normal procedure for balancing redox reactions, we find that the balanced reaction is:
$$5\ H_2O_2 + 2\ MnO_4^- + 6\ H^+ \rightarrow 5\ O_2 + 2\ Mn^{2+} + 8\ H_2O$$
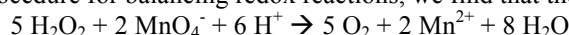
Question (b):

This is a redox titration question; all titration questions require us to use the M_AV_A formula to find concentrations:

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$$

We must first decide what is A and what is B. Let A = H_2O_2 and B = MnO_4^-. Question (b) asks us to find the concentration of hydrogen peroxide, so our unknown will be M_A:

$$M_A = \frac{M_BV_Bn_A}{n_BV_A}$$

Let us write what we know. M_B is the molarity of potassium permanganate. We know the gram concentration from the question: 3 g/dm^3. To convert from gram concentration to molarity, we divide the gram concentration by the molar mass of potassium permanganate, which is 158 g/mol.

$$\frac{3\ g/dm^3}{158\ g/mol} = 0.019\ mol/dm^3$$

So the molarity of potassium permanganate, M_B = 0.019 M. V_B = 15.77 cm^3 (the average titre from the burette). V_A = 25 cm^3 (the volume of the pipette). From the balanced equation, we see that n_A = 5 and n_B = 2. Substituting these values into the equation above, we find that M_A = 0.02996 M, or approximately M_A = 0.03 M. So the molar concentration of hydrogen peroxide is **0.03 mol/dm^3**.

To find the gram concentration, we multiply by the molar mass of H_2O_2, which is 34 g/mol.
$$0.03\ mol/dm^3 \times 34\ g/mol = 1.02\ g/dm^3$$
So the gram concentration of hydrogen peroxide is **1.02 g/dm^3**.

Question (c):

The formula for percentage purity is:

$$\%Purity = \frac{gram\ concentration\ pure}{gram\ concentration\ impure} x100\%$$

The gram concentration of impure BaO_2 is given by the question: 7.2 g of impure BaO_2 are used to prepare 1 dm^3 of solution, so the gram concentration is 7.2 g/dm^3.

What is the gram concentration of pure BaO_2? We must look at the mole ratio between H_2O_2 and BaO_2. We see from the balanced equation that one mole of BaO_2 reacts to form one mole of H_2O_2. We also know the molarity of H_2O_2 is 0.03 mol/dm^3. Because the mole ratio is 1:1, we know that 0.03

moles of BaO_2 react to form 0.03 moles H_2O_2 product. Therefore, the molarity of the (pure) BaO_2 solution is also 0.03 mol/dm^3.

To convert this molarity to gram concentration, we multiply by the molar mass of BaO_2, which is 169 g/mol.

$$0.03 \text{ mol/dm}^3 \text{ x } 169 \text{ g/mol} = 5.07 \text{ g/dm}^3$$

So the gram concentration of pure BaO_2 is 5.07 g/dm^3.

The percentage purity is given by:

$$\frac{5.07 g / dm^3}{7.2 g / dm^3} x 100\% = 70.4\%$$

So the percentage purity of the barium peroxide was **70.4%.**

Answers to Question 6 -- NECTA 1997

The completed results table and summary is below:

Titration	Pilot	1	2	3
Final volume (m^3)	19.4	19.1	19.4	19.2
Initial volume (cm^3)	0.0	0.0	0.0	0.0
Volume used (cm^3)	19.4	19.1	19.4	19.2

Summary:
25 cm^3 of FA1 liberated sufficient Iodine to oxidize 19.23 cm^3 of FA2

Before proceeding with the questions, we should be sure to understand what it happening in this practical. When you mix FA1 (potassium dichromate) with FA3 (potassium iodide), the dichromate ion reacts with the iodide ion to form free Iodine (I_2) in the following reaction:
$$Cr_2O_7^{2-} + 6 \text{ I}^- \text{ 14 H}^+ \rightarrow 2 \text{ Cr}^{3+} + 3 I_2 + 7 H_2O$$
The presence of the free Iodine causes the starch indicator to be blue in colour. The solution will retain this blue colour until all of the free Iodine (I_2) is again reduced to I^-. I_2 is reduced to I^- by the thiosulphate ion ($S_2O_3^{2-}$) in the following reaction:
$$I_2 + 2 S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2 \text{ I}^-$$
Thus, when you add thiosulphate from the burette, you are causing I_2 to be reduced to I^-. The end point of the titration is when the I_2 is finished and the solution loses its blue colour.
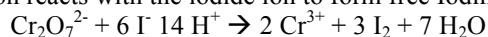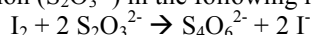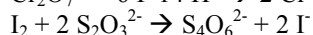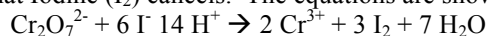
Question (a):
To find the molarity of the potassium dichromate, we use $M_A V_A$:
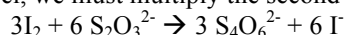
$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

Let A = $K_2Cr_2O_7$ and B = $Na_2S_2O_3$. We want to find M_A. Make M_A the subject:

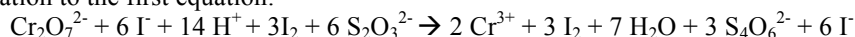$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

First we will find the mole ratio between $K_2Cr_2O_7$ and $Na_2S_2O_3$. To do this, add the two equations together so that Iodine (I_2) cancels. The equations are shown below:
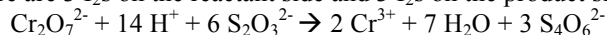$$Cr_2O_7^{2-} + 6 \text{ I}^- \text{ 14 H}^+ \rightarrow 2 \text{ Cr}^{3+} + 3 I_2 + 7 H_2O$$
$$I_2 + 2 S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2 \text{ I}^-$$
For I_2 to cancel, we must multiply the second equation by 3:
$$3I_2 + 6 S_2O_3^{2-} \rightarrow 3 S_4O_6^{2-} + 6 \text{ I}^-$$

Add this equation to the first equation:

$Cr_2O_7^{2-} + 6\ I^- + 14\ H^+ + 3I_2 + 6\ S_2O_3^{2-} \rightarrow 2\ Cr^{3+} + 3\ I_2 + 7\ H_2O + 3\ S_4O_6^{2-} + 6\ I^-$

Note that there are 3 I_2s on the reactant side and 3 I_2s on the product side: I_2 and also I^- cancel:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ S_2O_3^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ S_4O_6^{2-}$

The mole ratio from this equation is $n_A = 1$ and $n_B = 6$. M_B is the molarity of the sodium thiosulphate. The question tells us that solution FA2 is "12.40g of sodium thiosulphate pentahydrate ($Na_2S_2O_3 \cdot 5H_2O$) in 500 cm^3 of distilled water." The number of moles of sodium thiosulphate is given by

$$n = \frac{m}{M_r}$$

The mass of $Na_2S_2O_3 \cdot 5H_2O$ was 12.4 g, and the molar mass is 248 g/mol. 12.4g divided by 248g/mol = 0.05 moles. But molarity is mol/dm^3. The volume of the solution was 500 cm^3, or 0.5 dm^3. The molarity is 0.05 mol/0.5 dm^3 = 0.1 mol/dm^3. This is M_B. $V_B = 19.23$ cm^3 (volume added from the burette) and $V_A = 25$ cm^3 (volume added from the pipette). Substituting these values into M_AV_A, we find $M_A = 0.0128$ M, or approximately **$M_A = 0.013$ M**.

Question (b):

To find the concentration in g/dm^3, we multiply the molarity times the molar mass. The molar mass of $K_2Cr_2O_7$ is 294 g/mol. So the gram concentration is:

0.013 mol/dm^3 x 294 g/mol = 3.82 g/dm^3

So the gram concentration of the $K_2Cr_2O_7$ **is 3.82 g/dm^3**.

Question (c):

Percentage purity is equal to:

$$\%Purity = \frac{gram\ concentration\ of\ pure}{gram\ concentration\ of\ impure} x100\%$$

The pure concentration is 3.82 g/dm^3 (from Question (b)). The impure concentration can be found using information provided in the question: "FA1: A solution made by dissolving 3.25 g of impure Potassium dichromate ($K_2Cr_2O_7$) in 500cm^3 of distilled water." The impure concentration is 3.25 g divided by the volume, 500 cm^3 or 0.5 dm^3. 3.25g/0.5 dm^3 = 6.5 g/dm^3. Substituting these numbers into the equation, we find:

$$\frac{3.82g/dm^3}{6.5g/dm^3} x100\% = 58.8\%$$

So the percentage purity of the $K_2Cr_2O_7$ solution was **58.8%.**

Answers to Question 7 -- NECTA 2003

Here is the complete set of data:

The volume of the pipette used was 25 cm^3
The volume of the burette used was 50 cm^3

Titration number	Pilot	1	2	3
Final reading (cm^3)	23.0	44.4	21.3	42.4
Initial reading (cm^3)	0.0	23.0	0.0	21.3
Titre volume used (cm^3)	23.0	21.4	21.3	21.1

The average titre value, using experiments 1, 2, and 3 and ignoring the Pilot is 21.27 cm^3.

Question (a):

It is possible to find the molarity of A2 without looking at our experimental data. Everything we need is provided in the question: "A2: a solution made by dissolving 12.40g of $Na_2S_2O_3 \cdot 5H_2O$ in 500 cm^3 of distilled water." Molarity is concentration in mol/dm^3, so we must determine the number of moles of $Na_2S_2O_3 \cdot 5H_2O$, and the volume of the solution.

The number of moles, n, is given by

$$n = \frac{m}{M_r}$$

Where m = mass (grams) and M_r is molar mass (g/mol). The molar mass of $Na_2S_2O_3 \cdot 5H_2O$ is 248 g/mol, and the mass of $Na_2S_2O_3 \cdot 5H_2O$ is given to be 12.40 g. The number of moles is then 12.4 g/248 g/mol = 0.05 mol. The volume is 500 cm3, or 0.5 dm3. To find the molarity, we divide the number of moles by the volume in dm^3: 0.05 mol/0.5 dm^3 = **0.1 mol/dm^3**.

Question (b):

To find the molarity of A1 we must use $M_A V_A$:

$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

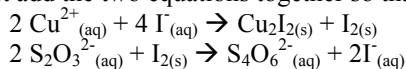There are two reactions taking place. First, $CuSO_4$ reacts with KI to liberate free Iodine (I_2):

$$2\ Cu^{2+}_{(aq)} + 4\ I^-_{(aq)} \rightarrow Cu_2I_{2(s)} + I_{2(s)}$$
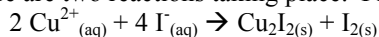
Then, the free Iodine reacts with sodium thiosulphate with starch as indicator:

$$2\ S_2O_3^{2-}_{(aq)} + I_{2(s)} \rightarrow S_4O_6^{2-}_{(aq)} + 2I^-_{(aq)}$$
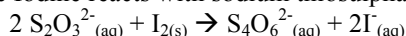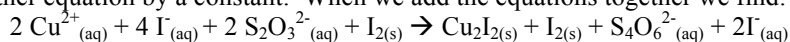
The disappearance of the blue colour shows us that the reaction is complete and the I_2 has been finished.

What is A and what is B? The question is asking for us to find the molarity of A1, $CuSO_4$, so we will say that A is $CuSO_4$. Then what is B? B will be sodium thiosulphate, $Na_2 S_2O_3$. Why? Because we already know the molarity of $Na_2 S_2O_3$ (from Question (a)) and we can find the mole ratio between $Na_2 S_2O_3$ and $CuSO_4$.

So let us write down what we know. M_A, the molarity of $CuSO_4$, is unknown. VA = 25 cm^3 (the volume of the pipette). M_B = 0.1 M (from Question (a)), and V_B = 21.27 cm^3. To find the mole ratio, we must add the two equations together so that I_2 cancels:

$$2\ Cu^{2+}_{(aq)} + 4\ I^-_{(aq)} \rightarrow Cu_2I_{2(s)} + I_{2(s)}$$
$$2\ S_2O_3^{2-}_{(aq)} + I_{2(s)} \rightarrow S_4O_6^{2-}_{(aq)} + 2I^-_{(aq)}$$

See that there is one I_2 on the product side of the first reaction, and one I_2 on the reactant side of the second reaction. This means that the reactions can be added together and I_2 will cancel without multiplying either equation by a constant. When we add the equations together we find:

$$2\ Cu^{2+}_{(aq)} + 4\ I^-_{(aq)} + 2\ S_2O_3^{2-}_{(aq)} + I_{2(s)} \rightarrow Cu_2I_{2(s)} + I_{2(s)} + S_4O_6^{2-}_{(aq)} + 2I^-_{(aq)}$$

Note that I_2 will cancel. And the mole ratio will be n_A = 2 and n_B = 2, because there are two moles of Cu^{2+} (the cation of $CuSO_4$) reacting with 2 moles of $S_2O_3^{2-}$.

Whenever you have to find the mole ratio between two compounds in different reactions and I_2 is present in both reactions, add the two equations together so that I_2 cancels. This is a very useful procedure and allows you to find the mole ratio between compounds that are not in the same reaction.

Our unknown is M_A:

$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

Substituting values into the equation, we find **M_A = 0.085 M**.

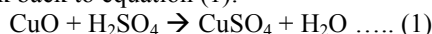**Question (c):**

To find the gram concentration, we multiply the molarity times the molar mass. The molar mass of $CuSO_4$ is 159.5 g/mol. 0.085 mol/dm^3 x 159.5 g/mol = **13.6 g/dm^3**.

Question (d):
　　　To find the percentage purity of Copper (II) oxide (CuO), first find the gram concentration of CuO. We know the gram concentration of $CuSO_4$ (from Question (c)), but we do not know the gram concentration of CuO. What is the relationship between CuO and $CuSO_4$?
　　　Look back to equation (1):
　　　　　$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ ….. (1)
We see from this equation that there is a 1:1 mole ratio between $CuSO_4$ and CuO. We know from Question (b) that the molarity of $CuSO_4$ is 0.085 mol/dm^3: In 1 dm^3 of solution, there are 0.085 moles of $CuSO_4$. This means that there were also 0.085 moles of CuO present in solution. In other words, 0.085 moles of CuO reacted with sulphuric acid to form 0.085 moles of $CuSO_4$ in 1 dm^3 of solution. The molarity of the CuO is therefore 0.085 mol/dm^3. The molar mass of CuO is 79.5 g/mol. The gram concentration of (pure) CuO is 0.085 mol/dm3 x 79.5 g/mol = 6.76 g/dm^3.
　　　Percentage purity is given by the formula:

$$\%Purity = \frac{gram\ concentration\ pure}{gram\ concentration\ impure} x100\%$$

The gram concentration of pure CuO is 6.76 g/dm^3. To find the gram concentration of impure CuO, look back to the question: "A1: a solution of copper(II) sulphate ($CuSO_4$), obtained by reacting 9 g of impure copper (II) oxide (CuO) with dilute sulphuric acid and making the solution to 1 dm^3 using distilled water after filtering." 9 g of impure CuO were used to prepare 1 dm3 of solution, so the gram concentration of impure CuO is 9 g/dm^3. The percentage purity is then:

$$\frac{6.76\ g/dm^3}{9\ g/dm^3} x100\% = 75.1\%$$

We see that the CuO was **75.1% pure**.

Answers to Question 8 -- NECTA 2005

The volume of the pipette used was 25 cm^3.

Titration Number	Pilot	1	2	3
Final reading (cm^3)	21.1	20.7	41.2	20.4
Initial Reading (cm^3)	0.00	0.00	20.7	0.00
Volume used (cm^3)	21.1	20.7	20.5	20.4

Summary: 25 cm^3 of solution D_1 required 20.53 cm^3 of solution D_2 for complete reaction.

　　　Before continuing with the questions, let's make sure we understand what is happening in this practical. We have an iodine solution of unknown molarity. We are going to titrate this unknown iodine solution against sodium thiosulphate, whose molarity we know. We can use the reaction between iodine and sodium thiosulphate to determine the mole ratio between these two compounds in the reaction.
　　　The indicator in this titration is starch. Starch indicates the presence of free iodine (I_2). If I_2 is present , starch indicator is blue. If I_2 is not present, the indicator is colourless. In this titration, solution D_1 in the conical flask will be blue after addition of starch indicator, because I_2 is present in solution D_1. The end point of this titration will be when the blue colour disappears. Why has the blue colour disappeared? Because $S_2O_3^{2-}$, being added from the burette, has reduced all of the I_2 to I^-. When the last of the I_2 is converted into I^-, the starch indicator will go from blue to colourless, because the I_2 has been finished.

Question (a):
　　　To find the concentration of D_2 in moles per dm^3, we simply take the number of moles, 0.025, and divide it by the volume in dm^3. We must first convert the volume to dm^3. 250 cm^3 = 0.25 dm^3. So the concentration in moles per dm^3 = 0.025 mol/0.25 dm^3 = 0.1 mol/dm^3.

Question (b):
 (i)
   To find the molarity of pure Iodine, we use the data from our titration and $M_A V_A$. The reaction that is being used to determine the molarity of pure Iodine is

$$I_2 + 2\ S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2\ I^-$$
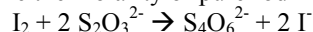
We know the molarity of $S_2O_3^{2-}$, so we will be able to determine the molarity of I_2. Let A = I_2 and B = $S_2O_3^{2-}$. The subject of the equation will be M_A, since we are searching for the molarity of I_2:

$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

M_B = 0.1 M (the molarity of the sodium thiosulphate), V_B = 20.53 cm^3 (the volume of sodium thiosulphate added from the burette), and V_A = 25 cm^3 (the volume of Iodine added from the pipette in solution D_1). From the balanced chemical equation, we see that the mole ratio is n_A = 1 and n_B = 2. Substituting these values into the equation above, we find M_A = 0.041 M. The molarity of pure I_2 is thus **0.041 M**. Remember, the value that you calculate with $M_A V_A$ is the pure molarity.

(ii)
To convert this molarity into gram concentration, we multiply the molarity times the molar mass of I_2. The molar mass of I_2 is 254 g/mol. The gram concentration is thus 10.414 g/dm^3.

(iii)
The percentage purity of Iodine is given by the equation:

$$\%Purity = \frac{gram\ concentration\ pure}{gram\ concentration\ impure} x100\%$$

The concentration of pure I_2, calculated in (ii), is 10.414 g/dm^3. The concentration of impure I_2 can be calculated using the information given in the question: "D1: is a solution containing 13 g of an impure iodine in 1 dm^3 of aqueous solution." This means the impure Iodine had a gram concentration of 13g in 1 dm^3, or 13 g/dm^3. Substituting these values into the equation, we find that the percentage purity is **80.1% pure**.

Question (c):
   To increase the solubility of I_2 in water, the I_2 is mixed with potassium iodide, KI. Iodine and the iodide ion (I^-) react to form a complex ion that increases the solubility of I_2.

$$I_{2(aq)} + I^-_{(aq)} \leftrightarrow I_3^-{}_{(aq)}$$
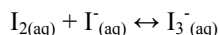
Answers to Question 9 -- NECTA 1990

The complete results are shown below:

   The volume of the pipette used was 20 cm^3
   The volume of the burette used was 50 cm^3

Titration Number	Pilot	1	2	3
Final Reading	11.4	11.3	22.7	33.7
Initial Reading	0.0	0.0	11.3	22.7
Titre (Volume used)	11.4	11.3	11.4	11.0

For complete reaction, 20 cm^3 of EF required 11.23 cm^3 of GH

   Question (a):
   To find the molarity in a titration question, we use $M_A V_A$:

$$\frac{M_A V_A}{n_A} = \frac{M_B V_B}{n_B}$$

Let A = oxalic acid ($H_2C_2O_4$) and B = potassium permanganate ($KMnO_4$). We need to find M_A. Making M_A the subject of the equation:

$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

Write down what we know. $M_B = 0.02$ (molarity of potassium permanganate), $V_B = 11.23$ cm^3 (volume added from burette), and $V_A = 20$ cm^3 (volume of pipette). To find the mole ratio, we look at the balanced chemical equation for the reaction between $KMnO_4$ and $H_2C_2O_4$. Two moles of permanganate (MnO_4^-) react with five moles of oxalate ion ($C_2O_4^{2-}$), so $n_A = 5$ and $n_B = 2$. Substituting these values into the equation, we find $M_A = 0.028$ M. So the molarity of oxalic acid was **0.028 M**.

Question (b):

The question does not specify whether you should calculate the gram concentration of hydrated or anhydrous oxalic acid, so we will find both. To find gram concentration, multiply the molarity times the molar mass. The molar mass of hydrated oxalic acid is 126 g/mol, while anhydrous oxalic acid has a molar mass 90 g/mol.

The gram concentration of hydrated is thus 0.028 mol/dm^3 x 126 g/mol = 3.53 g/dm^3. To convert to g/cm^3, we can use cross multiplication (remember that 1 dm^3 = 1000 cm^3):

$$\frac{1000 \; cm^3}{1 \; cm^3} = \frac{3.53g}{x}$$

We find x = **0.00353 g/cm^3**. This is the gram concentration of hydrated oxalic acid.

The gram concentration of anhydrous oxalic acid is 0.028 mol/dm^3 x 90 g/mol = 2.52 g/dm^3. Converting to g/cm3 gives **0.00252 g/cm^3**. This is the gram concentration of anhydrous oxalic acid.

Question (c):

To find the percentage purity of manganese dioxide (MnO_2), first consider the equation for percentage purity:

$$\%Purity = \frac{gram \; concentration \; pure}{gram \; concentration \; impure} x100\%$$

To find the gram concentration of impure MnO_2, we take the mass of impure MnO_2 divided by the volume in dm^3: 1g/0.25 dm^3 = 4 g/dm^3.

How can we find the concentration of pure MnO_2? We know that MnO_2 reacts with $H_2C_2O_4$ in the following reaction:

$$MnO_2 + 2 \; C_2H_2O_4 \rightarrow Mn^{2+} + 2 \; H_2O + 2 \; CO_2\uparrow$$

The question tells us that EF contains "2.2 g of oxalic acid crystals" made up in 250cm^3 of solution. We can convert this into molarity by first finding the number of moles. $n = m/M_r$, so n = 2.2/126 = 0.0175 mol. To find molarity, divide the number of moles by the volume in dm^3. 0.0175 mol/0.25 dm^3 = 0.07 M. So initially, solution EF was prepared so that the molarity of oxalic acid was 0.07 M, but we know from question (b) that the molarity obtained in the titration is only 0.028 M. What has caused the change?

The molarity of oxalic acid decreases because the oxalic acid reacts with MnO_2 in the reaction above, which causes the molarity of oxalic acid to decrease. In the reaction between MnO_2 and $C_2H_2O_4$, the oxalic acid is in excess; therefore the MnO_2 is finished (has a molarity of 0) when the reaction is complete, and the oxalic acid which remains has a molarity of 0.028 M.

We can construct a table showing how the molarities of the different reactants change in the reaction. Remember, our goal is to find the amount of pure MnO_2 present initially. The amount of pure MnO_2 will be the amount of MnO_2 that reacted with the oxalic acid and caused its molarity to decrease from 0.07M to 0.028M. Let A be the initial molarity of MnO_2. We know from our calculation that the initial molarity of $C_2H_2O_4$ was 0.07 M.

MnO_2 +	2 $C_2H_2O_4$ $\rightarrow$	Mn^{2+} + 2 H_2O + 2 $CO_2\uparrow$
Initial	A	0.07
Change	-x	-2x
Final	A-x	0.07-2x

But we know from the titration that the final molarity of oxalic acid is 0.028 M. Therefore, 0.07 - 2x = 0.028. So x = 0.021. We also know that the MnO_2 is finished at the end of the reaction, so the final

molarity of MnO_2 is 0. Therefore, A - x = 0. If x = 0.021, then A = 0.021. Thus, the initial molarity of (pure) MnO_2 was 0.021 M.

To convert this into gram concentration, we multiply molarity by the molar mass. The molar mass of MnO_2 is 87 g/mol. 0.021 mol/dm^3 x 87 g/mol = 1.83 g/dm^3. This is the gram concentration of pure MnO_2.

The percentage purity is thus:

$$\frac{1.83 g / dm^3}{4 g / dm^3} x100\% = 45.8\%$$

The MnO2 was **45.8% pure**.

Answers to Question 10 -- NECTA 1990

Here is the complete table of results:

| Temperature T/ ºC	Temperature T/K	Time of Reaction (t/sec)	1/T (K^{-1})	Log$_{10}$1/t /sec-1
50ºC	323	346.12	0.003096	-2.53923
60ºC	333	184.72	0.003003	-2.26651
70ºC	343	92.23	0.002915	-1.96483
80ºC	353	46.38	0.002832	-1.66633

Before beginning to answer the questions to a practical, it is always important to understand what is happening in the practical. In this practical, for example, what is it that is causing the white precipitate to form that eventually covers the mark on the piece of paper? If you look at the equation, you see:

$$S_2O_3^{2-} + 2 H_2O^+ \rightarrow 3 H_2O + S\downarrow + SO_2\uparrow$$
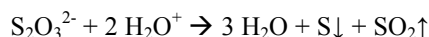
One of the products of this reaction is sulphur, which is formed as a precipitate. Precipitates are insoluble, and when they come out of the solution they cause the solution to become cloudy, or opaque. When enough precipitate has formed, the mark underneath the beaker is no longer visible. Thus, in this practical you are measuring the time it takes for enough precipitate to form in order to cover the mark on the paper beneath the beaker.

Question (i):
The graph of log$_{10}$(1/time) against 1/Temperature is shown below:

Plot of log (1/time) against 1/Temp
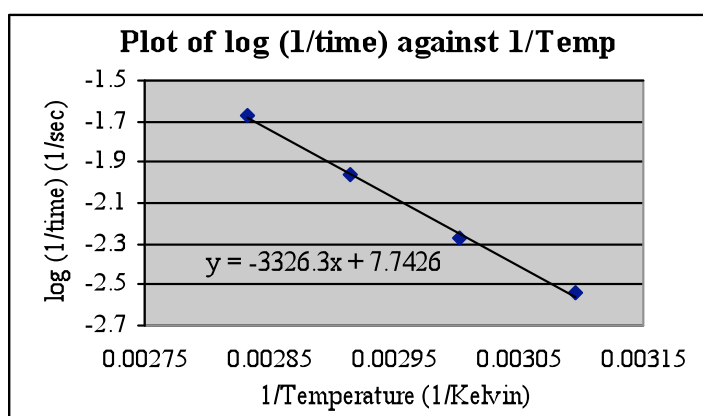
y = -3326.3x + 7.7426

Question (ii):
Your points should form a straight line. The equation of the line has been calculated by the computer. The graph shows the equation of the line is y = -3326.3x + 7.7426. This means the slope, m = -3326.3, while the y-intercept, c = 7.7426.

Question (iii):

In order to answer this question, we must first translate the Arrhenius Equation, $k = Ae^{-E/RT}$, into a form that can be compared to the equation of a straight line, $y = mx + c$. k is the rate constant, A is the Arrhenius Constant, E is activation energy, R is the universal gas constant (8.31 J mol^{-1} K^{-1}), and T is temperature in Kelvin.

We begin by taking the log of both sides:

$$\log k = \log Ae^{\frac{-E}{RT}}$$

Separate the right-hand side:

$$\log k = \log A + \log e^{\frac{-E}{RT}}$$

Bringing the exponent in front of the log:

$$\log k = \log A - \frac{E}{RT}\log e$$

e is a constant with value e = 2.718. R also is a constant, with R = 8.31 J mol^{-1} K^{-1}. If we substitute both these values into the equation, we find:

$$\log k = \log A - 0.052E\left(\frac{1}{T}\right)$$

Rearranging terms, we get:

$$\log k = -0.052E\left(\frac{1}{T}\right) + \log A$$

This equation has the same form as the equation of a straight line, $y = mx+c$. Compare the two equations:

$$\log k = -0.052E\left(\frac{1}{T}\right) + \log A$$

$$y \quad = \quad m \quad x \quad + \quad c$$

We see that if we set $y = \log k$ and $x = 1/T$, then $m = -0.052E$ and $c = \log A$. If you draw a graph with log k on the y-axis, and 1/T on the x-axis, then the slope of the line, $m = -0.052E$ and the y-intercept, c = log A. The units for E will be J/mol and A will be unitless. The rate constant, k, is directly proportional to 1/time. Therefore, instead of using log k on the y-axis it is also possible to use log (1/time). That is what we have done in this question. The resulting graph has the equation $y = -3326.3x + 7.7426$.

Now we can find the value of E and A. We will begin with E. We see in the relation above that $m = -0.052E$. Making E the subject, we find:

$$E = \frac{m}{-0.052}$$

Substituting m = -3326.3, we find E = 63,967 J/mol. Converting into KJ/mol, **E = 63.967 KJ/mol**.

Now we can find A. We know from the relation above that c = log A. To make A the subject, we take the antilog of c: $A = 10^c$. Substituting c = 7.7426, we find **A = 5.5x10^7**.

Note:

It is also possible to construct the graph using natural logarithms instead of base-10 logarithms. You would place ln k (or ln (/time)) on the y-axis and 1/Temp on the x-axis. If you choose to draw this graph, the slope, m = -E/R and the y-intercept, c = ln A. The value of R is 8.31 J mol^{-1} K^{-1}.

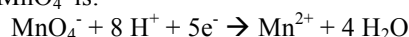# Answers to Question 11 -- NECTA 2000
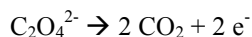
The complete table of results is given below:

Temp	Time	T	1/T	log t
50	37.1	323	0.003096	1.569374
60	14.2	333	0.003003	1.152288
70	6	343	0.002915	0.778151
80	3	353	0.002833	0.477121
90	1.8	363	0.002755	0.255273

Question (a):

The ionic half equation for MnO_4^- is:

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

For $C_2O_4^{2-}$ the half equation is:

$$C_2O_4^{2-} \rightarrow 2\ CO_2 + 2\ e^-$$

Adding them together, we find:

$$2\ MnO_4^- + 16\ H^+ + 5\ C_2O_4^{2-} \rightarrow 2\ Mn^{2+} + 8\ H_2O + 10\ CO_2$$
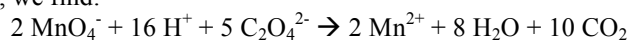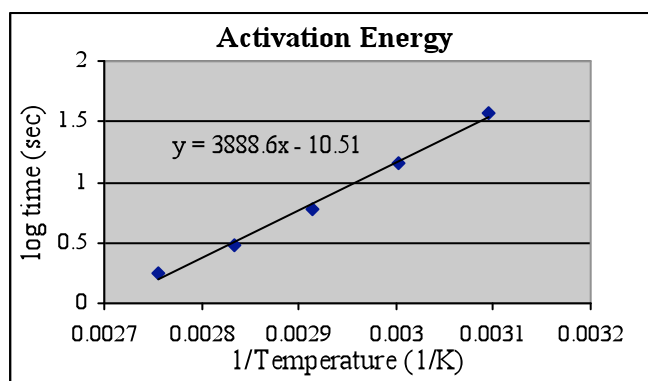
Manganese is reduced from a +7 state to a +2 state, while carbon is oxidized from a +3 state to a +4 state. The reduction of Manganese from +7 to +2 is what causes the solution to go from purple to colourless when the reaction is complete.

Question (b):

The graph of log t (sec) against 1/T (K^{-1}) is shown below:

Note that log time must be on the y-axis, while 1/Temperature must be on the x-axis. The points form a straight line. The equation of the line formed by the points is given on the graph: y = 3888.6x - 10.51. The slope of the line, m = 3888.6, while the y-intercept, c = -10.51.

There are a few things to remember about plotting your graph and drawing your line. First, when you draw your line, try to have the line come as close as possible to all of the points. It is *not* necessary for the line to pass through any of the points. In general, you want to have an equal number of points above the line as below the line. In the graph above, for example, the line passes through one point, above two points, and below two other points. If there is one point that does not fit the trend formed by the other points, it is possible that you have bad data for that one point, and that one point can be ignored. A large amount of human error is involved in drawing this graph, so a range of answers are expected. The slope of the line, for example, could range between 3700 and 4200; all of those answers would be correct.

Question (c):

The question tells us that the slope of the graph is equal to 0.052 Ea mol KJ^{-1} (Note: the units should be mol J^{-1}, not mol KJ^{-1}). For the derivation of this relationship, see the answers for Question 10. The reason m = 0.052 Ea and not m = -0.052 Ea, as in the previous question, is because the graph for this question has log time on the y-axis, while in the previous question it was log (1/time).

In mathematical terms:

$$m = 0.052xE_a$$

Making activation energy the subject:

$$E_a = \frac{m}{0.052}$$

The answer will have units of J/mol. From the graph, we know m = 3888.6. Therefore, E_a = 3888.6/0.052 = **74,780.76 J/mol.**

Answers to Question 12 -- NECTA 1992

Here is the incomplete table of results:

Time (Minutes) T	Vol. Of NN used	Concⁿ of Iodine (M)	1/Concn of Iodine (1/M)
5	12.2		
10	11.1		
15	10.3		
20	9		
25	7.9		
30	6.8		

Question (a):

How were the concentrations of Iodine calculated? We need to use the balanced chemical equation between $S_2O_3^{2-}$ and I_2, along with M_AV_A, to calculate the concentration (in mol/dm³) of Iodine, I_2. M_AV_A tells us:

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$$

Let A = I_2 and B = $S_2O_3^{2-}$. Making M_A the subject:

$$M_A = \frac{M_BV_Bn_A}{n_BV_A}$$

Write what you know. We will start with the data at 5 minutes. M_B = 0.01 M (given in the question), V_B = 12.2 cm³ (volume of sodium thiosulphate added from the burette for the first titration at five minutes), V_A = 10 cm³ (the volume of the pipette used to transfer the reaction mixture to the conical flask). From the balanced equation, n_A = 1 and n_B = 2. Substituting these values into the equation, we find M_A = 0.0061 M. This is the concentration (in mol/dm³) of Iodine in the first titration, which was done after 5 minutes.

You must repeat this calculation for each titration: at 10 minutes, 15 minutes, and so on until 30 minutes. The only thing that will change will be V_B, the volume of sodium thiosulphate. When you finish, your results should be the same as those shown in the table below:

| Time (Minutes) T	Vol. Of NN used	Concⁿ of Iodine (M)	1/Concn of Iodine (1/M)
5	12.2	0.0061	163.9344262
10	11.1	0.00555	180.1801802
15	10.3	0.00515	194.1747573
20	9	0.0045	222.2222222
25	7.9	0.00395	253.164557
30	6.8	0.0034	294.1176471

Question (b):

The two graphs are shown below. The first graph shows time on the x-axis in units of minutes and Concentration of Iodine on the y-axis, in units mol/dm^3:

Plot of Time against Molar Concentration of Iodine
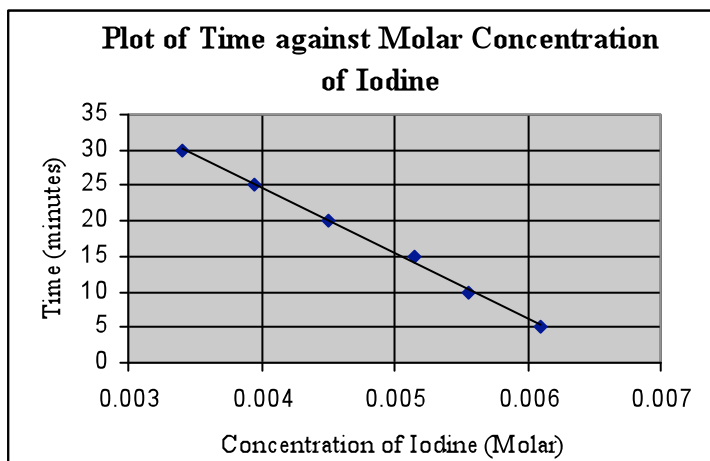
This graph shows a straight line with a negative slope.

The next graph shows a plot of time on the x-axis in units of minutes, and the y-axis is 1/Concentration of Iodine, units of inverse molarity (dm^3/mol):

Plot of Time against inverse Molar Concentration of Iodine (1/M)
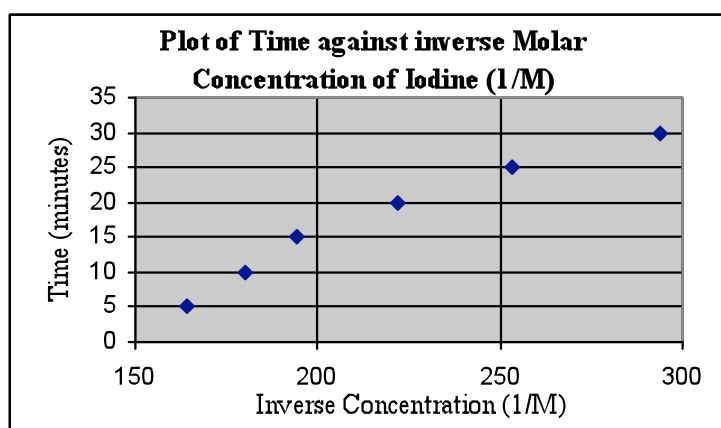
This graph shows a curve, not a straight line.

Question (c):

To determine the order of reaction with respect of Iodine, we must examine the graphs.
Here are some very important rules for the graphs drawn in chemical kinetics:

(i) If you plot a graph of time (y-axis) against concentration (x-axis) and the graph shows a straight line, the reaction is 0-order.

(ii) If you plot a graph of time (y-axis) against log concentration (x-axis) and the graph shows a straight line, then the reaction is 1st-order.

(iii) If you plot a graph of time (y-axis) against inverse concentration (x-axis) and the graph shows a straight line, then the reaction is 2nd-order.

Make sure to memorize these relationships described above!

To draw the above graphs, note that your choice of axes will not affect the shape of the graph. For a 0-order reaction, for example, you can plot a graph of time (y-axis) against concentration (x-axis), *or* concentration (y-axis) against time (x-axis). Both graphs would give a straight line and confirm that the reaction is 0-order.

What do our two graphs show us? We have a graph of time (x-axis) against concentration (y-axis) that shows a straight line with a negative slope. We also have a graph of time (x-axis) against inverse concentration (y-axis), which shows a parabolic curve. Because the first graph shows a straight line, we know that **the reaction is 0-order with respect to Iodine**.

Question (d):
The purpose of MM, $NaHCO_3$, is to stop the reaction by reacting with H^+ and thus removing the acidic condition. The reaction between Iodine and propanone requires acidic conditions in order to proceed. If the acid is removed, the reaction stops. We want to stop the reaction because we want to know the concentration of Iodine at an exact time (5 minutes or 10 minutes or so on). If we do not stop the reaction, our data will not be very accurate, since the reaction between Iodine and propanone will continue, even as we are titrating the Iodine against sodium thiosulphate.

# Answers to Question 13 -- NECTA 1999

The complete results tables are shown below:

Experiment	Sodium Thiosulphate (cm^3)	Water (cm^3)	0.1M HCl (cm^3)	Time (sec)
A	2	8	10	699
B	4	6	10	206
C	6	4	10	124
D	8	2	10	86
E	10	0	10	57

Sodium Thiosulphate (cm^3)	Time (t) Sec	1/Time (1/t) = Rate Sec^{-1}
2	699	0.00143
4	206	0.00485
6	124	0.00806
8	86	0.01163
10	57	0.01754

Question a (i):
The graph of 1/t (vertical axis) against volume of sodium thiosulphate is shown below

Rate against Volume of Sodium Thiosulphate
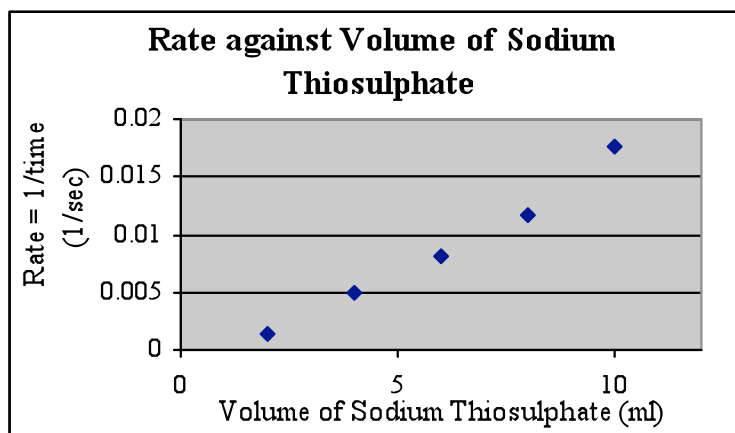
Question a (ii):

To determine the order of reaction from this graph, use the figure below:
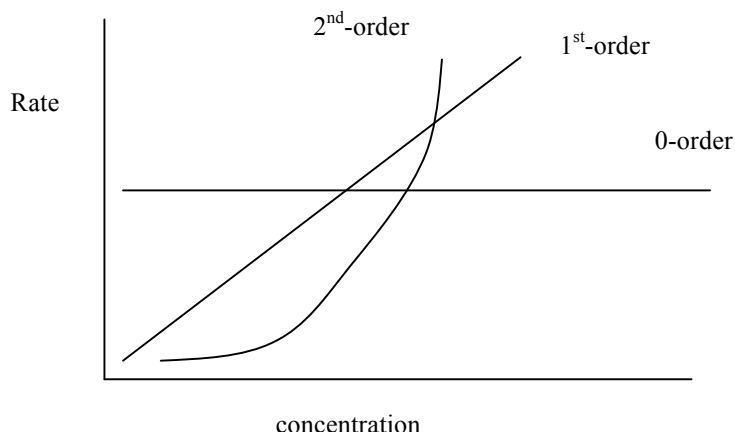
If you plot rate (y-axis) against concentration (x-axis), there are three possible results:

(i) A horizontal line shows that the reaction is 0-order with respect to whichever reactant you are studying

(ii) A straight line with a positive slope shows that the reaction is first-order with respect to whichever reactant you are studying

(iii) A parabolic curve shows that the reaction is 2nd-order with respect to whichever reactant you are studying.

Note that this is only true for graphs of rate against concentration! If you plot a graph of concentration against time, these rules are not true. See Question 12 for how to interpret graphs of concentration against time.

The graph we have plotted in this question is a graph of 1/time against volume of sodium thiosulphate. 1/time is directly proportional to rate, however, and we are told that the volume of sodium thiosulphate is directly proportional to its concentration. The graph we have drawn can therefore be interpreted as a graph of rate against concentration.

Which of the three curves does our graph most closely resemble? Certainly not 0-order. Differentiating between 1st and 2nd-order can be a little difficult. We see that the first four points form a straight line, but the final point deviates from this line and gives the appearance of a parabolic curve. So which will it be?

Because the first four lines shows a straight line, we will assume that there has been an error in the fifth point. The line should be straight, suggesting a 1st-order reaction. If the reaction were truly a 2nd-order reaction, we would expect to see *all* of the points forming a parabolic curve, not only the final point. Because the first four points show a straight line that only the final point deviates from, we say that the graph shows a straight line, indicating a 1st-order reaction.

In general, if your graph shows a general trend that only one point deviates from, it is okay to assume that there is some error in that point and ignore it. The reason we use five points is so that if there is an error in one, we will have four others to show us the correct trend.

Question (b):

The precipitate was caused by the production of solid sulphur, which is insoluble and thus comes out of solution as a precipitate.

Question (c):

The reaction between the acid and sodium thiosulphate is shown below:
$$2\,H^+ + S_2O_3^{2-} \rightarrow S\downarrow + SO_2\uparrow + H_2O$$
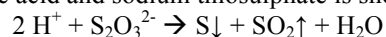
Answers -- Question 14, NECTA 1994

The complete table of results is shown below:

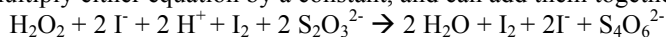| Total volume of EE added (cm^3)	Time, t, at which dark-blur colour appears (sec)	Total calculated volume of solution FF which has reacted (x cm^3)	$(a-x)$, where a = initial volume of FF in the mixture (cm^3)	$\log (a-x)$
2.50	148	1.25	8.75	0.942008
5.00	309	2.50	7.50	0.875061
7.50	509	3.75	6.25	0.79588
10.00	768	5.00	5.00	0.69897
12.50	1168	6.25	3.75	0.574031
15.00	1747	7.50	2.50	0.39794

How did we determine the total calculated volume of FF (H_2O_2) that had reacted? First note that the H_2O_2 and solution EE ($Na_2S_2O_3$) have the same molarity, 0.025 M. This means that equal volumes of the two solutions have equal number of moles, and that the volumes are directly proportional to the number of moles. 10 cm^3 of the H_2O_2 solution have the same number of moles of solute as 10 cm^3 of the $Na_2S_2O_3$ solution.

Now consider the mole ratio between H_2O_2 and $Na_2S_2O_3$. There is a problem: they are not in the same reaction. How can we find the mole ratio? *We add the two equations together so that I_2 cancels*. This will show us the mole ratio between H_2O_2 and $Na_2S_2O_3$.
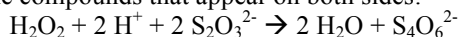
The two reactions are:

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

$I_2 + 2\ S_2O_3{}^{2-} \rightarrow 2I^- + S_4O_6{}^{2-}$
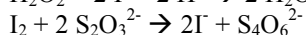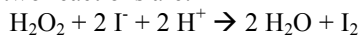
Now we add these equations together in such a way that I_2 cancels. We see that the first reaction has one I_2 on the product side, while the second equation has one I_2 on the reactant side. This means we do not need to multiply either equation by a constant, and can add them together as they are:

$H_2O_2 + 2\ I^- + 2\ H^+ + I_2 + 2\ S_2O_3{}^{2-} \rightarrow 2\ H_2O + I_2 + 2I^- + S_4O_6{}^{2-}$

Cancelling the compounds that appear on both sides:

$H_2O_2 + 2\ H^+ + 2\ S_2O_3{}^{2-} \rightarrow 2\ H_2O + S_4O_6{}^{2-}$

From this equation we can see that the mole ratio between H_2O_2: $S_2O_3{}^{2-}$ is 1:2.

In the table, the first column tells us that 2.50 cm^2 of solution EE ($S_2O_3{}^{2-}$), has been added. This will react with what volume of solution FF (H_2O_2)? Because the volumes are directly proportional to the number of moles, and we know the mole ratio, we can use cross multiplication:

$$\frac{2.5\ cm^3\ S_2O_3{}^{2-}}{2\ mol\ S_2O_3{}^{2-}} = \frac{x\ cm^3\ H_2O_2}{1\ mol\ H_2O_2}$$

So x = 1.25 cm^3 H_2O_2. In the first row of the table, the total volume of $S_2O_3{}^{2-}$ added was 2.50 cm^3, and this will react with 1.25 cm^3 of H_2O_2. We can use this same procedure to calculate the volume of H_2O_2 which has reacted in the following rows of the table, except instead of 2.5 cm3 $S_2O_3{}^{2-}$ you will use 5 cm^3, 7.50 cm^3, and so on. When you have finished, the third column of your table should be identical to the one shown above.

To calculate the values for the fourth column, $a-x$, remember that the initial volume of FF (H_2O_2) added was 10 cm^3. For the first row, this gives 10 cm^3 - 1.25 cm^3 = 8.75 cm^3. Use this same procedure for the rest of the column.

Question (a):
The graph of (a-x) against time, t, is shown below:

Plot of (a-x) against time
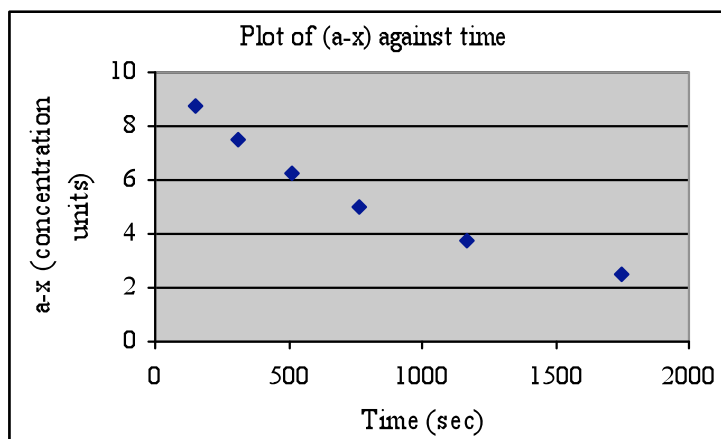
Remember that volume is directly proportional to number of moles, and is thus also directly proportional to concentration. This graph can therefore be interpreted as a graph of concentration (y-axis) against time (x-axis).

Question (b) (i):
The graph of log (a-x) against time, t, is shown below:

Plot of log (a-x) against time
$y = -0.0003x + 0.9759$
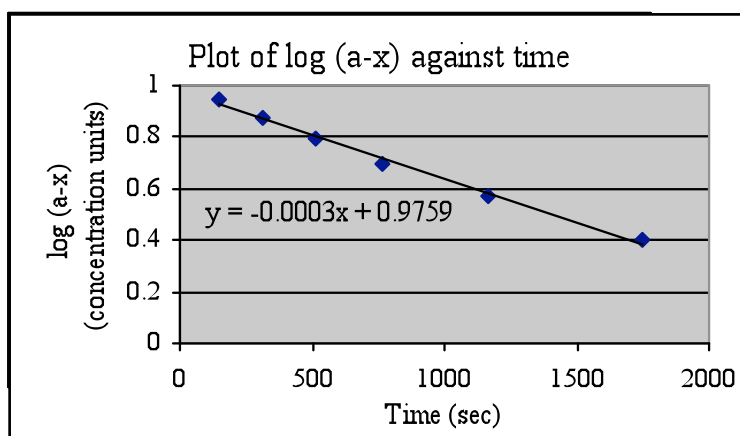
As with the previous graph, volume is directly proportional to concentration, so this graph can be interpreted as a graph of log concentration (y-axis) against time (x-axis).

(ii):
The equation of the line formed by the points is shown on the graph: $y = -0.0003x + 0.9759$. This means the slope, m = -.0003.

Question (c):
To answer this question, refer back to Question 12 and the explanation about how to interpret graphs of concentration against time (or time against concentration).

We know that if a graph of concentration against time gives a straight line, the reaction is 0-order. If a graph of log concentration against time gives a straight line, the reaction is 1st-order. Here we have both of these graphs: Which one shows us a straight line?

The first graph (concentration against time) shows an exponential curve, not a straight line. Note how the curve begins to slope upwards, like the left-hand side of the graph for $y=x^2$, as you move from right to left. This line is not straight. Because this line is not straight, we know the reaction is *not* 0-order.

The second graph (log concentration against time) shows a line that is relatively straight. True, it is not perfectly straight, but it is inevitable that data obtained in practicals will contain some

error. Because the graph of log concentration against time shows a straight line, **this reaction is 1st-order with respect to Hydrogen peroxide.**

Note that the exam essentially provides you with the answer. It asks you to find the slope of the second graph, but not the slope of the first graph. This tells you that the second graph must contain a straight line (only straight lines have a constant slope). *The fact that you can find the slope of the second graph tells you that the second graph will be a straight line*. You cannot find the slope of the line in the first graph because its slope is not constant; it is always changing.

Answers to Question 15 -- NECTA 2006

(a) Results

(i) The volume of the pipette used was 20 cm^3.
(ii) The volume of W$_1$ used is 20.9 cm^3.
(iii) The room temperature is 27 ºC

(b) Results:

(i) Volume of the aqueous layer taken was 20 cm^3.
(ii) Volume of W$_1$ used was 16 cm^3

Before continuing with the questions, it is important to understand what is happening in this practical. The two immiscible solvents are water and diethyl ether (the organic layer). The solute that it distributed between the two layers is succinic acid. Initially, all of the succinic acid is in the water layer. You know this because, at the beginning, you start with an (aqueous) solution of succinic acid: This is succinic acid dissolved in water. At the beginning, the diethyl ether is pure: There is nothing dissolved in it.

When you mix the two layers in the separating funnel, you cause a portion of the succinic acid to be extracted into the ether layer, while the rest remains in the water layer. Your goal in this practical is to determine how much succinic acid has gone into the ether layer (the concentration of succinic acid in the ether) and how much remains in the water layer (the concentration of succinic acid in the water layer). With this information, it will be possible to calculate the value of the partition coefficient, K$_d$.

Question (c):

What is the chemical reaction taking place in procedure 1.(iii)? Procedure 1.(iii) says: "Titrate very carefully solution W$_2$ against W$_1$ until there is a colour change." Solution W2 is the aqueous solution of succinic acid *before being mixed with ether*, while W1 is NaOH. This titration allows you to calculate the concentration of succinic acid in water before the extraction with ether.

Succinic acid is a di-carboxylic acid, which means that it has two acidic hydrogens. The balanced equation is:

2 NaOH + HOOC-CH$_2$-CH$_2$-COOH → NaOOC-CH$_2$-CH$_2$-COONa + 2 H$_2$O
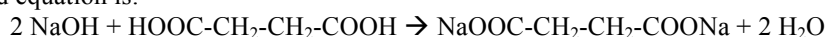
Question (d):

The initial concentration of W2 (succinic acid) in water is the concentration of succinic acid in water *before mixing the water with ether*. We can use the titration performed in Procedure 1 to find this concentration. We will use M$_A$V$_A$, and our subject will be M$_A$, since we are calculating the molarity of the succinic acid.

$$M_A = \frac{M_B V_B n_A}{n_B V_A}$$

M$_B$ = 0.1 M, V$_B$ = 20.9 cm^3 (the volume of NaOH added from the burette), and V$_A$ = 20cm^3 (the volume of acid pipetted into the conical flask). Note that the acid is added from the pipette and the base from the burette: you need to check to see where the acid and base are added from, since the pipette does *not* always hold the base and the burette does *not* always hold the acid. n$_A$ = 1 and n$_B$ = 2. Substituting these values into the equation above, we find M$_A$ = 0.05225 M, or approximately **M$_A$ = 0.052** M. We can convert this into gram concentration by multiplying the molarity by the molar mass. The molar mass of succinic acid is 118 g/mol, so so the gram concentration is 6.136 g/dm^3.

The final concentration of W_2 in the aqueous layer is the concentration of succinic acid in the aqueous layer after mixing with the ether. A portion of the succinic acid has been extracted into the ether layer, so we expect the concentration of succinic acid in the water to decrease.

We can use the titration performed in Procedure 2 to find the concentration of succinic acid in the aqueous layer after the extraction with ether. We will again use $M_A V_A$ with M_A as our subject. M_B = 0.1 M, V_B = 16 cm^3 (the volume of NaOH added from the burette in Procedure 2), and V_A = 20 cm^3 (the volume of the pipette). n_A = 1 and n_B = 2. Substituting these values into $M_A V_A$, we find M_A = **0.04 M**. Converting this to gram concentration gives **4.72 g/dm^3**.

Question (e):

The initial concentration of succinic acid in water was 6.136 g/dm^3, while the final concentration was 4.72 g/dm^3. What has caused the concentration to decrease? During the extraction, some of the succinic acid has been extracted from the water into the ether, so the concentration of succinic acid in the water has decreased.

The concentration of succinic acid in the ether layer is the difference between the initial and final concentration of succinic acid in water, since this difference is the amount that has moved into ether. It can be calculated by subtracting the final concentration from the initial concentration. 6.136 g/dm^3 - 4.72 g/dm^3 = **1.416 g/dm^3** is present in the ether layer after the extraction.

Question (f):

We now have all the information needed to find the partition coefficient. The concentration of succinic acid in water after the extraction is 4.72 g/dm^3. The concentration of succinic acid in ether after the extraction is 1.416 g/dm^3. Note that if we add 4.72 g/dm^3 + 1.416 g/dm^3 = 6.136 g/dm^3, the intitial concentration of succinic acid. The total concentration, or total amount, of succinic acid does not change; it just moves between the two solvents.

The question asks for the "partition coefficient of W_2 between water and diethyl ether." Because water is the solvent mentioned first, the concentration succinic acid in water will be the numerator. Because ether is the solvent mentioned second, the concentration of succinic acid in ether will be the denominator:

$$K_d = \frac{[succinic\ acid]_{water}}{[succinic\ acid]_{ether}}$$

You can use either molarity or gram concentration in this equation: the value of K_d will be the same. We will use gram concentration:

$$K_d = \frac{4.72 \frac{g}{dm^3}}{1.416 \frac{g}{dm^3}} = 3.33$$

So the value of **K_d = 3.33.**

The succinic acid is more soluble in the water than in the ether. Why do you think that is? Think of the principle like-dissolves-like: a polar solvent dissolves a polar solute well, while it does not dissolve a nonpolar solute well. Similarly, a nonpolar solvent dissolves a nonpolar solute well, but does not dissolve a polar solute well. Water is polar and ether is nonpolar. The solute, succinic acid, is polar. Therefore, the succinic acid is more soluble in the water, because both water and succinic acid are polar.