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CAMBRIDGE A LEVEL PROGRAMME

CHEMISTRY 9701

PRACTICAL MANUAL

(SEMESTER 3)

JULY 2012 INTAKES

BOOK 3

Name : ....

Class : ....

FOR INTERNAL USE ONLY

Practical 1

1 **FB 1** is an aqueous solution of sodium thiosulphate, $Na_2S_2O_3$.
FB 2 is dilute hydrochloric acid, HCl.

When aqueous sodium thiosulphate is mixed with hydrochloric acid, a fine suspension of solid sulphur is formed.

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

If a beaker containing the reaction mixture is placed over the printed insert provided, the sulphur slowly hides the printing from view.

If the **depth of solution is constant** the printing will always disappear when the same amount of sulphur has been formed.

You are to carry out experiments to determine the order of reaction with respect to sodium thiosulphate.

(a) Use a $50\,cm^3$ measuring cylinder to transfer $50\,cm^3$ of **FB 1** into a $250\,cm^3$ beaker. Dry the outside of the beaker containing **FB 1** and place it over the printing on the insert sheet.

Measure $5\,cm^3$ of **FB 2** using a small measuring cylinder or a test-tube graduated at $5\,cm^3$.

Pour the $5\,cm^3$ of **FB 2** from the measuring cylinder/test-tube into the beaker and at the same moment start a stop-clock or note the time on a clock with a seconds display.

Swirl the beaker to mix the solutions thoroughly and place it back over the insert. Then view the insert from above so that it is observed through the depth of the solution.

When the printing on the insert just disappears, record the time to the nearest second in Table 1.1.

Empty the beaker and thoroughly rinse with distilled water.

Repeat the experiment using the volumes of **FB 1**, water and **FB 2** shown in Table 1.1.

Table 1.1

expt	volume of **FB 1** / cm^3	volume of water / cm^3	volume of **FB 2** / cm^3	time / s	$\left(\frac{1000}{\text{time}}\right)$ / $s^{-1} \times 10^3$	lg(vol **FB 1**)	$\lg\left(\frac{1000}{\text{time}}\right)$
1	50	0	5			1.70	
2	40	10	5			1.60	
3	30	20	5			1.48	
4	20	30	5			1.30	
5	15	35	5			1.18	

[6]

Calculate $\left(\frac{1000}{\text{time}}\right)$ correct to one decimal place and $\lg\left(\frac{1000}{\text{time}}\right)$ correct to two decimal places for each experiment. Record your values in Table 1.1. [lg = $\log_{10}$] [2]

(b) Plot $\lg\left(\frac{1000}{\text{time}}\right)$ against lg(volume of **FB 1**) and draw the best-fit straight line.

Indicate clearly any point which represents an experimental error and was not considered when drawing the best-fit straight line.

[4]

The rate of reaction with respect to sodium thiosulphate is given by the equation

$$\text{rate} = k\,[Na_2S_2O_3]^{x}$$

where k is the rate constant and x is the order of reaction.

The rate equation can be converted into a log form.

$$\lg(\text{rate}) = x\,\lg[Na_2S_2O_3] + \lg k$$

The rate of reaction can be represented by $\left(\frac{1000}{\text{time}}\right)$, and $[Na_2S_2O_3]$ by the volume of **FB 1** used.

(c) How is the depth of solution kept constant during the experiment?

....

....

.... [1]

(d) Explain why the volume of **FB 1** in each experiment is a measure of $[Na_2S_2O_3]$.

....

....

.... [2]

(e) The equation

$$\lg(\text{rate}) = x\,\lg[Na_2S_2O_3] + \lg k$$

is the equation for a straight line and the order of reaction, x, is the gradient of the line.

Calculate the value of x from the graph you have drawn on page 3.

You must:

draw suitable construction lines on the graph,

show the numbers obtained from the graph and their use in the calculation.

The numerical value of x, the order of the reaction, is .... [3]

ASSESSMENT OF PLANNING SKILLS

The rate of reaction between aqueous sodium thiosulphate solution and hydrochloric acid may also be affected by the concentration of the hydrochloric acid.

(f) Enter, in Table 1.2, the time and the calculated rate, $\left(\frac{1000}{\text{time}}\right)$, from page 2, for *experiment 2*.

Data from two further experiments, *experiments 6 and 7*, together with that from *experiment 2*, can be used to investigate how the rate of reaction varies as the concentration of hydrochloric acid varies. Enter, in Table 1.2, the volumes of **FB 1**, water and **FB 2** that you would use in *experiments 6 and 7* to obtain this data.

Carry out *experiments 6 and 7*, and record the time and $\left(\frac{1000}{\text{time}}\right)$ for each in Table 1.2.

Table 1.2

expt	volume of **FB 1** / cm^3	volume of water / cm^3	volume of **FB 2** / cm^3	time / s	$\left(\frac{1000}{\text{time}}\right)$ / $s^{-1} \times 10^3$
2	40	10	5		
6					
7					

[3]

(g) Use your results to suggest how the rate of reaction depends on the concentration of hydrochloric acid.

....

....

.... [1]

[Total: 22]

2 ASSESSMENT OF PLANNING SKILLS

Copper has two oxides, CuO and Cu_2O.
Each oxide can be reduced to copper metal by heating it in a stream of hydrogen gas.
The oxide turns to copper metal powder in an exothermic reaction in which the powder may be seen to glow red hot.

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

$$Cu_2O(s) + H_2(g) \rightarrow 2Cu(s) + H_2O(g)$$

(a) Draw a diagram to show the assembled apparatus you could use in a school laboratory to carry out the reduction of one of the oxides. The apparatus should enable you to:

(i) weigh the oxide before heating and the metallic copper after heating,

(ii) condense, collect and weigh the steam/water produced,

(iii) burn any excess hydrogen after it has passed through the apparatus.

You may assume that a supply of hydrogen gas is available – you do not have to prepare the gas. You need not draw a balance.

[3]

(b) At the start of the experiment the apparatus is full of air. Hydrogen and air mixtures are explosive.

What precaution could you take to prevent explosion when igniting the excess hydrogen leaving the apparatus?

....

....

.... [1]

(c) Why is it necessary to continue passing hydrogen gas through the apparatus until the copper metal formed has cooled?

....

....

.... [1]

(d) How could you be certain that all of the copper oxide had been reduced to copper?

....

....

.... [1]

(e) Show how you would use the data you could obtain from the experiment in **(a)** to deduce the formula of the oxide used.
[A_r: Cu, 63.5; O, 16.0]

[2]

[Total: 8]

3. Chemical reactions occur more rapidly as the temperature of the reaction mixture increases. The mathematical relationship that summarises this is

$$\log_{10}(\text{rate of reaction}) = \frac{-E_A}{19T}$$

where $\mathbf{E_A}$ is the **activation energy** of the reaction and **T** is the **absolute temperature** in Kelvin and the **rate of reaction** can be taken as the reciprocal of the time taken in seconds **(1/time)**.

An experiment was carried out to investigate this relationship using dilute hydrochloric acid and aqueous sodium thiosulfate.

- $20\,cm^3$ of dilute hydrochloric acid was placed in a boiling tube contained in a water bath.
- $20\,cm^3$ of aqueous sodium thiosulfate was added to the dilute hydrochloric acid, while stirring and a stopwatch started.
- The temperature of the water bath was recorded.
- After a period of time the liquid became cloudy (opaque) due to the formation of a precipitate of sulfur.
- As soon as this cloudiness (opacity) appeared the time was recorded.
- The temperature of the water bath raised and the whole experiment repeated.

(a) The results of several such experiments are recorded below.

Process the results in the table to calculate $\mathbf{log_{10}}$ **(rate of reaction)**, the reciprocal of the absolute temperature (**1/T**) and the 'rate of reaction' (**1/time**). You should expect the values of $\mathbf{log_{10}}$ **(rate of reaction)** to be negative.

Record these values to **three significant figures** in the additional columns of the table.

Label the columns you use. For each column you use include units where appropriate and an expression to show how your values are calculated.

You may use the column headings A to F for these expressions (e.g. A–B). [3]

A	B	C	D	E	F
temperature /°C	absolute temperature /K	time /s			
20.0	293	60.3			
30.0	303	46.8			
40.0	313	41.6			
45.0	318	31.6			
50.0	323	28.8			
55.0	328	25.1			
60.0	333	21.0			
65.0	338	20.4			
70.0	343	18.1			
80.0	353	15.1			

(b) Plot a graph to show the relationship between **$\log_{10}$ (rate of reaction)** and the reciprocal of the absolute temperature. You are reminded that the values for **$\log_{10}$ (rate of reaction)** are negative.
Draw the line of best fit.

[3]

(c) Circle and label on the graph any point(s) you consider to be anomalous. For each anomalous point give a different reason why it is anomalous, clearly stating which point you are describing.

....

....

....

....

....

....

....

....[3]

(d) Comment on whether the results obtained can be considered as reliable.

....

....[1]

(e) Determine the slope of the graph. Mark clearly on the graph any construction lines and show clearly in your calculation how the values from the intercepts were used in the calculation of the slope.

[2]

(f) Using the value of the slope of your graph calculated in **(f)** calculate a value for the **activation energy, E_A**. Correct use of the equation will produce an answer in $kJ\,mol^{-1}$.

[1]

(g) By considering the movement of particles in the reaction explain why the rate of reaction increases with increasing temperature.

[2]

[Total: 15]

Practical 2

1 **FB 1** is 0.200 mol dm^{-3} propanoic acid, $C_2H_5CO_2H$.
FB 2 is 0.100 mol dm^{-3} sodium hydroxide, NaOH.
S is an organic liquid which is immiscible (forms two separate layers) with water.

On shaking **FB 1** with the solvent **S,** propanoic acid is transferred from the aqueous layer to the organic layer until equilibrium is reached.
You are to investigate this equilibrium in the following experiments.

(a) Preparation of the equilibrium mixtures

Experiment A

Use a measuring cylinder to place 50 cm^3 of **FB 1** into the stoppered conical flask labelled **A**.
Use the measuring cylinder to add 20 cm^3 of **S** to the flask and replace the stopper.

Experiment B

Use the measuring cylinder to place 50 cm^3 of **FB 1** into the stoppered conical flask labelled **B**.
Use the measuring cylinder to add 40 cm^3 of **S** into the flask and replace the stopper.

Shake the flasks vigorously for 3 minutes, then leave to stand. **START YOUR ANSWER TO QUESTION 2**, but shake the flasks vigorously for 1 minute after each 5 minutes, returning to Question 1 after a minimum of 15 minutes or when you have completed Question 2. You are attempting to establish an equilibrium mixture of propanoic acid dissolved in water and in solvent **S** in each flask.

One of the two following equilibrium mixtures is established.

I $C_2H_5CO_2H(\text{aqueous layer}) \rightleftharpoons C_2H_5CO_2H(\text{organic layer})$

For this equilibrium, $K_c = \dfrac{[C_2H_5CO_2H(\text{organic layer})]}{[C_2H_5CO_2H(\text{aqueous layer})]}$

II $C_2H_5CO_2H(\text{aqueous layer}) \rightleftharpoons \frac{1}{2}(C_2H_5CO_2H)_2\ (\text{organic layer})$

For this equilibrium, $K_c = \dfrac{\sqrt{[C_2H_5CO_2H(\text{organic layer})]}}{[C_2H_5CO_2H(\text{aqueous layer})]}$

You are to determine which of the two K_c expressions is supported by the results of your experiment.

(b) Titration of Flask A

Allow the layers to separate after the final shake. Fill the burette with **FB 2**. Tilt the flask and carefully pipette 10 cm^3 of the **lower** (aqueous) layer into a titration flask. Place your finger over the top of the pipette, or fit pipette filter if available, before lowering into the solution. This will minimise the amount of the top layer that enters the pipette. Withdraw 10 cm^3 of the lower layer.

Add two drops of phenolphthalein indicator and titrate with **FB 2** until a faint permanent pink colour is obtained.

Repeat the titration two more times and record the results of each titration in Table 1.1 below.

Table 1.1 Titration of Flask A

final burette reading / cm^3 initial burette reading / cm^3			
volume of **FB 2** used / cm^3			

[3] + [3]

Summary

10 cm^3 of the aqueous layer in Flask **A** react with .... cm^3 of **FB 2**.

Show which results you used to obtain this volume of **FB 2** by placing a tick (✔) under the readings in Table 1.1.

(c) Titration of Flask B

Titrate three 10 cm^3 portions of the **lower** (aqueous) layer in the same way as for Flask **A**.
Record the results of each titration in Table 1.2 below.

Table 1.2 Titration of Flask B

final burette reading / cm^3 initial burette reading / cm^3			
volume of **FB 2** used / cm^3			

[1] + [2]

Summary

10 cm^3 of the aqueous layer in Flask **B** react with .... cm^3 of **FB 2**.

Show which results you used to obtain this volume of **FB 2** by placing a tick (✔) under the readings in Table 1.2.

You are advised to show full working in all parts of the calculations.

(d) Calculate how many moles of propanoic acid were contained in 50 cm^3 of **FB 1**.

[1]

(e) For each flask, calculate how many moles of propanoic acid remain in 50 cm^3 of the aqueous layer after shaking with solvent **S**.

$$C_2H_5CO_2H + NaOH \rightarrow C_2H_5CO_2Na + H_2O$$

Flask A	**Flask B**

[1]

(f) For each flask, calculate how many moles of propanoic acid have transferred to the organic layer **S**.

Flask A	**Flask B**

[1]

(g) For each flask, calculate the concentration, in $mol\,dm^{-3}$, of propanoic acid in the aqueous layer.

Flask A	**Flask B**

[1]

(h) For each flask, calculate the concentration, in mol dm^{-3}, of propanoic acid in the organic layer **S**.

Flask A	Flask B

[2]

(i) Use your results to **(g)** and **(h)** to investigate which of the K_c expressions, **I** or **II** on page 2, is correct.

Equation .... is correct.

[2]

(j) Suggest **two** reasons why the calculated values for K_c may still vary even when the correct equilibrium expression is used.

Reason 1 ....

....

Reason 2 ....

....[2]

(k) The concentration of propanoic acid in the organic layer can be determined by direct titration. 10 cm^3 of the organic layer is pipetted into a titration flask. 15 cm^3 of distilled water is added and the mixture is titrated with **FB 2**, shaking between each addition of **FB 2**. Suggest why the 15 cm^3 of distilled water is added to the titration flask.

....

.... [1]

[Total: 20]

2 When a solute is added to two solvents, A and B, which do not mix, some of the solute dissolves in each of the solvents and an equilibrium is set up between the two solvents. At equilibrium the ratio of the two concentrations is a constant known as the **Partition Coefficient**, *K*.

$$\frac{\text{concentration in solvent A}}{\text{concentration in solvent B}} = K$$

An experiment was carried out to determine *K* for succinic acid, $HO_2CCH_2CH_2CO_2H$, between water (boiling point 100 °C) and diethyl ether, $(C_2H_5)_2O$, (boiling point 35 °C).

- $100\,cm^3$ of distilled water and $100\,cm^3$ of diethyl ether were transferred to a conical flask.
- A sample of succinic acid was added, the flask was stoppered and the mixture thoroughly shaken until all of the solid had dissolved.
- A $10.0\,cm^3$ sample of the water layer was removed and titrated with $0.10\,mol\,dm^{-3}$ aqueous sodium hydroxide using phenolphthalein as an indicator.
- A $25.0\,cm^3$ sample of the diethyl ether layer was removed and a small amount of water added. This was then titrated with $0.020\,mol\,dm^{-3}$ aqueous sodium hydroxide using phenolphthalein as an indicator.
- The experiment was repeated using the same volumes of water and diethyl ether but decreasing masses of succinic acid.

(a) The results of the series of titrations are recorded below.

A	B	C	D	E
expt. No.	volume of $0.10\,mol\,dm^{-3}$ NaOH reacting with $10.0\,cm^3$ of the water layer $/cm^3$	volume of $0.020\,mol\,dm^{-3}$ NaOH reacting with $25.0\,cm^3$ of the ether layer $/cm^3$		
1	24.3	18.6		
2	22.5	17.3		
3	20.3	15.6		
4	18.8	13.1		
5	16.3	12.5		
6	13.8	10.6		
7	10.3	7.9		
8	6.8	6.9		
9	5.0	3.8		
10	2.5	1.9		

Process the results in the table to calculate the concentration of the succinic acid in each layer.
Record these values to **three significant figures** in the additional columns of the table.
Label each column, including units and an expression to show how your values are calculated.
You may use the column headings A to E in your expression. [3]

(b) Present the concentration of the succinic acid in each layer in graphical form. Draw the line of best fit.

[3]

(c) Circle on the graph any point(s) you consider to be anomalous.
For any point circled on the graph suggest an error in the conduct of the experiment that might have led to this anomalous result.

....

....

....

....

....

....

....

.... [3]

(d) **(i)** Determine the value of K from your graph. Mark clearly on the graph any construction lines and show clearly in your calculation how the intercepts were used in the calculation of the slope.

(ii) By considering the data you have processed and the graph you have drawn, decide if the experimental procedure described is suitable for the determination of the Partition Coefficient, K. Explain your reasoning.

[3]

(e) In the experimental procedure a small volume of water was added to the diethyl ether prior to the titration with aqueous sodium hydroxide. The flask was constantly shaken during the titrations. What was the purpose of this technique?

....

....

.... [1]

(f) Using a burette, the error associated with a titration depends on the value of the titre. Comment on the magnitude of the titres recorded in the table in **(a)** and indicate, with reasons, which have the highest error.

....

....

.... [2]

[Total: 15]

Practical 3

1 **FB 1** is 2.00 mol dm^{-3} hydrochloric acid, HCl.
FB 2 is a solid carbonate, $\mathbf{X}_2CO_3$ in a stoppered tube.
FB 2 reacts with hydrochloric acid as shown in the equation below.

$$\mathbf{X}_2CO_3(s) + 2HCl\,aq) \rightarrow 2\mathbf{X}Cl\,(aq) + H_2O(l) + CO_2(g)$$

You are to determine the mass of carbon dioxide evolved in the reaction of the carbonate **FB2** with excess hydrochloric acid and to calculate from the results of the experiment the relative atomic mass, A_r, of **X**.

(a) Use a measuring cylinder to place 100 cm^3 of **FB 1** into a 250 cm^3 conical flask. Weigh the flask and acid. Record the mass in Table 1.2.

Weigh the stoppered tube containing **FB 2**. Record the mass in Table 1.1.

Add the weighed **FB 2**, **a little at a time** with swirling, to the acid in the conical flask. **N.B.** Take care to avoid excessive bubbling and loss of acid as 'spray'.

When all of the **FB 2** has been added from the tube, reweigh the empty tube (with its stopper) and record the mass in Table 1.1.

Leave the flask to stand for 2-3 minutes and then reweigh the flask and solution. Record the mass in Table 1.2.

mass of stoppered tube + **FB 2**	/ g	
mass of empty tube + stopper	/ g	

Table 1.1

mass of flask + acid	/ g	
mass of flask + solution after the reaction	/ g	

Table 1.2 [3]

(b) Calculate the mass of **FB 2** used.

.... g [1]

(c) Calculate the mass of carbon dioxide evolved.

.... g [1]

(d) Use your answer to **(c)** and the equation for the reaction to calculate the number of moles of $\mathbf{X}_2CO_3$ that reacted.
[A_r: C, 12.0; O, 16.0.]

.... moles [1]

(e) Calculate the relative molecular mass, M_r, of $\mathbf{X}_2CO_3$.

M_r = ... [1]

(f) Calculate the relative atomic mass, A_r, of **X**.
[A_r: C, 12.0; O, 16.0.]

A_r = ... [1]

[Total : 8]

2 **FB 3** is 1.50 mol dm^{-3} sodium hydroxide, NaOH.
FB 4 is an aqueous solution containing hydrochloric acid.

FB 4 has been prepared by dissolving 42.40 g of the carbonate **FB 2** in an excess of 3.00 mol dm^{-3} hydrochloric acid and making the solution up to 1 dm^3 in a graduated flask by adding more 3.00 mol dm^{-3} hydrochloric acid.

You are to perform a thermometric titration to determine the end-point for the reaction of **FB 3** and **FB 4**. In a thermometric titration the end-point is when the maximum temperature change occurs.

(a) Fill the burette with **FB 4**.

Support the plastic cup in a 250 cm^3 beaker and pipette into the cup 50.0 cm^3 of **FB 3**.

Record the steady temperature of **FB 3** in Table 2.1.

Read through the following instructions before starting the experiment.

Run 3.00 cm^3 of **FB 4** from the burette into the cup, stir the solution with the thermometer and record the new steady temperature. **Without delay** run a further 3.00 cm^3 of **FB 4** from the burette, stir and record the steady temperature as before. Continue the addition of **FB 4** in 3.00 cm^3 portions, taking and recording the steady temperature each time, until 48.00 cm^3 of solution **FB 4** have been run from the burette. Record all temperatures in Table 2.1.

The thermometer provided has a range from .... °C to .... °C

and has graduations at each .... °C.

volume of **FB 4** added /cm^3	temperature /°C	Δt (temperature – initial temperature) /°C	volume of **FB 4** added /cm^3	temperature /°C	Δt (temperature – initial temperature) /°C
0		0	27.00		
3.00			30.00		
6.00			33.00		
9.00			36.00		
12.00			39.00		
15.00			42.00		
18.00			45.00		
21.00			48.00		
24.00					

Table 2.1 [5]

(b) Plot a graph of Δ**t** against the volume of **FB 4** added.
Draw two smooth curves through the plotted points to find the end-point for the titration.
[3]

(c) Read from the graph the volume of hydrochloric acid, **FB4**, at the end-point of the titration.

.... cm^3 [1]

(d) Use your answer to **(c)** and the equation for the reaction to calculate the concentration of the hydrochloric acid in **FB 4**.

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

.... [1]

(e) The solution **FB 4** was prepared by dissolving 42.40 g of $\mathbf{X}_2CO_3$ in 1 dm^3 of 3.0 mol dm^{-3} HCl.

Use this information and your answer to **(d)** to calculate the number of moles of HCl that reacted with the dissolved $\mathbf{X}_2CO_3$.

.... [1]

(f) Calculate the relative molecular mass, M_r, of $\mathbf{X}_2CO_3$.

Calculate the relative atomic mass, A_r, of **X**.
[A_r: C, 12.0; O, 16.0.]

A_r = ... [1]

[Total : 12]

3 ANALYSIS AND EVALUATION

(a) Indicate the size of the error you would expect in making measurements with the thermometer in question 2.

....

....

....[1]

(b) Why is it **not** necessary to consider the errors in the measuring cylinder used in question 1?

....

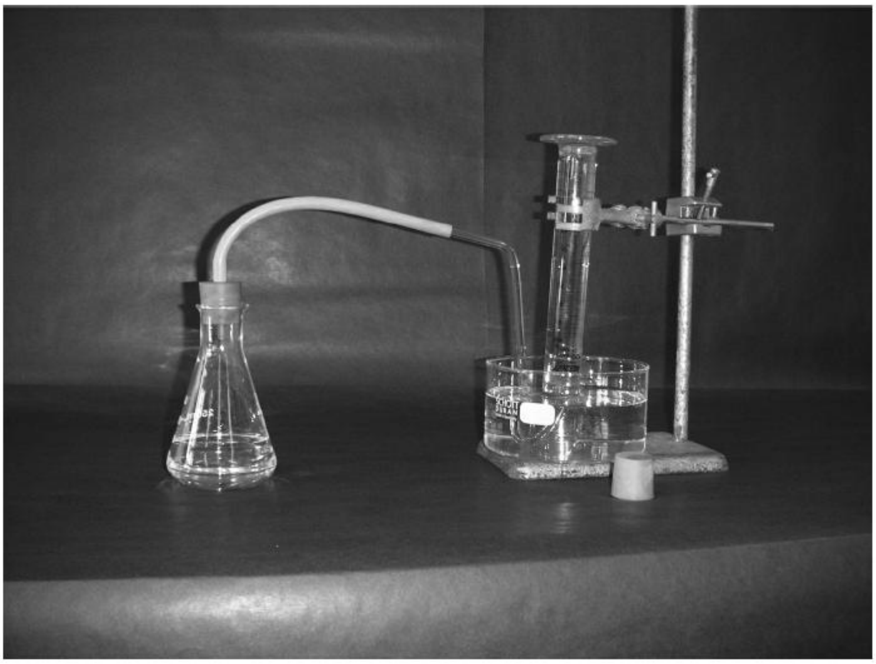
....

....[1]

ASSESSMENT OF PLANNING SKILLS

The relative atomic mass of **X** can also be determined by an experiment in which carbon dioxide is evolved in the reaction between a weighed sample of the carbonate of **X** and excess hydrochloric acid. The sample is added to the acid and the bung quickly re-inserted in the flask. The gas displaced from the apparatus is collected and its volume measured using the apparatus below.

The acid is placed in a 250 cm^3 conical flask and the gas collected by displacing water from an inverted 100 cm^3 measuring cylinder.



(c) In both Question 1 and the experiment above, the accuracy of the measured mass or volume can be improved by dissolving a small quantity of sodium carbonate in the acid before the start of the experiment.

Suggest the reason behind this.

....

....

....

....

....

....[2]

(d) For the experiment shown in the photograph, suggest **two other major** sources of error (not including that mentioned in part **(c)**) in the spaces provided on page 10.

For each source of error describe a method of reducing this error, explaining the reasoning for your method.

The data below may be relevant.

Data

Solubility of carbon dioxide gas at different temperatures

temperature / °C	solubility /g CO_2 per 100 g of water
0	0.348
25	0.145
40	0.097
60	0.058

The general gas equation $pV = nRT$

The specific heat capacity of water $= 4.18\ J\ g^{-1}\ K^{-1}$

Vapour pressure of water at different temperatures

temperature / °C	Vapour pressure /Pa
20	2388
25	3167
30	4243
35	5623
40	7376

Technical data for X_2CO_3

Minimum assay	99%
Substances insoluble in water	0.0025%
Water	0.35%
Arsenic	0.0001%
Lead	0.003%
Sulphate	0.02%
Iron	0.001%

(i) first major source of error

....

method for reducing this error

....

....

explanation

....

....[3]

(ii) second major source of error

....

....

method for reducing this error

....

....

explanation

....

....[3]

[Total : 10]

4. When potassium nitrate dissolves in water, the temperature of the solution goes down because the enthalpy of solution is endothermic.

You are to plan an experiment to investigate how the solubility of potassium nitrate varies with temperature. The units of solubility are grams per one hundred grams of water (g/100g water).

(a) **(i)** Predict how the solubility of potassium nitrate will change if the solution temperature is **increased**.

Explain your prediction using the fact that dissolving potassium nitrate is endothermic.

prediction ....

....

explanation ....

....

....

....

(ii) Display your prediction in the form of a sketch graph, labelling clearly the axes.

[3]

(b) In the experiment you are about to plan, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....

[2]

(c) Design a laboratory experiment to test your prediction in **(a)**.

In addition to the standard apparatus present in a laboratory you are provided with the following materials,

a boiling tube,
a looped wire stirrer,
a thermometer covering the temperature range 0 °C to 100 °C.

Describe how you would carry out the experiment. You should

- ensure a wide range of results suitable for analysis by graph,
- decide on the amounts of water and potassium nitrate to use,
- measure the amounts of the two reagents,
- heat the apparatus,
- decide at what point the temperature of the solution is to be taken.

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....[7]

(d) State a hazard that must be considered when planning the experiment and describe precautions that should be taken to keep risks to a minimum.

....

....[1]

(e) Draw a table with appropriate headings to show the data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in **(a)**. The headings **must** include the appropriate units.

[2]

[Total: 15]

Practical 4

1 If a container of gas has a tiny hole in it, the gas will gradually escape through the hole. This process is called **effusion** and the rate at which it occurs is called **the rate of effusion**.

> You are to plan an experiment to investigate how the **rate of effusion** depends on the **relative molecular mass**, M_r, of a gas.

(a) At a constant temperature, the rate of effusion of a gas depends on the kinetic energy of the molecules of the gas. So, for a series of gases all at the same temperature, as the M_r of a gas increases the speed of the molecules of the gas decreases.

(i) Predict how the rate of effusion will change as the M_r of the gas **increases**. Explain your prediction using the information in part **(a)** above.

prediction ....

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explanation ....

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(ii) Display your prediction in the form of a sketch graph below, clearly labelling the axes.

[3]

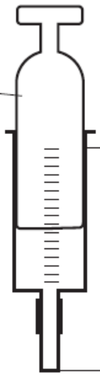
(b) In the experiment you are about to plan, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....

[2]

(c) Using the apparatus shown below design a laboratory experiment to test your prediction in **(a)**.



In addition to the standard apparatus present in a laboratory you are provided with the following materials,

- access to samples of the following gases; hydrogen, oxygen, carbon dioxide, butane and chlorine,
- a stop watch/clock.

Describe how you would carry out the experiment. You should

- ensure that the volume of gas measured is the same for each experiment,
- ensure that the syringe contains only the gas under investigation,
- ensure that the syringe is used under the same conditions throughout all of the experiments,
- measure the effusion time,
- produce reliable results.

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.... [6]

(d) State a hazard that must be considered when planning the experiment and describe precautions that should be taken to keep risks to a minimum.

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.... [2]

(e) Draw a table with appropriate headings to show the data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in **(a)**. The headings **must** include the appropriate units. Ensure that the table covers all the detail relating to the five gases listed in **(c)**.

[A_r: H, 1.0; C, 12.0; O, 16.0; C*l*, 35.5]

[2]

[Total: 15]

2 There are three oxides of lead, PbO, PbO_2 and Pb_3O_4 all of which can be reduced to metallic lead by hydrogen. A sample of one of these oxides is reduced to find out which of the three oxides it is.

An experiment was carried out as follows.

- An empty reduction tube was weighed and the mass recorded.
- A sample of the lead oxide was added to the reduction tube and the new mass recorded.
- The reduction tube and lead oxide was heated strongly for five minutes in a stream of hydrogen and then allowed to cool back to room temperature.
- The reduction tube and contents were then reweighed and the mass recorded.

(a) The results of several such experiments are recorded below.

[A_r: O, 16.0; Pb, 207.0]

Process the results in the table to calculate the number of moles of lead **atoms** and the number of moles of oxygen **atoms**.

Record these values in the additional columns of the table. You may use some or all of the columns. Label the columns you use.

Masses should be recorded to **two decimal places** while the number of moles should be recorded to **two significant figures**.

For each column you use include units where appropriate and an expression to show how your values are calculated.

You may use the column headings A to G for these expressions (e.g. A–B). [3]

A	B	C	D	E	F	G
mass of reduction tube /g	mass of reduction tube + lead oxide /g	mass of reduction tube + lead /g				
9.90	14.95	14.48				
10.05	16.17	15.60				
10.25	17.92	17.21				
9.80	18.12	17.43				
9.60	18.43	17.61				
10.30	20.27	19.34				
11.05	22.05	21.03				
10.00	21.46	20.26				
9.75	24.07	22.74				
10.15	26.15	24.66				

(b) Plot a graph to show the relationship between the number of moles of oxygen **atoms** and the number of moles of lead **atoms**.
Draw the line of best fit.

[3]

(c) Circle and label on the graph any point(s) you consider to be anomalous. For each anomalous point give a different reason why it is anomalous, clearly stating which point you are describing.

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(d) Comment on whether the results obtained can be considered as reliable.

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(e) Determine the slope of the graph. Mark clearly on the graph any construction lines and show clearly in your calculation how the values from the intercepts were used in the calculation of the slope.

[2]

(f) Comment on the value of the slope of the graph.
Deduce and explain the formula of the oxide investigated in this experiment.

comment ....

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deduction and explanation ....

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.... [3]

[Total: 15]

Practical 5

1. The carbonates of group II in the periodic table decompose on heating forming an oxide and carbon dioxide.
X is any group II cation (e.g. Mg^{2+})

$$XCO_3 \longrightarrow XO + CO_2$$

This decomposition occurs because the positively charged cations polarise (distort) the C—O bond in the carbonate ion causing the ion to break up. The charge density of the group II cations decreases down the group. This affects the decomposition rate.

You are to plan an experiment to investigate how the rate of decomposition of a group II carbonate varies as the group is descended. The rate can be conveniently measured by finding the time taken to produce the same volume of carbon dioxide from each carbonate.

(a) **(i)** Predict how the rate of decomposition of the group II carbonates will change as the group is descended.

Explain this prediction in terms of the charge density of the cation as the group is descended.

prediction ....

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explanation ....

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(ii) Display your prediction in the form of a sketch graph, clearly labelling the axes.

[3]

(b) In the experiment you are about to plan, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....

[2]

(c) Draw a diagram of the apparatus and experimental set up you would use to carry out this experiment. Your apparatus should use only standard items found in a school or college laboratory and show clearly the following.

(i) the apparatus used to heat the carbonate

(ii) how the carbon dioxide will be collected

Label each piece of apparatus used, indicating its size or capacity.

[2]

(d) Using the apparatus shown in **(c)** design a laboratory experiment to test your prediction in **(a)**.

In addition to the standard apparatus present in a laboratory you are provided with the following materials,

samples of the carbonates of magnesium, calcium, strontium and barium,
a stop-watch/clock with second hand.

Give a step-by-step description of how you would carry out the experiment by stating

(i) the gas volume you would collect from each carbonate,

(ii) how you would calculate the mass of each carbonate to ensure that this volume of carbon dioxide is produced,

(iii) how you would control the factors in the heating so that different carbonates can be compared.

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(e) State a hazard that must be considered when planning the experiment and describe precautions that should be taken to keep risks to a minimum.

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(f) Draw a table with appropriate headings to show the data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in **(a)**. The headings **must** include the appropriate units.

[2]

(g) This simple experiment is likely to produce only approximate results.
Suggest an improvement to your apparatus or an alternative apparatus that may improve the reliability of the results.

[1]

[Total: 16]

2. When sodium nitrate, $NaNO_3$, is heated, it decomposes into sodium nitrite, $NaNO_2$, and oxygen.
A suggested equation is:-

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

An experiment was carried out to attempt to confirm this.

- An empty boiling tube was weighed and the mass recorded.
- A sample of sodium nitrate was added to the boiling tube and the new mass recorded.
- The boiling tube and sodium nitrate was heated strongly for five minutes and then allowed to cool back to room temperature.
- The boiling tube and contents was then reweighed and the mass recorded.

(a) Calculate the relative molecular masses (M_r) of $NaNO_3$ and $NaNO_2$.
[A_r: N, 14.0; O, 16.0; Na, 23.0]

[1]

(b) The results of several such experiments are recorded below.

A	B	C	D	E	F	G
mass of boiling tube / g	mass of boiling tube + $NaNO_3$ / g	mass of boiling tube + $NaNO_2$ / g				
9.90	13.10	12.50				
10.05	14.73	13.91				
10.25	14.20	13.46				
9.80	12.67	12.65				
9.60	14.56	13.63				
10.30	15.80	14.76				
11.05	17.18	15.50				
10.00	17.00	15.68				
9.75	17.65	16.16				
10.15	18.48	16.84				

Process the results in the table to calculate the number of moles of sodium nitrate and the number of moles of sodium nitrite.

Record these values in the additional columns of the table. You may use some or all of the columns.

Masses should be recorded to **two decimal places**. Numbers of moles should be recorded to **two significant figures**.

Label the columns you use. For each column you use include units where appropriate and an expression to show how your values are calculated. You may use the column headings A to G for these expressions (e.g. A–B). [2]

(c) Plot a graph to show the relationship between the number of moles of sodium nitrate and the number of moles of sodium nitrite.
Draw the line of best fit.

[3]

(d) Circle and label on the graph any point(s) you consider to be anomalous.
For each anomalous point give a different reason why it is anomalous clearly indicating which point you are describing.

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....[3]

(e) Determine the slope of the graph. Mark clearly on the graph any construction lines and show clearly in your calculation how the intercepts were used in the calculation of the slope.

[3]

(f) **(i)** Does the value of the slope of your graph calculated in **(e)** confirm the equation given in **(a)** or not?

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(ii) Explain your answer in **(f)(i)** above.

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[2]

[Total: 14]

Practical 6

1 Reactions involving two aqueous solutions are dependent on collisions occurring between the particles of the two reagents.

As the temperature of the system is raised, the average kinetic energy of the particles increases.

You are to plan an experiment to investigate how the rate of the reaction between hydrochloric acid and sodium thiosulfate, $Na_2S_2O_3$, depends on the temperature of the reaction. When these two reagents react, after a short period they slowly produce a white or yellow precipitate of sulfur. As more sulfur is produced, the reaction mixture becomes more cloudy until it cannot be seen through (i.e. it is opaque). The time taken for the mixture to become opaque can be dependent on the relative concentrations of the reagents or the temperature of the reaction mixture.

(a) **(i)** Predict how the rate of reaction will change if the temperature of the reagents is **increased**. Using the idea of how the kinetic energy of the particles changes as the temperature of the reagents **increases**, explain your prediction in terms of particle collisions.

prediction ....

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explanation ....

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(ii) Display your prediction in the form of a sketch graph, clearly labelling the axes.

[3]

(b) In the experiment you are about to plan, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....
[2]

(c) Draw a diagram of the apparatus and experimental set up you would use to carry out this experiment. Your apparatus should use only standard items found in a school or college laboratory and show clearly the following.

(i) the apparatus used as the reaction vessel and how the thermometer will be positioned in order to measure the temperature of the solution as accurately as possible

(ii) how the solution will be heated

Label each piece of apparatus used, indicating its size or capacity and the temperature range that the thermometer should cover.

[2]

(d) Using the apparatus shown in **(c)** design a laboratory experiment to test your prediction in **(a)**.

In addition to the standard apparatus present in a laboratory you are provided with the following materials,

$0.100 \text{ mol dm}^{-3}$ aqueous sodium thiosulfate,
1.00 mol dm^{-3} hydrochloric acid,
A stop-watch/clock with second hand.

Give a step-by-step description of how you would carry out the experiment by stating

(i) the number of experiments you would do, and their temperature range (minimum and maximum temperatures),

(ii) what you would keep constant in all the experiments,

(iii) what temperature measurements you would make,

(iv) how you would use the cloudiness (opacity) of the reaction mixture to measure the time taken for each reaction.

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(e) State a hazard that must be considered when planning the experiment and describe precautions that should be taken to keep risks to a minimum.

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.... [1]

(f) Draw a table with appropriate headings to show the data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in section **(a)**. The headings **must** include the appropriate units.

[2]

[Total: 15]

2 The solubility of potassium chlorate(V) in water increases with temperature. The units of solubility are grams per one hundred grams of water (g/100g water).

An experiment is carried out to investigate this solubility.

- An empty boiling tube was weighed and the mass recorded.
- Some distilled water was added to the boiling tube and the new mass recorded.
- A small sample of potassium chlorate(V) was added and this new mass recorded.
- The boiling tube was carefully heated with stirring until all the solid had dissolved.
- The apparatus was allowed to cool slowly while constantly stirring and the temperature recorded when the first crystals appeared in the tube.

(a) The results of several such experiments are recorded below.

A	B	C	D	E	F	G
crystallising temperature / °C	mass of boiling tube /g	mass of boiling tube and water /g	mass of boiling tube, water and solid /g			
20.0	10.10	35.10	36.85			
25.0	10.20	35.20	37.45			
30.0	9.80	29.20	31.20			
40.0	9.95	32.95	36.55			
45.0	10.35	30.35	33.45			
50.0	9.90	34.90	39.40			
60.0	9.70	30.70	35.53			
65.0	9.95	33.95	40.07			
70.0	10.45	30.45	36.15			
75.0	10.35	35.35	42.75			
80.0	10.05	35.05	44.05			
90.0	10.10	40.10	53.90			

Process the results in the table to calculate the solubility in g/100g of the potassium chlorate(V) for each of the temperatures listed.

Record these values to **two decimal places** in the additional columns of the table. You may use some or all of the columns.

Label the columns you use.

For each column you use include units where appropriate and an expression to show how your values are calculated.

Use the column headings A to G for these expressions (e.g. A–B). [3]

(b) Plot a graph to show the variation of solubility with temperature.
Draw the line of best fit.

[4]

(c) Circle and label on the graph any point(s) you consider anomalous.
For each anomalous point give a different reason why it is anomalous clearly indicating which point you are describing.

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(d) A solution of potassium chlorate(V) is made up using 50 g of water. This is found to be saturated at 85 °C. The solution is then cooled to 35 °C. Using your graph calculate the mass of solid deposited as a result of this temperature change.

[2]

(e) From the pattern of solubility demonstrated by your graph predict and explain whether the dissolving of potassium chlorate(V) in water is an exothermic or an endothermic reaction.

prediction

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explanation

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[Total: 15]

Practical 7

1 When aqueous sodium chloride, NaCl, is added to aqueous lead nitrate, $Pb(NO_3)_2$, a white precipitate of lead chloride, PbCl_2, is produced. A suggested stoichiometric equation is

$$Pb(NO_3)_2(aq) \quad + \quad 2NaCl(aq) \longrightarrow PbCl_2(s) \quad + \quad 2NaNO_3(aq)$$

In separate experiments, different volumes of 0.20 mol dm^{-3} aqueous sodium chloride are added to a fixed volume of 0.10 mol dm^{-3} aqueous lead nitrate. In each case, the precipitate is filtered, washed with distilled water and thoroughly dried. The mass of the precipitate is recorded.

You are to plan an experiment to investigate this reaction in order to confirm or reject the stoichiometry of the equation.

(a) By considering the suggested stoichiometric equation, predict and explain how the number of moles of the precipitate, PbCl_2, will change as the number of moles of NaCl added increases.

Prediction ....

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Explanation ....

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.... [2]

(b) State a limiting factor that must be taken into account when increasing the volume of the aqueous sodium chloride added.

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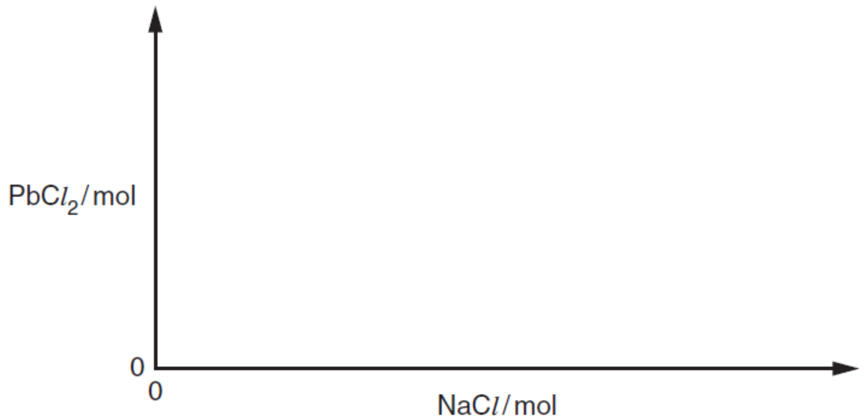
Sketch the graph which would result if, after some of the experiments, the NaCl is in excess. Start your graph with no NaCl added.



[3]

(c) In the experiment you are about to plan, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....

(iii) another variable to be controlled ....

[2]

(d) Design a laboratory experiment to test your prediction in **(a)**.

You are provided with 250 cm^3 of 0.20 mol dm^{-3} aqueous sodium chloride.

(i) Outline how you would prepare 250 cm^3 of 0.10 mol dm^{-3} aqueous lead nitrate.

[A_r: N, 14; O, 16; Pb, 207]

(ii) Give a step by step description of how you would carry out **one** experiment.
You should state

- the volumes of each solution to be used,
- how the volumes will be measured,
- how you would dry the precipitate.

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(e) In the table below

- enter appropriate headings to show additional data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in **(a)**. The headings should include the appropriate units,
- enter the volumes from your plan in **(d)**,
- enter suitable volumes for four further experiments.

[2]

(f) How would you ensure that at the end of each experiment the precipitate was thoroughly dried?

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.... [1]

[Total: 16]

2 The melting point of solid water is 0 °C. This is the same as the freezing point of water. This freezing point can be lowered (depressed) by the addition of a solute, such as glucose. The extent of the freezing point depression depends on the **number of particles of solute dissolved** in the solution.

The freezing point depression, ΔT_f, is proportional to the molal concentration, c_m, of the solution.

$$\Delta T_f = K_f c_m$$

where K_f is the freezing point depression constant.

> ***The molal concentration (molality) of a solution is defined as the number of moles of a solute dissolved in one kilogram of water e.g. a one molal solution has one mole of solute dissolved in one kilogram of water.***

An experiment was carried out to investigate the relationship between ΔT_f and c_m.

- A weighed sample of distilled water was placed in a boiling tube.
- A weighed sample of glucose was added.
- The mixture was stirred until a solution was obtained.
- The tube was placed in a freezing apparatus to lower the temperature.
- The freezing point of the solution was measured precisely and the freezing point depression calculated.

(a) Calculate the M_r of glucose $C_6H_{12}O_6$.

$[A_r$: H, 1.0; C, 12.0; O, 16.0] [1]

(b) The results of the experiment are recorded below.

A	B	C	D	E	F
mass of water / g	mass of glucose / g	freezing point depression ΔT_f / °C			
100	10.0	1.03			
100	12.2	1.26			
100	18.0	2.09			
100	23.3	2.40			
100	27.7	2.86			
100	30.9	3.22			
100	33.1	3.31			
100	38.6	3.98			
100	42.3	4.37			

Process the results in the table to calculate the molality of the glucose solution. This will enable you to plot a graph to show how the freezing point depression, ΔT_f, varies with the molality of the solution.
Record these values to **three significant figures** in the additional columns of the table.
You may use some or all of the columns.
Label the columns you use.
For each column you use include units where appropriate and an expression to show how your values are calculated. You may use the column headings A to F for this purpose. [2]

(c) Present the data calculated in **(b)** in graphical form. Draw the line of best fit.

[3]

(d) Circle on the graph any point(s) you consider to be anomalous.
For any point circled on the graph suggest an error in the conduct of the experiment that might have led to this anomalous result.

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.... [3]

(e) **(i)** Determine the value of $\Delta T_f / c_m$ from your graph. This is the freezing point depression constant K_f. Mark clearly on the graph any construction lines and show clearly in your calculation how the intercepts were used in the calculation of the slope.

(ii) By considering the data you have processed and the graph you have drawn, decide if the experimental procedure described is suitable for the determination of the freezing point depression constant K_f. Explain your reasoning.

[3]

(f) When the experiment was repeated using sodium chloride instead of glucose as the solute, the freezing point depressions were found to be twice the value obtained in the glucose experiment for each molality.
Using the information given at the start of the question suggest a reason for this.

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.... [1]

(g) Using your suggestion from **(f)** predict the effect on the freezing point depression if a weak acid such as ethanoic acid was used instead of glucose or sodium chloride as the solute.

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.... [1]

[Total: 14]

Practical 8

1 When ammonium nitrate(V), NH_4NO_3, is heated it decomposes completely into nitrogen(I) oxide, N_2O, and water vapour, H_2O, which if allowed to cool will condense to liquid water.

The stoichiometric equation for this decomposition is

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l).$$

The following information gives some of the hazards associated with ammonium nitrate(V).

Ammonium nitrate(V) NH_4NO_3
Oxidising: Contact with combustible material may cause fire. Explosive when mixed with combustible material.
Do not allow the salt to become contaminated with organic matter and do not grind it.

You are to plan an experiment to investigate the molar ratio of nitrogen(I) oxide and ammonium nitrate(V) at 25 °C, and confirm that it remains unchanged as the mass of ammonium nitrate(V) changes.

(a) (i) Predict quantitatively how the number of moles of nitrogen(I) oxide varies as the number of moles of ammonium nitrate(V) increases, if the products are measured at room temperature 25 °C.

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(ii) Predict quantitatively how the sum of the number of moles of water vapour and nitrogen(I) oxide varies as the number of moles of ammonium nitrate(V) increases, if the products are measured at 110 °C.

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(iii) Display both your predictions in the form of sketch graphs on the axes below. Label clearly each axis and each graph line.

[3]

(b) In the experiment you are about to plan to test your prediction in **(a)(i)** at 25 °C, identify the following.

(i) the independent variable ....

(ii) the dependent variable ....

[2]

(c) Draw a diagram of the apparatus and the experimental set up you would use to carry out this experiment. Your apparatus should use only standard items found in a school or college laboratory and show clearly

(i) how the solid will be heated,

(ii) how the water vapour will be condensed into a liquid and collected. Ice is available,

(iii) how the nitrogen(I) oxide will be collected.

Label each piece of apparatus used, indicating its size or capacity.

[3]

(d) Using the apparatus shown in **(c)** design a laboratory experiment to test your prediction in **(a)(i)** for an experiment at 25 °C.

In addition to the standard apparatus present in a laboratory you are provided with the following materials.

a sample of solid ammonium nitrate(V)
crushed ice

Give a step-by-step description of how you would carry out the experiment,

(i) to produce enough results to give sufficient data to plot a graph as in **(a)(iii)**,
(ii) by stating the volumes of nitrogen(I) oxide you would collect,
(iii) by calculating the mass of ammonium nitrate(V) needed to produce one of the volumes of nitrogen(I) oxide suggested in **(ii)**,
(iv) by stating how you would ensure that decomposition was complete.
[A_r : H, 1.0; N, 14.0; O, 16.0; the molar volume of a gas at 25 °C, 24.0 dm^3]

[4]

(e) State one hazard that must be considered when planning the experiment and describe a precaution that should be taken to minimise the risk from this hazard.

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....[1]

(f) Draw a table with appropriate headings to show the data you would record when carrying out your experiments and the values you would calculate in order to construct a graph to support or reject your prediction in **(a)(i)**. The headings should include the appropriate units.

[2]

[Total: 15]

2 The variation of the volume with pressure of a fixed mass of any ideal gas at constant temperature may be represented by a relationship known as Boyle's law,

$$PV = \text{constant}$$

where P is the pressure of the gas, V is the volume of the gas.

A gas such as carbon dioxide, under certain conditions of temperature and pressure, does not always behave as an ideal gas.

An experiment was carried out on carbon dioxide to investigate its behaviour.

- A calibrated glass tube was filled with a sample of carbon dioxide.
- The tube was attached to a calibrated pressure pump.
- The pressure and the volume of the gas sample were recorded.
- The measured pressure on the gas was increased and the new volume recorded.

(a) The results of the experiment are recorded in the table below.

A graph of V against P has been plotted for you.

Process the results in the table to enable you to plot two further graphs:

- PV against P
- 1/V against P

Record these values to **three significant figures** in the additional columns of the table.

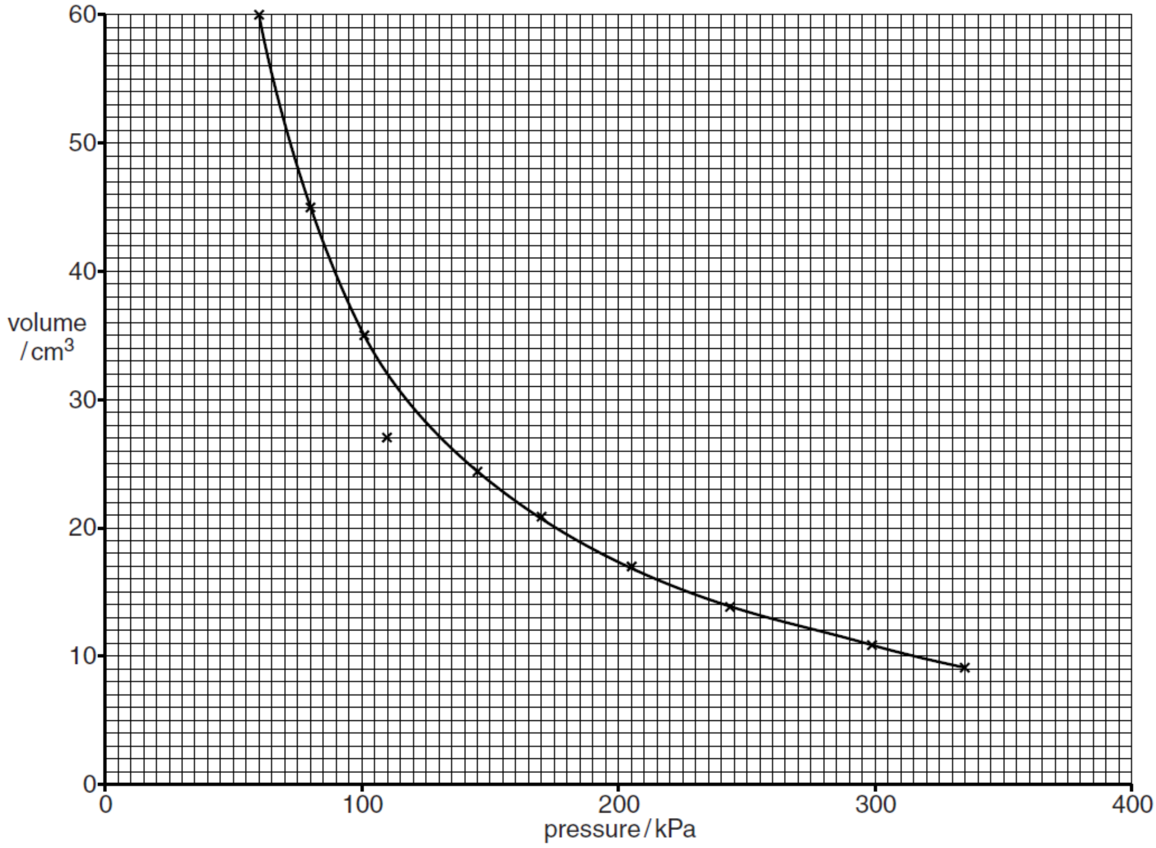
Label the columns you use. For each column you use include units where appropriate and an expression to show how your values are calculated.

You may use the column headings A to D for these expressions (e.g. A–B).

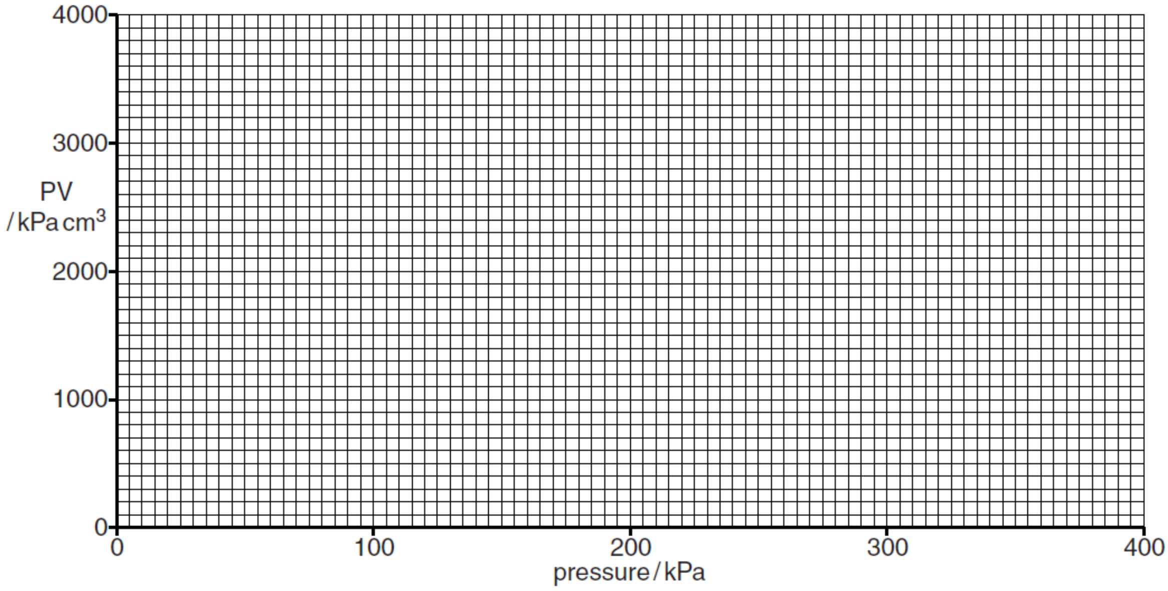
A	B	C	D
pressure of the gas/kPa	volume of the gas/cm^3		
335	9.09		
298	10.9		
243	13.9		
205	17.0		
170	20.8		
145	24.4		
110	27.0		
101	35.0		
80.0	45.5		
60.0	60.0		

[2]

This graph shows the relationship between the volume of the gas and the pressure of the gas.



(b) Plot a graph to show the relationship between the product of the pressure and volume, PV, of the gas and the pressure, P, of the gas. Draw the line of best fit.



[1]

(c) Plot a graph to show the relationship between the reciprocal of the volume of the gas and the pressure (1/V against P) of the gas. Begin the scales on both axes at 0. Draw the line or curve of best fit.

[3]

(d) Circle and label on the graph in **(c)** any point(s) you consider to be anomalous.
For each anomalous point give a different reason why it is anomalous, clearly indicating which point(s) you are describing.

....

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....

....

.... [2]

(e) Determine the initial slope of the graph in **(c)**. Mark clearly on the graph any construction lines and show in your calculation how the intercepts were used in the calculation of the slope.

[2]

(f) **(i)** Does the initial shape of your graph in **(c)** confirm the equation PV = constant?

....

(ii) Explain your answer in **(i)** above.

....

....

....

....

(iii) Why is the graph of volume against pressure provided inappropriate for the verification of Boyle's law?

[3]

(g) **(i)** Explain why it was important to measure the initial slope of the graph in **(e)**.

....

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....

(ii) What is the significance of the value of the initial slope?

....

....

.... [2]

[Total: 15]

Practical 9

1 The particles of a gas are considered to be in constant random motion. For a fixed mass of gas in a container, the particles collide with the container walls and it is these collisions that are responsible for the pressure exerted by the gas.

The more frequent the collisions, the greater the pressure. If the temperature is kept constant, the average kinetic energy of the gas remains constant. A temperature increase means an increase in this average kinetic energy.

(a) **(i)** Predict how the pressure of a fixed mass of gas held at a constant temperature varies as the volume of the gas decreases. Explain this prediction in terms of the frequency of the gas collisions with the container walls.

Predict how the pressure will change ....

....

Explanation ....

....

....

(ii) Display your prediction in the form of a sketch graph, labelling clearly the axes.

(iii) Sketch a second graph showing how your projected relationship will change if the temperature of the fixed mass of gas is increased. Label this second graph clearly.

[3]

(b) If you were to carry out an experiment to investigate how the **pressure** of a fixed mass of gas varies as the **volume decreases**, name

(i) the independent variable ....

(ii) the dependent variable ....

[2]

[Total: 5]

2 Copper has two oxides, Cu_2O and CuO.
Copper(II) carbonate, $CuCO_3$, decomposes on heating to form one of these oxides. Separate equations can be written showing the two possible decompositions.

2.1 $2CuCO_3(s) \rightarrow Cu_2O(s) + 2CO_2(g) + \frac{1}{2}O_2(g)$

2.2 $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

The following information gives some of the hazards associated with copper(II) carbonate.

Copper(II) carbonate; Harmful if swallowed. Dispose of by reacting no more than 60 g in 1 dm^3 of warm 1 mol dm^{-3} ethanoic acid before pouring down a foul-water drain. This procedure should be kept to a minimum.

You are to plan an experiment to investigate the decomposition of copper(II) carbonate on heating and hence decide which of the two equations represents the actual decomposition. The volume of gas liberated per mole of copper(II) carbonate depends on which equation is correct and this should be used as the basis of the plan.

(a) Draw a diagram of the apparatus and experimental set up you would use in the experiment. Your apparatus should use only standard items found in a school or college laboratory and should show clearly

(i) how the copper(II) carbonate will be heated,

(ii) how the volume of the gas evolved will be collected and its volume measured.

Label each piece of apparatus used, indicating its size or capacity.

[3]

(b) Using the apparatus shown in **(a)**, design a laboratory experiment which will enable you to determine the way in which copper(II) carbonate decomposes on heating.

In addition to the standard apparatus present in a laboratory, you are provided with the following material.

a sample of copper(II) carbonate

Give a step-by-step description of how you would carry out the experiment by

(i) calculating the volume of gas evolved by decomposing 1 mole of copper(II) carbonate according to equation **2.1**.

(ii) calculating the volume of gas evolved by decomposing 1 mole of copper(II) carbonate according to equation **2.2**.

(iii) using these results to calculate a mass of copper(II) carbonate that would give volumes of gas that could be collected using the apparatus proposed, under either decomposition.

(iv) stating how you would ensure that the decomposition is complete.

(v) stating how you would use your results to reach a conclusion.

[A_r: C, 12.0; O, 16.0; Cu, 63.5; the molar gas volume at 25 °C is 24.0 dm^3]

[5]

(c) State one hazard that must be considered when planning the experiment and describe a precaution that should be taken to keep risks to a minimum.

....

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.... [2]

[Total: 10]

3 (a) The concentration of aqueous sulfuric acid can be determined by measuring the density of the solution.

In order to establish a calibration graph, various solutions of known percentage composition by mass were set up using very accurate apparatus.

The preparation of aqueous sulfuric acid is exothermic. Once each of the solutions had cooled, the actual density was determined using a hydrometer.

The table below shows some of the calculations involved, together with the measured densities.

To calculate the actual volumes of sulfuric acid and water mixed, the mass was used in conjunction with the appropriate density.

The density of pure water is 0.997 $g\,cm^{-3}$ and the density of pure sulfuric acid is 1.826 $g\,cm^{-3}$.

All measurements were carried out at 25 °C.

percentage by mass of sulfuric acid	mass of sulfuric acid / g	mass of water / g	volume of sulfuric acid / cm^3	volume of water / cm^3	total volume of 100 g of solution / cm^3	calculated density of the solution / $g\,cm^{-3}$	measured density of the solution / $g\,cm^{-3}$
0	0.000	100.000	0.000	100.301	100.301	0.997	0.997
10	10.000	90.000	5.476	90.271	95.747	1.044	1.064
20	20.000	80.000		80.241	91.194		1.137
30	30.000	70.000	16.429	70.211	86.640	1.154	1.215
40	40.000	60.000			82.087	1.218	1.299
50	50.000		27.382		77.532		1.391
60	60.000	40.000	32.859	40.120		1.370	1.494
70	70.000	30.000	38.335	30.090			1.606
80	80.000	20.000	43.812		63.872	1.566	1.722
90	90.000	10.000		10.030		1.686	1.809
100	100.000	0.000	54.765	0.000	54.765	1.826	1.826

Complete the table. You may use the space below for any working for your calculations.

[4]

(b) On the grid below, using appropriate scales and labelled axes, plot **two** curves to show the variation of **both** the calculated densities **and** the measured densities against the percentage of sulfuric acid. Label each curve.

[4]

(c) **(i)** Using your curves, deduce the difference between the calculated and measured densities for the 45 % by mass of sulfuric acid mixture. State which of the two has the larger density.

(ii) By considering the molecular structures of the two liquids, suggest an explanation for the difference.

....

....

....

....

[3]

(d) You have 100.000 g of a 60 % by mass sulfuric acid mixture of a **calculated density** of 1.370 g cm^{-3}. Calculate the mass of water you would need to add to the 100.000 g in order to give a final **calculated density** of 1.154 g cm^{-3}.

[2]

(e) Attempting to determine the measured density of aqueous sulfuric acid in a school / college laboratory would involve using less accurate apparatus than in **(a)**.
Calculate the errors that could arise when measuring both the mass and volume of 100 g of sulfuric acid using a balance accurate to the nearest 0.01 g and a 100 cm^3 measuring cylinder accurate to the nearest 0.25 cm^3.

mass error

volume error

[2]

[Total: 15]