1. Hydrogen forms compounds with most non-metallic elements and with some metals.

(a) Calculate the empirical formula of the compound used in the manufacture of artificial rubber which has the following composition by mass.

Hydrogen 11.1% Carbon 88.9%

(3)

(b) The boiling temperatures of hydrogen chloride and hydrogen iodide are:

Hydrogen chloride –85ºC

Hydrogen iodide –35ºC

Explain why hydrogen iodide has a higher boiling temperature than hydrogen chloride.

....

....

....

(2)

(c) Draw and explain the shapes of:

(i) the PH_3 molecule;

....

.....

(2)

(ii) the AlH_4^- ion.

.....

.....

(2)

(d) Calculate the number of **molecules** in 8.0 cm^3 of gaseous phosphine, PH_3, at room temperature and pressure.

(The molar volume of a gas at room temperature and pressure should be taken as 2.4×10^4 cm^3 mol^{-1}. The Avogadro constant is 6.0×10^{23} mol^{-1}.)

(2)
(Total 11 marks)

2. In the Periodic Table, where elements are arranged by atomic number, chlorine is a p-block element whereas manganese, a transition element, is in the d-block.

(a) (i) Define the term *atomic number*.

.....

.....

(1)

(ii) Define the term *d-block element*

.....

....

(1)

(iii) Define the term *transition element.*

.....

....

.....

(1)

(b) The electron configuration of chlorine is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$. Write the electron configuration for manganese in a similar manner.

....

(1)

(c) (i) Define the terrm *first ionisation energy of chlorine.*

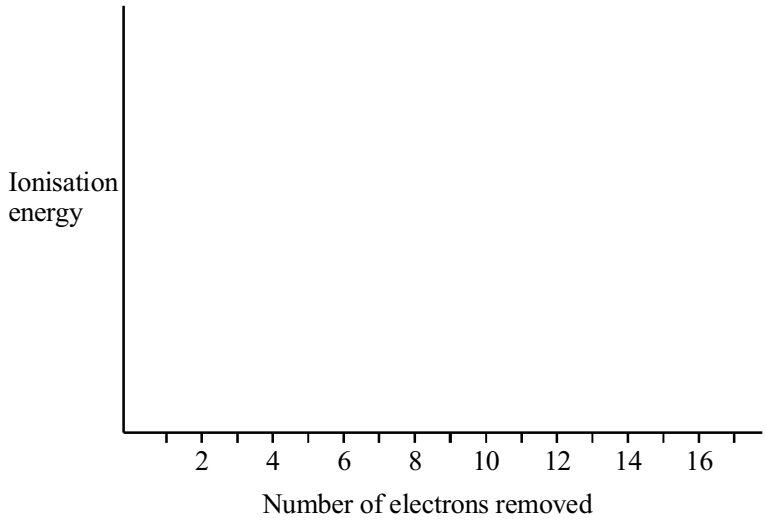
.....

.....

....

(2)

(ii) Sketch the **pattern** you would expect to see in a plot of successive ionisation energies of chlorine against the number of electrons removed.



(3)

(d) Manganese(IV) oxide, MnO_2, reacts with concentrated hydrochloric acid to produce chlorine, water and a salt. The salt has a composition of 43.7% manganese and 56.3% chlorine by mass.

Determine the empirical formula of the salt.

(3)
(Total 12 marks)

3. (a) Sodium reacts with cold water.

(i) What would you **see** as the reaction proceeds?

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.....

.....

(2)

(ii) Write the balanced chemical equation for this reaction.

.....

(2)

(b) Calculate the volume of gas produced if 3.0 g of sodium reacts with an excess of water.

(One mole of any gas at the temperature and pressure of the experiment occupies 24 dm^3.)

(3)
(Total 7 marks)

4. (a) (i) What is meant by the **mass number** of an atom?

.....

(1)

(ii) Define the term **relative atomic mass**.

.....

(2)

(iii) What are isotopes?

.....

(2)

(b) Magnesium has three isotopes. The mass spectrum of magnesium shows peaks at *m/e* 24 (78.60%), 25 (10.11%), and 26 (11.29%). Calculate the relative atomic mass of magnesium to 4 significant figures.

(2)
(Total 7 marks)

5. (a) A compound of sodium, chlorine and oxygen contains, by mass, 21.6% Na, 33.3% Cl and 45.1% O. Show that this is consistent with the formula $NaClO_3$.

(2)

(b) $NaClO_3$ can be obtained from NaOCl(aq) by a disproportionation reaction on heating.

(i) Give the **ionic** equation for this disproportionation reaction.

.....

(2)

(ii) By a consideration of the oxidation numbers of the **chlorine** in the various species, show why the reaction in (i) is disproportionation.

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....

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(4)

(c) Chlorine is used in the extraction of bromine from seawater.

(i) Give the half-equation for the reduction of chlorine.

.....

(1)

(ii) Give the half-equation for the oxidation that is occurring given that the overall equation for the reaction is:

$$Cl_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$$

.....

....

(1)

(Total 10 marks)

6. (a) (i) Define the term **standard enthalpy of formation**, $\Delta H_f^{\ominus}$.

.....

.....

.....

.....

(3)

(ii) The following table shows some values of standard enthalpy of formation.

Name	Formula	$\Delta H_f^{\ominus}$/kJ mol^{-1}
ethene	$C_2H_4(g)$	+52.3
hydrogen bromide	HBr(g)	–36.2
bromoethane	$C_2H_5Br(g)$	–60.4

Use the data in the table above to calculate the standard enthalpy change for the following reaction.

$$C_2H_4(g) + HBr(g) \rightarrow C_2H_5Br(g)$$

(2)

(iii) State the significance of the sign of the value obtained in part (a)(ii) above.

.....

....

(1)

(b) Enthalpy changes can also be calculated using average bond enthalpy data.

Bond	Average bond enthalpy/kJ mol^{-1}
C == C	+612
C — C	+348
C — H	+412
C — Br	+276
H— Br	+366

Use the data in the table above to recalculate the enthalpy change for the reaction in part (a)(ii).

$$C_2H_4(g) + HBr(g) \rightarrow C_2H_5Br(g)$$

(3)

(c) Suggest why the value obtained in part (b) above is likely to be less accurate than that obtained in part (a)(ii).

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....

....

(2)
(Total 11 marks)

7. (a) In an experiment to standardise an aqueous solution of sodium hydroxide 0.25 g ofsolid sulphamic acid, NH_2SO_3H, was dissolved in distilled water in a conical flask. When the aqueous sodium hydroxide was run into the flask from a burette 23.45 cm^3 was required to exactly react with the sulphamic acid solution. The equation for the reaction is:

$$NH_2SO_3H(aq) + NaOH(aq) \rightarrow NH_2SO_3Na(aq) + H_2O(l)$$

(i) Calculate the amount (number of moles) of sulphamic acid in 0.25 g. $M_r(NH_2SO_3H) = 97.0$.

(1)

(ii) State the amount (number of moles) of sodium hydroxide in 23.45 cm^3 of solution and hence calculate the concentration of the solution in mol dm^{-3}.

(3)

(b) The balance used to weigh the sulphamic acid is accurate to –0.01 g. Calculate the percentage error in the mass of the sulphamic acid weighed.

(1)

(c) An alternative method to that described in (a) involves making an aqueous solution of sulphamic acid of **accurately known concentration**.

Describe a procedure bv which you would prepare 250 cm^3 of aqueous sulphamic acid of accurately known concentration. Assume that you are provided with a weighing bottle containing between 2.40 g and 2.50 g of sulphamic acid and that this is a suitable mass to use.
In your answer give full practical details including the name of each piece of apparatus used, how each would be prepared for the procedure and how you would calculate the concentration (in mol dm^{-3}) of the sulphamic acid solution. State, with a reason, one appropriate safety precaution that should be taken.

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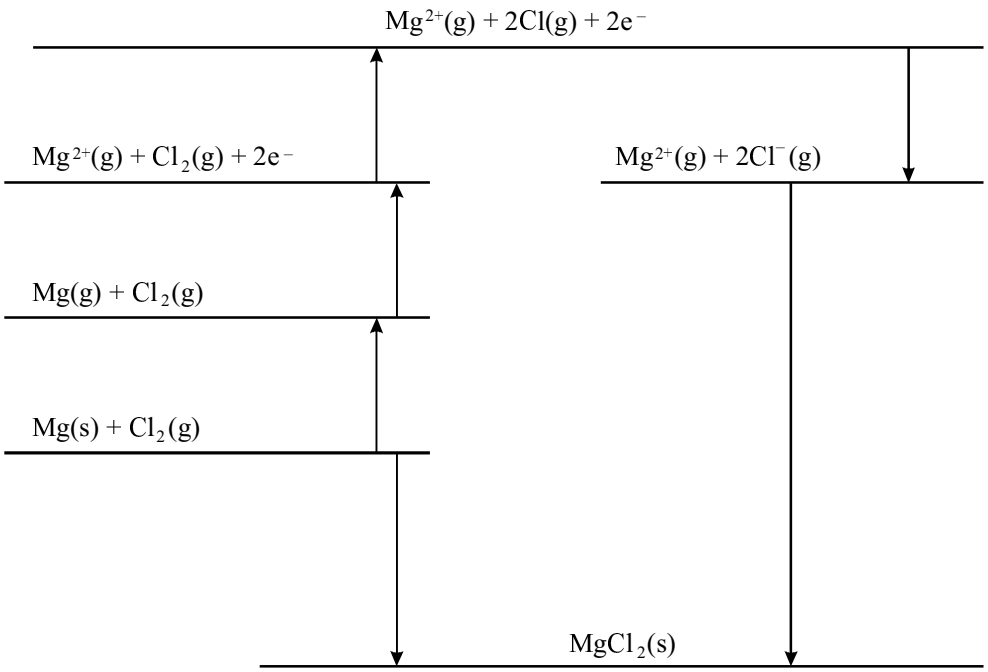
(Allow one lined page)

(8)
(Total 13 marks)

8. The formation of magnesium chloride from magnesium and chlorine may be represented by the following Born-Haber cycle:



(a) Define the terms:

Lattice enthalpy.

....

....

....

(3)

Enthalpy of atomisation.

....

....

....

(2)

(b) (i) Identify **on the diagram** the chance representing the enthalpy of atomisation of magnesium.

(1)

(ii) Use the data below to calculate the first electron affinity of chlorine.

Enthalpy change	Value of the enthalpy change / kJ mol^{-1}
Enthalpy of atomisation of magnesium	+150
1st Ionisation energy of magnesium	+736
2nd Ionisation energy of magnesium	+1450
Enthalpy of formation of magnesium chloride	–642
Enthalpy of atomisation of chlorine	+121
Lattice enthalpy of magnesium chloride	–2493

(2)

(c) Hydrogen gas reacts with sodium metal to form an ionic solid, NaH, which contains sodium cations.

Draw a Born-Haber cycle which could be used to determine the electron affinity of

hydrogen.

(3)
(Total 11 marks)

9. (a) (i) Calculate the number of moles of potassium nitrate, KNO_3, in 10.1 g of KNO_3.

(1)

(ii) Potassium nitrate, KNO_3, can be prepared from potassium hydroxide solution as shown in the following equation:

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

Calculate the minimum volume, in cm^3, of 2.00 mol dm^{-3} KOH required to produce 10.1 g of KNO_3.

(2)

(iii) Potassium nitrate decomposes, when heated, to produce oxygen.

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

Calculate the volume of oxygen gas, in dm^3, produced when 10.1 g of potassium nitrate decomposes in this way.
(1 mole of gas has a volume of 24 dm^3 under the conditions of the experiment.)

(2)

(b) A compound of potassium and oxygen contains 70.9% potassium.

(i) Calculate the empirical formula of this compound, using the data above and the periodic table.

(3)

(ii) 0.200 moles of this compound has a mass of 22.0 g. Use this information to help you deduce the molecular formula of this compound.

(2)
(Total 10 marks)

10. The reaction of an acid with a base to give a salt is an exothermic reaction. In an experiment to determine the enthalpy of neutralisation of hydrochloric acid with sodium hydroxide, 50.0cm^3 of 1.00 mol dm^{-3} HCl was mixed with 50.0 cm^3 of 1.10 mol dm^{-3} NaOH. The temperature rise obtained was 6.90 °C.

(a) Define the term **enthalpy of neutralisation**.

....

....

(1)

(b) Assuming that the density of the final solution is 1.00 g cm^{-3} and that its heat capacity is 4.18 J K^{-1} g^{-1}, calculate the heat evolved during the reaction.

(3)

(c) 0.0500 mol of acid was neutralised in this reaction; calculate $\Delta H_{neutralisation}$ in kJ mol^{-1}.

(2)

(d) Suggest why sodium hydroxide is used in slight excess in the experiment.

....

....

(1)
(Total 7 marks)

11. A student was required to determine the enthalpy change for the reaction between iron and copper sulphate solution.

The student produced the following account of their experiment.

A piece of iron, mass about 3 g, was placed in a glass beaker. 50 cm^3 of 0.5 mol dm^{-3} aqueous copper sulphate solution was measured using a measuring cylinder and added to the beaker. The temperature of the mixture was measured immediately before the addition and every minute afterwards until no further change took place.

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

Timing	before addition	1 min	2 mins	3 mins	4 mins	5 mins
Temperature/°C	22	27	29	26	24	22

(a) Suggest **two** improvements you would make to this experiment. Give a reason for each of the improvements suggested.

Improvement 1 .....

....

Reason 1 .....

....

Improvement 2 .....

....

Reason 2 .....

....

(4)

(b) In an improved version of **the same experiment** a maximum temperature rise of 15.2 °C occurred when reacting excess iron with 50.0 cm^3 of 0.500 mol dm^{-3} aqueous copper sulphate solution.

(i) Using this data and taking the specific heat capacity of all aqueous solutions as 4.18 Jg^{-1} deg^{-1} calculate the heat change.

(1)

(ii) Calculate the number of moles of copper sulphate used.

(1)

(iii) Calculate the enthalpy change of this reaction in kJ mol^{-1}.

(2)
(Total 8 marks)

12. (a) The first ionisation energy of potassium is +419 kJ mol^{-1} and that of sodium is +496 kJ mol^{-1}.

(i) Define the term **first ionisation energy**.

……..

……..

……..

(3)

(ii) Explain why the first ionisation energy of potassium is only a little less than the first ionisation energy of sodium.

……..

……..

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……..

(3)

(b) Potassium forms a superoxide, KO_2. This reacts with carbon dioxide according to the equation:

$$4KO_2(s) + 2CO_2(g) \rightarrow 2K_2CO_3(s) + 3O_2(g)$$

Carbon dioxide gas was reacted with 4.56 g of potassium superoxide.

(i) Calculate the amount, in moles, of KO_2 in 4.56 g of potassium superoxide.

(2)

(ii) Calculate the amount, in moles, of carbon dioxide that would react with 4.56 g of potassium superoxide.

(1)

(iii) Calculate the volume of carbon dioxide, in dm^3, that would react with 4.56 g of potassium superoxide. Assume that 1.00 mol of a gas occupies 24 dm^3 under the conditions of the experiment.

(1)

(iv) What volume of oxygen gas, in dm^3, measured under the same conditions of pressure and temperature, would be released?

(1)
(Total 11 marks)

13. (a) Define the term **standard enthalpy of combustion**, making clear the meaning of **standard** in this context.

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……..

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……..

(3)

(b) Use the enthalpies of combustion given below to find the enthalpy change for the reaction:

$$2C(graphite) + 2H_2(g) + O_2(g) \rightarrow CH_3COOH(l)$$

	$\Delta H_{combustion}$/kJ mol^{-1}
C(graphite)	–394
$H_2(g)$	–286
$CH_3COOH(l)$	–874

(3)

(c) With reference to ethanoic acid, CH_3COOH, what is the enthalpy change obtained in (b) called?

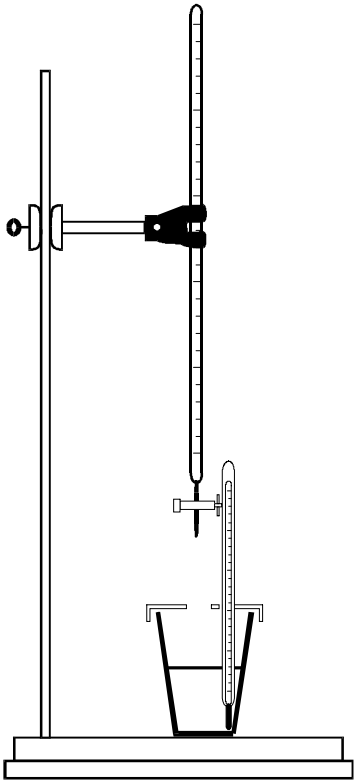
……..

(1)

(d) Draw an enthalpy level diagram to represent the enthalpy change for the combustion of graphite. Show both the enthalpy levels of the reactants and products and an energy profile which represents the activation energy for the reaction.

(3)
(Total 10 marks)

14. (a) This question is about finding the formula of copper hydroxide. The method is as follows:



20.0 cm^3 of an aqueous solution of a copper salt of concentration 1.00 mol dm^{-3} was placed in a polystyrene cup and its temperature measured using a thermometer graduated in 0.1 °C intervals.

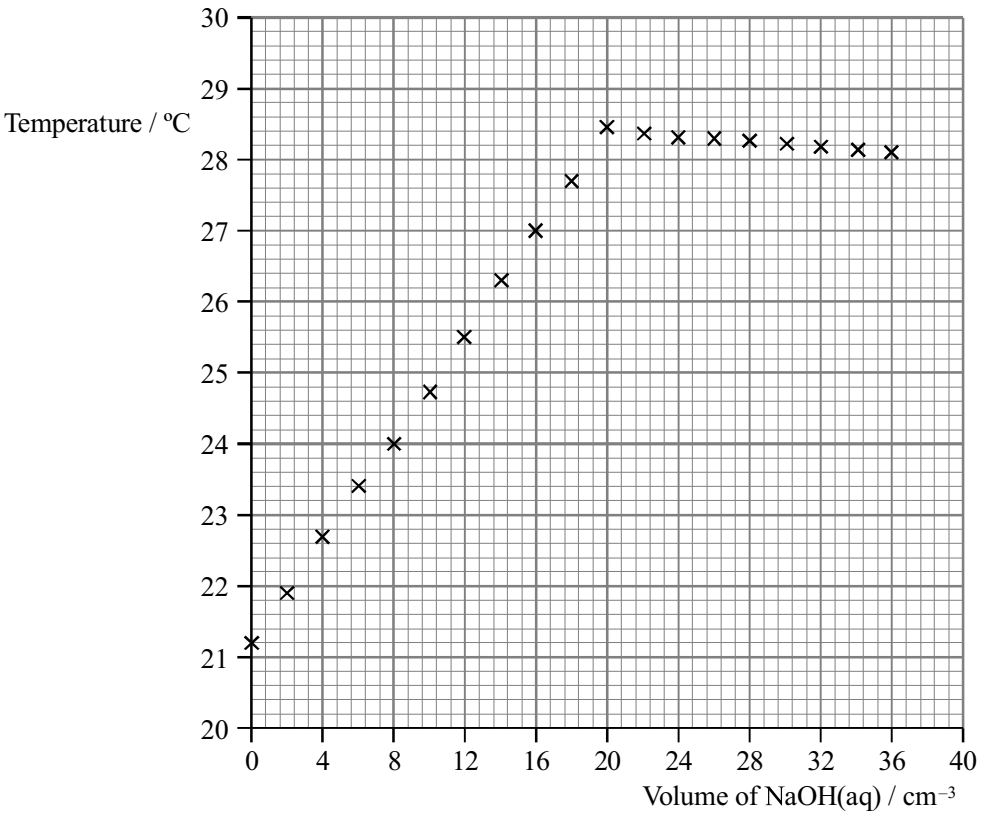
A burette was filled with aqueous sodium hydroxide, of concentration 2.00 mol dm^{-3}.

2.00 cm^3 of sodium hydroxide solution was run into the solution of the copper salt and the temperature was measured immediately.

As soon as possible a further 2.00 cm^3 of sodium hydroxide solution was run in and the temperature measured again.

This process of adding 2.00 cm^3 portions of sodium hydroxide solution and measuring the temperature was continued until a total of 36.0 cm^3 of the sodium hydroxide solution had been added.

The temperature readings are shown in the graph below.



(i) Explain why the temperature reaches a maximum and then falls slightly on addition of further sodium hydroxide solution.

……..……..……..

……..……..……..

……..……..……..

(2)

(ii) From the graph, what volume of the aqueous sodium hydroxide was required for complete reaction?

……..……..……..

(1)

(iii) Calculate the amount (number of moles) of sodium hydroxide in this volume of solution.

(1)

(iv) Calculate the amount (number of moles) of copper ions that have reacted.

(1)

(v) Write the ratio of moles of copper ions to hydroxide ions reacting.

(1)

(vi) Write the formula of the copper hydroxide that is produced.

(1)

(b) The data call be used to find the enthalpy change for the reaction between sodium hydroxide and the copper salt.

(i) Use the graph to find the temperature rise that occurs for complete reaction.

……..……..……..

(1)

(ii) Find the heat change, q, that occurs in the polystyrene cup for complete reaction. Use the formula

$$\boldsymbol{q = 168 \times \Delta T \text{ joules}}$$

(1)

(iii) Use your results from (a)(iv) and (b)(ii) above, to find the molar enthalpy change, ΔH, for the reaction. Give the correct sign and units to the answer.

(3)

(c) Identify one potential source of error in this experiment, and say what you would do to reduce its effect.

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……...……..……..

……...……..…… ……..

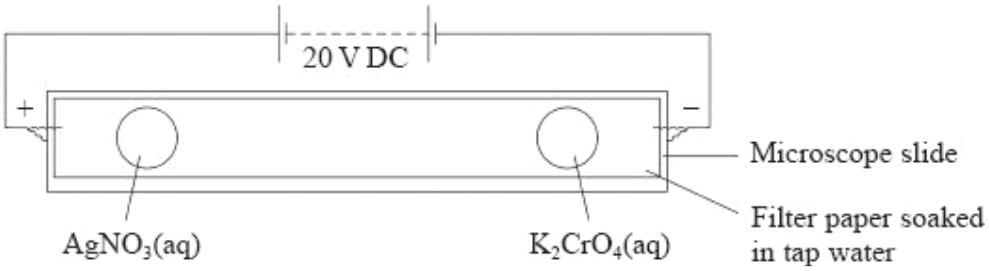
……...……..……..

……...……..……..

(2)

(Total 14 marks)

15. In an experiment to show the migration of ions, silver(I) nitrate solution, $AgNO_3(aq)$, and potassium chromate(VI) solution, $K_2CrO_4(aq)$, were used. The experiment was set up as shown in the diagram below. After a short while, a red precipitate formed in the centre of the filter paper.



(a) Write the formulae, including charges, of the following ions:

(i) Nitrate ions ....

(ii) Chromate (VI) ions .....

(2)

(b) Which ions would meet in the centre of the filter paper?

....

(1)

(c) Write a balanced ionic equation, including state symbols, for the formation of the red precipitate.

(2)
(Total 5 marks)

16. An excess of zinc powder was added to 20.0 cm^3 of a solution of copper(II) sulphate of concentration 0.500 mol dm^{-3}. The temperature increased by 26.3 °C.

(a) How many moles of copper(II) sulphate were used in this experiment?

(1)

(b) Calculate the enthalpy change, ΔH, in kJ mol^{-1} for this reaction given that:

energy change =	specific heat capacity	× mass of solution	× temperature change
/J	/J g^{-1} K^{-1}	/g	/K

Assume that the mass of solution is 20.0 g and the specific heat capacity of the solution is 4.18 J $g^{-1}K^{-1}$.

(2)
(Total 5 marks)

17. A sample of titanium (atomic number 22) is made up of five isotopes. The sample has the following percentage composition:

Mass number	% composition
46	8.0
47	7.3
48	74.0
49	5.5
50	5.2

(a) (i) What is the average relative atomic mass of titanium? Give your answer to **three** significant figures.

(2)

(ii) What instrument would have been used to find this percentage composition?

.....

(1)

(b) (i) Give the electronic configuration of a titanium atom, using s p d notation.

....

(2)

(ii) Name the part of the Periodic Table where titanium appears.

....

(1)

(c) Titanium occurs naturally as rutile, TiO_2. One possible method of obtaining pure titanium is to heat rutile with carbon.

$$TiO_2(s) + 2C(s) \rightarrow Ti(s) + 2CO(g)$$

(i) What type of reaction is this?

.....

(1)

(ii) Calculate ΔH for this reaction given that

$$\Delta H_f^{\ominus} [TiO_2(s)] = -940 \text{ kJ mol}^{-1}$$

$$\Delta H_f^{\ominus} [CO(g)] = -110 \text{ kJ mol}^{-1}$$

Include a sign and units in your answer.

(3)

(iii) Name the law you have used in your calculation.

.....

(1)

(iv) When titanium is manufactured by this method, explain what pollution problem arises.

.....

....

.....

(2)
(Total 13 marks)

18. This question is about nitrogen trifluoride, NF_3, and nitrogen trichloride, NCl_3, which are covalent compounds. Van der Waals attractions and permanent dipole–dipole attractions exist between molecules of both compounds in the liquid state.

(a) (i) Describe how van der Waals attractions are caused.

.....

.....

.....

.....

(1)

(ii) In which of the two compounds would you expect there to be greater van der Waals attractions? Justify your answer.

.....

....

(1)

(iii) In which of the two compounds would you expect there to be the greater permanent dipole–dipole attractions? Justify your answer.

.....

.....

(1)

(iv) The boiling point of nitrogen trichloride is much higher than that of nitrogen trifluoride. How might this be explained in terms of the two types of intermolecular attractions?

....

.....

(1)

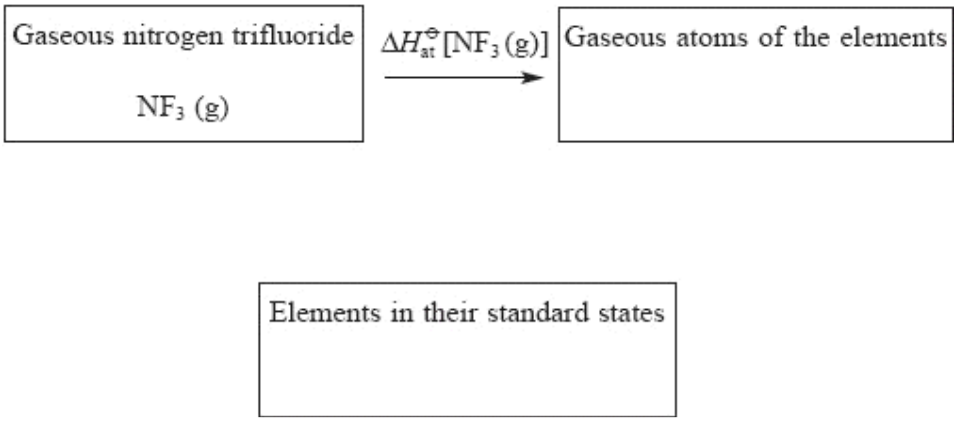
(b) The standard enthalpy change for the formation of gaseous nitrogen trifluoride is -125 kJ mol^{-1}.

$$\tfrac{1}{2}N_2\,(g) + 1\tfrac{1}{2}F_2\,(g) \rightarrow NF_3\,(g) \quad \Delta H^{\ominus}_f = -125 \text{ kJ mol}^{-1}$$

The standard molar enthalpy changes of atomisation of nitrogen, N_2, and of fluorine, F_2, are given below.

	ΔH_{at}/kJ mol^{-1}
$\tfrac{1}{2}N_2\,(g) \rightarrow N\,(g)$	+ 473
$\tfrac{1}{2}F_2\,(g) \rightarrow F\,(g)$	+ 79

This information can be represented on a Hess cycle in the following way, and then used to calculate bond energies.



(i) Insert formulae, showing the correct quantities of each element, into the appropriate boxes.

(1)

(ii) Insert arrows between the boxes and write the correct numerical data alongside the appropriate arrows.

(2)

(iii) Use the cycle to calculate the N — F bond energy in nitrogen trifluoride.

(2)
(Total 9 marks)

19. (a) State the meaning of the terms

(i) relative atomic mass

.....

.....

.....

(2)

(ii) mass number

....

.....

(1)

(iii) isotopes

....

.....

....

(2)

(b) The isotopic composition of a sample of sulphur is found using a mass spectrometer.

(i) Explain how atoms of the sample of sulphur are ionised.

.....

.....

.....

(2)

(ii) State the type of charge on the sulphur ions formed in the mass spectrometer.

....

(1)

(iii) State how the resulting sulphur ions are then accelerated.

....

(1)

(c) For a particular sample of sulphur atoms the following isotopic composition was recorded.

Isotope	Percentage composition
^{32}S	95.00
^{33}S	0.76
^{34}S	4.24

Calculate the relative atomic mass of this sample of sulphur. Give your answer to two decimal places.

(2)

(d) Predict the electronic configuration of a ^{34}S atom, using *s*, *p* and *d* notation.

$1s^2$ .....

(1)
(Total 12 marks)

20. A compound **A** is formed when chlorine is bubbled through hot concentrated potassium hydroxide solution.

(a) Analysis of **A** shows that it contains 31.84% potassium, 28.98% chlorine and the remainder is oxygen.

Show that the empirical formula of **A** is $KClO_3$

(3)

(b) On being heated strongly solid **A** decomposes completely to give oxygen gas and solid potassium chloride.

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

If 1.00 g of solid **A** is decomposed completely in this way, calculate the volume of oxygen gas produced at room temperature and pressure.

(One mole of a gas at room temperature and pressure occupies 24 dm^3)

(3)
(Total 6 marks)

21. Urea, which is used as a fertillser in much of mainland Europe, Asia and Africa, is manufactured by the reaction of ammonia and carbon dioxide.

$$2NH_3(g) + CO_2(g) \rightarrow NH_2CONH_2(s) + H_2O(l)$$

(a) Define the term **standard enthalpy of formation**, $\Delta H_f^{\ominus}$, of urea.

.....

.....

.....

.....

.....

(3)

(b) Calculate the enthalpy change, $\Delta H^{\ominus}$ for the reaction above, given the following standard enthalpies of formation.

Substance	$\Delta H_f^{\ominus}$ / kJ mol^{-1}
$NH_3(g)$	–46.2
$CO_2(g)$	–393.5
$NH_2CONH_2(s)$	–632.2
$H_2O(l)$	–285.8

(3)
(Total 6 marks)

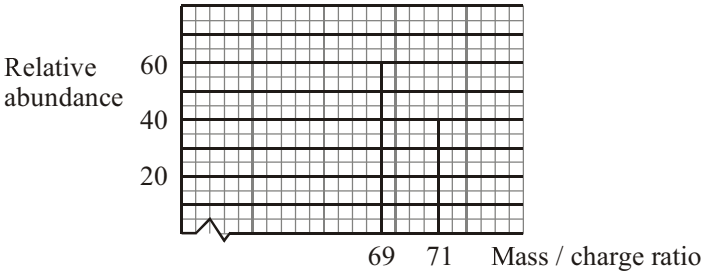
22. (a) An atom of gallium has mass number 69.

Complete the table to show the number of sub-atomic particles in this gallium atom.

Electrons	Neutrons	Protons

(2)

(b) The mass spectrum of a sample of gallium is shown below.



What is the average relative atomic mass of gallium in this sample? Give your answer to **three** significant figures.

(2)

(c) What type of bonding would you expect to find in gallium?

....

(1)
(Total 5 marks)

23. (a) (i) Calculate the number of moles of sulphur atoms in 4.00 g of sulphur.

(1)

(ii) When sulphur burns in air, it forms sulphur dioxide gas, SO_2.

What volume of sulphur dioxide, measured at room temperature and pressure, would be produced when 4.00 g of sulphur is burnt in air?
[Molar volume is 24 dm^3 mol^{-1} at room temperature and pressure.]

(2)

(b) Sulphur dioxide reacts with sodium hydroxide solution to form a solution of sodium sulphite, Na_2SO_3.

The equation for the reaction is

$$SO_2(g) + 2NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$$

Rewrite this equation as an ionic equation, omitting the spectator ions.

....

(2)
(Total 5 marks)

24. (a) The equation below shows the reaction which occurs when ammonia is dissolved in water.

$$NH_3(g) + H_2O(1) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

(i) Explain why water is classified as an acid in this reaction.

.....

.....

(1)

(ii) The ammonia is acting as a weak base in this reaction.

What is the difference between a weak base and a strong base?

.....

.....

.....

(1)

(b) Ammonia reacts with oxygen to form the gases nitrogen(II) oxide and steam.

(i) Complete the Hess cycle below so that $\Delta H_{reaction}$ can be calculated using standard enthalpy changes of formation. Include state symbols.

$$4NH_3(g) + 5O_2(g) \xrightarrow{\Delta H_{reaction}} 4NO(g) + 6H_2O(g)$$

.....

(2)

(ii) Calculate $\Delta H_{reaction}$ for this reaction using the following data.

ΔH_f [NH_3(g)] = – 46.1 kJ mol^{-1}

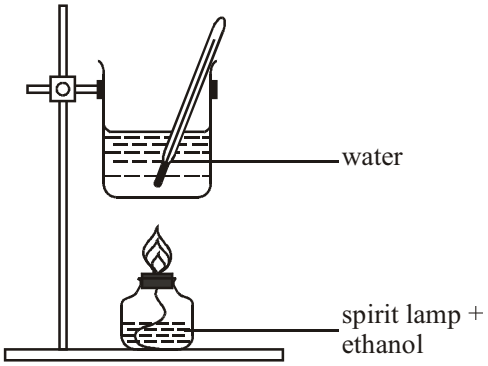
ΔH_f [NO(g)] = + 90.2 kJ mol^{-1}

ΔH_f [H_2O(g)] = – 241.8 kJ mol^{-1}

Include a sign and units in your answer and give your answer to **three** significant figures.

(3)
(Total 7 marks)

25. The apparatus shown in the diagram below may be used to find the enthalpy of combustion of alcohols.



Using the apparatus, a student recorded the results included in the table below.

Alcohol = ethanol, C_2H_5OH

Molar Mass (C_2H_5OH) = 46.0 g mol^{-1}

Volume of water in beaker = 200 cm^3

∴ mass of water in beaker = 200 g

Weighings

Spirit lamp + ethanol before combustion = 198.76 g

Spirit lamp + ethanol after combustion = 197.68 g

Temperatures

Water before heating = 19.5 °C

Water after heating = 38.1 °C

Specific heat capacity of water = 4.18 J g^{-1} $°C^{-1}$

(a) What assumption is the student making about water to be able to state that its mass is numerically equal to its volume?

....

(1)

(b) Calculate the heat gained by the water. Give your answer in kJ.

(2)

(c) Calculate the amount (number of moles) of ethanol used.

(2)

(d) Using your values from (b) and (c), calculate the enthalpy of combustion of ethanol. Give your answer to a number of significant figures consistent with the readings in the table. Include a sign and units in your answer.

(3)

(e) The student's evaluation of the experiment is given below.

> My calculated value of the enthalpy of combustion was numerically much less than the data book value. The reasons for my low value include:
>
> 1 heat losses to the surrounding air;
> 2 when I re-checked the mass of the spirit lamp and ethanol after combustion, I noticed that it had lost mass even when it was not being used;
> 3 a black solid which formed on the base of the beaker.

(i) Explain why the spirit lamp and ethanol **lost mass** even when not in use.

....

....

(1)

(ii) Suggest the identity of the black solid. Explain why its formation will lead to a low value for the enthalpy of combustion.

Identity .....

Explanation .....

.....

(2)

(Total 11 marks)

26. How many nitrogen molecules, N^2, are present in 12 dm^3 of nitrogen gas at room temperature and pressure?

[Molar volume is 24 dm^3 mol^{-1} at room temperature and pressure;
Avogadro constant is 6×10^{23} mol^{-1}]

(Total 2 marks)

27. Copper(II) sulphate solution can be prepared from solid copper(II) carbonate by reaction with hot dilute sulphuric acid.

(a) Write the balanced equation for this reaction, including state symbols.

(2)

(b) This experiment was carried out using 25 cm^3 of 1.0 mol dm^{-3} sulphuric acid.

(i) How would you measure out 25 cm^3 of 1.0 mol dm^{-3} sulphuric acid?

....

(1)

(ii) Calculate the number of moles of sulphuric acid used.

(1)

(c) (i) It is usual to react the sulphuric acid with a slight excess of copper(II) carbonate. Calculate the mass of copper(II) carbonate needed if a 10 % excess is required. [Molar mass of copper(II) carbonate = 123.5 g mol^{-1}]

(2)

(ii) The sulphuric acid is heated to boiling and the copper(II) carbonate is added in small portions.

Suggest why the copper(II) carbonate is added in small portions.

.....

....

(1)

(iii) What would be the next step needed to obtain pure copper(II) sulphate solution?

....

.....

(1)

(d) 3.98 g of copper(II) sulphate-5-water, $CuSO_4.5H_2O$, was obtained from this solution.

(i) Calculate the molar mass of copper(II) sulphate-5-water.
Use the Periodic Table as a source of data.

(1)

(ii) Calculate the percentage yield of the reaction.

(2)

(e) Suggest why copper compounds are hazardous.

....

(1)
(Total 12 marks)

28. This question is about a self-heating can of coffee.

The bottom of the can has a compartment containing copper(II) nitrate solution. When a button on the bottom of the can is pressed, magnesium powder is released into the compartment where it reacts with the copper(II) nitrate solution.

(a) (i) Write an ionic equation for the reaction between magnesium powder and copper(II) ions. Include state symbols, but omit any spectator ions.

(2)

(ii) Show how the standard enthalpy change for this reaction could be calculated from the standard enthalpies of formation of copper(II) ions and magnesium ions. You should include a Hess cycle in your answer.

(3)

(b) The can contains 150 g of a solution of coffee in water.

The temperature of the solution needs to increase by 60 °C to produce a hot drink.

(i) Calculate the energy change needed to produce a temperature increase of 60 °C in the coffee, using the relationship

Energy change = 4.2 × mass of solution × temperature change.

Remember to include a unit in your answer.

(2)

(ii) The standard enthalpy change for this reaction is –530 kJ mol^{-1}.

Calculate the number of moles of reactants needed to produce the energy change in (i).

(1)

(iii) A solution of copper(II) nitrate of concentration 8.0 mol dm^{-3} is used.

Use your answer to (ii) to calculate the volume, in cm^3, of copper(II) nitrate solution needed.

Your answer should be given to two significant figures.

(1)

(c) Suggest TWO reasons why the temperature of the coffee may **not** increase by as much as 60 °C.

....

....

....

....

(2)
(Total 11 marks)

29. The reaction between chlorine and methane, in the presence of ultraviolet light, involves the formation of free radicals and includes the following steps:

A	$Cl_2 \rightarrow 2Cl^{\bullet}$	$\Delta H^{\ominus} = +242$ kJ mol^{-1}
B	$CH_4 + Cl^{\bullet} \rightarrow HCl + CH_3^{\bullet}$	$\Delta H^{\ominus} = +4$ kJ mol^{-1}
C	$Cl_2 + CH_3^{\bullet} \rightarrow CH_3Cl + Cl^{\bullet}$	$\Delta H^{\ominus} = -97$ kJ mol^{-1}
D	$Cl^{\bullet} + Cl^{\bullet} \rightarrow Cl_2$	
E	$CH_3^{\bullet} + CH_3^{\bullet} \rightarrow CH_3CH_3$	
F	$Cl^{\bullet} + CH_3^{\bullet} \rightarrow CH_3Cl$	$\Delta H^{\ominus} = -339$ kJ mol^{-1}

(a) (i) What is meant by a **free radical**? .....

.....

.....

(1)

(ii) Draw a 'dot-and-cross' diagram, showing outer shell electrons only, for a chlorine free radical.

(1)

(iii) What type of bond breaking occurs in step **A**?

....

(1)

(b) Which of the steps, **A** to **F**, are chain propagation steps?

.....

(1)

(c) (i) Write the equation for the overall reaction between one mole of chlorine and one mole of methane molecules.

(1)

(ii) Calculate the standard enthalpy change, $\Delta H^{\ominus}$, for this reaction.

(2)

(d) (i) What is the value of $\Delta H^{\ominus}$ for step **D**? ....

(1)

(ii) Would you expect step **E** to be exothermic or endothermic? Justify your answer.

.....

.....

.....

(1)

(e) The overall reaction was repeated using bromine gas instead of chlorine gas.

Would you expect step **A** for bromine to be more or less endothermic than step **A** for chlorine? Justify your answer.

....

....

....

(2)
(Total 11 marks)

30. In two similar, separate experiments the enthalpy changes for the reactions of sodium hydrogencarbonate and sodium carbonate with excess dilute hydrochloric acid were determined.

(a) The first experiment was to find the enthalpy change, ΔH_1, for the reaction

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Measurement	Reading
Mass of solid sodium hydrogencarbonate added to hydrochloric acid.	5.00 g
Volume of hydrochloric acid	50.0 cm^3
Temperature of hydrochloric acid before addition of solid sodium hydrogencarbonate	22.0 °C
Final temperature of solution	15.5 °C
Molar mass of sodium hydrogencarbonate	84.0 g mol^{-1}
Specific heat capacity of solution	4.18 J g^{-1} $°C^{-1}$

(i) Calculate the amount (moles) of sodium hydrogencarbonate used.

(1)

(ii) Calculate the heat absorbed in the reaction in kJ.
[Assume that 1 cm^3 of solution has a mass of 1 g]

(2)

(iii) Calculate the value of ΔH_1 in kJ mol^{-1}. Include a sign in your answer expressing it to a number of significant figures suggested by the data in the table.

(2)

(b) In the second experiment the enthalpy change for the reaction between sodium carbonate and dilute hydrochloric acid was measured.

$$Na_2CO_3(s) + 2HCl\,(aq) \rightarrow 2NaCI(aq) + CO_2(g) + H_2O(l)$$

The molar enthalpy change, ΔH_2, was calculated to be –35.6 kJ mol^{-1}

(i) Give TWO ways in which the temperature change differs when equal molar amounts of sodium hydrogencarbonate and sodium carbonate react separately with the same volume of hydrochloric acid.

....

....

....

(2)

(ii) Give ONE assumption that has been made in calculating the values of ΔH_1, and ΔH_2 from experimental results.

....

....

(1)

(Total 8 marks)

31. In the manufacture of beer, brewers often add small amounts of salts of Group 2 elements to the water used. These salts influence the chemical reactions during the brewing process.
Two such salts are calcium sulphate and magnesium sulphate.

(a) A flame test can be used to confirm that a sample of a salt contains calcium ions.

(i) Describe how you would carry out a flame test.

.....

....

.....

....

.....

....

.....

(3)

(ii) A positive test results in a brick-red flame colour. Describe the changes that occur in calcium ions to produce a colour.

.....

....

.....

....

(2)

(iii) Impurities in the salt may lead to other colours being observed in the flame.
What metal ion is likely to be present if a yellow flame is seen?

....

(1)

(b) Magnesium sulphate can be used in its anhydrous form, $MgSO_4(s)$, or in its hydrated form, $MgSO_4.7H_2O(s)$.

An experiment was carried out to find the enthalpy change when hydrated magnesium sulphate dissolved completely in water.

$$MgSO_4.7H_2O(s) \xrightarrow{\text{excess water}} MgSO_4(aq) + 7H_2O(l)$$

12.3 g of hydrated magnesium sulphate was added to 100 g of water in a simple calorimeter and the temperature was found to fall by 1.1 °C.

(i) Calculate the energy change, in joules, that occurred in the experiment, using the relationship

Energy change (J) = 4.18 × mass of water × temperature change

(2)

(ii) Calculate the number of moles of hydrated magnesium sulphate used in the experiment. Use the Periodic Table as a source of data.

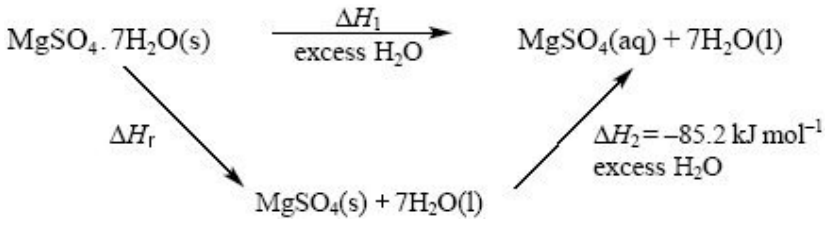
(2)

(iii) Use your answers to (i) and (ii) to calculate the enthalpy change for the reaction. Include a sign and units in your final answer, which should be given to 2 significant figures.

(2)

(c) The enthalpy change as hydrated magnesium sulphate is converted to anhydrous magnesium sulphate is very difficult to measure. The Hess Cycle below can be used to

find this enthalpy change, ΔH_r.



(i) Use the cycle to write an expression for ΔH_r using ΔH_1 and ΔH_2.

(1)

(ii) Use your expression in (c)(i) and your answer from (b)(iii) to calculate ΔH_r.

Include a sign and units in your final answer, which should be given to 2 significant figures.

(2)
(Total 15 marks)

32. Phosphine, PH_3, is a hydride of the Group 5 element, phosphorus.

(a) (i) Draw a 'dot-and-cross' diagram of a phosphine molecule. You should include only outer shell electrons.

(1)

(ii) Draw the shape you would expect for the phosphine molecule, suggesting a value for the HPH bond angle.

HPH bond angle .....

(2)

(iii) Explain the shape of the phosphine molecule you have given in your answer in (ii).

Justify your value for the HPH bond angle.

.....

.....

.....

.....

(2)

(b) (i) Write a balanced equation, including state symbols, for the atomisation of phosphine gas.

.....

(1)

(ii) Use your answer to (i) and the data below to calculate the standard enthalpy change of atomisation of phosphine at 298 K. Include a sign and units in your answer.

$\Delta H^{\ominus}_{f}[PH_3(g)] = +\ 5.4\ kJ\ mol^{-1}$

$\Delta H^{\ominus}_{at}[\frac{1}{2}H_2(g)] = +218.0\ kJ\ mol^{-1}$

$\Delta H^{\ominus}_{at}[P(s)] = +314.6\ kJ\ mol^{-1}$

(3)

(iii) Calculate a value for the bond energy of the bond between phosphorus and hydrogen, using your answer to (ii).

(1)
(Total 10 marks)

33. (a) Define the term **reduction**.

....

....

(1)

(b) Sodium iodide, NaI, contains iodide ions which reduce Fe^{3+} ions to Fe^{2+} ions in aqueous solution.

(i) Write the ionic half-equation for the reduction of Fe^{3+} ions to Fe^{2+} ions.

....

(ii) Write the ionic half-equation for the oxidation of iodide ions.

....

(1)

(iii) Hence write the overall ionic equation for the reduction of Fe^{3+} ions by iodide ions.

....

(1)

(c) Sodium chlorate, $NaClO_3$, is used as a weedkiller and is manufactured by the following disproportionation reaction

$$3NaClO \rightarrow 2NaCl + NaClO_3$$

(i) State the oxidation numbers of chlorine in

NaClO .....

$NaClO_3$ .....

NaCl .....

(3)

(ii) Explain why this reaction is classified as a disproportionation reaction.

....

....

....

(1)

(d) The decomposition of sodium azide, NaN_3, takes place when an airbag, which is used as a safety feature in cars, inflates. An airbag requires a large volume of a cool gas to be produced in a few milliseconds. The gas is produced by the rapid decomposition of the sodium azide.

$$2NaN_3(s) \rightarrow 2Na(l) + 3N_2(g)$$

When the airbag is fully inflated, 54 dm^3 of nitrogen gas are produced.

(i) Calculate the amount (in moles) of nitrogen gas produced.

[The molar volume of nitrogen gas under the conditions in the airbag is 20 $dm^3\ mol^{-1}$].

(1)

(ii) Calculate the amount (in moles) of sodium azide, NaN_3, that would produce 54 dm^3 of nitrogen gas.

(1)

(iii) Calculate the mass of sodium azide that would produce 54 dm^3 of nitrogen gas.

(2)
(Total 12 marks)

34. Methane, CH_4, is used as a domestic and industrial fuel and as a reagent in the petrochemical industry.

(a) Define the term **standard enthalpy of combustion**.

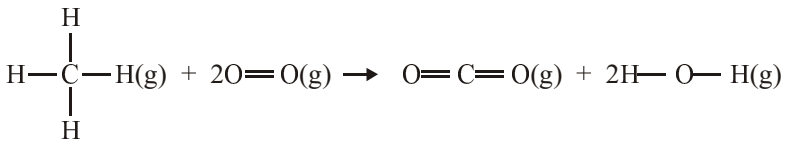
....

....

....

(3)

(b) Methane burns in oxygen according to the equation:



Use the average bond enthalpy data shown below to calculate the enthalpy change of this reaction.

Bond	Bond enthalpy/kJ mol^{-1}
C—H	+435
O==O	+498
C==O	+805
H—O	+464

(3)

(e) Methane is the feedstock in the manufacture of hydrogen according to the equation:

$$CH_4(g) + 2H_2O(g) \rightarrow CO_2(g) + 4H_2(g)$$

Given the enthalpy of formation data below, draw a **labelled** Hess's law cycle and use it to calculate the enthalpy change of this reaction.

Substance	Enthalpy of formation/kJ mol^{-1}
$CH_4(g)$	–75
$CO_2(g)$	–394
$H_2O(g)$	–242

(4)
(Total 10 marks)

35. In an experiment to find the enthalpy of neutralisation of a monobasic acid, HX, with an alkali, the following procedure was followed:

Step 1 25.0 cm^3 of 1.00 mol dm^{-3} dilute aqueous acid, HX, was measured into a polystyrene cup.

Step II A 0-100 °C thermometer was placed in the acid. The temperature of the acid was immediately read and recorded.

Step III 5.00 cm^3 portions of aqueous sodium hydroxide were added to the acid from a burette. After each addition, the temperature of the solution was read and recorded. The thermometer was removed and rinsed with water between each addition. A total of 50.0 cm^3 of aqueous sodium hydroxide was added.

(a) Suggest ONE change that could be made at **Step II** and ONE change that could be made at **Step III** to improve the accuracy of the experiment.

Step II ....

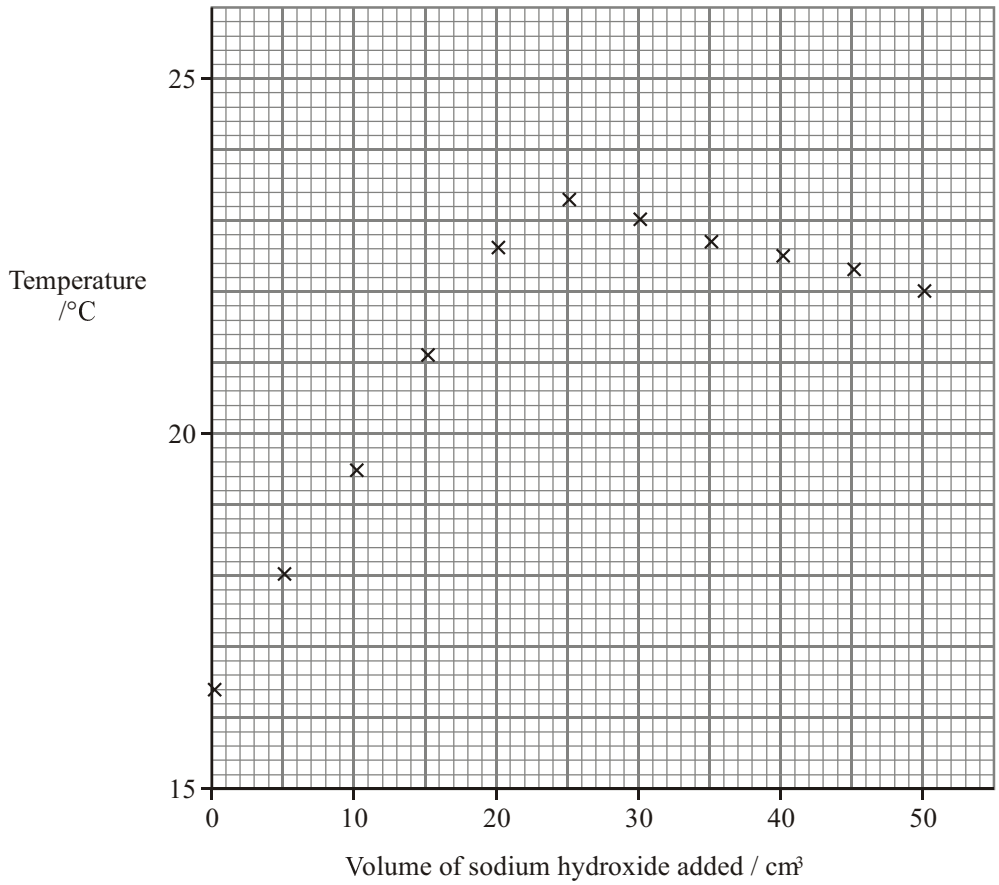
....

Step III ....

....

(2)

(b) The readings of temperature and volume are plotted on the grid. Draw two separate straight lines of best fit, extending the two lines so that they intersect.



(2)

(c) From the graph, read off the maximum temperature rise, ΔT, and the volume of aqueous sodium hydroxide added at neutralisation, V_N.

ΔT = ... °C V_N = ... cm^3

(2)

(d) (i) Use the formula below to calculate the heat evolved in the neutralisation.

$$\text{Heat evolved} = \frac{(V_N + 25) \times \Delta T \times 4.18}{1000} \text{ kJ}$$

(1)

(ii) Given that the amount (moles) of acid neutralised was 0.025 mol, calculate the enthalpy of neutralisation, ΔH_{neut}, in units of kJ mol^{-1}.

ΔH_{neut} = ... kJ mol^{-1}

(2)
(Total 9 marks)

36. Calculate the volume, measured at room temperature and pressure, of the following masses of gases.

Use the Periodic Table as a source of data.

[Molar volume of a gas is 24 dm^3 at room temperature and pressure.]

(a) 4 g of helium, He.

.....

(1)

(b) 4 g of hydrogen, H_2.

.....

(1)
(Total 2 marks)

37. This question is about ammonium dichromate(VI), $(NH_4)_2Cr_2O_7(s)$, which is slightly soluble in water. Ammonium dichromate(VI) can be prepared by the following reaction:

$$Na_2Cr_2O_7(aq) + 2NH_4NO_3(aq) \rightarrow (NH_4)_2Cr_2O_7(s) + 2NaNO_3(aq)$$

(a) What type of reaction is this?

....

(1)

(b) (i) Give the formulae of the ions present in ammonium dichromate(VI).

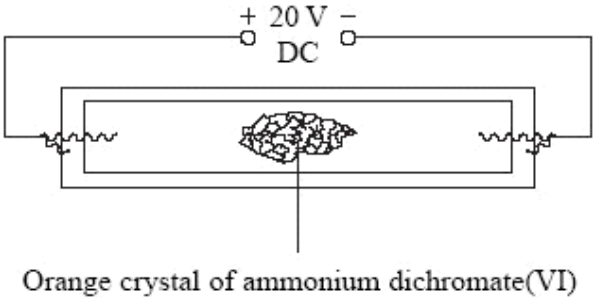
....

(2)

(ii) Re-write the equation as an ionic equation, but omit any spectator ions.

(1)

(iii) An orange crystal of ammonium dichromate(VI) was placed in the centre of a piece of moist filter paper supported on a glass microscope slide.



Describe what you would expect to **see** when electricity is passed through the circuit.

....

(1)

(c) (i) Calculate the mass of 1 mole of ammonium dichromate(VI). Use the Periodic Table as a source of data.

(1)

(ii) In an experiment to prepare ammonium dichromate(VI), 0.1 mol of sodium dichromate(VI) was used.

What is the maximum mass of ammonium dichromate(VI) which could be obtained?

(1)

(iii) What volume of ammonium nitrate solution of concentration 2 mol dm^{-3} would react exactly with 0.1 mol of sodium dichromate(VI)?

(1)

(iv) Describe how you would obtain pure, dry crystals of ammonium dichromate(VI) from the resultant mixture.

.....

.....

.....

.....

.....

(3)

(v) Suggest TWO reasons why the mass of ammonium dichromate(VI), prepared in this way, is likely to be less than the mass calculated in (ii).

.....

.....

.....

.....

(2)
(Total 13 marks)

38. A reaction of ammonium dichromate(VI) is shown by the following equation.

$$(NH_4)_2Cr_2O_7(s) \xrightarrow{\text{heat}} N_2(g) + 4H_2O(g) + Cr_2O_3(s)$$

(a) What type of reaction is this?

.....

(1)

(b) The enthalpy change for this reaction can be calculated from standard enthalpy changes of formation.

(i) State fully what is meant by the **standard enthalpy change of formation**, $\Delta H_f^{\ominus}$, of a compound.

....

....

....

....

(3)

(ii) Complete the Hess cycle for the reaction so that you can calculate the enthalpy change of the reaction from standard enthalpy changes of formation.

$$\boxed{(NH_4)_2Cr_2O_7(s)} \xrightarrow{\Delta H_r^{\ominus}} \boxed{N_2(g) + 4H_2O(g) + Cr_2O_3(s)}$$

$$\boxed{N_2(g) + \qquad\qquad\qquad\qquad}$$

(3)

(iii) What is the value of $\Delta H_f^{\ominus}[N_2(g)]$? .....

(1)

(iv) Calculate $\Delta H^{\ominus}_r$ for the reaction using the following data. Remember to include a sign and units in your answer.

$\Delta H_f^{\ominus}[(NH_4)_2Cr_2O_7(s)]$ = –1810 kJ mol^{-1}

$\Delta H_f^{\ominus}[H_2O(g)]$ = –242 kJ mol^{-1}

$\Delta H_f^{\ominus}[Cr_2O_3(s)]$ = –1140 kJ mol^{-1}

(3)

(c) In this reaction, water vapour is formed which condenses to liquid water on cooling. Is this reaction $H_2O(g) \rightarrow H_2O(l)$ exothermic or endothermic?

Justify your answer.

.....

.....

.....

.....

(2)
(Total 13 marks)

39. (a) (i) Potassium superoxide contains 54.9 % potassium by mass.
Show that the empirical formula of this compound is KO_2.

(3)

(ii) Give the oxidation number of oxygen in the compound KO_2.

.....

(1)

(b) Which of potassium nitrate or lithium nitrate has the higher thermal stability? Explain your answer.

.....

.....

.....

.....

.....

.....

.....

(3)
(Total 7 marks)

40. (a) Enthalpy changes can be calculated using average bond enthalpy data.

(i) The enthalpy change to convert methane into gaseous atoms is shown below.

$$CH_4(g) \rightarrow C(g) + 4H(g) \qquad \Delta H = +1664 \text{ kJ mol}^{-1}$$

Calculate the average bond enthalpy of a C—H bond in methane.

(1)

(ii) Use the data in the table below and your answer to (a)(i) to calculate the enthalpy change for

$$2C(g) + 2H_2(g) + Br_2(g) \rightarrow CH_2BrCH_2Br(g)$$

Bond	Average bond enthalpy / kJ mol^{-1}	Bond	Average bond enthalpy / kJ mol^{-1}
C—C	+348	H—H	+436
Br—Br	+193	C—Br	+276

(3)

(b) The standard enthalpy of formation of 1,2-dibromoethane, CH_2BrCH_2Br, is –37.8 kJ mol^{-1}.

Suggest the main reason for the difference between this value and your calculated value in (a)(ii).

....

....

....

(1)
(Total 5 marks)

41. In an experiment to find the enthalpy change for the reaction

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

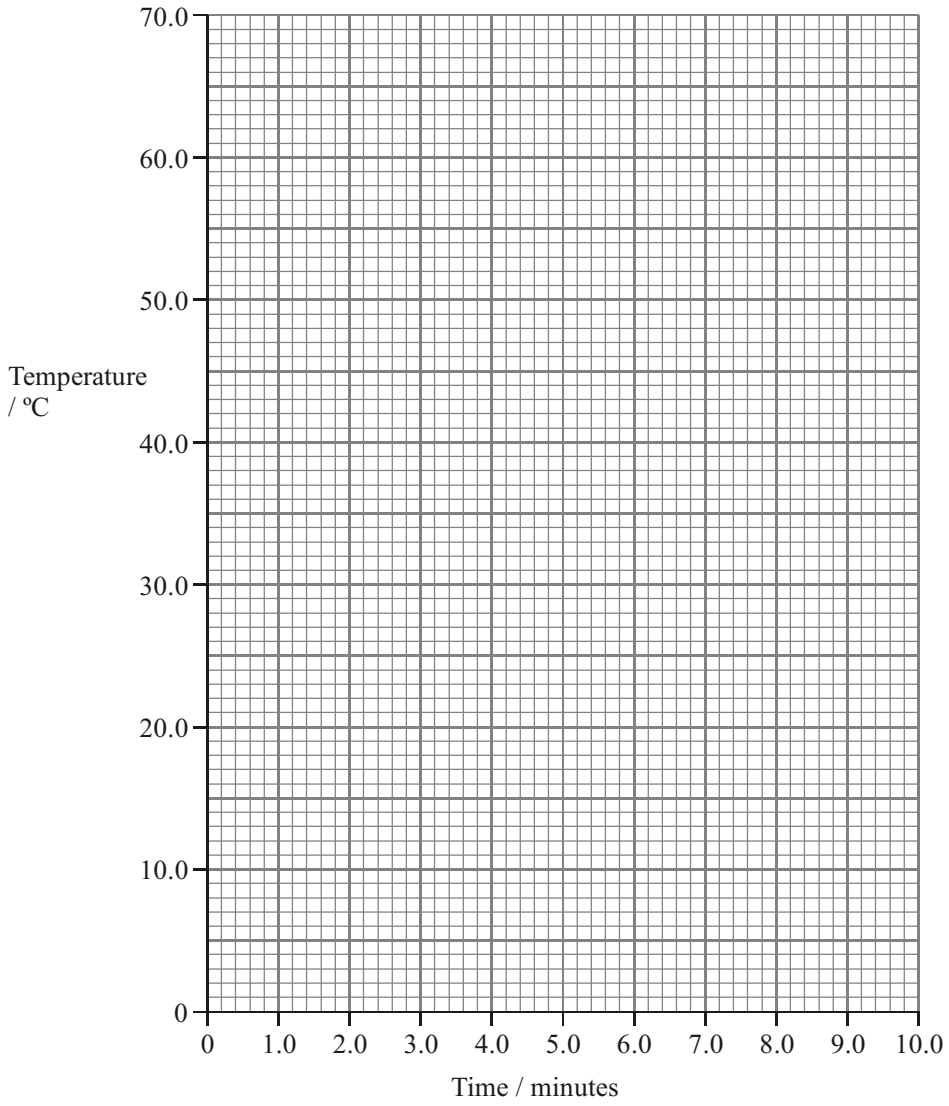
a student was given the following list of instructions:

- weigh out 5.0 g of zinc powder into a weighing bottle
- use a measuring cylinder to transfer 50 cm^3 of 1.0 mol dm^{-3} aqueous copper(II) sulphate into a polystyrene cup, firmly held in a 250 cm^3 beaker
- stir the solution with the thermometer and record the temperature to the nearest 0.5 °C
- continue to stir the solution, recording its temperature every minute
- at exactly 3.5 minutes, add the zinc powder to the aqueous copper(II) sulphate, stirring continuously
- record the temperature of the solution every minute from 4.0 to 9.0 minutes.

The temperature readings obtained are shown in the table below.

Time/min	0	1.0	2.0	3.0	4.0	5.0	6.0	7.0	8.0	9.0
Temperature/°C	20.0	20.0	20.0	20.0	63.0	60.5	59.0	57.0	55.5	53.0

(a) (i) Plot a graph of temperature against time on the grid below.



(2)

(ii) Use the graph to calculate the maximum temperature change, ΔT. Show clearly on the graph how you obtained your answer.

ΔT = ...°C

(2)

(iii) Give ONE reason why a series of temperature readings is obtained instead of just the starting and maximum temperatures.

.....

.....

(1)

(b) (i) Calculate the heat change, in joules.

The specific heat capacity of the solution is 4.18 J g^{-1} °C^{-1}.

(1)

(ii) What assumption have you made about the solution in your calculation in (i)?

....

.....

(1)

(iii) Calculate the amount (moles) of copper (II) sulphate, $CuSO_4$, in 50 cm^3 of a 1.0 mol dm^{-3} solution.

(1)

(iv) Calculate the enthalpy change for this reaction in kJ mol^{-1}.

(2)

(c) Suggest TWO improvements that could be made to the experimental procedure.

Give a reason for each.

Improvement 1 .....

....

....

Reason ....

....

....

Improvement 2 .....

....

....

Reason ....

....

....

(4)
(Total 14 marks)

42. (a) Barium carbonate can be converted into barium chloride solution by a reaction with hydrochloric acid. In a particular experiment, an excess of barium carbonate was added to 25 cm^3 of hydrochloric acid of concentration 1.0 mol dm^{-3}.

(i) Describe how you would obtain dry crystals of hydrated barium chloride, $BaCl_2.2H_2O$, from the reaction mixture.

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(4)

(ii) Write a balanced equation, including state symbols, for this reaction.

(2)

(iii) Calculate the number of moles of hydrochloric acid used in the experiment.

(1)

(iv) Calculate the mass of one mole of hydrated barium chloride, $BaCl_2.2H_2O$. Use the Periodic Table as a source of data.

(1)

(v) Calculate the theoretical mass of crystals which could be obtained.

(1)

(vi) Suggest a reason why this mass of crystals is unlikely to be obtained in practice.

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(1)

(b) (i) What colour do barium compounds produce in a flame test?

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(1)

(ii) When carrying out a flame test on a solid, state a suitable material on which it can be supported in the flame.

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(1)

(Total 12 marks)

43. When solutions of potassium carbonate and calcium chloride are mixed together, the following reaction takes place

$$CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2KCl(aq)$$

(a) Re-write the above equation as an ionic equation. Include state symbols, but omit any spectator ions.

(2)

(b) An experiment was carried out to measure the enthalpy change for this reaction. 50 cm^3 of a 1.00 mol dm^{-3} solution of potassium carbonate was added to 50 cm^3 of a 1.00 mol dm^{-3} solution of calcium chloride. The temperature fell by 1.5 °C.

(i) Calculate the energy taken in from the surroundings using the relationship

energy =	mass of solution	×	specific heat capacity of solution	×	temperature change
/J	/g		/J g^{-1} °C^{-1}		/°C

You may assume that

- 1.0 cm^3 of solution has a mass of 1.0 g.
- The specific heat capacity of the solution is 4.2 J g^{-1} °C^{-1}.

Energy taken in = ... J

(1)

(ii) How many moles of calcium chloride are used in this experiment?

(1)

(iii) Calculate the enthalpy change for the reaction, giving your answer to two significant figures. Include a sign and units in your answer.

(2)

(iv) Which measurement is likely to have caused the major source of error in this experiment? Explain your answer.

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(1)

(v) What apparatus should be used to contain the reaction mixture during this experiment?

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(1)

(c) If the experiment in (b) was repeated, but using only 25 cm^3 of each solution, predict what the fall in temperature would be.

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(1)

(Total 9 marks)

44. (a) Sodium chloride, NaCl, can be made by the reaction of sodium with chlorine.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

(i) Calculate the maximum mass of sodium chloride which could be obtained from 92 g of sodium.

(2)

(ii) Calculate the concentration of the solution obtained when this mass of sodium chloride is dissolved in water and made up to a volume of 10 dm^3 with distilled water.

(1)

(iii) Calculate the volume of chlorine gas required to react with 92 g of sodium.

[1 mol of gas occupies 24 dm^3 under the conditions of the experiment]

(2)

(b) Describe the structure of solid sodium metal and explain why it conducts electricity.

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(3)

(c) (i) Define the term **first ionisation energy.**

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(3)

(ii) Explain why the first ionisation energy of chlorine is higher than that of sodium.

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....

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(2)
(Total 13 marks)

45. (a) State Hess's Law.

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(2)

(b) Methane burns in oxygen.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

(i) Calculate the enthalpy change for this reaction, using the bond enthalpies given below.

	Bond enthalpy / kJ mol^{-1}
C – H	+435
O = O	+498
C = O	+805
H – O	+464

(3)

(ii) State the name of this enthalpy change.

.....

(1)

(iii) The value of this enthalpy change, under standard conditions, is –890 kJ mol^{-1}. State the meaning of **standard conditions.**

.....

....

.....

(2)

(iv) Suggest, with a reason, why the enthalpy change calculated in (i) is different from the standard value quoted in (iii).

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(2)

(c) Although the reaction between methane and oxygen is exothermic, it does not occur unless the mixture is ignited.

Use these facts to explain the difference between thermodynamic and kinetic stability.

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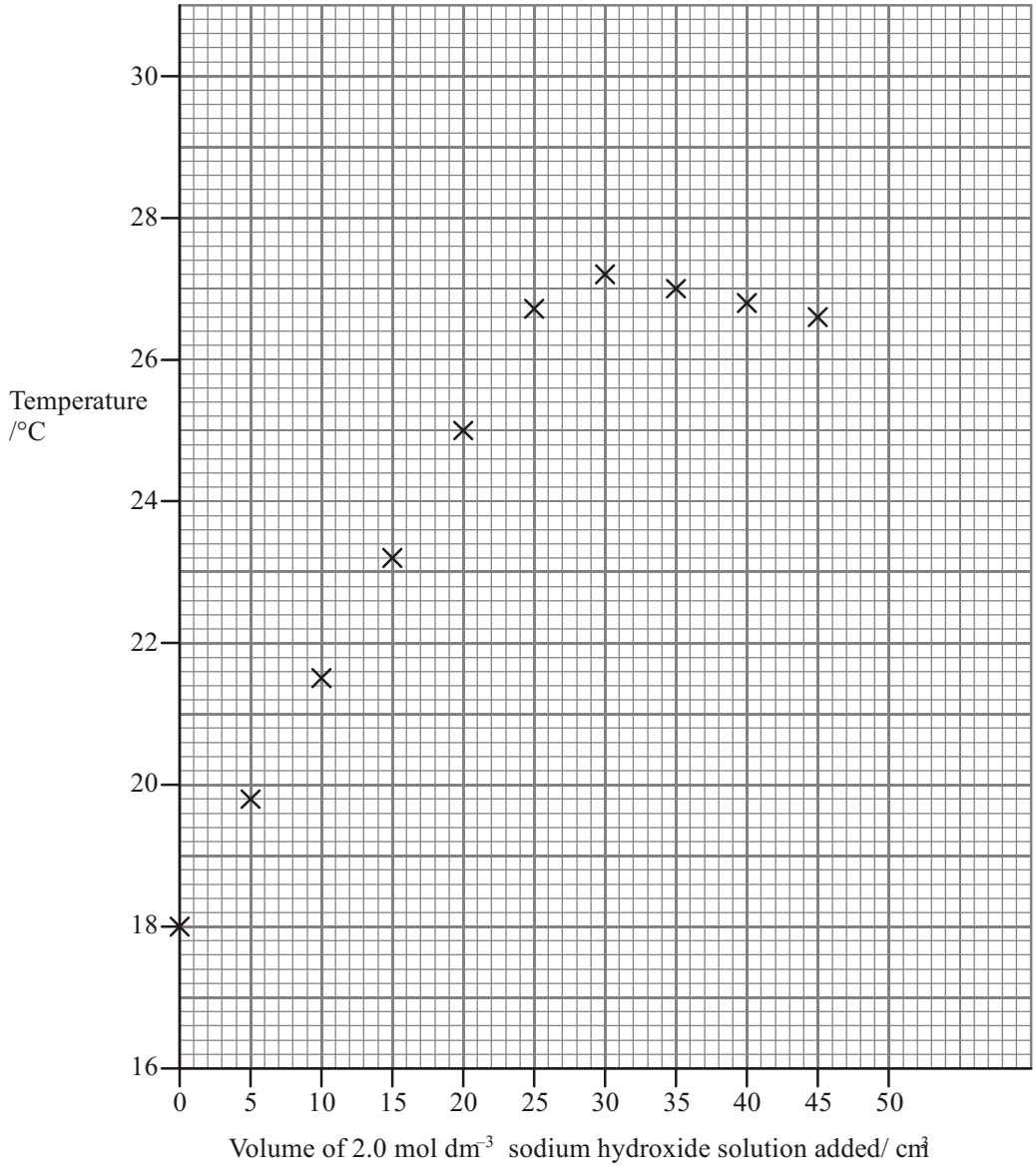
(4)
(Total 14 marks)

46. A student carried out an experiment to find the concentration of a solution of nitric acid and also its enthalpy of neutralisation.

- The solutions of nitric acid and sodium hydroxide were allowed to reach the same temperature.
- 50.0 cm^3 of the nitric acid was pipetted into a polystyrene cup.
- A burette was filled with a solution of 2.0 mol dm^3 sodium hydroxide, NaOH.
- The initial temperature of the acid was recorded.
- The sodium hydroxide was added to the acid in 5.0 cm^3 portions.
- After each addition, the mixture was stirred and the maximum temperature recorded.
- This was repeated until 45 cm^3 of the sodium hydroxide solution had been added.

The student plotted the results, as shown below.



(a) Complete the graph by drawing two intersecting straight **lines** of best fit.

(1)

(b) The point where the lines cross represents the neutralisation of the nitric acid by the sodium hydroxide solution.

Use the graph to find:

(i) the volume of 2.0 mol dm^3 sodium hydroxide solution, NaOH, that reacts exactly with the 50 cm^3 of the nitric acid.

....

(1)

(ii) the maximum temperature **change,** ΔT, in the reaction.

.....

(1)

(c) The information in (b)(i) and the equation below can be used to calculate the concentration of the nitric acid.

$$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

(i) Calculate the amount (moles) of sodium hydroxide used to neutralise the 50 cm^3 of nitric acid.

(1)

(ii) Write the amount (moles) of nitric acid in 50.0 cm^3 of the solution.

.....

(1)

(iii) Hence calculate the concentration of nitric acid, HNO_3, in mol dm^3.

(2)

(d) (i) Use the data from (b) to calculate the heat change for this reaction.

The density of the mixture produced at neutralisation is 1.0g cm^{-3} and the specific heat capacity of the mixture is 4.2 J g^{-1} °C^{-1}.

Heat change = mass × specific heat capacity × ΔT

(2)

(ii) Use your answer from (d)(i) and (c)(iii) to calculate the enthalpy of neutralisation per mole of nitric acid, HNO_3. Include a sign and units with your answer.

(3)

(e) The enthalpy of neutralisation found by this method may be **less exothermic** than the data book value because of heat loss.

Suggest ONE way to reduce the error due to heat loss.

....

....

(1)
(Total 13 marks)

47. (a) Calculate the number of atoms in 3.50 g of lithium.

Use the Periodic Table as a source of data.

[The Avogadro constant, $L = 6.02 \times 10^{23}\ mol^{-1}$]

(2)

(b) The equation for the reaction of lithium with hydrochloric acid is shown below.

$$2Li(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2(g)$$

(i) Rewrite this equation as an ionic equation, omitting the spectator ions.

(1)

(ii) Draw a 'dot and cross' diagram of lithium chloride showing **all** the electrons. Indicate charges clearly on your diagram.

(2)

(iii) The value of the standard enthalpy change for the reaction, $\Delta H^{\ominus}$, is –557 kJ mol^{-1}. State TWO of the reaction conditions necessary for this enthalpy change to be **standard**.

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(2)
(Total 7 marks)

48. (a) An atom of argon has mass number 40. Complete the table below showing the composition of this argon atom.

Use the Periodic Table as a source of data.

Protons	
Electrons	
Neutrons	

(2)

(b) An atom of potassium has mass number 39. Explain, in terms of atomic structure, why argon comes before potassium in the Periodic Table.

....

....

....

....

(1)

(c) A sample of argon contains a mixture of isotopes as shown below.

Calculate the relative atomic mass of argon in the sample. Give your answer to **three** significant figures.

Isotopic mass	% abundance
36.0	1.34
38.0	0.160
40.0	98.5

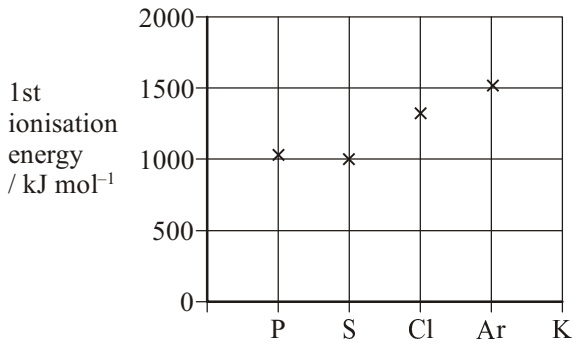
(2)

(d) Write the electron configuration of argon in s, p notation.

....

(1)

(e) The chart shows the first ionisation energy of some elements in the third period of the Periodic Table.



(i) Write the chemical equation, with state symbols, which corresponds to the first ionisation energy of argon.

.....

(1)

(ii) On the chart, add a cross to show the first ionisation energy of potassium. Justify your choice of the position of the cross.

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(2)

(iii) Explain why there is a small decrease in first ionisation energy going from phosphorus to sulphur.

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(2)

(iv) Explain why there is an increase in first ionisation energy going from sulphur to chlorine.

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(2)

(f) Suggest why argon is used to fill some types of light bulbs.

....

....

(1)
(Total 14 marks)

49. (a) Sodium iodide reacts with chlorine to produce sodium chloride and iodine.

(i) State the oxidation numbers of the iodine and chlorine species in the spaces provided.

$2NaI$	+	Cl_2	→	$2NaCl$	+	I_2
.....		.....		.....		.....

(2)

(ii) Use these oxidation numbers to explain why this reaction is a redox reaction.

.....

....

(2)

(iii) Calculate the maximum mass of iodine that could be produced from 30.0 g of sodium iodide.

(3)

(iv) Calculate the volume of chlorine gas required to produce this amount of iodine.

[1 mol of gas occupies 24 dm^3 under the conditions of the experiment]

(1)

(b) (i) Give the colour of iodine and its physical state at room temperature and pressure.

Colour .....

Physical state .....

(2)

(ii) Write an equation, including state symbols, to represent the process occurring when the first ionisation energy of iodine atoms is measured.

.....

(2)

(Total 12 marks)

50. (a) (i) Explain why a water molecule does **not** have a linear shape.

.....

.....

.....

(2)

(ii) State the HOH bond angle in water and explain why it has this value.

.....

....

.....

(2)

(b) (i) Draw the boron trichloride molecule, BCl_3, making its shape clear. Mark in the bond angle on your diagram.

(2)

(ii) Explain why a B–Cl bond is polar.

.....

.....

(1)

(iii) Explain why a BCl_3 molecule is non-polar.

....

.....

....

.....

(1)

(iv) Name the strongest intermolecular force between boron trichloride molecules.

.....

.....

(1)

(c) A compound of phosphorus and chlorine has the composition by mass shown below.

Element	% by mass
P	14.9
Cl	85.1

Calculate the empirical formula of this compound.

(2)
(Total 11 marks)

51. (a) Define the term **standard enthalpy of formation**.

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....

....

....

(3)

(b) The dissociation of phosphorus pentachloride is a reversible reaction.

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

(i) Use the values of enthalpy of formation given to calculate ΔH for the forward reaction.

	ΔH_f / kJ mol^{-1}
$PCl_5(g)$	– 399
$PCl_3(g)$	– 306

(1)

(ii) Explain, with reasons, the effect that raising the temperature would have on the composition of the equilibrium mixture.

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(2)

(iii) Other than by changing the temperature, suggest how the amount of PCl_5 present at equilibrium could be increased. Give a reason for your answer.

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....

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(2)
(Total 8 marks)

52. A metal carbonate decomposes on heating to give an oxide and carbon dioxide.

$$XCO_3(s) \rightarrow XO(s) + CO_2(g)$$

where **X** is the metallic element.

In an experiment to find the identity of **X**, 5.75 g of the solid XCO_3 was heated until there was no further change in mass; 3.55 g of solid XO was produced.

(a) Explain why it was necessary to heat the carbonate until there was no further change in mass.

....

....

(1)

(b) (i) Calculate the mass of carbon dioxide gas given off.

(1)

(ii) Calculate the amount (moles) of carbon dioxide gas given off.

(1)

(iii) Use the answer from (ii) to state the amount (moles) of XCO_3 in 5.75 g of the solid.

.....

(1)

(iv) Calculate the molar mass of XCO_3.

(1)

(v) Use your answer from (iv) and the relative atomic masses of carbon and oxygen to calculate the relative atomic mass of **X** in XCO_3.

(1)

(c) All measurements of mass have some uncertainty. In this case, this leads to an error of ±0.91 % in the molar mass of XCO_3.

(i) Use this information and your answer to (b)(iv) to calculate the error in the molar mass of XCO_3.

(1)

(ii) Hence suggest the range of possible values for the molar mass of XCO_3.

.....

(1)

(iii) Hence give the range of possible values of the relative atomic mass of **X**.

(1)

(iv) Use the Periodic Table and your answer to (c)(iii) to suggest possible identities of

metal **X**.

.....

....

(1)
(Total 10 marks)

53. (a) The mass of one atom of the isotope $^{79}_{35}Br$ is 1.31×10^{-22} g. The molar mass of $^{79}_{35}Br$ is 79.0 g mol^{-1}.

Use this information to calculate a value for the Avogadro constant. Give your answer to **three** significant figures.

(2)

(b) According to the Periodic Table, the relative atomic mass of naturally occurring bromine is 80.

What information can you deduce from this about naturally occurring bromine? (No calculation is expected.)

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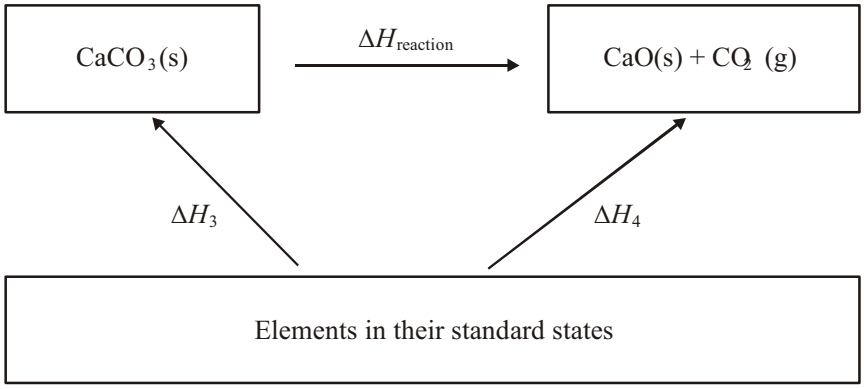
(1)
(Total 3 marks)

54. The enthalpy change for the thermal decomposition of calcium carbonate cannot be measured directly, but can be found by carrying out two reactions as shown in the Hess cycle below.



(a) Suggest ONE reason why it is difficult to measure $\Delta H_{reaction}$ directly by experiment.

....

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....

....

(1)

(b) In an experiment to find ΔH_1 a student added 2.00 g of finely powdered calcium carbonate to 20.0 cm^3 of 2.50 mol dm^{-3} hydrochloric acid solution (an excess) in a polystyrene container. The temperature rose from 20.5 °C to 23.0 °C.

(i) Why is the calcium carbonate used in this experiment finely powdered, rather than in lumps? Explain why this is important for an accurate result.

....

.....

....

(2)

(ii) Calculate the energy change using the relationship below.

Energy change	=	4.2	×	mass of solution	×	temperature change
/J		/J g^{-1} K^{-1}		/g		/K

Assume that the mass of the solution is 20 g.

(1)

(iii) Calculate the enthalpy change, ΔH_1. Include a sign and units in your answer.
[The molar mass of $CaCO_3$ is 100 g mol^{-1}]

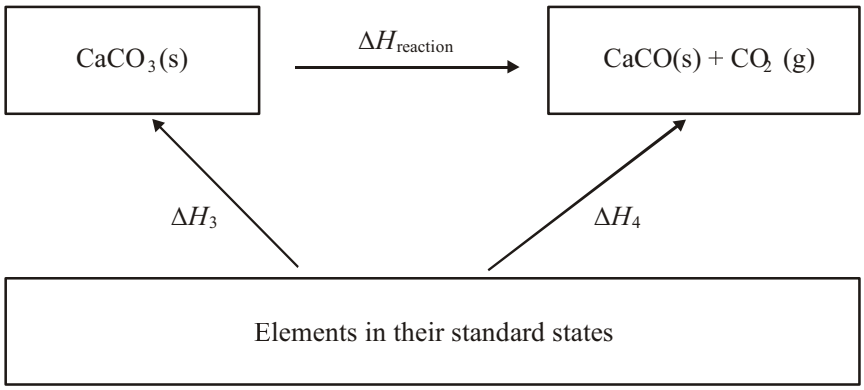
(3)

(iv) In another experiment, the value of ΔH_2 was found to be –181 kJ mol^{-1}.

Use this result and your answer to (iii) to calculate the value of $\Delta H_{reaction}$.

(2)

(c) The student checked the experimental results using information from the *Book of data* in another Hess cycle.



Name the enthalpy change represented by ΔH_3.

....

(1)

(Total 10 marks)

55. (a) Complete the table below which is about isotopes and an ion of magnesium.

	Numbers of		
	Protons	Neutrons	Electrons
$^{24}_{12}Mg$	12	12	
$^{26}_{12}Mg$	12		12
$^{24}_{12}Mg^{2+}$	12	12	

(3)

(b) Complete the electronic configurations of magnesium and chlorine atoms.

	1s	2s	2p			3s	3p		
Mg	↓↑	↓↑							

(1)

	1s	2s	2p			3s	3p		
Cl	↓↑	↓↑							

(1)

(c) Write the equation, including state symbols, for the reaction of magnesium with chlorine.

....

(2)

(d) The mass spectrum of a sample of chlorine molecules shows three molecular peaks. These are formed from the molecules shown below.

Molecule	Percentage abundance
^{35}Cl-^{35}Cl	56.25
^{35}Cl-^{37}Cl	37.50
^{37}Cl-^{37}Cl	6.25

Calculate the relative molecular mass of chlorine in this sample.

(2)

(e) Calculate the volume of 4.73 g of chlorine gas at 100 °C.

[The molar volume of a gas at 100 °C = 30.6 dm^3 mol^{-1}]

(2)

(f) State and explain the type of bond that exists in solid magnesium.

Type .....

Explanation ....

....

....

(3)

(g) State the type of bond that exists in magnesium chloride. Draw a dot and cross diagram showing the **outer** shell electrons.

Type .....

Dot and cross diagram

(3)
(Total 17 marks)

56. Sodium azide, NaN_3, is used to inflate air bags in cars because, in a collision, the sodium azide decomposes rapidly to produce nitrogen gas.

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

(a) (i) Calculate the mass of one mole of sodium azide, NaN_3.
Use the Periodic Table as a source of data.

(1)

(ii) What mass of sodium azide would be needed to liberate 48 dm^3 of nitrogen?
[Molar volume of a gas is 24 dm^3 mol^{-1} at room temperature and pressure]

(2)

(b) What safety problem would arise when disposing of the air bag after a collision? Justify your answer.

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(2)
(Total 5 marks)

57. An experiment was carried out to find the enthalpy change for the reaction of zinc powder with copper(II) sulphate solution.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

50cm^3 of copper(II) sulphate solution, of concentration 1.0 mol dm^{-3}, was put into a polystyrene cup and the temperature of the solution measured. After one minute, 5.0 g of zinc powder was added, the mixture stirred with a thermometer and the temperature measured every 30 s.

(a) (i) What is meant by a spectator ion?

.....

.....

(1)

(ii) Give the formula of the spectator ion in this reaction.

.....

(1)

(iii) Write the equation for this reaction, omitting the spectator ion.

(1)

(b) How would you measure the 50 cm^3 of copper(II) sulphate solution?

....

....

(1)

(c) Give TWO reasons why it is better to use a polystyrene cup, rather than a metal container, to obtain more accurate results.

....

....

....

....

(2)

(d) Calculate the number of moles of each of the reactants and hence deduce which reactant is completely used up.
Use the Periodic Table as a source of data.

Moles of zinc powder

Moles of copper(II) sulphate

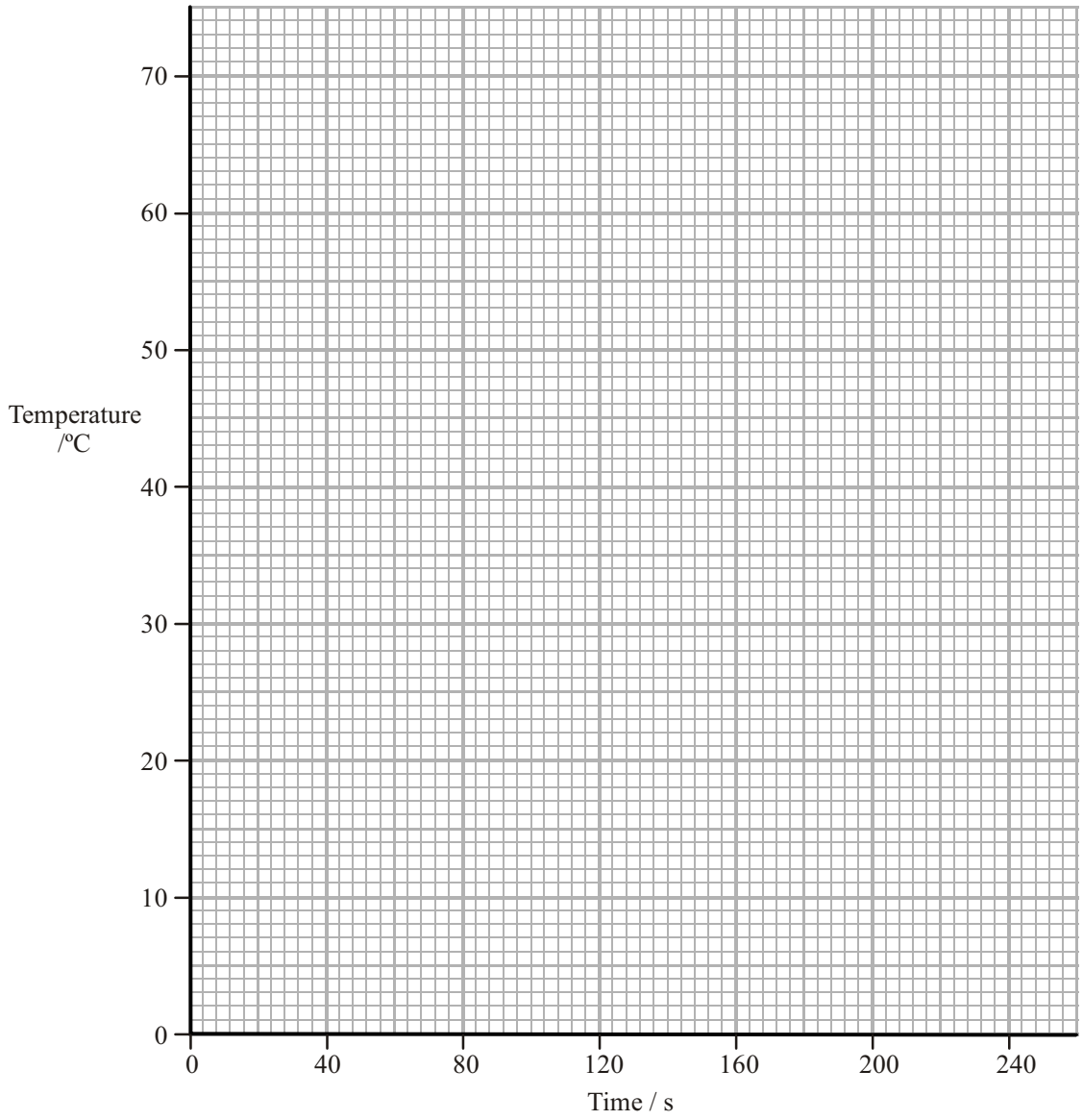
Reactant used up .....

(3)

(e) The following results were obtained.

Time /s	0	60	90	120	150	180	210
Temperature /°C	22	22	60	65	63	61	59

(i) On the graph paper below, plot the results of this experiment.



(2)

(ii) Explain the shape of your graph

.....

....

.....

....

(2)

(iii) The maximum recorded temperature in this experiment was 65°C. Use your graph to estimate a more accurate maximum temperature.

....

(1)

(f) (i) Calculate the energy change in this experiment using your answer to (e)(iii) and the relationship

energy change =	mass of solution ×	specific heat capacity of solution ×	temperature **rise**
/J	/g	/J °C^{-1} g^{-1}	/°C

You may assume that

- 1.0 cm^3 of solution has a mass of 1.0 g
- The specific heat capacity of the solution is 4.2 J °C^{-1} g^{-1}

(1)

(ii) Use your answers to (d) and (f)(i) to calculate ΔH for this reaction. Include a sign and units in your answer.

(3)
(Total 18 marks)

58. (a) Complete the table below which is about the isotopes and an ion of bromine.

	Number of		
	protons	neutrons	electrons
$^{79}_{35}Br$	35		35
$^{81}_{35}Br$		46	35
$^{81}_{35}Br^{-}$	35	46	

(3)

(b) Complete the electronic configurations of

Na $1s^2$ .....

Br $1s^2$ ....

(2)

(c) Explain why the isotopes of bromine have identical chemical reactions.

....

....

....

(1)

(d) What instrument could be used to measure the abundance and mass of the isotopes of bromine?

....

(1)

(e) The isotopic abundance of bromine is shown below.

Relative isotopic mass	Percentage abundance
78.93	50.54
80.91	49.46

Calculate the relative atomic mass of bromine. Give your answer to **four** significant figures.

(2)

(f) State the types of bonding present in bromine **liquid**

between the atoms .....

between the molecules ....

(2)
(Total 11 marks)

59. (a) Define the term **Avogadro constant**.

....

.....

....

(2)

(b) **Z** is a Group 0 element.

(i) 1.907g of **Z** contains 2.87×10^{22} atoms of **Z**.

Calculate the relative atomic mass of **Z**.
[Avogadro constant = 6.02×10^{23} mol^{-1}]

(2)

(ii) Suggest the identity of **Z**.

....

(1)

(c) Potassium superoxide, KO_2, reacts with water as follows:

$$2KO_2(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2O_2(aq) + O_2(g)$$

(i) Calculate the mass of potassium superoxide needed to produce 3.09 g of hydrogen peroxide.

[Molar mass of potassium superoxide, KO_2: 71g mol^{-1}. Molar mass of hydrogen peroxide, H_2O_2 : 34 g mol^{-1}]

(3)

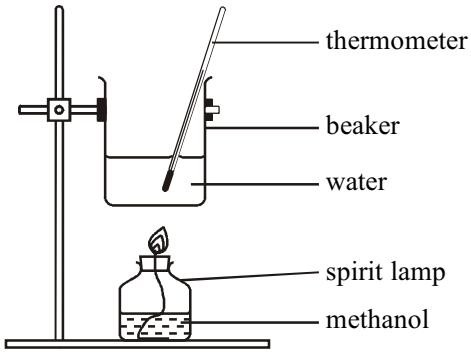
(ii) Calculate the volume of oxygen produced from the reaction in (i).

[Molar volume of oxygen under the conditions of the reaction = 24.0 dm^3 mol^{-1}]

(1)
(Total 9 marks)

60. The apparatus used and the recordings made by a student, carrying out an experiment to determine the enthalpy of combustion of methanol, are shown below.

Diagram



Results

Molar mass (methanol) = 32 g mol^{-1}
Volume of water in beaker = 50 cm^3
Mass of water in beaker = 50 g

Weighings
Spirit lamp + methanol before combustion = 163.78 g
Spirit lamp + methanol after combustion = 163.44 g

Temperatures
Water before heating = 22.0 °C
Water after heating = 43.5 °C

Specific heat capacity of water = 4.18 J g^{-1} $°C^{-1}$

Observations

- When the spirit lamp was being weighed its mass was continually falling.
- A black substance formed on the bottom of the beaker as the methanol burned.

(a) (i) Calculate the amount (moles) of methanol, CH_3OH, burned.

(2)

(ii) Calculate the heat gained by the water. Give your answer in kJ.

(2)

(iii) Use your values from (i) and (ii) to calculate the enthalpy of combustion of methanol in kJ mol^{-1}. Include a sign with your answer.

ΔH = ... kJ mol^{-1}

(2)

(b) (i) The thermometer used in the experiment can be read to an accuracy of ±0.5 °C. Calculate the percentage error in the temperature change.

(1)

(ii) Calculate the maximum temperature change that could have occurred during the experiment.

(1)

(c) (i) Give a reason why the mass of the spirit lamp fell as it was being weighed.

....

.....

(1)

(ii) Suggest the identity of the black substance that forms on the beaker. State the effect on the value of the enthalpy of combustion obtained.

.....

....

.....

(2)
(Total 11 marks)

61. 25.0 cm^3 of a solution contains 0.020 mol of ethanoic acid.

Calculate its concentration in mol dm^{-3}.

(Total 1 mark)

62. Nickel is one of the elements in the d-block of the Periodic Table.

(a) Complete the electron configuration of a nickel atom using the s, p, d notation.

Use the Periodic Table as a source of data.

$1s^2$ ....

(2)

(b) A sample of nickel consists of three isotopes. Their percentage abundances are shown in the table below.

Isotope	Percentage Abundance
^{58}Ni	69.02
^{60}Ni	27.32
^{62}Ni	3.66

Calculate the average relative atomic mass of nickel.

(2)

(c) Nickel reacts with carbon monoxide to give the compound nickel carbonyl, $Ni(CO)_4$.

$$Ni(s) + 4CO(g) \rightarrow Ni(CO)4(g)$$

Calculate the volume of carbon monoxide required to react completely with 5.9 g of nickel.

Use the Periodic Table as a source of data.
[The molar volume of a gas is 24 dm^3 mol^{-1} at room temperature and pressure.]

(2)

(d) Finely powdered nickel reacts slowly with dilute sulphuric acid to form a solution of nickel(II) sulphate.

$$Ni(s) + H_2SO_4(aq) \rightarrow NiSO_4(aq) + H_2(g)$$

(i) Re-write this equation in an ionic form, omitting the spectator ion. State symbols are not required.

(1)

(ii) A sample of finely powdered nickel is added to a slight excess of dilute sulphuric acid.

Describe the practical steps you would take to obtain dry crystals of hydrated nickel(II) sulphate from the mixture, after the reaction is complete.

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(3)

(iiii) Calculate the maximum mass of hydrated nickel(II) sulphate, $NiSO_4.7H_2O$, which could be formed from 2.95 g of nickel.

(2)
(Total 12 marks)

63. Calcium hydroxide decomposes on strong heating to form calcium oxide and water.

$$Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(l)$$

Two samples of **calcium hydroxide** were taken, each weighing exactly 1.00 g.

The first sample was cautiously added to 25.0 cm^3 of dilute hydrochloric acid contained in a glass beaker. The temperature rise was measured and found to be 16.5 °C.

The other sample was heated for some time. It was then allowed to cool and then added to another 25.0 cm3 portion of hydrochloric acid as before. In this case the temperature rose by 25.5 °C.

In both cases, the acid used was an excess.

(a) (i) Calculate the energy produced by the reaction of each solid with the acid.

Use the relationship

Energy produced	=	mass of solution	×	4.2	×	temperature rise
/ J		/ g		/ J °C^{-1} g^{-1}		/°C

You may assume that 1.0 cm^3 of solution has a mass of 1.0 g. Ignore the mass of the solid.

For the solid calcium hydroxide

For the solid calcium oxide

(1)

(ii) How many moles of calcium hydroxide were used in each experiment?
[Molar mass of $Ca(OH)_2$ = 74.0 g mol^{-1}]

(1)

(iii) Using your answers to (a)(i) and (ii), calculate the enthalpy changes for each reaction.

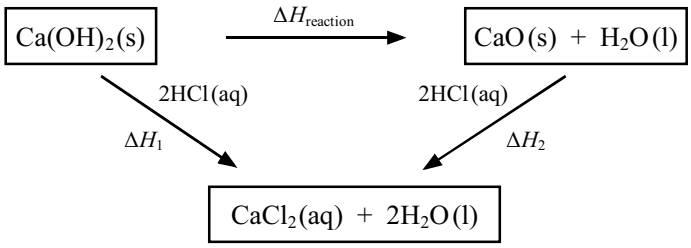
Give your answers to **two significant figures**. Include the sign and units for each answer.

For the solid calcium hydroxide, ΔH_1

For the solid calcium oxide, ΔH_2

(2)

(b) A Hess cycle for all these reactions is shown below.



(i) Use this Hess cycle and your answers in (a)(iii) to calculate $\Delta H_{reaction}$. Include a sign and units.

(2)

(ii) Apart from the approximations involved in using the equation given in (a)(i), give TWO other potential sources of error which are likely to affect the accuracy of the results.

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(2)

(iii) Suggest why $\Delta H_{reaction}$ is difficult to determine directly by experiment.

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(1)
(Total 9 marks)

64. A gaseous hydrocarbon, **W**, is a product formed in the cracking of eicosane, $C_{20}H_{42}$. **W** decolourises bromine, forming compound **X**.

When **X** is reacted with aqueous potassium hydroxide, compound **Y** is formed.

When a solution of **Y** is refluxed with an excess of acidified potassium dichromate(VI), compound **Z** is formed.

Compound **Z** contains carbon, hydrogen and oxygen only.

(a) (i) On complete combustion, 0.10 g of **Z** produced 53 cm^3 of carbon dioxide and 0.020 g of water at room temperature and pressure.

Calculate the empirical formula of compound **Z**.
[Molar volume of a gas is 24 000 cm^3 mol^{-1} at room temperature and pressure]

(3)

(ii) The molar mass of **Z** is 90 g mol^{-1}. Find the molecular formula of **Z**.

(1)

(iii) A solution made by dissolving 0.900 g of compound **Z** in water is titrated with sodium hydroxide solution. 20.0 cm^3 of sodium hydroxide solution of concentration 1.00 mol dm^{-3} is required for complete neutralisation.

Deduce the structural formula of compound **Z**.

(2)

(iv) Deduce the structural formulae of compounds **W**, **X**, and **Y**.

(3)

(v) Suggest a balanced equation for the cracking of eicosane.

(1)

(b) Compound **Y** can be made in one step from compound **W**.

State the reagents needed for this reaction.

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(2)
(Total 12 marks)

65. (a) Complete the electronic configuration of a copper atom and a bromide ion.

(i) Copper atom, Cu $1s^2 2s^2 2p^6 3s^2 3p^6$ .....

(1)

(ii) Bromide ion, Br^- $1s^2 2s^2 2p^6 3s^2 3p^6$ .....

(1)

(b) Define the term **relative atomic mass**.

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(2)

(c) The following data were obtained for a mass spectrum of a sample of copper.

Relative isotopic mass	Percentage abundance
62.93	69.17
64.93	30.83

Calculate the relative atomic mass of this sample of copper. Give your answer to two decimal places.

(2)

(d) Copper occurs naturally as the mineral malachite. The composition, by mass, of malachite is as follows:

Cu = 57.5% C = 5.4% O = 36.2% H = 0.9%

(i) Calculate its empirical formula.

(2)

(ii) The molar mass of malachite is 221 g mol^{-1}. Calculate its **formula**.

(1)

(e) Copper forms a chloride, $CuCl_2$. Use the data below to calculate the maximum and the minimum values for the molar mass of $CuCl_2$.

Data : Relative isotopic masses of chlorine are 35 and 37.
Relative isotopic masses of copper are 63 and 65.

(2)
(Total 11 marks)

66. This question is about ammonia, NH3, which is produced as shown in the following equation.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(a) Use oxidation numbers to explain why this is a redox reaction.

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(2)

(b) (i) Use the average (mean) bond enthalpy data to calculate a value for the enthalpy change for this reaction. You are reminded to show **all** your working.

Bond	Average bond enthalpy / kJ mol^{-1}
N≡N	944
H—H	436
N—H	388

(3)

(ii) The actual standard enthalpy change for this reaction is –92 kJ mol^{-1}. Explain why the value you calculated in (b)(i) is not the same as this.

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(1)

(iii) At room temperature, a mixture of nitrogen and hydrogen is thermodynamically unstable with respect to ammonia, but is kinetically stable.

Use the data in (b)(i) and (ii) to help you explain why this mixture is

thermodynamically unstable

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kinetically stable

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(3)

(c) The manufacturer of ammonia would like to achieve a high rate of reaction and a high equilibrium yield of product.

(i) State and explain, in terms of collision theory, TWO ways to increase the rate of the reaction. An increase in pressure does **not** alter the rate in this process.

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(6)

(ii) State and explain TWO ways to increase the equilibrium yield of ammonia.

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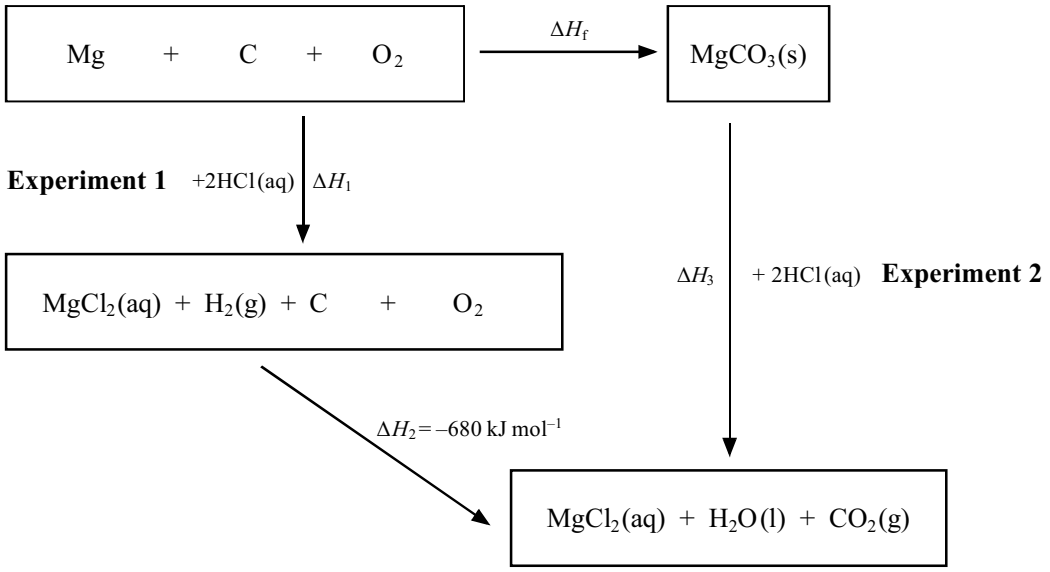
(4)
(Total 19 marks)

67. Two experiments were carried out in order to calculate the enthalpy change of formation of magnesium carbonate, $MgCO_3$.

A Hess cycle for these reactions is shown below.



(a) Complete the Hess cycle above for the formation of magnesium carbonate from its elements by balancing the equations and adding state symbols.

(2)

(b) In **Experiment 1** the temperature of 100 cm^3 of hydrochloric acid was measured. After one minute, 0.100 g of magnesium was added to the excess acid and the temperature measured every minute. The following results were obtained:

Time / min	0	1	2	3	4	5	6
Temp / °C	21.0	21.0	25.3	25.1	24.9	24.8	24.7

(i) How many moles of magnesium were used in this experiment?

Use the Periodic Table as a source of data.

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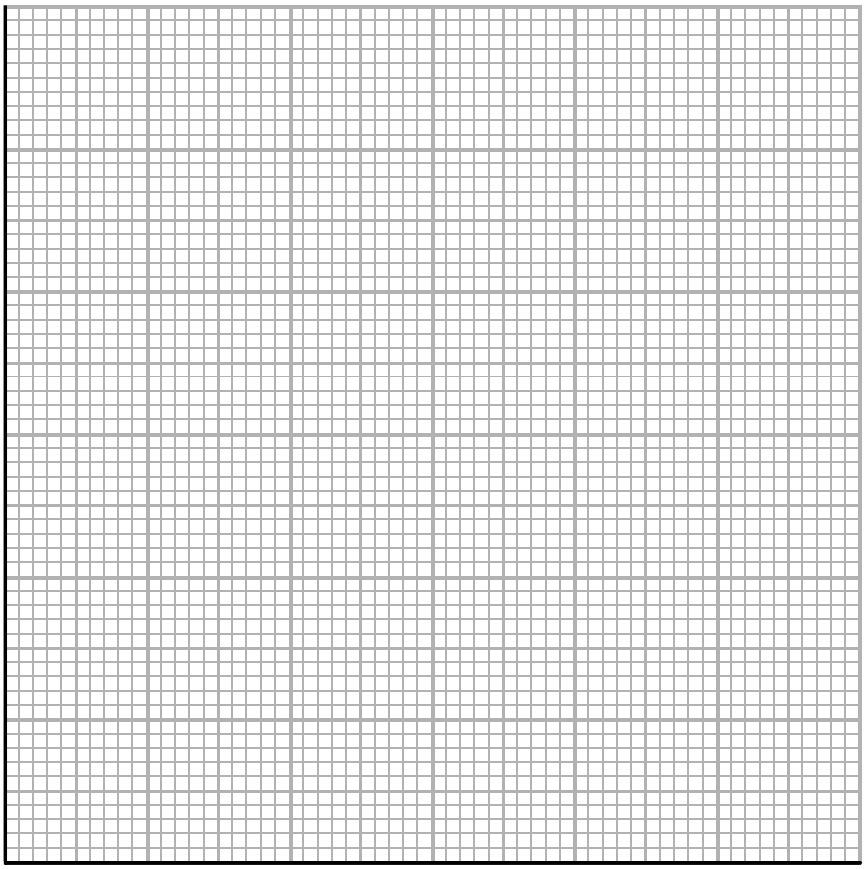
(1)

(ii) The initial concentration of the hydrochloric acid was 2.00 mol dm^{-3}.

Calculate the number of moles of hydrochloric acid at the start and hence the number remaining at the end of the experiment.

(3)

(iii) Plot the graph of temperature against time.



(2)

(iv) Calculate the energy change in this experiment assuming the temperature rise is 4.5 °C. Use the expression

Energy change (J) = 4.2 × mass of solution × temperature change

[Assume that 1 cm^3 of solution has a mass of 1 g]

(1)

(v) Use your answer to (iv) to calculate ΔH_1 for one mole of magnesium reacting with hydrochloric acid. Include a sign and units in your answer.

(2)

(vi) Suggest why a temperature rise of 4.5 °C was used in the calculation in (iv).

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(1)

(c) 2.2 g of magnesium carbonate was added to 100 cm^3 of the same acid in **Experiment 2**.

The temperature changed from 21.0 °C to 23.5 °C resulting in an energy change of 1.05 kJ.

(i) Calculate the mass of one mole of magnesium carbonate, $MgCO_3$ and hence the number of moles of magnesium carbonate used in this experiment.
Use the Periodic Table as a source of data.

(2)

(ii) Using the method in part (b)(v), calculate ΔH_3.

(1)

(d) Using your answers to (b)(v) and (c)(ii), calculate the enthalpy change of formation, ΔH_f, of magnesium carbonate, $MgCO_3$.
Include a sign and units in your answer.

(2)

(e) Why is it impossible to measure ΔH_f of $MgCO_3(s)$ directly?

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(1)
(Total 18 marks)

68. Phosphorus reacts with excess chlorine to produce phosphorus pentachloride, PCl_5.

$$2P(s) + 5Cl_2(g) \rightarrow 2PCl_5(s)$$

(i) Calculate the mass of phosphorus needed to produce 7.19 g of phosphorus pentachloride.

(2)

(ii) Calculate the volume of chlorine needed to produce 7.19 g of phosphorus pentachloride.

[molar volume of chlorine under the conditions of this experiment = 24.0 dm^3 mol^{-1}]

(2)
(Total 4 marks)

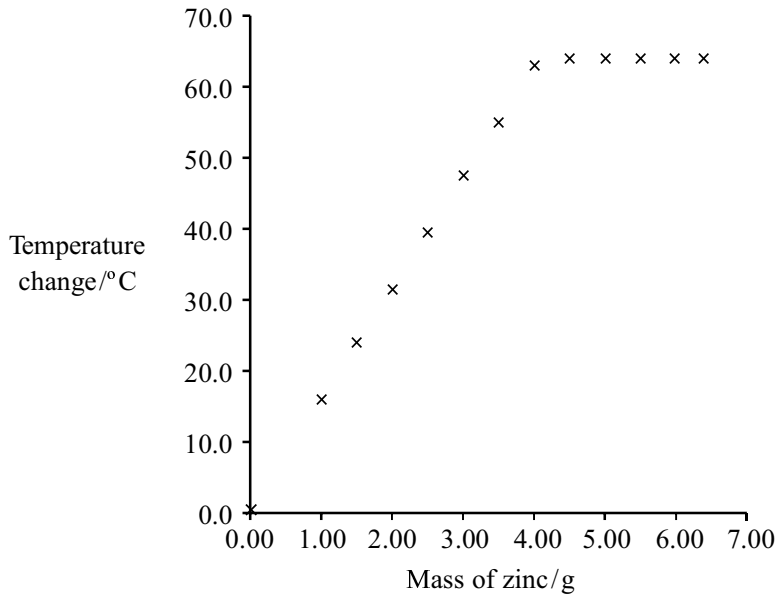
69. An experiment was carried out to measure the enthalpy change for the reaction of zinc with aqueous copper(II) sulphate.

The equation for the reaction is

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- A measuring cylinder was used to transfer separate 50 cm3 samples of 1.25 mol dm–3 copper(II) sulphate solution into polystyrene cups.
- Weighed amounts of zinc powder were added to each sample in turn.
- Each mixture was stirred thoroughly and the temperature rise noted with a thermometer accurate to 0.5 °C.

The results of this experiment are summarised on the graph below.



(a) Explain why the graph initially shows a rise in temperature and then levels off.

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(2)

(b) (i) Suggest why the mass of metal is **not** used in the calculation of the heat change.

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(1)

(ii) The graph shows that the maximum temperature change is 63.5 °C. Use this value to calculate the maximum heat change, in joules, in this reaction.

You should assume that the density of the solution is 1.00 g cm^{-3} and its heat capacity is the same as water, 4.18 J g^{-1} $°C^{-1}$.

(1)

(iii) From the heat change calculated in (b)(ii) calculate the enthalpy change, in kJ mol^{-1}, for the reaction. Include the appropriate sign and give your answer to **three** significant figures.

(4)

(c) (i) It is suggested that the precision of the experiment would be improved by using a thermometer accurate to 0.1 °C.

Explain why this suggestion is incorrect.

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(1)

(ii) Suggest a simple practical change to the **method** that would make the experiment more accurate.

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(1)
(Total 10 marks)

70. A sample of the element osmium, atomic number 76, is made up of four isotopes. The sample has the following percentage composition.

Relative Atomic Mass of Isotope	% Composition
188	15.20
189	17.40
190	26.40
192	41.00

(i) What is the minimum number of neutrons present in any single atom of osmium in the sample?

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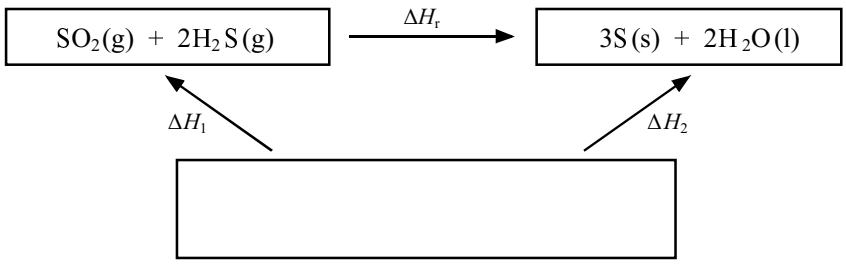
(1)

(ii) Calculate the average relative atomic mass of osmium in the sample. Give your answer to **four** significant figures.

(2)
(Total 3 marks)

71. The Hess cycle below can be used to find the enthalpy change, ΔH_r, for the reaction between hydrogen sulphide and sulphur dioxide, using standard enthalpy changes offormation.



(i) Complete the cycle by filling in the empty box.

(2)

(ii) What is meant by the **standard enthalpy change of formation**, $\Delta H_f^{\ominus}$, of a compound?

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(3)

(iii) Use the cycle and the data below to calculate the enthalpy change of the reaction, ΔH_r.

	$\Delta H_f^\ominus$ / kJ mol^{-1}
SO_2 (g)	–296.8
H_2S (g)	–20.6
H_2O (l)	–285.8

(2)
(Total 7 marks)

72. Use the data about four fuels given below to answer this question.

Fuel	**Formula**	**Name**	**Enthalpy change of combustion / kJ mol^{-1}**	**Molar mass / g mol^{-1}**
A	CH_4	methane	–890	16
B	CH_3OH	methanol	–726	32
C	C_3H_8	propane	–2219	44
D	C_4H_{10}	butane	–2877	58

(a) Which fuel, A, B, C or D, produces most energy per gram on complete combustion?

A

B

C

D

(1)

(b) Scientists give governments advice on technical issues. What information would scientists use when advising governments on the choice of one of these fuels, if the aim was to minimise carbon dioxide production?

A mass of carbon per gram of fuel

B mass of carbon per kilojoules produced

C number of kilojoules produced per gram

D number of kilojoules produced per mole

(1)
(Total 2 marks)

73. This question is about four hydrocarbons with molecular formulae as shown.

A C_2H_2

B C_3H_6

C C_3H_8

D C_4H_{10}

(a) Which hydrocarbon has the same empirical formula as its molecular formula?

A

B

C

D

(1)

(b) Which has a molecular ion in the mass spectrum at mass/charge ratio = 58?

A

B

C

D

(1)

(c) Which is neither an alkane nor an alkene?

A

B

C

D

(1)

(d) Which could be 2-methylpropane?

A

B

C

D

(1)
(Total 4 marks)

74. Copper(II) sulfate solution can be prepared from solid copper(II) carbonate by reaction with hot dilute sulfuric acid.

(a) Write the equation for the reaction, including state symbols.

(1)

(b) The experiment was carried out using 0.025 moles of sulfuric acid of concentration 2.0 mol dm^{-3}. What volume of this sulfuric acid was used?

(1)

(c) (i) It is usual to react the sulfuric acid with a slight excess of copper(II) carbonate.

Calculate the mass of copper(II) carbonate needed if a 10% excess is required.

[Molar mass of copper(II) carbonate = 123.5 g mol^{-1}]

(2)

(ii) A student doing this experiment chose to use a balance reading to 0.01g in an attempt to work accurately.

Was this choice of balance necessary from the point of view of accuracy? Explain your answer.

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(1)

(d) The sulfuric acid is heated to boiling and the copper(II) carbonate is added in small portions.

State the next step needed to prepare pure copper(II) sulfate solution. Justify your answer.

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(1)

(e) When the solution of copper(II) sulfate is allowed to crystallise, the crystals which are produced have the formula $CuSO_4.5H_2O$.

(i) What is the molar mass of $CuSO_4.5H_2O$?

(1)

(ii) 3.98 g of $CuSO_4.5H_2O$ crystals were obtained. Calculate the percentage yield in this experiment.

(2)
(Total 9 marks)

75. This question is about magnesium and magnesium oxide.

(a) Describe the bonding in magnesium and explain why it is a good conductor of electricity.

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(3)

(b) Draw a diagram (using dots or crosses) for the ions in magnesium fluoride showing all the electrons and the ionic charges on:

(i) the magnesium ion

(1)

(ii) the fluoride ion.

(1)

(c) Under what conditions does magnesium fluoride conduct electricity?

Explain your answer.

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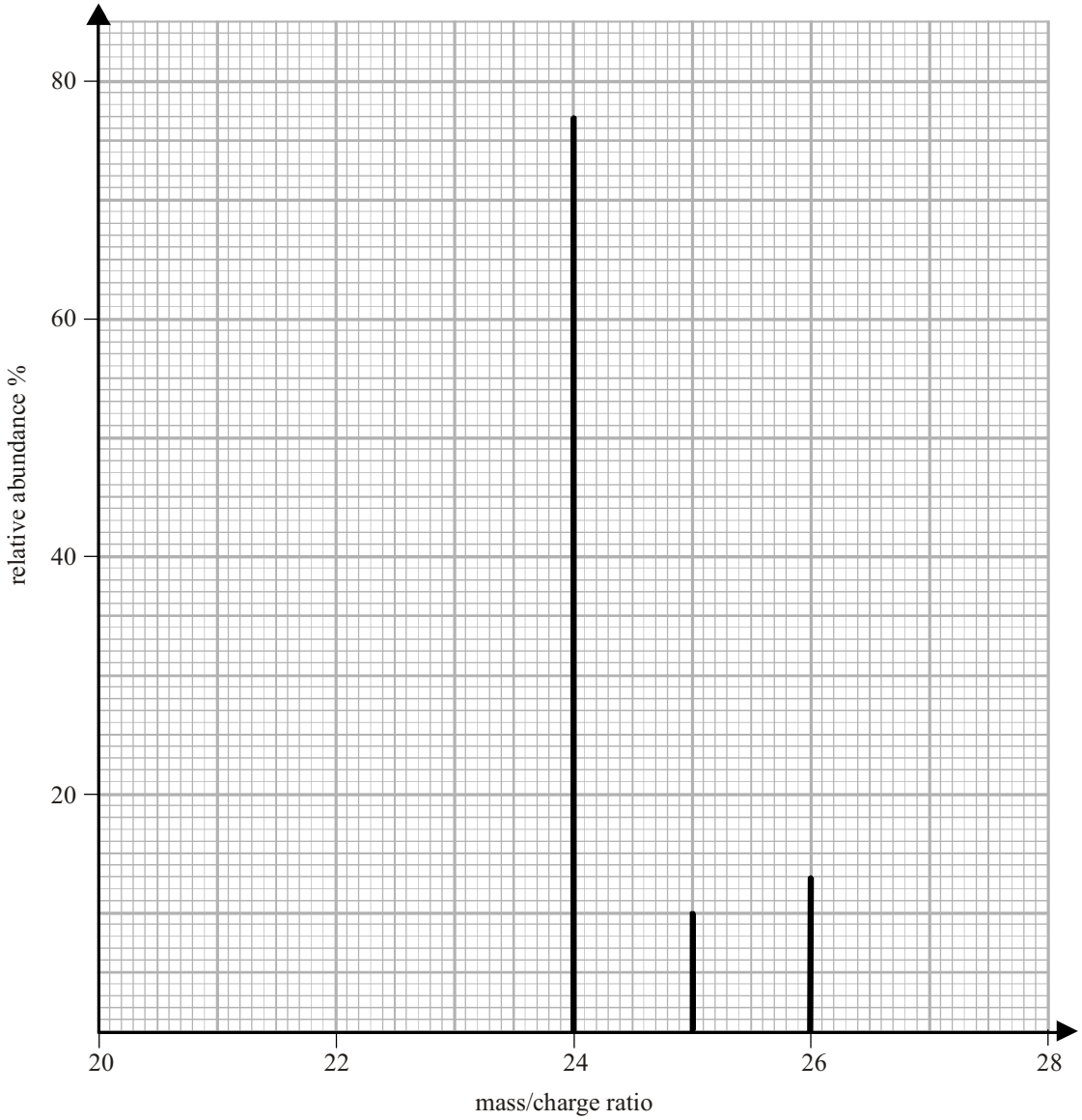
(1)

(d) The mass spectrum of a sample of magnesium is shown below.



(i) Use the data above to estimate the percentage isotopic composition of the sample of magnesium. Hence calculate the average atomic mass of the sample of magnesium.

(2)

(ii) Why do the three isotopes have the same chemical properties?

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(1)

(e) (i) Oceanographers studying plankton found that a sample of seawater contained 1.20 nanomol dm^{-3} of chlorophyll, $C_{55}H_{77}MgN_4O_5$. (1 nanomol = 1×10^{-9} mol)

What mass of magnesium would be present in 1.00 cm^3 of this sample of seawater? Give your answer to three significant figures.

(2)

(ii) X-ray diffraction can be used to locate atoms or ions in molecules like chlorophyll. X-rays are scattered by the electrons in atoms and ions. In chlorophyll the atoms of one of the elements still cannot be located with certainty by this technique.

Suggest which element is most difficult to locate.

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(1)
(Total 12 marks)

76. Airbags, used as safety features in cars, contain sodium azide, NaN3. An airbag requires a large volume of gas to be produced in a few milliseconds. The gas is produced in this reaction:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g) \qquad \Delta H \text{ is positive}$$

When the airbag is fully inflated, 50 dm3 of nitrogen gas is produced.

(a) Calculate the number of molecules in 50 dm^3 of nitrogen gas under these conditions.

[The Avogadro constant = 6.02×10^{23} mol^{-1}. The molar volume of nitrogen gas under the conditions in the airbag is 24 dm^3 mol^{-1}].

(2)

(b) Calculate the mass of sodium azide, NaN_3, that would produce 50 dm^3 of nitrogen gas.

(3)

(c) What will happen to the temperature in the airbag when the reaction occurs?

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(1)

(d) The airbag must be strong enough not to burst in an accident. An airbag which has burst in an accident is hazardous if the sodium azide in it has decomposed.

Explain why this is so.

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(2)
(Total 8 marks)

77. A student investigated a reaction which could be used to warm up coffee in self-heating cans.

$$Mg(s) + Cu(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + Cu(s)$$

In the self-heating cans, the bottom has a compartment containing copper(II) nitrate solution. When a button on the bottom of the can is pressed, the magnesium powder is released into the compartment where it reacts with the copper(II) nitrate solution.

(a) A student investigated the enthalpy change for this reaction by measuring

50.0 cm^3 of 0.300 mol dm^{-3} copper(II) nitrate solution into a 100 cm3 beaker and adding 1g (an excess) of magnesium powder.

The results are shown below.

Temperature of copper(II) nitrate solution at start	= 22 °C
Temperature of mixture after reaction	= 43 °C

(i) Calculate the energy change which took place. The specific heat capacity of the solution is 4.20 J g^{-1} K^{-1}.

Which is the correct value for the energy change in joules?

(1)

(ii) How many moles of copper(II) nitrate were used in the experiment?

(1)

(iii) Calculate the enthalpy change for the reaction. You should include a sign and units in your answer.

(2)

(iv) Suggest two changes you would make to the equipment used in order to improve the accuracy of the result.

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(2)

(b) The ionic equation for the reaction is shown below:

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$ $\Delta H = -532$ kJ mol^{-1}

Would the following affect the value of the experimental result?

Explain your answer, stating the effect, if any, on the value of the enthalpy change obtained.

(i) The student used 2 g rather than 1g of magnesium.

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(2)

(ii) The heat losses that occurred from the student's beaker.

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(2)

(c) The temperature in the self-heating can needs to increase by 60 °C to produce a hot drink.

Suggest a change you could make to the mixture in the experiment in (a) to produce a greater temperature rise. You are **not** expected to do a calculation.

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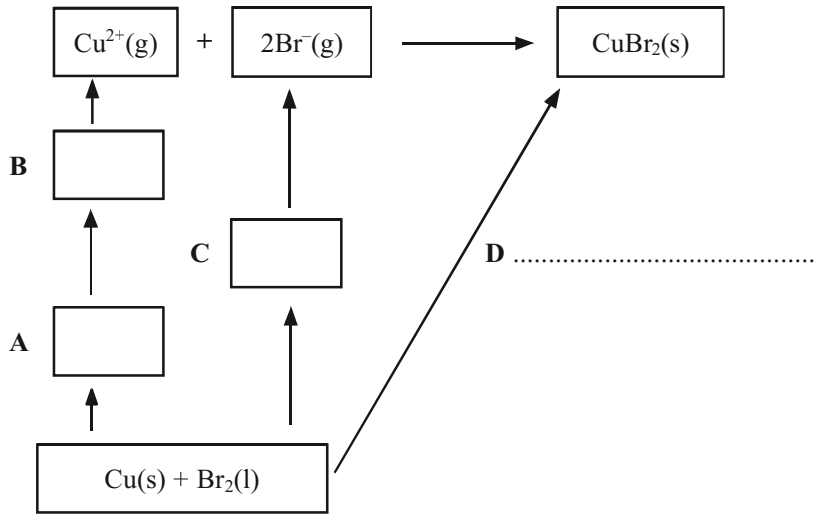
(1)
(Total 11 marks)

78. The following data can be used in a Born-Haber cycle for copper(II) bromide, $CuBr_2$.

Enthalpy change of atomisation of bromine $\Delta H^{\ominus}_{at}[\frac{1}{2}Br_2(l)]$	+111.9 kJ mol^{-1}
Enthalpy change of atomisation of copper, $\Delta H^{\ominus}_{at}[Cu(s)]$	+338.3 kJ mol^{-1}
First ionisation energy of copper, $E_{m1}[Cu(g)]$	+746.0 kJ mol^{-1}
Second ionisation energy of copper, E_{m2} [Cu(g)]	+1958.0 kJ mol^{-1}
Electron affinity of bromine, $E_{aff}[Br(g)]$	–342.6 kJ mol^{-1}
Enthalpy change of formation of $CuBr_2(s)$, $\Delta H^{\ominus}_{f}$ [$CuBr_2(s)$]	–141.8 kJ mol^{-1}

(a) On the following outline of a Born-Haber cycle complete the boxes **A**, **B**, and **C** by putting in the formula and state symbol for the appropriate species and writing the name of the enthalpy change **D**.



(3)

(b) Use the data to calculate a value for the lattice energy of copper(II) bromide.

Give a sign and units in your answer.

(3)

(c) When the lattice energy of copper(II) bromide is calculated from ionic radii and charges, the result is a value numerically about 10% less than the one obtained from the Born-Haber cycle.

(i) What does this suggest about the nature of the bonding in copper(II) bromide?

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(1)

(ii) Draw a diagram to show how the smaller copper ion alters the shape of the larger bromide ion.

(1)
(Total 8 marks)