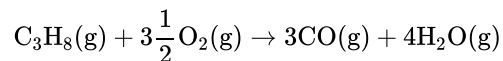


SL Paper 2

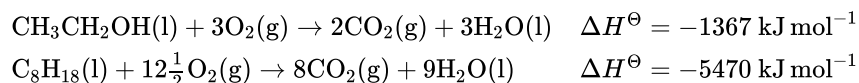
The following equation represents a combustion reaction of propane, $C_3H_8(g)$ when the oxygen supply is limited.



- a. Define the term *average bond enthalpy*. [2]
- b. (i) Determine ΔH , the enthalpy change of the reaction, in kJ mol^{-1} , using average bond enthalpy data from Table 10 of the Data Booklet. [4]
The bond enthalpy for the carbon-oxygen bond in carbon monoxide, CO, is 1072 kJ mol^{-1} .
- (ii) The CO molecule has dative covalent bonding. Identify a nitrogen-containing positive ion which also has this type of bonding.

Ethanol is used as a component in fuel for some vehicles. One fuel mixture contains 10% by mass of ethanol in unleaded petrol (gasoline). This mixture is often referred to as Gasohol E10.

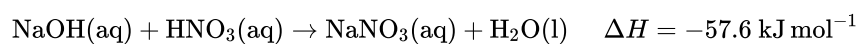
Assume that the other 90% by mass of Gasohol E10 is octane. 1.00 kg of this fuel mixture was burned.



- a.i. Calculate the mass, in g, of ethanol and octane in 1.00 kg of the fuel mixture. [1]
- a.ii. Calculate the amount, in mol, of ethanol and octane in 1.00 kg of the fuel mixture. [1]
- a.iii. Calculate the total amount of energy, in kJ, released when 1.00 kg of the fuel mixture is completely burned. [3]
- b. If the fuel blend was vaporized before combustion, predict whether the amount of energy released would be greater, less or the same. Explain your answer. [2]

Ammonia, NH_3 , is a base according to both the Brønsted–Lowry and the Lewis theories of acids and bases.

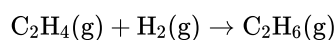
The equation for the reaction between sodium hydroxide, NaOH, and nitric acid, HNO_3 , is shown below.



- a. Distinguish between the terms *strong base* and *weak base*, and state one example of each. [3]

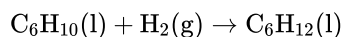
- b.i.State the equation for the reaction of ammonia with water. [1]
- b.ii.Explain why ammonia can act as a Brønsted–Lowry base. [1]
- b.iii.Explain why ammonia can also act as a Lewis base. [1]
- c. (i) When ammonium chloride, $\text{NH}_4\text{Cl}(\text{aq})$, is added to excess solid sodium carbonate, $\text{Na}_2\text{CO}_3(\text{s})$, an acid–base reaction occurs. Bubbles of gas are produced and the solid sodium carbonate decreases in mass. State **one** difference which would be observed if nitric acid, $\text{HNO}_3(\text{aq})$, was used instead of ammonium chloride. [5]
- (ii) Deduce the Lewis structures of the ammonium ion, NH_4^+ , and the carbonate ion, CO_3^{2-} .
- Ammonium ion Carbonate ion
- (iii) Predict the shapes of NH_4^+ and CO_3^{2-} .
- NH_4^+ :
- CO_3^{2-} :
- d. (i) Sketch and label an enthalpy level diagram for this reaction. [6]
- (ii) Deduce whether the reactants or the products are more energetically stable, stating your reasoning.
- (iii) Calculate the change in heat energy, in kJ, when 50.0 cm^3 of 2.50 mol dm^{-3} sodium hydroxide solution is added to excess nitric acid.
- e. When 5.35 g ammonium chloride, $\text{NH}_4\text{Cl}(\text{s})$, is added to 100.0 cm^3 of water, the temperature of the water decreases from $19.30 \text{ }^\circ\text{C}$ to $15.80 \text{ }^\circ\text{C}$. Determine the enthalpy change, in kJ mol^{-1} , for the dissolving of ammonium chloride in water. [3]

Two students were asked to use information from the Data Booklet to calculate a value for the enthalpy of hydrogenation of ethene to form ethane.

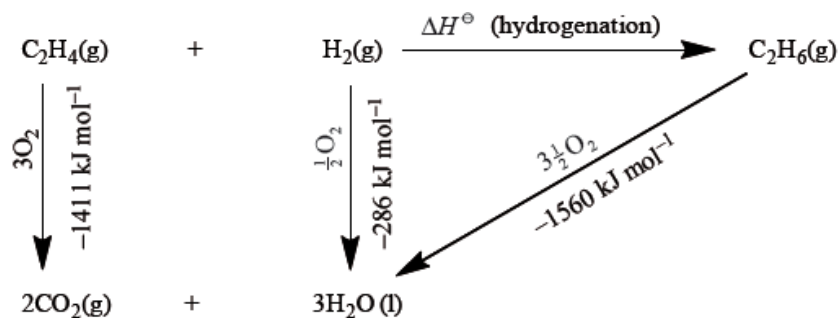


John used the average bond enthalpies from Table 10. Marit used the values of enthalpies of combustion from Table 12.

John then decided to determine the enthalpy of hydrogenation of cyclohexene to produce cyclohexane.



- a. Calculate the value for the enthalpy of hydrogenation of ethene obtained using the average bond enthalpies given in Table 10. [2]
- b. Marit arranged the values she found in Table 12 into an energy cycle. [1]



Calculate the value for the enthalpy of hydrogenation of ethene from the energy cycle.

c. Suggest **one** reason why John's answer is slightly less accurate than Marit's answer. [1]

d.i. Use the average bond enthalpies to deduce a value for the enthalpy of hydrogenation of cyclohexene. [1]

d.ii. The percentage difference between these two methods (average bond enthalpies and enthalpies of combustion) is greater for cyclohexene than it was for ethene. John's hypothesis was that it would be the same. Determine why the use of average bond enthalpies is less accurate for the cyclohexene equation shown above, than it was for ethene. Deduce what extra information is needed to provide a more accurate answer. [2]

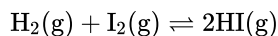
a. Define the term *average bond enthalpy*. [2]

b. Deduce the balanced chemical equation for the complete combustion of butan-1-ol. [1]

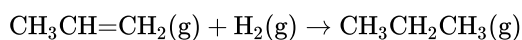
c. Determine the standard enthalpy change, in kJ mol^{-1} , for the complete combustion of butan-1-ol, using the information from Table 10 of the Data Booklet. [3]

d. Based on the types of intermolecular force present, explain why butan-1-ol has a higher boiling point than butanal. [2]

An example of a homogeneous reversible reaction is the reaction between hydrogen and iodine.



Propane can be formed by the hydrogenation of propene.



a.i. Outline the characteristics of a homogeneous chemical system that is in a state of equilibrium. [2]

a.ii. Deduce the expression for the equilibrium constant, K_c . [1]

a.iii. Predict what would happen to the position of equilibrium and the value of K_c if the pressure is increased from 1 atm to 2 atm. [2]

a.iv. The value of K_c at 500 K is 160 and the value of K_c at 700 K is 54. Deduce what this information tells us about the enthalpy change of the forward reaction. [1]

a.v. The reaction can be catalysed by adding platinum metal. State and explain what effect the addition of platinum would have on the value of the equilibrium constant. [2]

b.i. State the conditions necessary for the hydrogenation reaction to occur. [2]

b.ii. Enthalpy changes can be determined using average bond enthalpies. Define the term *average bond enthalpy*. [2]

b.iii. Determine a value for the hydrogenation of propene using information from Table 10 of the Data Booklet. [2]

b.iv. Explain why the enthalpy of hydrogenation of propene is an exothermic process. [1]

c.i. Describe a chemical test that could be used to distinguish between propane and propene. In **each** case state the result of the test. [2]

c.ii. Under certain conditions propene can polymerize to form poly(propene). State the type of polymerization taking place and draw a section of the polymer to represent the repeating unit. [2]

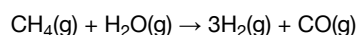
c.iii. Other than polymerization, state **one** reaction of alkenes which is of economic importance. [1]

Enthalpy changes depend on the number and type of bonds broken and formed.

The table lists the standard enthalpies of formation, ΔH_f^\ominus , for some of the species in the reaction above.

	CH₄(g)	H₂O(g)	CO(g)	H₂(g)
$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	-74.0	-242	-111	

a. Hydrogen gas can be formed industrially by the reaction of natural gas with steam. [3]



Determine the enthalpy change, ΔH , for the reaction, in kJ, using section 11 of the data booklet.

Bond enthalpy for C≡O: 1077 kJ mol⁻¹

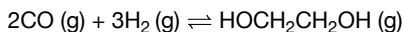
b.i. Outline why no value is listed for H₂(g). [1]

b.ii. Determine the value of ΔH^\ominus , in kJ, for the reaction using the values in the table. [1]

b.iii. Outline why the value of enthalpy of reaction calculated from bond enthalpies is less accurate. [1]

Ethane-1,2-diol, HOCH₂CH₂OH, has a wide variety of uses including the removal of ice from aircraft and heat transfer in a solar cell.

a. Ethane-1,2-diol can be formed according to the following reaction. [7]



(i) Deduce the equilibrium constant expression, K_c , for this reaction.

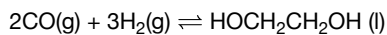
(ii) State how increasing the pressure of the reaction mixture at constant temperature will affect the position of equilibrium and the value of K_c .

Position of equilibrium:

K_c :

(iii) Calculate the enthalpy change, ΔH^\ominus , in kJ, for this reaction using section 11 of the data booklet. The bond enthalpy of the carbon–oxygen bond in CO(g) is 1077 kJ mol⁻¹.

(iv) The enthalpy change, ΔH^\ominus , for the following similar reaction is -233.8 kJ.



Deduce why this value differs from your answer to (a)(iii).

- b. Determine the average oxidation state of carbon in ethene and in ethane-1,2-diol. [2]

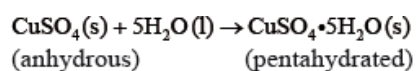
Ethene:

Ethane-1,2-diol:

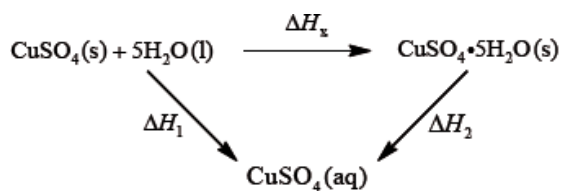
- c. Explain why the boiling point of ethane-1,2-diol is significantly greater than that of ethene. [2]

- d. Ethane-1,2-diol can be oxidized first to ethanedioic acid, $(\text{COOH})_2$, and then to carbon dioxide and water. Suggest the reagents to oxidize ethane-1,2-diol. [1]

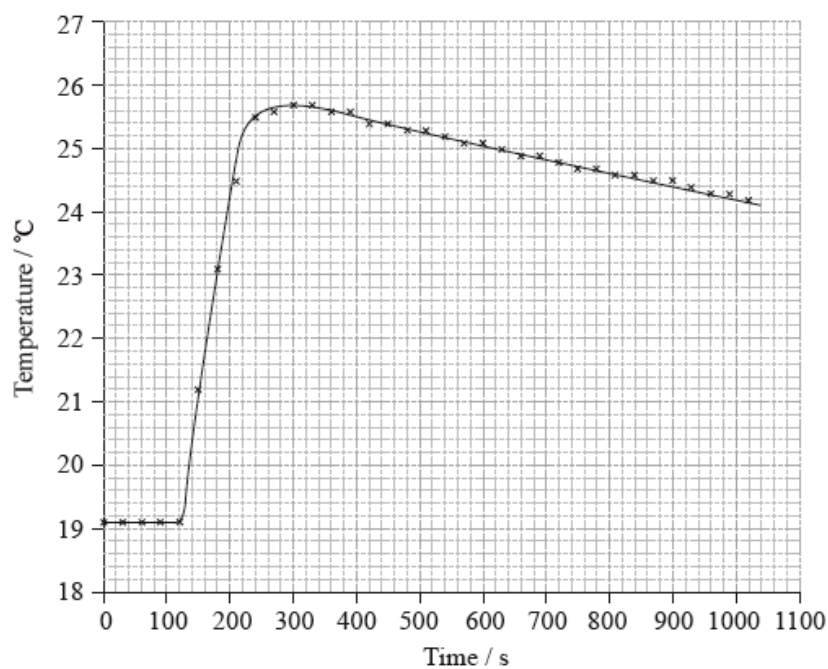
If white anhydrous copper(II) sulfate powder is left in the atmosphere it slowly absorbs water vapour giving the blue pentahydrated solid.



It is difficult to measure the enthalpy change for this reaction directly. However, it is possible to measure the heat changes directly when both anhydrous and pentahydrated copper(II) sulfate are separately dissolved in water, and then use an energy cycle to determine the required enthalpy change value, ΔH_x , indirectly.



To determine ΔH_1 a student placed 50.0 g of water in a cup made of expanded polystyrene and used a data logger to measure the temperature. After two minutes she dissolved 3.99 g of anhydrous copper(II) sulfate in the water and continued to record the temperature while continuously stirring. She obtained the following results.



To determine ΔH_2 , 6.24 g of pentahydrated copper(II) sulfate was dissolved in 47.75 g of water. It was observed that the temperature of the solution decreased by 1.10 °C.

The magnitude (the value without the + or – sign) found in a data book for ΔH_x is 78.0 kJ mol⁻¹.

a.i. Calculate the amount, in mol, of anhydrous copper(II) sulfate dissolved in the 50.0 g of water. [1]

a.ii. Determine what the temperature rise would have been, in °C, if no heat had been lost to the surroundings. [2]

a.iii. Calculate the heat change, in kJ, when 3.99 g of anhydrous copper(II) sulfate is dissolved in the water. [2]

a.iv. Determine the value of ΔH_1 in kJ mol⁻¹. [1]

b.i. Calculate the amount, in mol, of water in 6.24 g of pentahydrated copper(II) sulfate. [2]

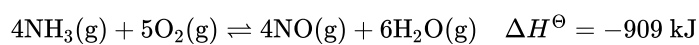
b.ii. Determine the value of ΔH_2 in kJ mol⁻¹. [2]

b.iii. Using the values obtained for ΔH_1 in (a) (iv) and ΔH_2 in (b) (ii), determine the value for ΔH_x in kJ mol⁻¹. [1]

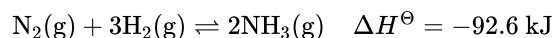
c.i. Calculate the percentage error obtained in this experiment. (If you did not obtain an answer for the experimental value of ΔH_x then use the value 70.0 kJ mol⁻¹, but this is **not** the true value.) [1]

c.ii. The student recorded in her qualitative data that the anhydrous copper(II) sulfate she used was pale blue rather than completely white. Suggest a reason why it might have had this pale blue colour and deduce how this would have affected the value she obtained for ΔH_x . [2]

Consider the following equilibrium:



Nitrogen reacts with hydrogen to form ammonia in the Haber process, according to the following equilibrium.



a.i. Deduce the equilibrium constant expression, K_c , for the reaction. [1]

a.ii. Predict the direction in which the equilibrium will shift when the following changes occur. [4]

The volume increases.

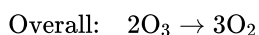
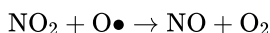
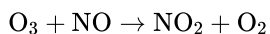
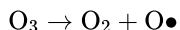
The temperature decreases.

H₂O(g) is removed from the system.

A catalyst is added to the reaction mixture.

b. Define the term *activation energy*, E_a . [1]

c. Nitrogen monoxide, NO, is involved in the decomposition of ozone according to the following mechanism. [2]



State and explain whether or not NO is acting as a catalyst.

d.i. Define the term *endothermic reaction*. [1]

d.ii. Sketch the Maxwell-Boltzmann energy distribution curve for a reaction with and without a catalyst, and label both axes. [3]

e.i. Define the term *rate of reaction*. [1]

f. Iron, used as the catalyst in the Haber process, has a specific heat capacity of $0.4490 \text{ J g}^{-1}\text{K}^{-1}$. If 245.0 kJ of heat is supplied to 8.500 kg of iron, initially at a temperature of 15.25°C , determine its final temperature in K. [3]

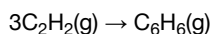
This question is about ethene, C_2H_4 , and ethyne, C_2H_2 .

a.i. Ethyne, like ethene, undergoes hydrogenation to form ethane. State the conditions required. [2]

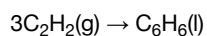
a.ii. Outline the formation of polyethene from ethene by drawing three repeating units of the polymer. [1]

b.i. Under certain conditions, ethyne can be converted to benzene. [2]

Determine the standard enthalpy change, ΔH^\ominus , for the reaction stated, using section 11 of the data booklet.

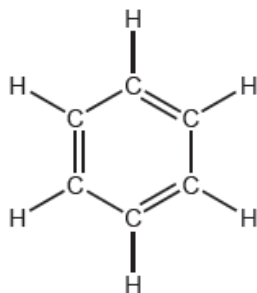


b.ii. Determine the standard enthalpy change, ΔH^\ominus , for the following similar reaction, using ΔH_f^\ominus values in section 12 of the data booklet. [2]



b.iii. Explain, giving two reasons, the difference in the values for (b)(i) and (ii). If you did not obtain answers, use -475 kJ for (i) and -600 kJ for (ii). [2]

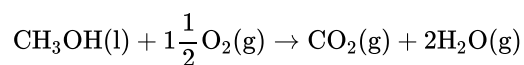
c. One possible Lewis structure for benzene is shown. [1]



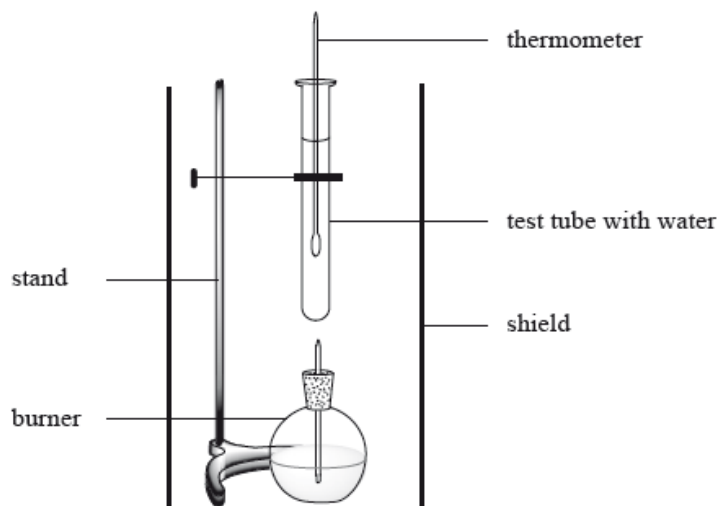
State one piece of physical evidence that this structure is **incorrect**.

d. State the characteristic reaction mechanism of benzene. [1]

Methanol is made in large quantities as it is used in the production of polymers and in fuels. The enthalpy of combustion of methanol can be determined theoretically or experimentally.



The enthalpy of combustion of methanol can also be determined experimentally in a school laboratory. A burner containing methanol was weighed and used to heat water in a test tube as illustrated below.



The following data were collected.

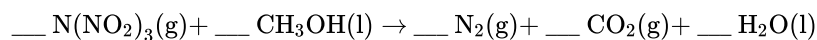
Initial mass of burner and methanol / g	80.557
Final mass of burner and methanol / g	80.034
Mass of water in test tube / g	20.000
Initial temperature of water / °C	21.5
Final temperature of water / °C	26.4

The Data Booklet value for the enthalpy of combustion of methanol is -726 kJ mol^{-1} . Suggest why this value differs from the values calculated in parts (a) and (b).

- a. Using the information from Table 10 of the Data Booklet, determine the theoretical enthalpy of combustion of methanol. [3]
- b.i. Calculate the amount, in mol, of methanol burned. [2]
- b.ii. Calculate the heat absorbed, in kJ, by the water. [3]
- b.iii. Determine the enthalpy change, in kJ mol^{-1} , for the combustion of 1 mole of methanol. [2]
- c.i. Part (a) [1]
- c.ii. Part (b) [1]

In December 2010, researchers in Sweden announced the synthesis of N,N-dinitronitramide, $\text{N}(\text{NO}_2)_3$. They speculated that this compound, more commonly called trinitramide, may have significant potential as an environmentally friendly rocket fuel oxidant.

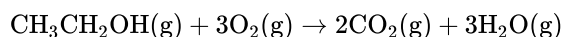
- a. Methanol reacts with trinitramide to form nitrogen, carbon dioxide and water. Deduce the coefficients required to balance the equation for this reaction. [1]



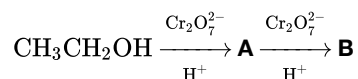
- c. Calculate the enthalpy change, in kJ mol^{-1} , when one mole of trinitramide decomposes to its elements, using bond enthalpy data from Table 10 of the Data Booklet. Assume that all the N–O bonds in this molecule have a bond enthalpy of 305 kJ mol^{-1} . [3]
- d. Outline how the length of the N–N bond in trinitramide compares with the N–N bond in nitrogen gas, N_2 . [2]
- e. Deduce the N–N–N bond angle in trinitramide and explain your reasoning. [3]
- f. Predict, with an explanation, the polarity of the trinitramide molecule. [2]
- g.i. Methanol can also be burnt as a fuel. Describe an experiment that would allow the molar enthalpy change of combustion to be calculated from the results. [3]
- g.ii. Explain how the results of this experiment could be used to calculate the molar enthalpy change of combustion of methanol. [3]
- g.iii. Predict, with an explanation, how the result obtained would compare with the value in Table 12 of the Data Booklet. [2]

In some countries, ethanol is mixed with gasoline (petrol) to produce a fuel for cars called gasohol.

- a.i. Define the term *average bond enthalpy*. [2]
- a.ii. Use the information from Table 10 of the Data Booklet to determine the standard enthalpy change for the complete combustion of ethanol. [3]

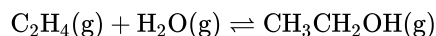


- a.iii. The standard enthalpy change for the complete combustion of octane, C_8H_{18} , is $-5471 \text{ kJ mol}^{-1}$. Calculate the amount of energy produced in kJ when 1 g of ethanol and 1 g of octane is burned completely in air. [2]
- a.iv. Ethanol can be oxidized using acidified potassium dichromate, $\text{K}_2\text{Cr}_2\text{O}_7$, to form two different organic products. [4]



State the structural formulas of the organic products **A** and **B** and describe the conditions required to obtain a high yield of each of them.

- a.v. Deduce and explain whether ethanol or **A** has the higher boiling point. [2]
- a.vi. Ethene can be converted into ethanol by direct hydration in the presence of a catalyst according to the following equation. [2]



For this reaction identify the catalyst used and state **one** use of the ethanol formed other than as a fuel.

- b.i. State the name of **one** structural isomer of pentane. [1]
-

To determine the enthalpy change of combustion of methanol, CH_3OH , 0.230 g of methanol was combusted in a spirit burner. The heat released increased the temperature of 50.0 cm^3 of water from $24.5 \text{ }^\circ\text{C}$ to $45.8 \text{ }^\circ\text{C}$.

The manufacture of gaseous methanol from CO and H_2 involves an equilibrium reaction.

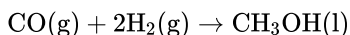


State and explain the effect of the following changes on the equilibrium position of the reaction in part (c).

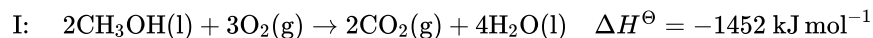
a.i. Calculate the enthalpy change of combustion of methanol. [4]

a.ii. Using the theoretical value in Table 12 of the Data Booklet, discuss the experimental results, including **one** improvement that could be made. [3]

b. Methanol can be produced according to the following equation. [3]



Calculate the standard enthalpy change of this reaction using the following data:



c.i. Outline the characteristics of a chemical equilibrium. [2]

c.ii. Deduce the equilibrium constant expression, K_c , for this reaction. [1]

d.i. Increase in temperature. [2]

d.ii. Increase in pressure. [2]

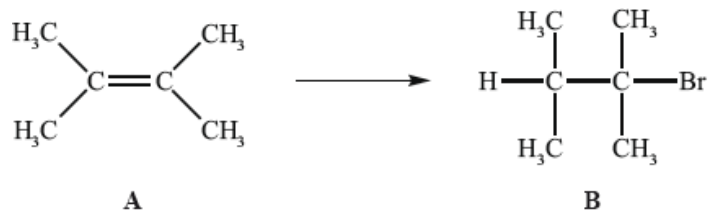
d.iii. Addition of a catalyst. [2]

a. Define the term *activation energy*, E_a . [1]

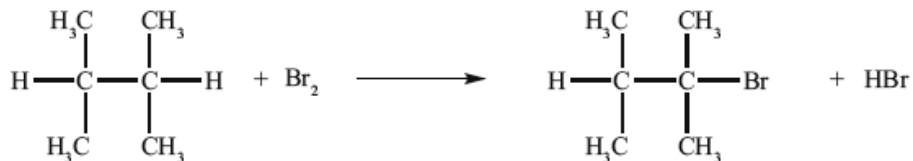
b. State **two** conditions necessary for a reaction to take place between two reactant particles. [2]

c. Sketch an enthalpy level diagram to describe the effect of a catalyst on an *exothermic* reaction. [3]

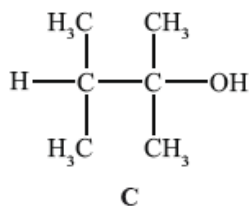
Alkenes, such as **A** (shown below), are important intermediates in the petrochemical industry because they undergo addition reactions to produce a wide variety of products, such as the conversion shown below.



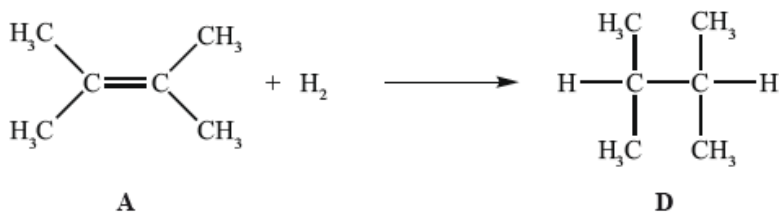
Another way to make **B** is the reaction shown below.



B can be converted into **C**.



In the gas phase, **A** reacts with hydrogen to form **D**.



- a. Applying IUPAC rules, state the name of **A**. [1]
- b. State the reagent required to convert **A** into **B**. [1]
- c. (i) State the conditions required for this reaction to occur. [2]
 - (ii) Outline why it would give a poor yield of the desired product.
- d. (i) State the reagent required. [4]
 - (ii) Explain the mechanism of this reaction, using curly arrows to represent the movement of electron pairs.
- e. **A** can also be converted into **C** without going via **B**. State the reagent and conditions required. [2]
- f. (i) State why **C** is **not** readily oxidized by acidified potassium dichromate(VI). [2]
 - (ii) Deduce the structural formula of an isomer of **C** that could be oxidized to a carboxylic acid by this reagent.
- g.i. State the conditions required for this reaction to occur. [1]

g.ii State the homologous series to which **D** belongs. [1]

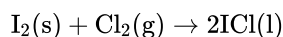
g.iii Determine the enthalpy change, in kJ mol^{-1} , for the reaction of **A** with hydrogen, using Table 10 of the Data Booklet, and state whether the reaction is exothermic or endothermic. [4]

g.iv The standard enthalpy change of combustion of **A** is $-4000 \text{ kJ mol}^{-1}$. Calculate the amount of **A**, in mol, that would have to be burned to raise the temperature of 1 dm^3 of water from $20 \text{ }^\circ\text{C}$ to $100 \text{ }^\circ\text{C}$. [2]

Two groups of students (Group A and Group B) carried out a project* on the chemistry of some group 7 elements (the halogens) and their compounds.

* Adapted from J Derek Woollins, (2009), *Inorganic Experiments* and Open University, (2008), *Exploring the Molecular World*.

In the first part of the project, the two groups had a sample of iodine monochloride (a corrosive brown liquid) prepared for them by their teacher using the following reaction.



The following data were recorded.

Mass of $\text{I}_2(\text{s})$	10.00 g
Mass of $\text{Cl}_2(\text{g})$	2.24 g
Mass of $\text{ICl}(\text{l})$ obtained	8.60 g

The students reacted $\text{ICl}(\text{l})$ with $\text{CsBr}(\text{s})$ to form a yellow solid, $\text{CsICl}_2(\text{s})$, as one of the products. $\text{CsICl}_2(\text{s})$ has been found to produce very pure $\text{CsCl}(\text{s})$ which is used in cancer treatment.

To confirm the composition of the yellow solid, Group A determined the amount of iodine in 0.2015 g of $\text{CsICl}_2(\text{s})$ by titrating it with $0.0500 \text{ mol dm}^{-3} \text{ Na}_2\text{S}_2\text{O}_3(\text{aq})$. The following data were recorded for the titration.

Mass of $\text{CsICl}_2(\text{s})$ taken (in $\text{g} \pm 0.0001$)	0.2015
Initial burette reading of $0.0500 \text{ mol dm}^{-3} \text{ Na}_2\text{S}_2\text{O}_3(\text{aq})$ (in $\text{cm}^3 \pm 0.05$)	1.05
Final burette reading of $0.0500 \text{ mol dm}^{-3} \text{ Na}_2\text{S}_2\text{O}_3(\text{aq})$ (in $\text{cm}^3 \pm 0.05$)	25.25

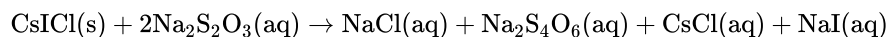
a. (i) State the number of significant figures for the masses of $\text{I}_2(\text{s})$ and $\text{ICl}(\text{l})$. [6]

$\text{I}_2(\text{s})$:

$\text{ICl}(\text{l})$:

- (ii) The iodine used in the reaction was in excess. Determine the theoretical yield, in g, of ICl(l).
- (iii) Calculate the percentage yield of ICl(l).
- (iv) Using a digital thermometer, the students discovered that the reaction was exothermic. State the sign of the enthalpy change of the reaction, ΔH .
- b. Although the molar masses of ICl and Br₂ are very similar, the boiling point of ICl is 97.4 °C and that of Br₂ is 58.8 °C. Explain the difference in [2] these boiling points in terms of the intermolecular forces present in each liquid.

- c. (i) Calculate the percentage of iodine by mass in CsICl₂(s), correct to **three** significant figures. [6]
- (ii) State the volume, in cm³, of 0.0500 mol dm⁻³ Na₂S₂O₃(aq) used in the titration.
- (iii) Determine the amount, in mol, of 0.0500 mol dm⁻³ Na₂S₂O₃(aq) added in the titration.
- (iv) The overall reaction taking place during the titration is:



Calculate the amount, in mol, of iodine atoms, I, present in the sample of CsICl₂(s).

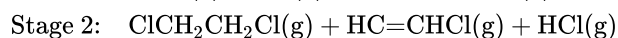
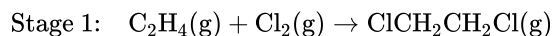
- (v) Calculate the mass of iodine, in g, present in the sample of CsICl₂
- (vi) Determine the percentage by mass of iodine in the sample of CsICl₂(s), correct to **three** significant figures, using your answer from (v).

Chlorine occurs in Group 7, the halogens.

Two stable isotopes of chlorine are ³⁵Cl and ³⁷Cl with mass numbers 35 and 37 respectively.

Chlorine has an electronegativity value of 3.2 on the Pauling scale.

Chloroethene, H₂C=CHCl, the monomer used in the polymerization reaction in the manufacture of the polymer poly(chloroethene), PVC, can be synthesized in the following two-stage reaction pathway.



a.i. Define the term *isotopes of an element*. [2]

a.ii. Calculate the number of protons, neutrons and electrons in the isotopes ³⁵Cl and ³⁷Cl. [2]

Isotope	Number of protons	Number of neutrons	Number of electrons
³⁵ Cl			
³⁷ Cl			

a.iii. Using the mass numbers of the two isotopes and the relative atomic mass of chlorine from Table 5 of the Data Booklet, determine the percentage abundance of each isotope. [2]

Percentage abundance ^{35}Cl :

Percentage abundance ^{37}Cl :

- b.i. Define the term *electronegativity*. [1]
- b.ii. Using Table 7 of the Data Booklet, explain the trends in electronegativity values of the Group 7 elements from F to I. [2]
- b.iii. State the balanced chemical equation for the reaction of potassium bromide, $\text{KBr}(\text{aq})$, with chlorine, $\text{Cl}_2(\text{aq})$. [1]
- b.iv. Describe the colour change likely to be observed in this reaction. [1]
- c.ii. Determine the enthalpy change, ΔH , in kJ mol^{-1} , for stage 1 using average bond enthalpy data from Table 10 of the Data Booklet. [3]
- c.iii. State whether the reaction given in stage 1 is exothermic or endothermic. [1]
- c.iv. Draw the structure of poly(chloroethene) showing **two** repeating units. [1]
- c.v. Suggest why monomers are often gases or volatile liquids whereas polymers are solids. [2]
-

Consider the following list of organic compounds.

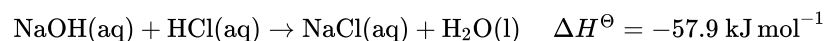
Compound 1: $\text{CH}_3\text{CH}_2\text{CH}(\text{OH})\text{CH}_3$

Compound 2: $\text{CH}_3\text{CH}_2\text{COCH}_3$

Compound 3: $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$

Compound 4: $\text{CH}_3\text{CH}_2\text{CH}_2\text{CHO}$

Hydrochloric acid neutralizes sodium hydroxide, forming sodium chloride and water.



- a. Apply IUPAC rules to state the name of compound 1. [1]
- b. (i) Define the term *structural isomers*. [2]
- (ii) Identify the two compounds in the list that are structural isomers of each other.
- c. Determine the organic product formed when each of the compounds is heated under reflux with excess acidified potassium dichromate(VI). If [4]
no reaction occurs write NO REACTION in the table.

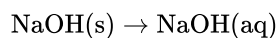
Compound	Organic product
CH ₃ CH ₂ CH(OH)CH ₃	
CH ₃ CH ₂ COCH ₃	
CH ₃ CH ₂ CH ₂ OH	
CH ₃ CH ₂ CH ₂ CHO	

d. Explain the mechanism for the substitution reaction of bromoethane with sodium hydroxide. Use curly arrows to represent the movement of electron pairs. [4]

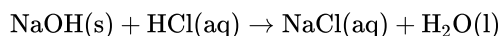
e. (i) Define the term *standard enthalpy change of reaction*, ΔH^\ominus . [9]

(ii) Determine the amount of energy released, in kJ, when 50.0 cm³ of 1.00 mol dm⁻³ sodium hydroxide solution reacts with 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid solution.

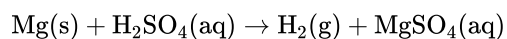
(iii) In an experiment, 2.50 g of solid sodium hydroxide was dissolved in 50.0 cm³ of water. The temperature rose by 13.3 °C. Calculate the standard enthalpy change, in kJ mol⁻¹, for dissolving one mole of solid sodium hydroxide in water.



(iv) Using relevant data from previous question parts, determine ΔH^\ominus , in kJ mol⁻¹, for the reaction of solid sodium hydroxide with hydrochloric acid.



0.100 g of magnesium ribbon is added to 50.0 cm³ of 1.00 mol dm⁻³ sulfuric acid to produce hydrogen gas and magnesium sulfate.

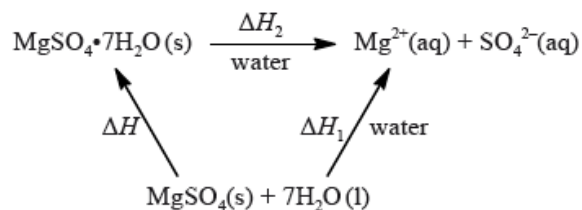


Magnesium sulfate can exist in either the hydrated form or in the anhydrous form. Two students wished to determine the enthalpy of hydration of anhydrous magnesium sulfate. They measured the initial and the highest temperature reached when anhydrous magnesium sulfate, MgSO₄(s), was dissolved in water. They presented their results in the following table.

mass of anhydrous magnesium sulfate / g	3.01
volume of water / cm ³	50.0
initial temperature / °C	17.0
highest temperature / °C	26.7

The students repeated the experiment using 6.16 g of solid hydrated magnesium sulfate, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}(\text{s})$, and 50.0 cm^3 of water. They found the enthalpy change, ΔH_2 , to be $+18 \text{ kJ mol}^{-1}$.

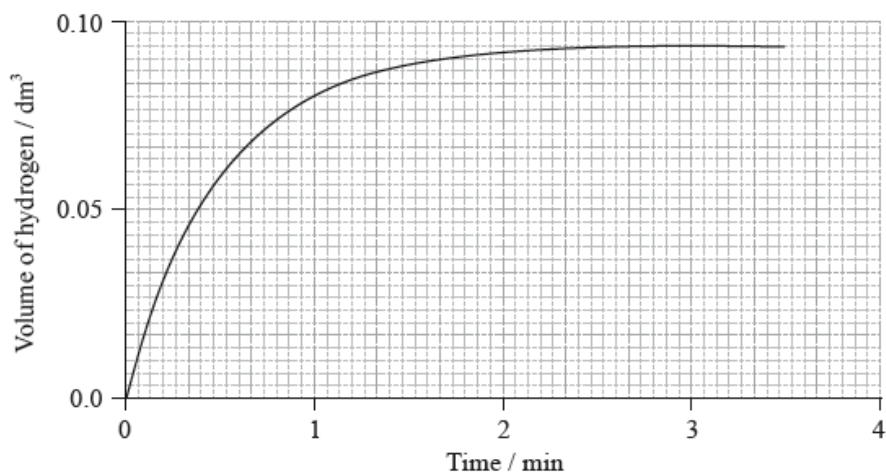
The enthalpy of hydration of solid anhydrous magnesium sulfate is difficult to determine experimentally, but can be determined using the diagram below.



Magnesium sulfate is one of the products formed when acid rain reacts with dolomitic limestone. This limestone is a mixture of magnesium carbonate and calcium carbonate.

a. (i) The graph shows the volume of hydrogen produced against time under these experimental conditions.

[3]



Sketch two curves, labelled I and II, to show how the volume of hydrogen produced (under the same temperature and pressure) changes with time when:

- I. using the same mass of magnesium powder instead of a piece of magnesium ribbon;
- II. 0.100 g of magnesium ribbon is added to 50 cm^3 of $0.500 \text{ mol dm}^{-3}$ sulfuric acid.

(ii) Outline why it is better to measure the volume of hydrogen produced against time rather than the loss of mass of reactants against time.

b. (i) Calculate the amount, in mol, of anhydrous magnesium sulfate.

[3]

(ii) Calculate the enthalpy change, ΔH_1 , for anhydrous magnesium sulfate dissolving in water, in kJ mol^{-1} . State your answer to the correct number of significant figures.

c. (i) Determine the enthalpy change, ΔH , in kJ mol^{-1} , for the hydration of solid anhydrous magnesium sulfate, MgSO_4 .

[2]

(ii) The literature value for the enthalpy of hydration of anhydrous magnesium sulfate is -103 kJ mol^{-1} . Calculate the percentage difference between the literature value and the value determined from experimental results, giving your answer to **one** decimal place. (If you did not obtain an answer for the experimental value in (c)(i) then use the value of -100 kJ mol^{-1} , but this is **not** the correct value.)

d. Another group of students experimentally determined an enthalpy of hydration of -95 kJ mol^{-1} . Outline two reasons which may explain the variation between the experimental and literature values. [2]

e. (i) State the equation for the reaction of sulfuric acid with magnesium carbonate. [1/N/A]

(ii) Deduce the Lewis (electron dot) structure of the carbonate ion, giving the shape and the oxygen-carbon-oxygen bond angle.

Lewis (electron dot) structure:

Shape:

Bond angle:

Ethanol has many industrial uses.

a. (i) State an equation for the formation of ethanol from ethene and the necessary reaction conditions. [4]

Equation:

Conditions:

(ii) Deduce the volume of ethanol, in dm^3 , produced from 1.5 dm^3 of ethene, assuming both are gaseous and at the same temperature and pressure.

b.i. Define the term *average bond enthalpy*. [2]

b.ii. Ethanol can be used as a fuel. Determine the enthalpy of combustion of ethanol at 298 K, in kJ mol^{-1} , using the values in table 10 of the data booklet, assuming all reactants and products are gaseous. [4]

b.iii. Suggest why the value of the enthalpy of combustion of ethanol quoted in table 12 of the data booklet is different to that calculated using bond enthalpies. [1]

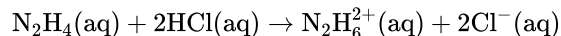
b.iv. Explain why the reaction is exothermic in terms of the bonds involved. [1]

c. Identify the homologous series to which ethanol belongs and state **two** features of a homologous series. [3]

Ethene, C_2H_4 , and hydrazine, N_2H_4 , are hydrides of adjacent elements in the periodic table.

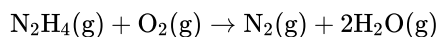
The polarity of a molecule can be explained in terms of electronegativity.

The reaction between $\text{N}_2\text{H}_4(\text{aq})$ and $\text{HCl}(\text{aq})$ can be represented by the following equation.



- a. (i) Draw Lewis (electron dot) structures for C_2H_4 and N_2H_4 showing all valence electrons. [7]
- (ii) State and explain the H–C–H bond angle in ethene and the H–N–H bond angle in hydrazine.
- b. (i) Define the term *electronegativity*. [4]
- (ii) Compare the relative polarities of the C–H bond in ethene and the N–H bond in hydrazine.
- (iii) Hydrazine is a polar molecule and ethene is non-polar. Explain why ethene is non-polar.
- c. The boiling point of hydrazine is much higher than that of ethene. Explain this difference in terms of the intermolecular forces in each compound. [2]
- d. Hydrazine is a valuable rocket fuel. [3]

The equation for the reaction between hydrazine and oxygen is given below.



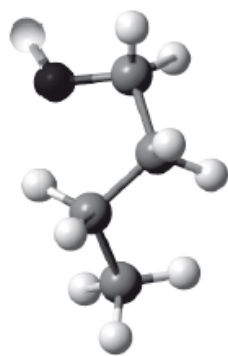
Use the bond enthalpy values from Table 10 of the Data Booklet to determine the enthalpy change for this reaction.

- e. State the name of the product and identify the type of reaction which occurs between ethene and hydrogen chloride. [2]
- f. (i) Identify the type of reaction that occurs. [2]
- (ii) Predict the value of the H–N–H bond angle in $\text{N}_2\text{H}_6^{2+}$.

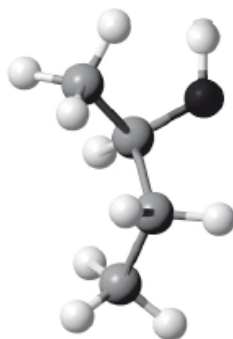
In an experiment to measure the enthalpy change of combustion of ethanol, a student heated a copper calorimeter containing 100 cm^3 of water with a spirit lamp and collected the following data.

Initial temperature of water:	20.0 °C
Final temperature of water:	55.0 °C
Mass of ethanol burned:	1.78 g
Density of water:	1.00 g cm ⁻³

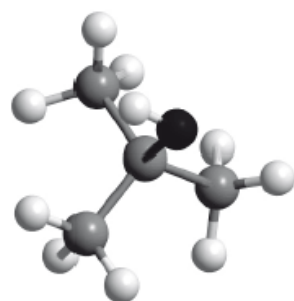
- a. (i) Use the data to calculate the heat evolved when the ethanol was combusted. [6]
- (ii) Calculate the enthalpy change of combustion per mole of ethanol.
- (iii) Suggest two reasons why the result is not the same as the value in the Data Booklet.
- b. Ethanol is part of the homologous series of alcohols. Describe **two** features of a homologous series. [2]
- c. (i) Below are **four structural** isomers of alcohols with molecular formula $\text{C}_4\text{H}_{10}\text{O}$. State the name of each of the isomers **a**, **b**, **c** and **D**. [8]



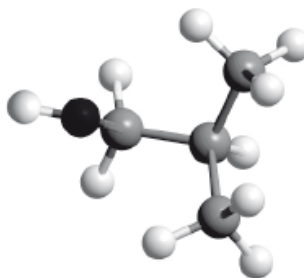
A



B

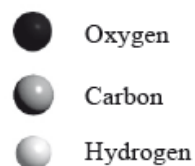


C



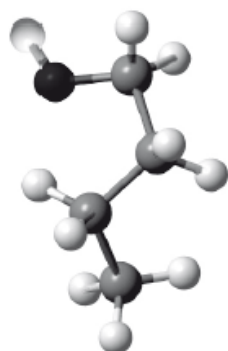
D

Key:

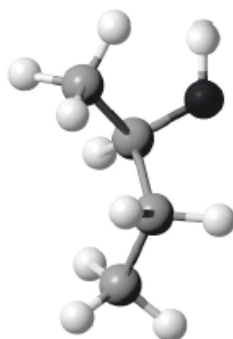


- (ii) Determine the isomer that cannot be oxidized by acidified potassium dichromate(VI), $K_2Cr_2O_7$.
- (iii) Determine the isomer which can be oxidized to butanal.
- (iv) Determine the isomer which can be oxidized to butanone.
- (v) Suggest the structural formula of another isomer of $C_4H_{10}O$.

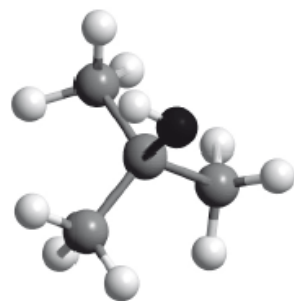
d.



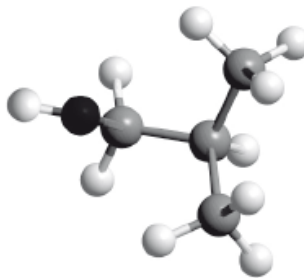
A



B

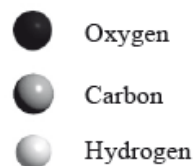


C



D

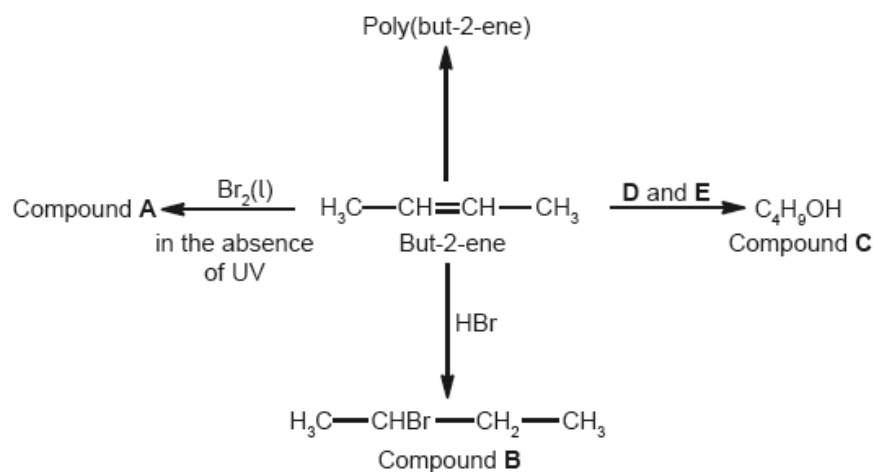
Key:



- (i) Isomer a is formed by reacting 1-bromobutane with aqueous sodium hydroxide. State whether the reaction would proceed via an S_N1 or S_N2 mechanism.
- (ii) Explain the mechanism named in part (d) (i) using curly arrows to represent the movement of electron pairs.

[4]

Some reactions of but-2-ene are given below.

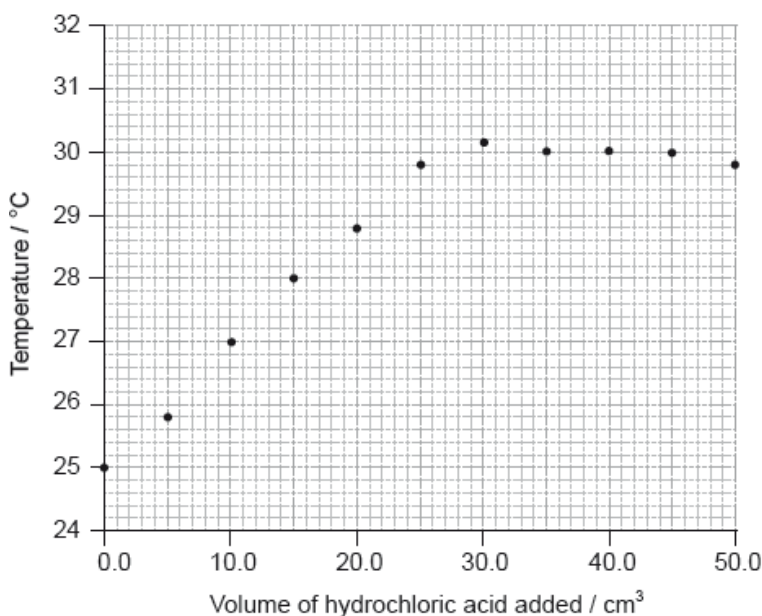


- a.i. Deduce the full structural formula of compound **A**. [1]
- a.ii. Apply IUPAC rules to name compound **A**. [1]
- a.iii. Describe the colour change observed when excess but-2-ene reacts with bromine to form compound **A**. [1]
- b. State the names of the reagents **D** and **E**. [2]
- c. (i) Outline **two** reasons why the polymerization of alkenes is of economic importance. [3]
- (ii) Identify the structure of the repeating unit of poly(but-2-ene).
- d. Compound **C**, $\text{C}_4\text{H}_9\text{OH}$, can also be formed directly from compound **B**, $\text{CH}_3\text{CHBrCH}_2\text{CH}_3$. [2]
- (i) State the reagent and the conditions required for this reaction.
- (ii) State the name of the type of reaction occurring in this conversion.
- e. Compound **C** can be oxidized by acidified potassium dichromate(VI) to form compound **F**. [2]
- (i) State the name of the functional group present in compound **F**.
- (ii) Deduce the structural formula of an alcohol which is a structural isomer of compound **C** and **cannot** be oxidized by acidified potassium dichromate(VI).
- f. Explain why but-2-ene is more volatile than compound **C**, $\text{C}_4\text{H}_9\text{OH}$. [2]
- g.i. Define the term *average bond enthalpy*. [2]
- g.ii. Deduce the equation for the complete combustion of compound **C**. [1]

- g.iii Determine the enthalpy change, ΔH , in kJ mol^{-1} , for the complete combustion of compound **C** when all reactants and products are in the gaseous state, using table 10 of the data booklet. [3]

A student carried out an experiment to determine the concentration of a hydrochloric acid solution and the enthalpy change of the reaction between aqueous sodium hydroxide and this acid by thermometric titration.

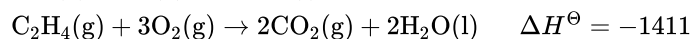
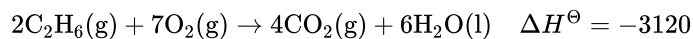
She added 5.0 cm^3 portions of hydrochloric acid to 25.0 cm^3 of 1.00 mol dm^{-3} sodium hydroxide solution in a glass beaker until the total volume of acid added was 50.0 cm^3 , measuring the temperature of the mixture each time. Her results are plotted in the graph below.



The initial temperature of both solutions was the same.

- a.i. By drawing appropriate lines, determine the volume of hydrochloric acid required to completely neutralize the 25.0 cm^3 of sodium hydroxide solution. [2]
- a.ii. Determine the concentration of the hydrochloric acid, including units. [2]
- b.i. Determine the change in temperature, ΔT . [1]
- b.ii. Calculate the enthalpy change, in kJ mol^{-1} , for the reaction of hydrochloric acid and sodium hydroxide solution. [3]
- b.iii. The accepted theoretical value from the literature of this enthalpy change is -58 kJ mol^{-1} . Calculate the percentage error correct to **two** significant figures. [1]
- b.iv. Suggest the major source of error in the experimental procedure **and** an improvement that could be made to reduce it. [2]

The standard enthalpy change of three combustion reactions is given below in kJ.

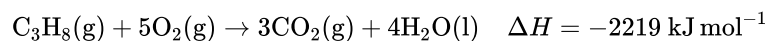


Based on the above information, calculate the standard change in enthalpy, ΔH^\ominus , for the following reaction.

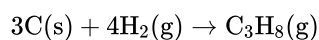


a. The standard enthalpy change of three combustion reactions are given below.

[4]



Determine the change in enthalpy, ΔH , in kJ mol^{-1} , for the formation of propane in the following reaction.



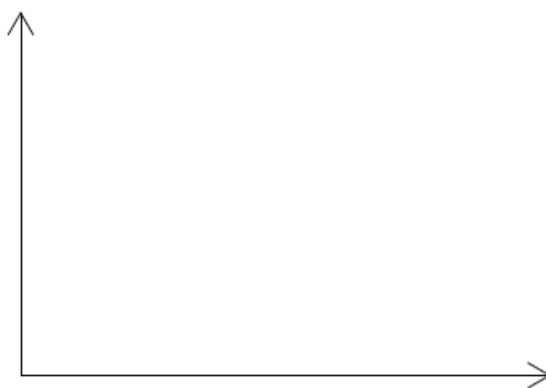
b. A catalyst provides an alternative pathway for a reaction, lowering the activation energy, E_a . Define the term *activation energy*, E_a .

[1]

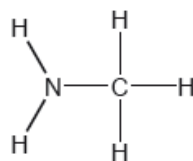
c. Sketch **two** Maxwell–Boltzmann energy distribution curves for a fixed amount of gas at two different temperatures, T_1 and T_2 ($T_2 > T_1$) and

[3]

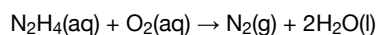
label **both** axes.



Two hydrides of nitrogen are ammonia and hydrazine, N_2H_4 . One derivative of ammonia is methanamine whose molecular structure is shown below.



Hydrazine is used to remove oxygen from water used to generate steam or hot water.

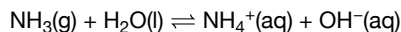


The concentration of dissolved oxygen in a sample of water is $8.0 \times 10^{-3} \text{ g dm}^{-3}$.

a. Estimate the H–N–H bond angle in methanamine using VSEPR theory.

[1]

b. Ammonia reacts reversibly with water. [2]

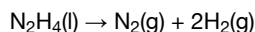


Explain the effect of adding $\text{H}^+(\text{aq})$ ions on the position of the equilibrium.

c. Hydrazine reacts with water in a similar way to ammonia. Deduce an equation for the reaction of hydrazine with water. [1]

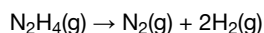
d. Outline, using an ionic equation, what is observed when magnesium powder is added to a solution of ammonium chloride. [2]

e. Hydrazine has been used as a rocket fuel. The propulsion reaction occurs in several stages but the overall reaction is: [1]

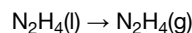


Suggest why this fuel is suitable for use at high altitudes.

f. Determine the enthalpy change of reaction, ΔH , in kJ, when 1.00 mol of gaseous hydrazine decomposes to its elements. Use bond enthalpy values in section 11 of the data booklet. [3]



g. The standard enthalpy of formation of $\text{N}_2\text{H}_4(\text{l})$ is $+50.6 \text{ kJ mol}^{-1}$. Calculate the enthalpy of vaporization, ΔH_{vap} , of hydrazine in kJ mol^{-1} . [2]



(If you did not get an answer to (f), use -85 kJ but this is not the correct answer.)

h.i. Calculate, showing your working, the mass of hydrazine needed to remove all the dissolved oxygen from 1000 dm^3 of the sample. [3]

h.ii. Calculate the volume, in dm^3 , of nitrogen formed under SATP conditions. (The volume of 1 mol of gas = 24.8 dm^3 at SATP.) [1]

Impurities cause phosphine to ignite spontaneously in air to form an oxide of phosphorus and water.

a. (i) 200.0 g of air was heated by the energy from the complete combustion of 1.00 mol phosphine. Calculate the temperature rise using section 1 of the data booklet and the data below. [5]

Standard enthalpy of combustion of phosphine, $\Delta H_c^\ominus = -750 \text{ kJ mol}^{-1}$

Specific heat capacity of air = $1.00 \text{ J g}^{-1} \text{ K}^{-1} = 1.00 \text{ kJ kg}^{-1} \text{ K}^{-1}$

(ii) The oxide formed in the reaction with air contains 43.6% phosphorus by mass. Determine the empirical formula of the oxide, showing your method.

(iii) The molar mass of the oxide is approximately 285 g mol^{-1} . Determine the molecular formula of the oxide.

b. (i) State the equation for the reaction of this oxide of phosphorus with water. [5]

(ii) Predict how dissolving an oxide of phosphorus would affect the pH and electrical conductivity of water.

pH:

Electrical conductivity:

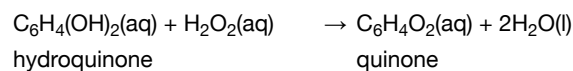
(iii) Suggest why oxides of phosphorus are not major contributors to acid deposition.

(iv) The levels of sulfur dioxide, a major contributor to acid deposition, can be minimized by either pre-combustion and post-combustion methods. Outline **one** technique of each method.

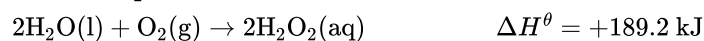
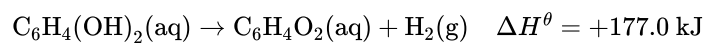
Pre-combustion:

Post-combustion:

The Bombardier beetle sprays a mixture of hydroquinone and hydrogen peroxide to fight off predators. The reaction equation to produce the spray can be written as:



a.i. Calculate the enthalpy change, in kJ, for the spray reaction, using the data below. [2]



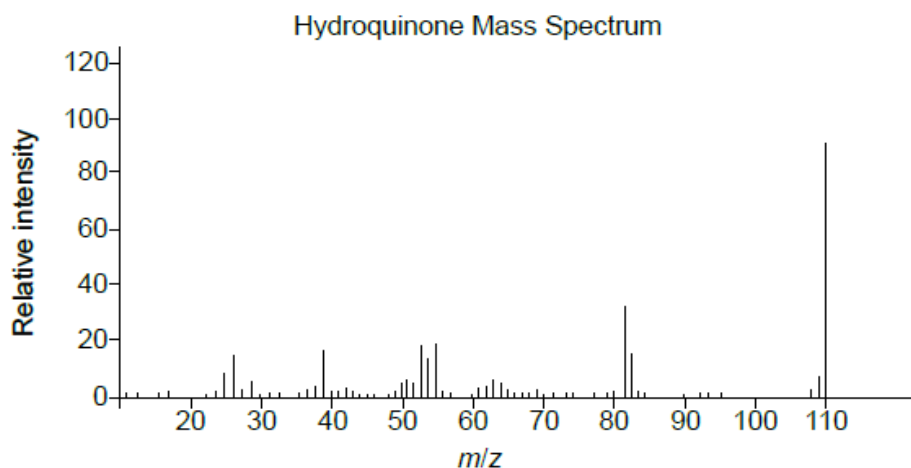
a.ii. The energy released by the reaction of one mole of hydrogen peroxide with hydroquinone is used to heat 850 cm³ of water initially at 21.8°C. [2]

Determine the highest temperature reached by the water.

Specific heat capacity of water = 4.18 kJ kg⁻¹ K⁻¹.

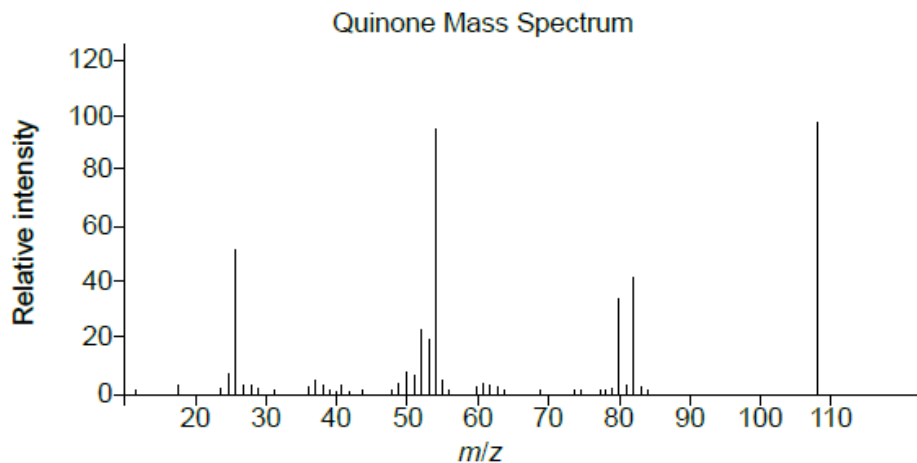
(If you did not obtain an answer to part (i), use a value of 200.0 kJ for the energy released, although this is not the correct answer.)

b. Identify the species responsible for the peak at $m/z = 110$ in the mass spectrum of hydroquinone. [1]



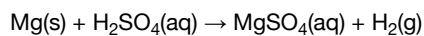
[Source: <http://webbook.nist.gov>]

c. Identify the highest m/z value in the mass spectrum of quinone. [1]

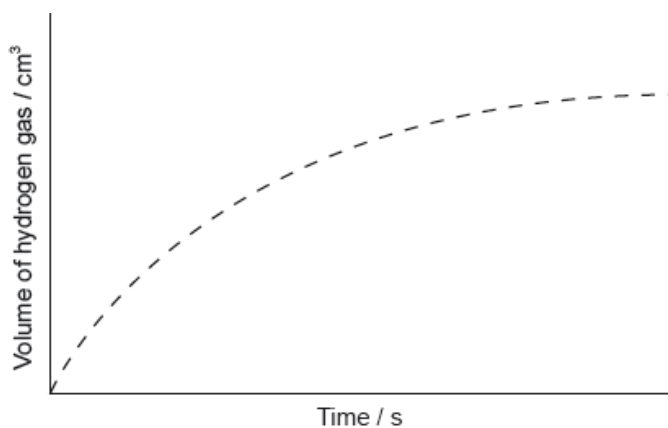


[Source: <http://webbook.nist.gov>]

Magnesium reacts with sulfuric acid:



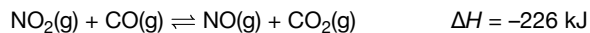
The graph shows the results of an experiment using excess magnesium ribbon and dilute sulfuric acid.

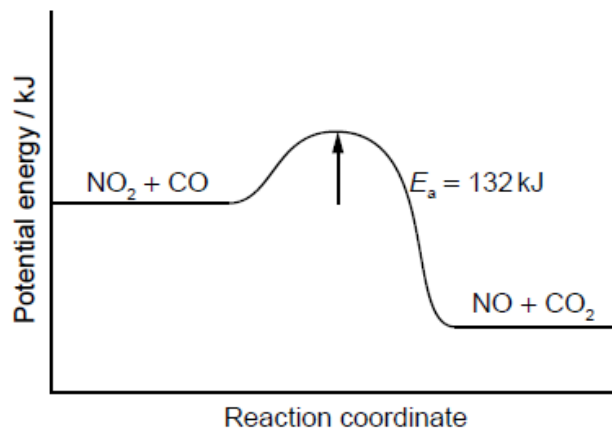


a.i. Outline why the rate of the reaction decreases with time. [1]

a.ii. Sketch, on the same graph, the expected results if the experiment were repeated using powdered magnesium, keeping its mass and all other variables unchanged. [1]

b. Nitrogen dioxide and carbon monoxide react according to the following equation: [1]





Calculate the activation energy for the reverse reaction.

- c. State the equation for the reaction of NO_2 in the atmosphere to produce acid deposition.

[1]

Trends in physical and chemical properties are useful to chemists.

The Activity series lists the metal in order of reactivity.

Mn ↑ Most reactive
 Ni ↑
 Ag | Least reactive

- a. Explain the general increasing trend in the first ionization energies of the period 3 elements, Na to Ar.

[2]

- b. Explain why the melting points of the group 1 metals ($\text{Li} \rightarrow \text{Cs}$) decrease down the group.

[2]

- c. State an equation for the reaction of phosphorus (V) oxide, $\text{P}_4\text{O}_{10}(\text{s})$, with water.

[1]

- d. Describe the emission spectrum of hydrogen.

[2]

- e.i. Identify the strongest reducing agent in the given list.

[1]

- e.ii. A voltaic cell is made up of a Mn^{2+}/Mn half-cell and a Ni^{2+}/Ni half-cell.

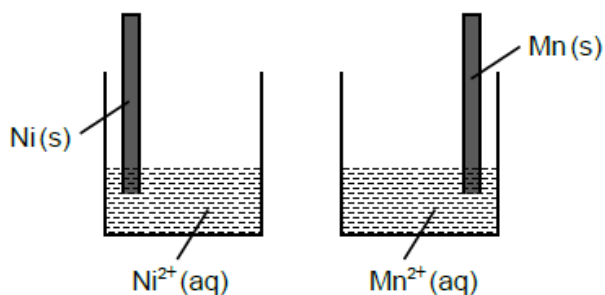
[1]

Deduce the equation for the cell reaction.

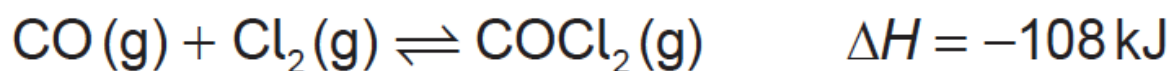
- e.iii. The voltaic cell stated in part (ii) is partially shown below.

[2]

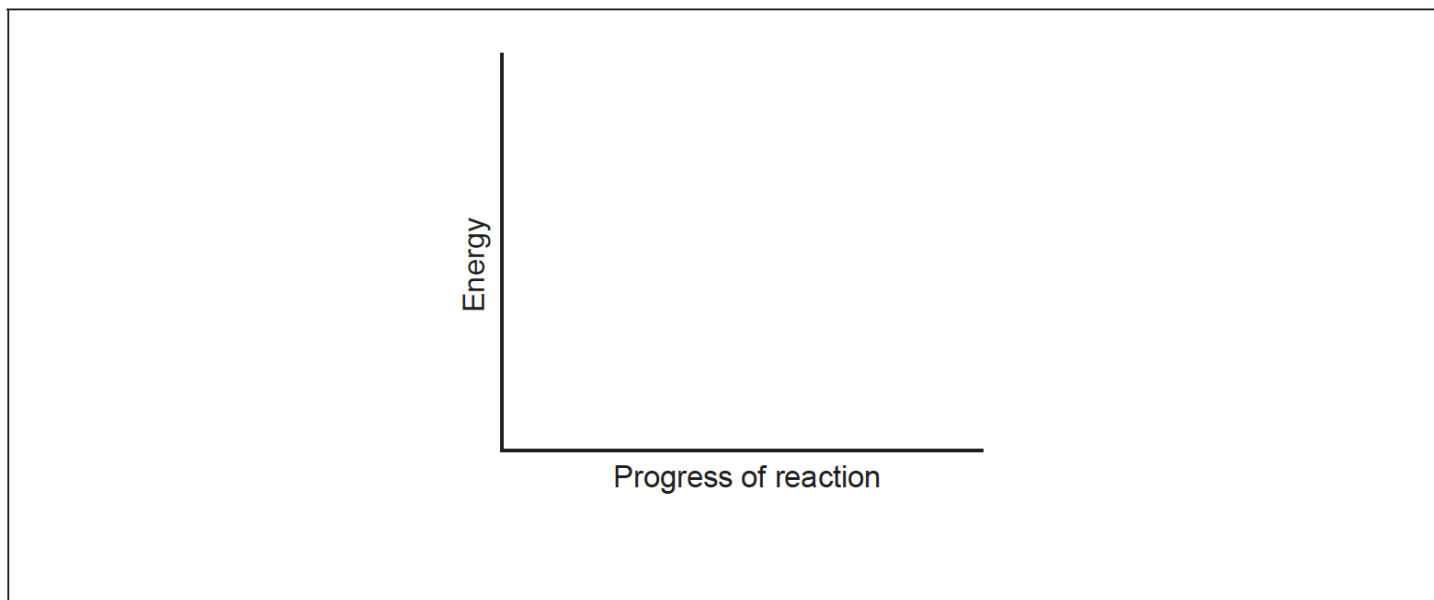
Draw and label the connections needed to show the direction of electron movement and ion flow between the two half-cells.



Phosgene, COCl_2 , is usually produced by the reaction between carbon monoxide and chlorine according to the equation:

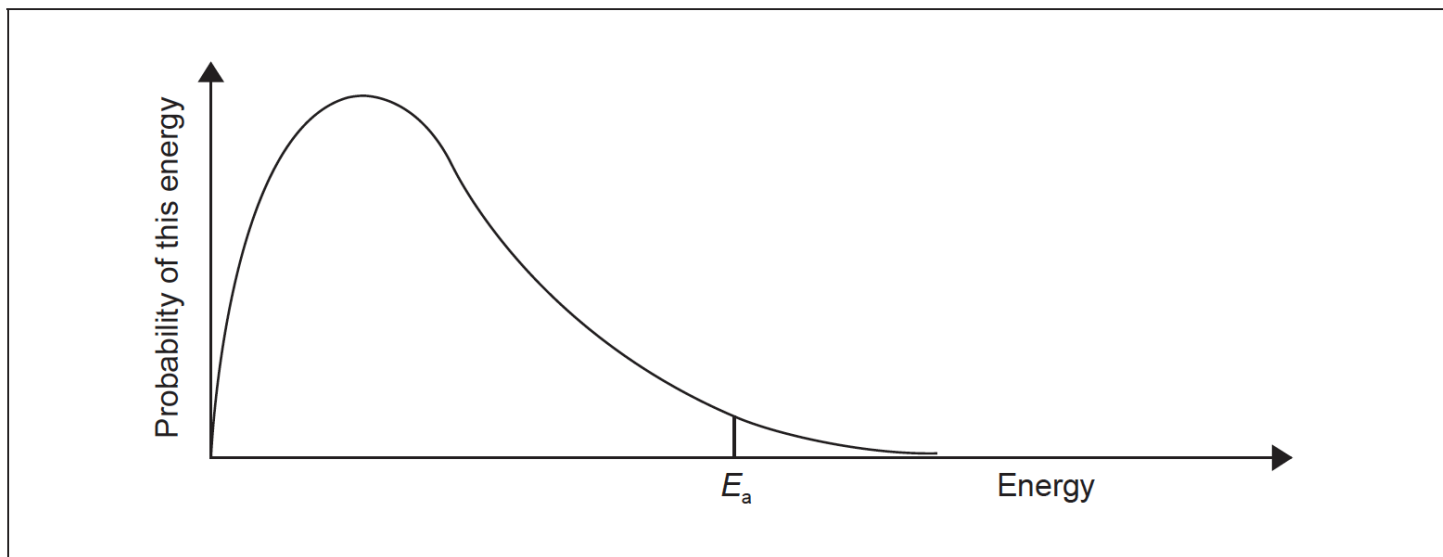


- a. (i) Deduce the equilibrium constant expression, K_c , for this reaction. [2]
- (ii) State the effect of an increase in the total pressure on the equilibrium constant, K_c .
- b. (i) Sketch the potential energy profile for the synthesis of phosgene, using the axes given, indicating both the enthalpy of reaction and activation energy. [6]



(ii) This reaction is normally carried out using a catalyst. Draw a dotted line labelled "Catalysed" on the diagram above to indicate the effect of the catalyst.

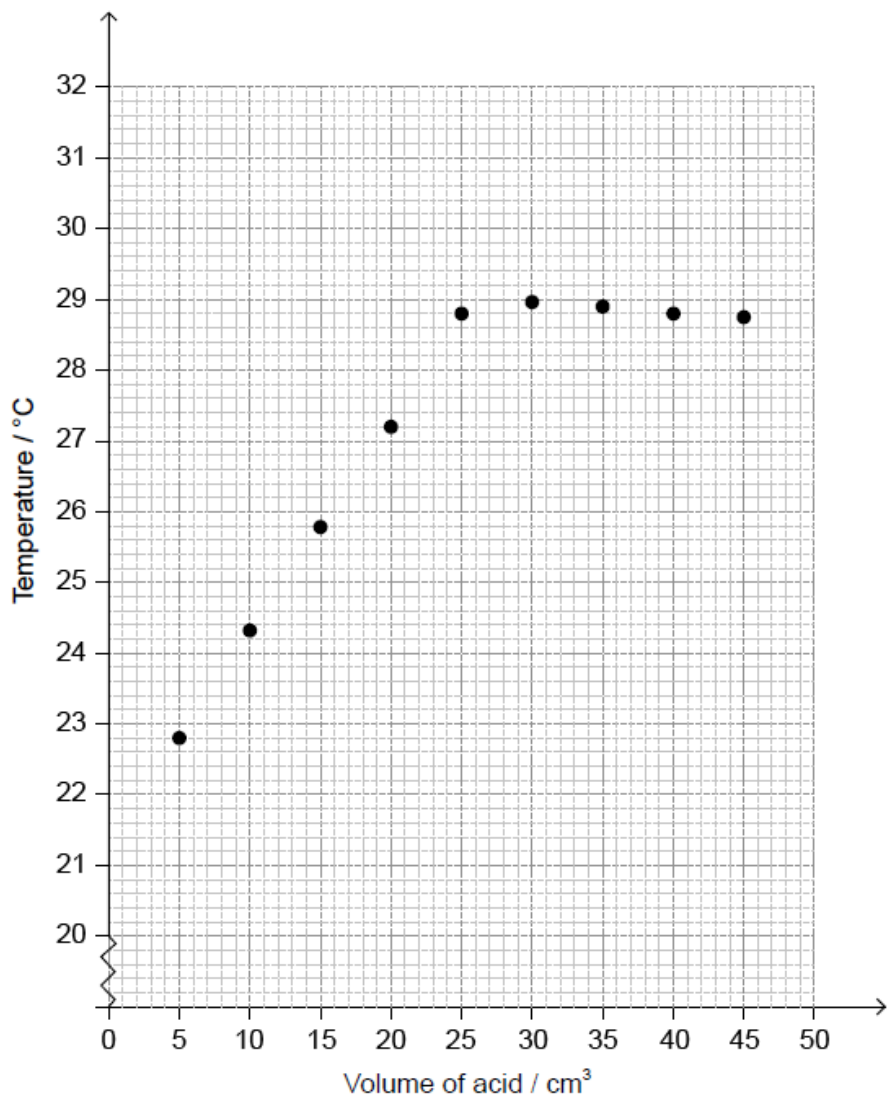
(iii) Sketch and label a second Maxwell-Boltzmann energy distribution curve representing the same system but at a higher temperature, T_{higher} .



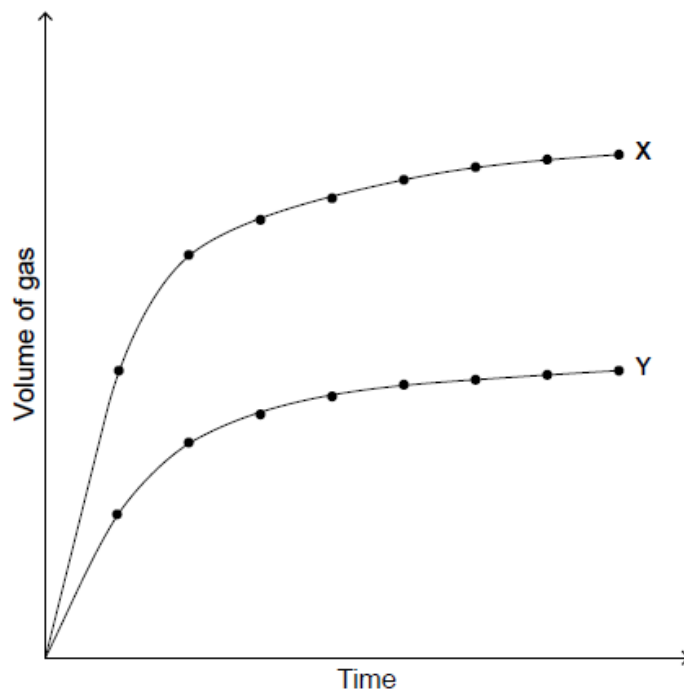
(iv) Explain why an increase in temperature increases the rate of this reaction.

A student titrated an ethanoic acid solution, $\text{CH}_3\text{COOH}(\text{aq})$, against 50.0 cm^3 of $0.995 \text{ mol dm}^{-3}$ sodium hydroxide, $\text{NaOH}(\text{aq})$, to determine its concentration.

The temperature of the reaction mixture was measured after each acid addition and plotted against the volume of acid.



Curves X and Y were obtained when a metal carbonate reacted with the same volume of ethanoic acid under two different conditions.



- a. Using the graph, estimate the initial temperature of the solution. [1]
- b. Determine the maximum temperature reached in the experiment by analysing the graph. [1]
- c. Calculate the concentration of ethanoic acid, CH_3COOH , in mol dm^{-3} . [2]
- d.i. Determine the heat change, q , in kJ, for the neutralization reaction between ethanoic acid and sodium hydroxide. [2]
- Assume the specific heat capacities of the solutions and their densities are those of water.
- d.ii. Calculate the enthalpy change, ΔH , in kJ mol^{-1} , for the reaction between ethanoic acid and sodium hydroxide. [2]
- e.i. Explain the shape of curve **X** in terms of the collision theory. [2]
- e.ii. Suggest **one** possible reason for the differences between curves **X** and **Y**. [1]
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