SL Paper 2

a.i. Define the term *isotopes*. [1]

a.ii. A sample of silicon contains three isotopes. [2]

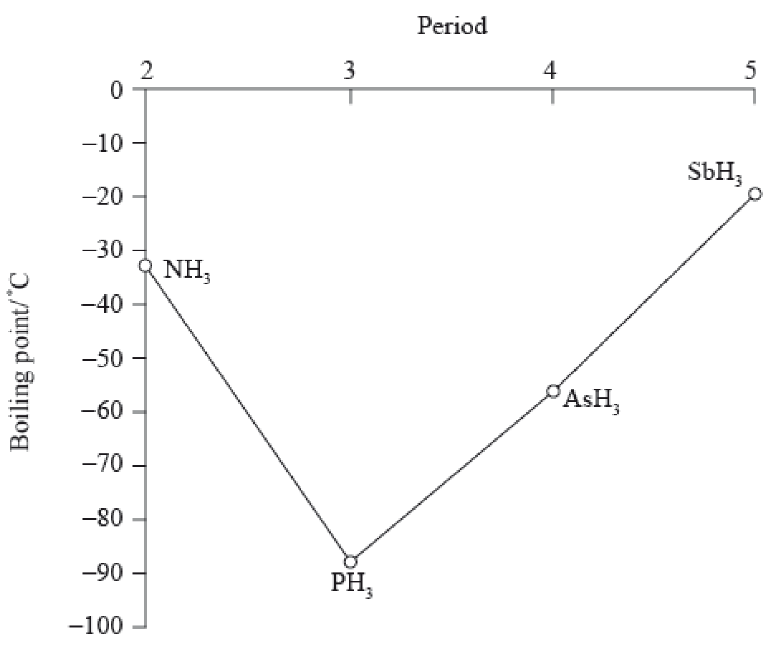
Isotope	Percentage abundance / %
^{28}Si	92.23
^{29}Si	4.68
^{30}Si	3.09

Calculate the relative atomic mass of silicon using this data.

a.iii. Describe the structure and bonding in silicon dioxide and carbon dioxide. [4]

b.i. Draw the Lewis structure of NH3, state its shape and deduce and explain the H–N–H bond angle in NH_3. [4]

b.ii. The graph below shows the boiling points of the hydrides of group 5. Discuss the variation in the boiling points. [4]



c. Explain, using diagrams, why CO and NO_2 are polar molecules but CO_2 is a non-polar molecule. [5]

Boron is most often encountered as a component in borosilicate glass (heat resistant glass).

The naturally occurring element contains two stable isotopes, $^{10}_{5}B$ and $^{11}_{5}B$.

a. State the number of protons, neutrons and electrons in an atom of $^{11}_{5}B$. [1]

	Protons	Neutrons	Electrons
${}^{11}_{5}B$			

b. The relative atomic mass of boron is 10.8, to three significant figures. Calculate the percentage of ${}^{10}_{5}B$ in the naturally occurring element. [2]

c. Isotopes of boron containing 7 and 8 neutrons also exist. Suggest why releasing isotopes containing more neutrons than the stable isotope into the environment can be dangerous. [1]

d. (i) State the formula of the compound that boron forms with fluorine. [3]

(ii) Explain why this compound acts as a Lewis acid.

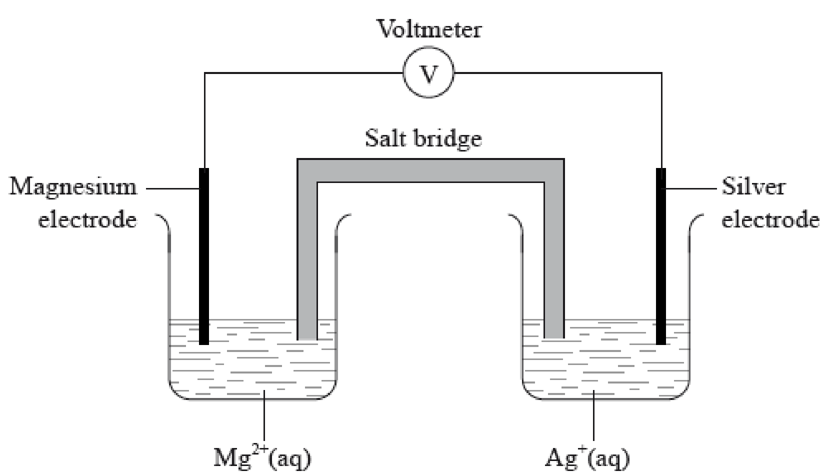
Iron rusts in the presence of oxygen and water. Rusting is a redox process involving several steps that produces hydrated iron(III) oxide, $Fe_2O_3 \bullet nH_2O$, as the final product.

The half-equations involved for the first step of rusting are given below.

Half-equation 1: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Half-equation 2: $O_2(aq) + 4e^- + 2H_2O(l) \rightarrow 4OH^-(aq)$

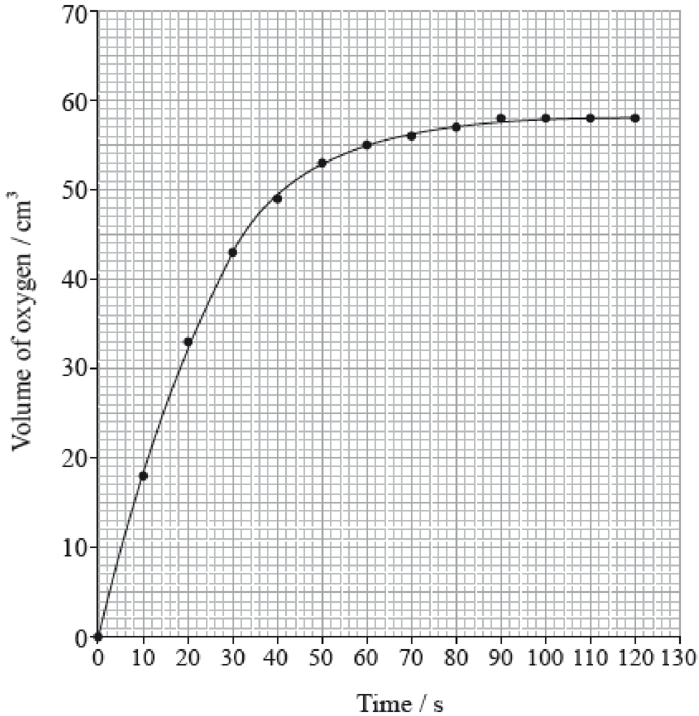
A voltaic cell is made from a half-cell containing a magnesium electrode in a solution of magnesium nitrate and a half-cell containing a silver electrode in a solution of silver(I) nitrate.



Hydrogen peroxide decomposes according to the equation below.

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

The rate of the decomposition can be monitored by measuring the volume of oxygen gas released. The graph shows the results obtained when a solution of hydrogen peroxide decomposed in the presence of a CuO catalyst.



a. (i) Identify whether half-equation 1 represents oxidation or reduction, giving a reason for your answer. [5]

(ii) Identify the oxidation number of each atom in the three species in half-equation 2.

$$O_2(aq) + 4e^- + 2H_2O(l) \rightarrow 4OH^-(aq)$$

$\square \qquad \square\square \qquad \square\square$

(iii) Deduce the overall redox equation for the first step of rusting by combining half-equations 1 and 2.

(iv) Identify the reducing agent in the redox equation in part (iii).

b. The oxygen in half-equation 2 is atmospheric oxygen that is found dissolved in water in very small concentrations. Explain, in terms of intermolecular forces, why oxygen is not very soluble in water. [2]

c. (i) Given that magnesium is more reactive than silver, deduce the half-equations for the reactions occurring at each electrode, including state symbols. [3]

Negative electrode (anode):

Positive electrode (cathode):

(ii) Outline **one** function of the salt bridge.

d. (i) State the property that determines the order in which elements are arranged in the periodic table. [3]

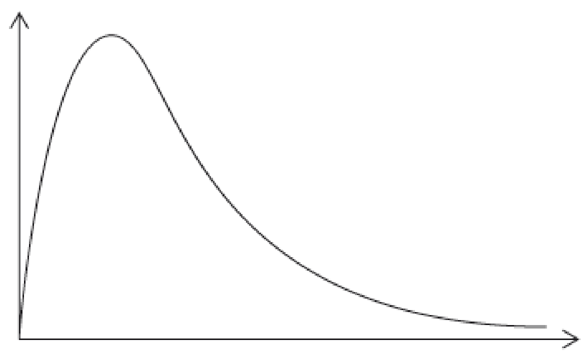
(ii) State the relationship between the electron arrangement of an element and its group and period in the periodic table.

e. (i) The experiment is repeated with the same amount of a more effective catalyst, MnO_2, under the same conditions and using the same concentration and volume of hydrogen peroxide. On the graph above, sketch the curve you would expect. [7]

(ii) Outline how the initial rate of reaction can be found from the graph.

(iii) Outline a different experimental procedure that can be used to monitor the decomposition rate of hydrogen peroxide.

(iv) A Maxwell–Boltzmann energy distribution curve is drawn below. Label both axes and explain, by annotating the graph, how catalysts increase the rate of reaction.



A sample of magnesium contains three isotopes: magnesium-24, magnesium-25 and magnesium-26, with abundances of 77.44%, 10.00% and 12.56% respectively.

Phosphorus(V) oxide, P_4O_{10} $(M_r = 283.88)$, reacts vigorously with water $(M_r = 18.02)$, according to the equation below.

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

a.i. Calculate the relative atomic mass of this sample of magnesium correct to **two** decimal places. [2]

a.iii Predict the relative atomic radii of the three magnesium isotopes, giving your reasons. [2]

b. Describe the bonding in magnesium. [2]

c. State an equation for the reaction of magnesium oxide with water. [1]

d.i. A student added 5.00 g of P_4O_{10} to 1.50 g of water. Determine the limiting reactant, showing your working. [2]

d.ii. Calculate the mass of phosphoric(V) acid, H_3PO_4, formed in the reaction. [2]

d.iii State a balanced equation for the reaction of aqueous H_3PO_4 with excess aqueous sodium hydroxide, including state symbols. [2]

d.iv State the formula of the conjugate base of H_3PO_4. [1]

e. (i) Deduce the Lewis structure of PH_4^+. [4]

(ii) Predict, giving a reason, the bond angle around the phosphorus atom in PH_4^+.

(iii) Predict whether or not the P–H bond is polar, giving a reason for your choice.

The element boron has two naturally occurring isotopes, ^{10}B and ^{11}B.

a.i. Define the term *isotopes of an element*. [1]

a.ii. Calculate the percentage abundance of **each** isotope, given that the relative atomic mass of B is 10.81. [2]

c.i. Deduce the Lewis structures of NH_3 and BF_3. [2]

NH_3 BF_3

c.ii. Describe how covalent bonds are formed. [1]

c.iii Compare the shapes of the two molecules and explain the difference using valence shell electron pair repulsion theory (VSEPR). [4]

c.iv Predict and explain whether the molecules NH_3 and BF_3 are polar molecules. [2]

Lithium and boron are elements in period 2 of the periodic table. Lithium occurs in group 1 (the alkali metals) and boron occurs in group 3. Isotopes exist for both elements.

Every element has its own unique line emission spectrum.

a. (i) Define the terms *atomic number*, *mass number* and *isotopes of an element*. [10]

Atomic number:

Mass number:

Isotopes of an element:

(ii) Distinguish between the terms *group* and *period*.

(iii) Deduce the electron arrangements of the lithium ion, Li^+, and the boron atom, B.

Li^+:

B:

(iv) Naturally occurring boron exists as two isotopes with mass numbers of 10 and 11. Calculate the percentage abundance of the lighter isotope, using this information and the relative atomic mass of boron in Table 5 of the Data Booklet.

v) Lithium exists as two isotopes with mass numbers of 6 and 7. Deduce the number of protons, electrons and neutrons for each isotope.

Mass number (A)	Number of protons	Number of electrons	Number of neutrons
6			
7			

b. (i) Distinguish between a *continuous spectrum* and a *line spectrum*. [6]

(ii) Draw a diagram to show the electron transitions between energy levels in a hydrogen atom that are responsible for the two series of lines in the ultraviolet and visible regions of the spectrum. Label your diagram to show **three** transitions for each series.

c. (i) Explain why metals are good conductors of electricity and why they are malleable. [4]

(ii) Iron is described as a transition metal. Identify the **two** most common ions of iron.

iii) Deduce the chemical formulas of lithium oxide and iron(II) oxide.

Lithium oxide:

Iron(II) oxide:

Iron has three main naturally occurring isotopes which can be investigated using a mass spectrometer.

b. A sample of iron has the following isotopic composition by mass. [2]

Isotope	^{54}Fe	^{56}Fe	^{57}Fe
Relative abundance / %	5.95	91.88	2.17

Calculate the relative atomic mass of iron based on this data, giving your answer to **two decimal places**.

c. Calculate the number of electrons in the ion $^{56}Fe^{2+}$. [1]

d. Describe the bonding in iron and explain the electrical conductivity and malleability of the metal. [4]

c. The relative atomic mass of naturally occurring copper is 63.55. Calculate the abundances of $^{63}\mathrm{Cu}$ and $^{65}\mathrm{Cu}$ in naturally occurring copper. [2]

d. The isotopes of some elements are radioactive. State a radioisotope used in medicine. [1]

e. State a balanced equation for the reaction of sodium with water. Include state symbols. [2]

f. With reference to electronic arrangements, suggest why the reaction between rubidium and water is more vigorous than that between sodium and water. [2]

g. Describe and explain what you will see if chlorine gas is bubbled through a solution of [3]

(i) potassium iodide.

(ii) potassium fluoride.

Calcium carbide, CaC_2, is an ionic solid.

a. Describe the nature of ionic bonding. [1]

b. State the electron configuration of the Ca^{2+} ion. [1]

c. When calcium compounds are introduced into a gas flame a red colour is seen; sodium compounds give a yellow flame. Outline the source of the colours and why they are different. [2]

d.i. Suggest **two** reasons why solid calcium has a greater density than solid potassium. [2]

d.ii. Outline why solid calcium is a good conductor of electricity. [1]

e. Calcium carbide reacts with water to form ethyne and calcium hydroxide. [1]

$CaC_2(s) + H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$

Estimate the pH of the resultant solution.

$^{131}\mathrm{I}$ is a radioactive isotope of iodine.

a.i. Define the term *isotope*. [1]

a.ii. Determine the number of neutrons in one atom of iodine-131. [1]

2-methylbutan-2-ol, $(CH_3)_2C(OH)CH_2CH_3$, is a liquid with a smell of camphor that was formerly used as a sedative. One way of producing it starts with 2-methylbut-2-ene.

2-chloro-2-methylbutane contains some molecules with a molar mass of approximately $106\ \text{g mol}^{-1}$ and some with a molar mass of approximately $108\ \text{g mol}^{-1}$.

a. Draw the structure of 2-methylbut-2-ene. [1]

b. State the other substances required to convert 2-methylbut-2-ene to 2-methylbutan-2-ol. [2]

c. Explain whether you would expect 2-methylbutan-2-ol to react with acidified potassium dichromate(VI). [2]

d. Explain why 2-methylbut-2-ene is less soluble in water than 2-methylbutan-2-ol. [2]

f.i. Outline why there are molecules with different molar masses. [1]

Rubidium contains two stable isotopes, ^{85}Rb and ^{87}Rb. The relative atomic mass of rubidium is given in Table 5 of the Data Booklet.

a. Calculate the percentage of each isotope in pure rubidium. State your answers to **three** significant figures. [2]

c. State the number of electrons and the number of neutrons present in an atom of ^{87}Rb. [2]

Number of electrons:

Number of neutrons:

Isotopes are atoms of the same element with different mass numbers. Two isotopes of cobalt are Co-59 and Co-60.

Deduce the missing information and complete the following table.

Symbol	$^{59}Co^{3+}$	^{60}Co	
Number of protons	27		53
Number of neutrons		33	72
Number of electrons		27	53

Chlorine occurs in Group 7, the halogens.

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

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

Chloroethene, $H_2C{=}CHCl$, 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.

$$\begin{array}{ll}\text{Stage 1:} & C_2H_4(g) + Cl_2(g) \rightarrow ClCH_2CH_2Cl(g) \\ \text{Stage 2:} & ClCH_2CH_2Cl(g) + HC{=}CHCl(g) + HCl(g)\end{array}$$

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

a.ii. Calculate the number of protons, neutrons and electrons in the isotopes ^{35}Cl and ^{37}Cl. [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, KBr(aq), with chlorine, Cl_2(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]

Carbon and silicon belong to the same group of the periodic table.

Both silicon and carbon form oxides.

b. State the period numbers of both carbon and silicon. [1]

c. Describe and compare **three** features of the structure and bonding in the three allotropes of carbon: diamond, graphite and C_{60} fullerene. [6]

d.i. Draw the Lewis structure of CO_2 and predict its shape and bond angle. [2]

d.ii. Describe the structure and bonding in SiO_2. [2]

d.iii. Explain why silicon dioxide is a solid and carbon dioxide is a gas at room temperature. [2]

e. Describe the bonding within the carbon monoxide molecule. [2]

f. Silicon has three stable isotopes, ^{28}Si, ^{29}Si and ^{30}Si. The heaviest isotope, ^{30}Si, has a percentage abundance of 3.1%. Calculate the percentage abundance of the lightest isotope to one decimal place. [2]

Draw and label an energy level diagram for the hydrogen atom. In your diagram show how the series of lines in the ultraviolet and visible regions of its emission spectrum are produced, clearly labelling each series.

A sample of vaporized elemental magnesium is introduced into a mass spectrometer.

One of the ions that reaches the detector is $^{26}Mg^{+}$.

a. Calculate the number of protons, neutrons and electrons in the $^{26}Mg^{+}$ ion. [2]

d. The sample contained the three isotopes ^{24}Mg, ^{25}Mg and ^{26}Mg. The relative percentage abundances of ^{25}Mg and ^{26}Mg are 10.00% and 11.01% respectively. Calculate the relative atomic mass (A_r) of magnesium, accurate to **two** decimal places. [2]

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a. State the relative mass and charge of the subatomic particles of an atom. [2]

	Relative mass	Relative charge
Proton	....	+1
Electron	5×10^{-4}	....
Neutron	....	....

b. (i) Calculate the number of neutrons and electrons in one atom of $^{65}\mathrm{Cu}$. [3]

Neutrons:

Electrons:

(ii) State one difference in the physical properties of the isotopes $^{63}\mathrm{Cu}$ and $^{65}\mathrm{Cu}$ and explain why their chemical properties are the same.

Physical:

Chemical:

c. Describe the bonding in solid copper. [2]

d. Suggest **two** properties of copper that make it useful and economically important. [1]

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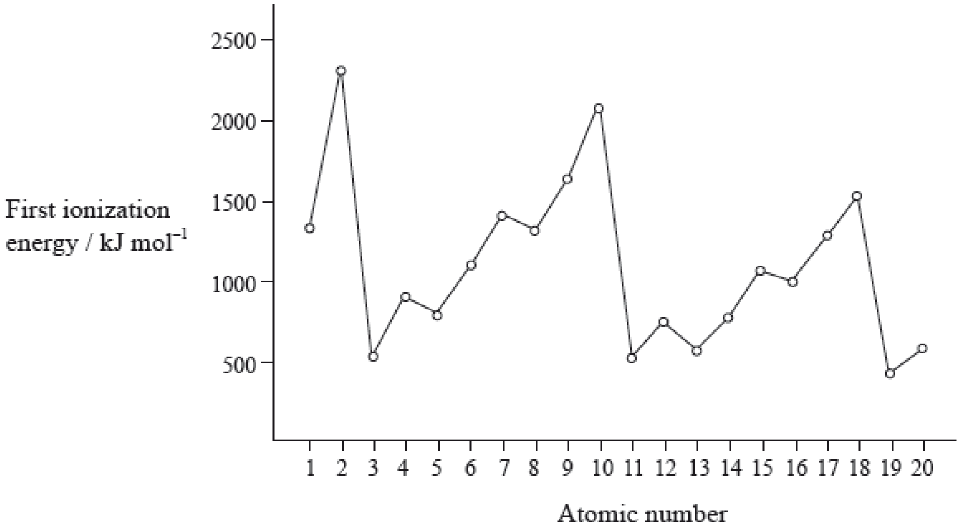
a. Explain why the relative atomic mass of argon is greater than the relative atomic mass of potassium, even though the atomic number of potassium is greater than the atomic number of argon. [1]

b. Deduce the numbers of protons and electrons in the K^{+} ion. [1]

The graph of the first ionization energy plotted against atomic number for the first twenty elements shows periodicity.



a.iii Strontium exists as four naturally-occurring isotopes. Calculate the relative atomic mass of strontium to two decimal places from the following data. [2]

Isotope	Percentage abundance
Sr-84	0.56
Sr-86	9.90
Sr-87	7.00
Sr-88	82.54

b.i. Define the term *first ionization energy* and state what is meant by the term *periodicity*. [2]

b.ii. State the electron arrangement of argon and explain why the noble gases, helium, neon and argon show the highest first ionization energies for their respective periods. [3]

b.iii. A graph of atomic radius plotted against atomic number shows that the atomic radius decreases across a period. Explain why chlorine has a smaller atomic radius than sodium. [1]

b.iv. Explain why a sulfide ion, $\mathrm{S^{2-}}$, is larger than a chloride ion, $\mathrm{Cl^{-}}$. [1]

b.v. Explain why the melting points of the Group 1 metals $(\mathrm{Li} \rightarrow \mathrm{Cs})$ decrease down the group whereas the melting points of the Group 7 elements $(\mathrm{F} \rightarrow \mathrm{I})$ increase down the group. [3]

Magnesium is a group 2 metal which exists as a number of isotopes and forms many compounds.

a. State the nuclear symbol notation, ${}^{A}_{Z}X$, for magnesium-26. [1]

b. Mass spectroscopic analysis of a sample of magnesium gave the following results: [2]

	% abundance
Mg-24	78.60
Mg-25	10.11
Mg-26	11.29

Calculate the relative atomic mass, A_r, of this sample of magnesium to two decimal places.

c. Magnesium burns in air to form a white compound, magnesium oxide. Formulate an equation for the reaction of magnesium oxide with water. [1]

d. Describe the trend in acid-base properties of the oxides of period 3, sodium to chlorine. [2]

e. In addition to magnesium oxide, magnesium forms another compound when burned in air. Suggest the formula of this compound [1]

f. Describe the structure and bonding in solid magnesium oxide. [2]

g. Magnesium chloride can be electrolysed. [2]

Deduce the half-equations for the reactions at each electrode when **molten** magnesium chloride is electrolysed, showing the state symbols of the products. The melting points of magnesium and magnesium chloride are 922 K and 987 K respectively.

Anode (positive electrode):

Cathode (negative electrode):

There are many oxides of silver with the formula Ag_xO_y. All of them decompose into their elements when heated strongly.

a.i. After heating 3.760 g of a silver oxide 3.275 g of silver remained. Determine the empirical formula of Ag_xO_y. [2]

a.ii. Suggest why the final mass of solid obtained by heating 3.760 g of Ag_xO_y may be greater than 3.275 g giving one design improvement for your proposed suggestion. Ignore any possible errors in the weighing procedure. [2]

b. Naturally occurring silver is composed of two stable isotopes, ^{107}Ag and ^{109}Ag. [1]

The relative atomic mass of silver is 107.87. Show that isotope ^{107}Ag is more abundant.

c.i. Some oxides of period 3, such as Na_2O and P_4O_{10}, react with water. A spatula measure of each oxide was added to a separate 100 cm^3 flask containing distilled water and a few drops of bromothymol blue indicator. [3]

The indicator is listed in section 22 of the data booklet.

Deduce the colour of the resulting solution and the chemical formula of the product formed after reaction with water for each oxide.

Flask containing	Colour of solution	Product formula
Na_2O	. . . .	. . . .
P_4O_{10}	. . . .	. . . .

c.ii. Explain the electrical conductivity of molten Na_2O and P_4O_{10}. [2]

d. Outline the model of electron configuration deduced from the hydrogen line emission spectrum (Bohr's model). [2]

Titanium is a transition metal.

$TiCl_4$ reacts with water and the resulting titanium(IV) oxide can be used as a smoke screen.

a. Describe the bonding in metals. [2]

b. Titanium exists as several isotopes. The mass spectrum of a sample of titanium gave the following data: [2]

Mass number	% abundance
46	7.98
47	7.32
48	73.99
49	5.46
50	5.25

Calculate the relative atomic mass of titanium to two decimal places.

c. State the number of protons, neutrons and electrons in the $^{48}_{22}\mathrm{Ti}$ atom. [1]

Protons:

- -

Neutrons:

- -

Electrons:

- -

d.i. State the full electron configuration of the $^{48}_{22}\mathrm{Ti}^{2+}$ ion. [1]

d.ii. Explain why an aluminium-titanium alloy is harder than pure aluminium. [2]

e.i. State the type of bonding in potassium chloride which melts at 1043 K. [1]

e.ii. A chloride of titanium, $TiCl_4$, melts at 248 K. Suggest why the melting point is so much lower than that of KCl. [1]

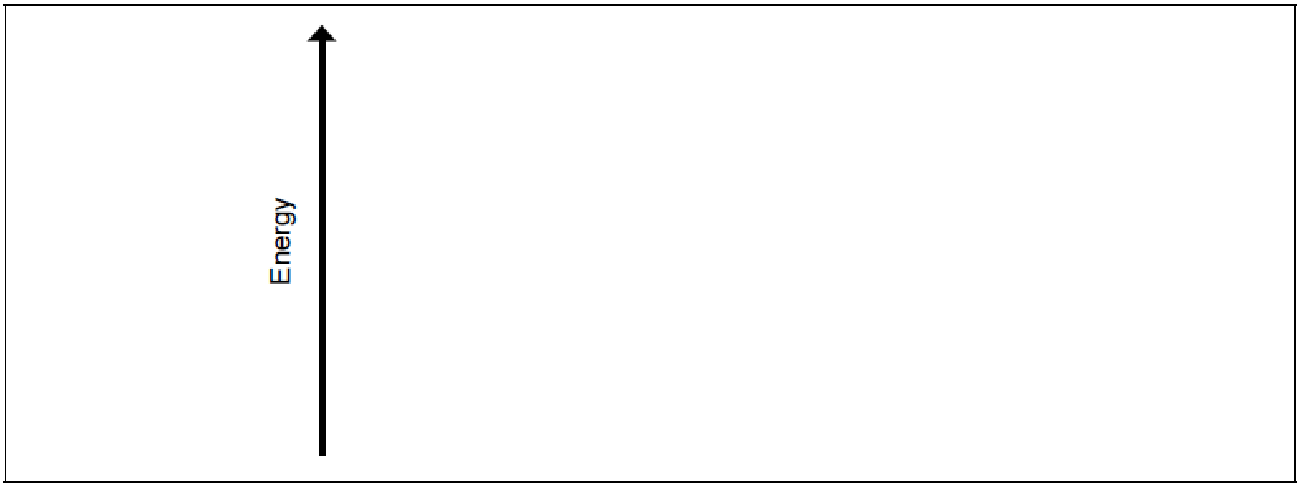
f.i. Formulate an equation for this reaction. [2]

f.ii. Suggest **one** disadvantage of using this smoke in an enclosed space. [1]

The emission spectrum of an element can be used to identify it.

Elements show trends in their physical properties across the periodic table.

a.i. Draw the first four energy levels of a hydrogen atom on the axis, labelling n = 1, 2, 3 and 4. [1]



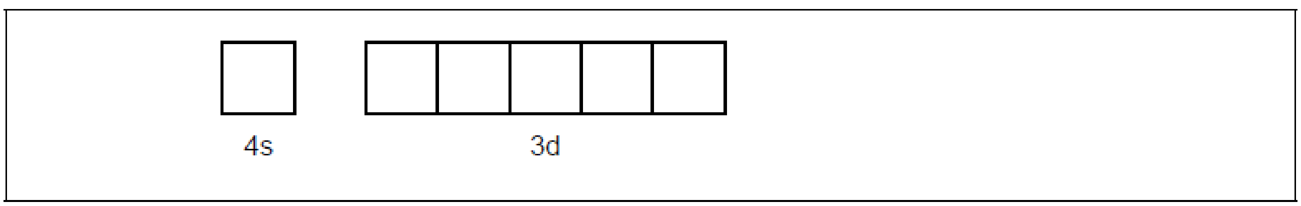
a.ii. Draw the lines, on your diagram, that represent the electron transitions to n = 2 in the emission spectrum. [1]

b.i. Outline why atomic radius decreases across period 3, sodium to chlorine. [1]

b.ii. Outline why the ionic radius of K^+ is smaller than that of Cl^-. [2]

c.i. Copper is widely used as an electrical conductor. [1]

Draw arrows in the boxes to represent the electronic configuration of copper in the 4s and 3d orbitals.



c.ii. Impure copper can be purified by electrolysis. In the electrolytic cell, impure copper is the anode (positive electrode), pure copper is the cathode (negative electrode) and the electrolyte is copper(II) sulfate solution. [2]

Formulate the half-equation at each electrode.

Anode (positive electrode):

. . . .

Cathode (negative electrode):

. . . .

c.iii Outline where and in which direction the electrons flow during electrolysis. [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 (Li → Cs) decrease down the group. [2]

c. State an equation for the reaction of phosphorus (V) oxide, P_4O_{10} (s), with water. [1]

d. Describe the emission spectrum of hydrogen. [2]

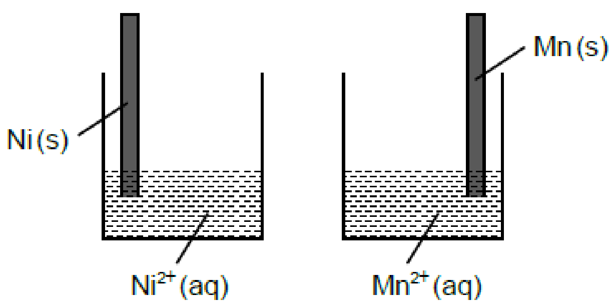
e.i. Identify the strongest reducing agent in the given list. [1]

e.ii. A voltaic cell is made up of a Mn^{2+}/Mn half-cell and a Ni^{2+}/Ni half-cell. [1]

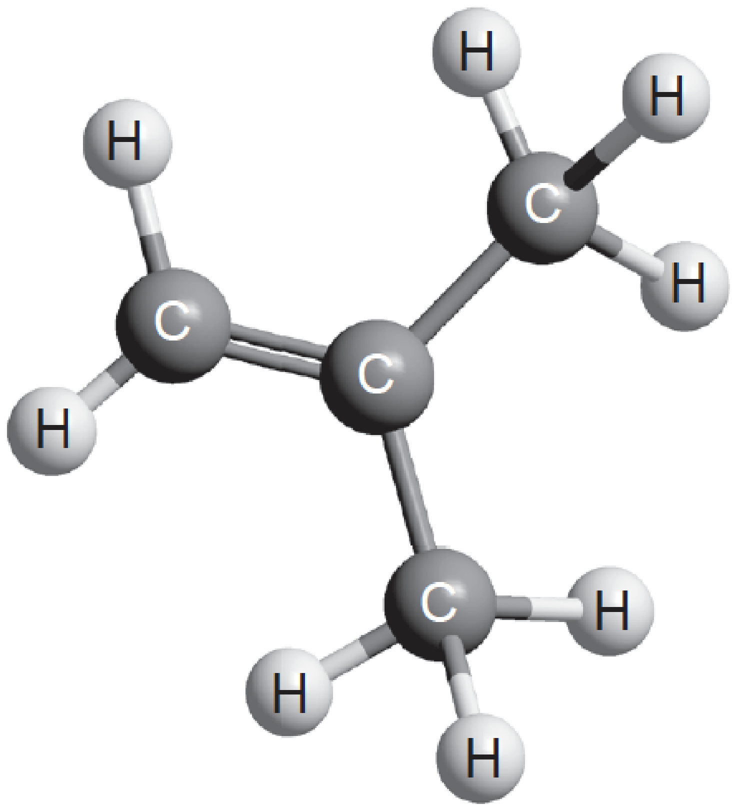
Deduce the equation for the cell reaction.

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.



Alkenes are widely used in the production of polymers. The compound **A**, shown below, is used in the manufacture of synthetic rubber.



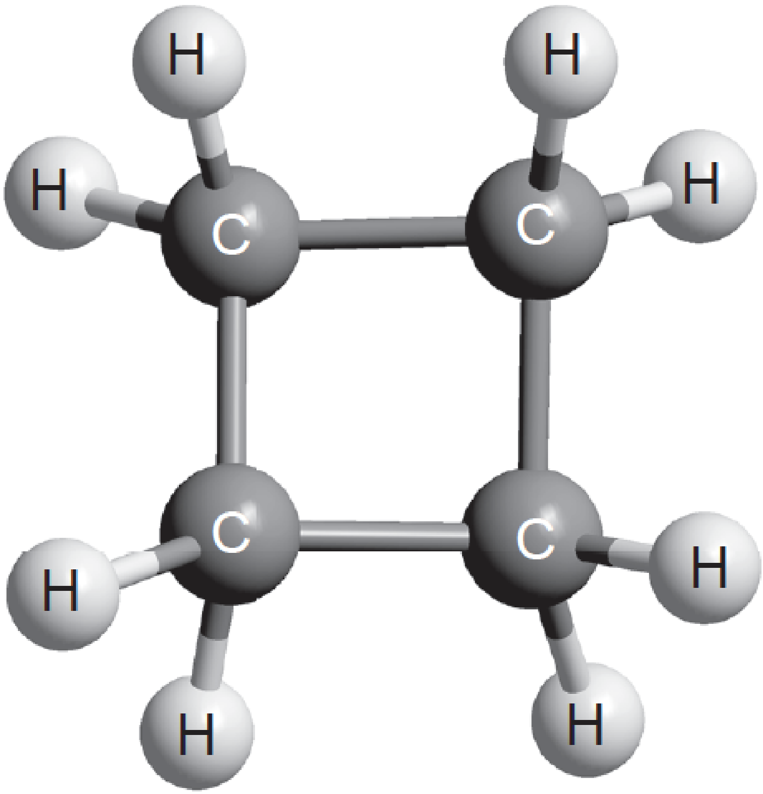
A

a. (i) State the name, applying IUPAC rules, of compound **A**. [3]

(ii) Draw a section, showing three repeating units, of the polymer that can be formed from compound **A**.

(iii) Compound **A** is flammable. Formulate the equation for its complete combustion.

b. Compound **B** is related to compound **A**. [5]



B

(i) State the term that is used to describe molecules that are related to each other in the same way as compound **A** and compound **B**.

(ii) Suggest a chemical test to distinguish between compound **A** and compound **B**, giving the observation you would expect for each.

Test:

Observation with **A**:

Observation with **B**:

(iii) Spectroscopic methods could also be used to distinguish between compounds **A** and **B**.

Predict one difference in the IR spectra **and** one difference in the ^{1}H NMR spectra of these compounds, using sections 26 and 27 of the data booklet.

IR spectra:

^{1}H NMR spectra:

c. A sample of compound **A** was prepared in which the ^{12}C in the CH_2 group was replaced by ^{13}C. [3]

(i) State the main difference between the mass spectrum of this sample and that of normal compound **A**.

(ii) State the structure of the nucleus and the orbital diagram of ^{13}C in its ground state.

No. protons . . . . No. neutrons . . . .

Orbital diagram

1s 2s 2p

d. Draw a 1s atomic orbital and a 2p atomic orbital. [1]

1s:

2p: