5th Edition

AS CHEMISTRY P2

Questions with Mark Scheme

From 2014 to 2022

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CONTENT
A LEVEL CHEMISTRY TOPICAL P2

Physical Chemistry

Topic 1

Atomic Structure

1 **9701/22/O/N/14/Q1**

(a) Successive ionisation energies for the elements magnesium to barium are given in the table.

element	1st ionisation energy / kJ mol^{-1}	2nd ionisation energy / kJ mol^{-1}	3rd ionisation energy / kJ mol^{-1}
Mg	736	1450	7740
Ca	590	1150	4940
Sr	548	1060	4120
Ba	502	966	3390

(i) Explain why the first ionisation energies decrease down the group.

....

....

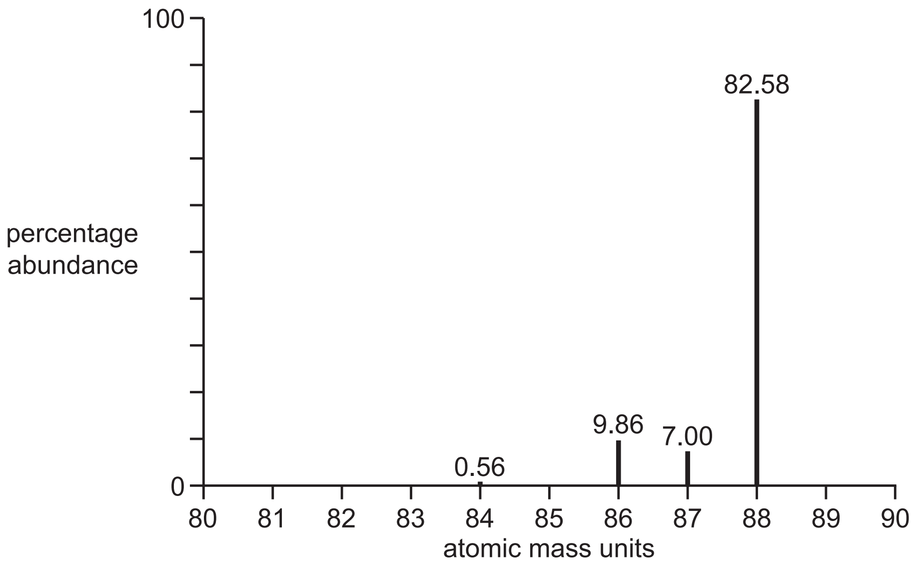
.... [3]

(ii) Explain why, for each element, there is a large increase between the 2nd and 3rd ionisation energies.

....

.... [2]

(b) A sample of strontium, atomic number 38, gave the mass spectrum shown. The percentage abundances are given above each peak.



(i) Complete the full electronic configuration of strontium.

1s^2 2s^2 2p^6 .... [1]

(ii) Explain why there are four different peaks in the mass spectrum of strontium.

.... [1]

(c) A compound of barium, **A**, is used in fireworks as an oxidising agent and to produce a green colour. Explain, in terms of electron transfer, what is meant by the term *oxidising agent*.

.... [1]

2 **9701/21/M/J/14/Q1**

(a) Explain what is meant by the term *ionisation energy*.

....

....

.... [3]

(b) The first seven ionisation energies of an element, **A**, in $kJ\,mol^{-1}$, are

1012 1903 2912 4957 6274 21269 25398.

(i) State the group of the Periodic Table to which **A** is most likely to belong. Explain your answer.

....

....

.... [2]

(ii) Complete the electronic configuration of the element in Period 2 that is in the same group as **A**.

$1s^2$ .... [1]

3 **9701/21/M/J/16/Q1**

(a) Complete the table to show the composition and identity of some atoms and ions.

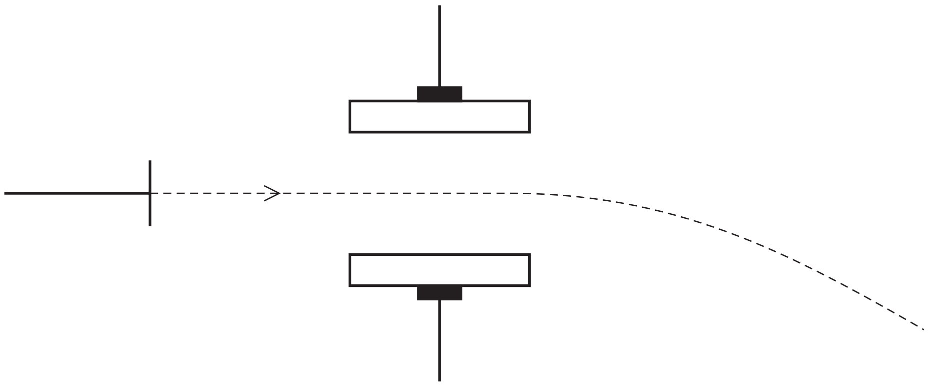
name of element	nucleon number	atomic number	number of protons	number of neutrons	number of electrons	overall charge
lithium	6	3	....	....	....	+1
oxygen	....	....	....	9	10	....
....	54	26	26	....	24	....
....	....	....	17	18	....	0

[4]

(b) Beams of protons, neutrons and electrons behave differently in an electric field due to their differing properties.

The diagram shows the path of a beam of electrons in an electric field.

Add and label lines to represent the paths of beams of protons and neutrons in the same field.



[3]

4 **9701/22/M/J/16/Q1**

(a) Complete the table to show the composition and identity of some atoms and ions.

name of element	nucleon number	atomic number	number of protons	number of neutrons	number of electrons	overall charge
boron	10	5	....	....	....	0
nitrogen	....	....	....	8	10	....
....	208	82	82	....	80	....
....	....	....	3	3	....	+1

[4]

(b) The fifth to eighth ionisation energies of three elements in the third period of the Periodic Table are given. The symbols used for reference are **not** the actual symbols of the elements.

	ionisation energies, $kJ\,mol^{-1}$			
	fifth	sixth	seventh	eighth
X	7012	8496	27 107	31 671
Y	6542	9362	11 018	33 606
Z	7238	8781	11 996	13 842

(i) State and explain the group number of element **Y**.

group number ....

explanation ....

....[1]

(ii) State and explain the general trend in **first** ionisation energies across the third period.

.... [2]

(iii) Complete the electronic configuration of element **X**.

$1s^2$ .... [1]

(c) Complete the electronic configuration of aluminium and of titanium, proton number 22.

Al	$1s^2$
Ti	$1s^2$

[1]

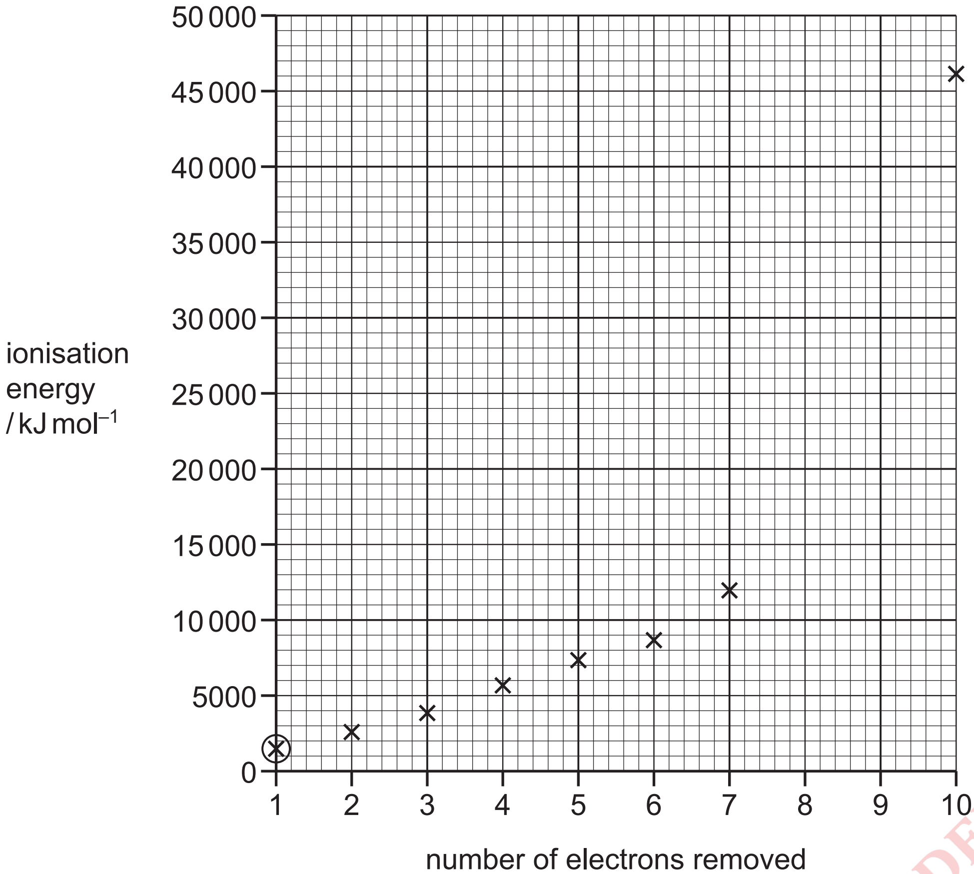
5 9701/21/M/J/19/Q3

(a) Construct an equation for the **second** ionisation energy of argon.

.... [1]

(b) The graph shows successive ionisation energies for the element argon.

Complete the graph with predictions for the eighth and ninth ionisation energies of argon. Use a cross (×) for each data point. [2]



(c) The energy value required to remove the first electron from an atom of argon is circled on the graph.

Sketch the shape of the orbital that contains this electron.

[1]

6 9701/21/O/N/20/Q1(c)

(c) The second ionisation energy of Be is 1757 kJ mol^{-1}.

Explain why the second ionisation energy of Be is higher than the first ionisation energy of Be.

....

....

....

....

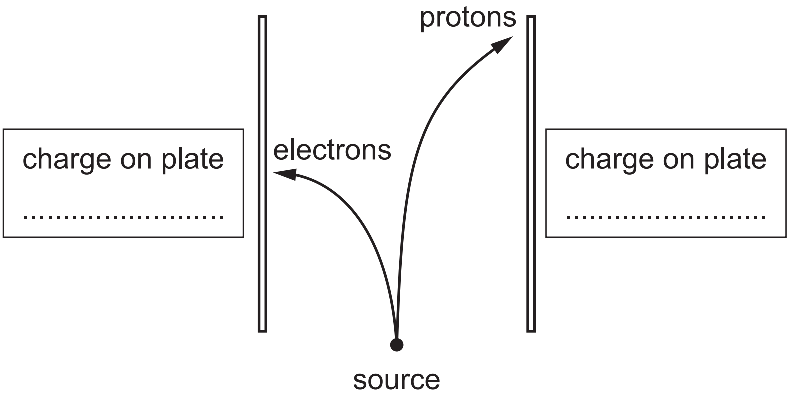
.... [2]

---

7 9701/22/O/N/20/Q1

Atoms contain the subatomic particles electrons, protons and neutrons. Protons and electrons were discovered by observations of their behaviours in electric fields.

(a) The diagram shows the behaviour of separate beams of electrons and protons in an electric field.



(i) Complete the diagram with the relative charge of each of the electrically charged plates. [1]

(ii) On the diagram, draw a line to show how a separate beam of neutrons from the same source behaves in the same electric field. [1]

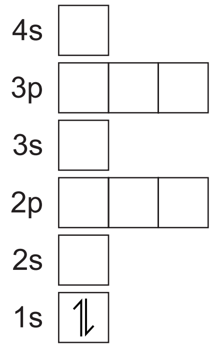
(b) Electrons in atoms up to $_{36}$Kr are distributed in s, p and d orbitals.

(i) State the number of occupied orbitals in an isolated atom of $_{36}$Kr.

type of orbital	s	p	d
number of orbitals			

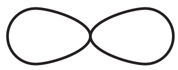
[3]

(ii) Complete the diagram to show the number and relative energies of the electrons in an isolated atom of $_{14}Si$.



[2]

(iii) The diagram shows a type of orbital.



State the total number of electrons that exist in all orbitals of this type in an atom of $_{9}F$.

.... [1]

(iv) The first ionisation energies of elements in the first row of the d block ($_{21}Sc$ to $_{29}Cu$) are very similar. For all these elements, it is a 4s electron that is lost during the first ionisation.

Suggest why the first ionisation energies of these elements are very similar.

....

....

....

.... [3]

(c) *Hydron* is a general term used to represent the ions $^{1}_{1}H^{+}$, $^{2}_{1}H^{+}$ and $^{3}_{1}H^{+}$.

State, in terms of subatomic particles in the nucleus, what is the same about each of these ions and what is different.

same ....

different ....

[1]

8 9701/22/F/M/22/Q1

Fig. 1.1 shows how **first** ionisation energies vary across Period 2.

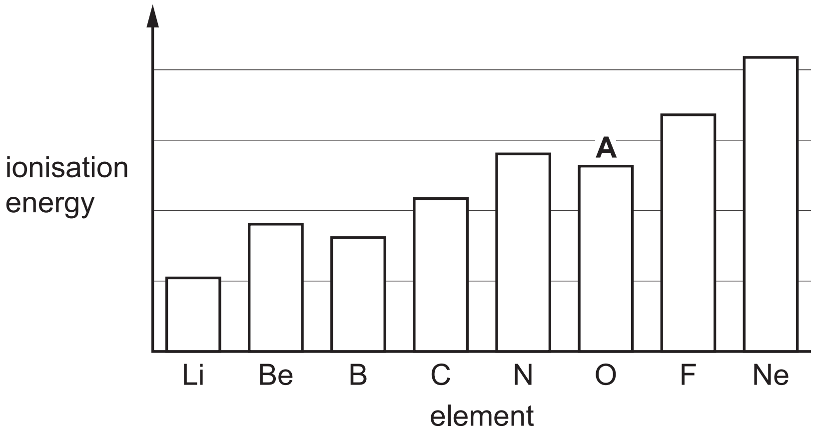


Fig. 1.1

(a) Construct an equation to represent the **first** ionisation energy of oxygen.
Include state symbols.

.... [1]

(b) (i) State and explain the general trend in first ionisation energies across Period 2.

....

....

....

....

.... [3]

(ii) Explain why ionisation energy **A** in Fig. 1.1 does **not** follow the general trend in first ionisation energies across Period 2.

....

....

....

.... [2]

(c) Element **E** is in Period 3 of the Periodic Table.
The first eight ionisation energy values of **E** are shown in Table 1.1.

Table 1.1

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy / $kJ\,mol^{-1}$	577	1820	2740	11600	14800	18400	23400	27500

Deduce the full electronic configuration of **E**.
Explain your answer.

full electronic configuration of **E** = ...

explanation ....

....

.... [3]

9 9701/21/M/J/22/Q1(a)

Calcium, magnesium and radium are Group 2 elements. Radium follows the same trends as the other members of Group 2.

(a) Identify the highest energy orbital which contains electrons in a calcium atom. Sketch the shape of this orbital.

identity of highest energy orbital in Ca ....

shape

[1]

10 9701/22/M/J/22/Q4(b-iii)

(b) Halothane is an anaesthetic.

(iii) Complete Fig. 4.2 to show the arrangement of electrons in a bromine atom using the electrons in boxes notation.

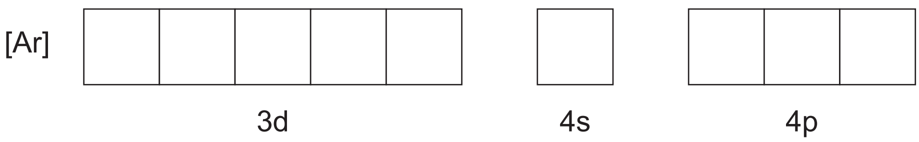


Fig. 4.2

[1]

1 **(a)** **(i)** increasing **distance** of (outer) electron(s) from nucleus 1
OR increasing distance of outer / valence shell from nucleus

increased **shielding** / screening (from inner shells) 1

reduces **attraction** 1 [3]

(ii) (3^{rd} electron for each in) inner / lower energy level / **shell** / closer to nucleus (than first two) / less shielding 1

(large) increase in nuclear attraction 1 [2]

(b) **(i)** ($1s^2$ $2s^2$ $2p^6$) $3s^2$ $3p^6$ $3d^{10}$ $4s^2$ $4p^6$ $5s^2$ 1 [1]

(ii) four isotopes owtte 1 [1]

(c) **(i)** (a species that) gains / takes electron(s) 1 [1]

---

2 **(a)** The amount of energy required / energy change / enthalpy change when one electron is removed
from each atom / (cat)ion in one mol 1
of gaseous atoms / (cat)ions 1
OR energy change when 1 mole of electrons is removed from one mole of gaseous atoms / ions
$X(g) \rightarrow X^+(g) + e^-$ gains 2 marks 3

(b) **(i)** Group V / 5 / 15 1
Big difference between fifth and sixth ionisation energies 1 2

(ii) $1s^2$ $2s^2$ $2p^3$ 1 1
ecf from **(b)(i)** if period 2

---

3 **(a)**

name of element	nucleon no.	atomic no.	no. of protons	no. of neutrons	no. of electrons	overall charge
lithium	6	3	3	3	2	+1
oxygen	17	8	8	9	10	–2
iron	54	26	26	28	24	+2
chlorine	35	17	17	18	17	0

line straight on labelled ‘neutrons’
line (curving) up labelled ‘protons’
proton line clearly shows less (overall) deflection than electron curve

(b) **(i)** Group 16 / 6 / VI
AND
Big (owtte) increase / big difference / big gap / big jump / jump in increase / jump in difference after 6th IE

(ii) increases (across period) due to increasing attraction (of nucleus for electrons)

due to increasing nuclear charge / atomic / proton number AND
constant/similar shielding / same (outer / number of) shell / energy level

(iii) electron (pair) repulsion
(**Y** has a) pair of electrons in a (3)<u>p orbital</u> / a (3)<u>p orbital</u> is full ORA

(iv) ($1s^2$)$2s^2 2p^6 3s^2 3p^5$

---

4 **(a)**

name of element	**nucleon number**	**atomic number**	**number of protons**	**number of neutrons**	**number of electrons**	**overall charge**
boron	10	5	5	5	5	0
nitrogen	15	7	7	8	10	–3
lead	208	82	82	126	80	+2
lithium	6	3	3	3	2	+1

(b) **(i)** Group 17 / VII / 7
AND big (owtte) increase / big difference / big gap / big jump / jump in increase / jump in difference after 7th IE

(ii) increases across period due to increasing attraction (of nucleus for electrons)
due to increasing nuclear charge / atomic / proton number AND constant / similar shielding / same (outer) shell / energy level

(iii) $1s^22s^22p^63s^23p^4$

(c) A*l* $1s^2\ 2s^22p^6\ 3s^23p^1$ (1)

Ti $1s^2\ 2s^22p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$ **or**

$1s^2\ 2s^22p^6\ 3s^2\ 3p^6\ 4s^23d^2$ penalise any error (1) [2]

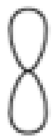
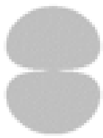
---

5 (a) $Ar^+(g) \rightarrow Ar^{2+}(g) + e^{(-)}$ **OR** $Ar^+(g) - e^{(-)} \rightarrow Ar^{2+}(g)$ **1**

(b) at *x* = 8, within range 13000–20000 **1**

at *x* = 9, within range 35000–45000 **1**

(c) **1**

 **OR** 

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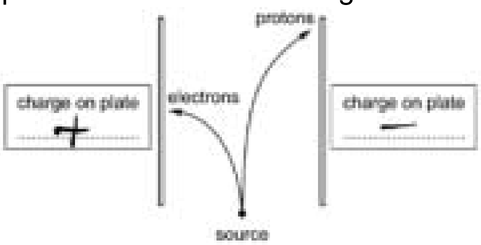
6 (c) **M1**: **greater nuclear attraction** **2**

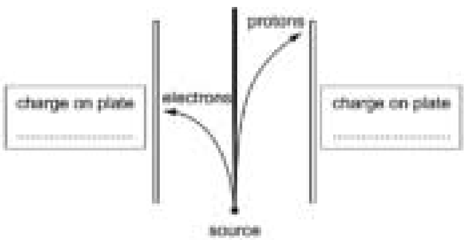
M2: (2nd / 2s) electron being removed from **smaller** (ion)

---

7 (a)(i) positive / + on left **AND** negative / – on right **1**



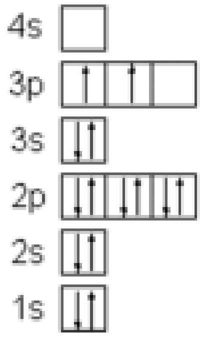
(a)(ii) straight line vertically upwards from the source **1**



(b)(i) **3**

type of orbital	s	p	d
number of orbitals	4	9	5

(b)(ii) **2**



(b)(iii) 5 **1**

(b)(iv) Award one mark for each correct bullet point – max 3 marks **3**

- nuclear charge increases
- extra electron(s) in inner shell / n=3 /d-subshell / d- orbital
- increased shielding (of 4s electrons by electrons in n=3 / 3[rd] shell / 3d)
- (overall) **similar** nuclear attraction (for outer electron)

(c) *answer in terms of **subatomic particles** in the **nucleus*** same (number of) protons **AND** different (number of) neutrons **1**

---

8 (a) $O(g) \rightarrow O^{+}(g) + e^{-}$ **1**

(b)(i) increase across period **AND** increased nuclear attraction **for** (valence / outer) **electrons** [1] **3**

increase in (positive) nuclear charge / number of protons (in the nucleus) [1]
similar shielding (of outer electrons) [1]

(b)(ii) spin-pair repulsion (of electrons) in (2)p <u>orbital [1]</u> **2**
outweighs increased nuclear charge [1]

(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ [1] **3**
greatest jump between 3rd and 4th ionisations [1]
indicates three electrons in outer shell [1]

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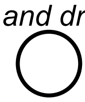
9 (a) *Identify and draw the shape of highest energy orbital of Ca* **1**
4s AND 

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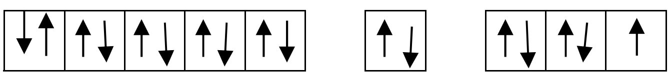
10 (b)(iii) [Ar]  **1**

Physical Chemistry

Topic 2

Atoms, Molecules and Stoichiometry

1 **9701/21/M/J/14/Q2**

The commonest form of iron(II) sulfate is the heptahydrate, $FeSO_4.7H_2O$. On heating at 90 °C this loses **some** of its water of crystallisation to form a different hydrated form of iron(II) sulfate, $FeSO_4.xH_2O$.

3.40 g of $FeSO_4.xH_2O$ was dissolved in water to form 250 cm^3 of solution.

A 25.0 cm^3 sample of this solution was acidified and titrated with 0.0200 mol dm^{-3} potassium manganate(VII).

In this titration 20.0 cm^3 of this potassium manganate(VII) solution was required to react fully with the Fe^{2+} ions present in the sample.

(a) The MnO_4^- ions in the potassium manganate(VII) *oxidise* the Fe^{2+} ions in the acidified solution.

Complete and balance the ionic equation for the reaction between the manganate(VII) ions and the iron(II) ions.

$MnO_4^-(aq)$ + $5Fe^{2+}(aq)$ + ...$H^+(aq)$ → ....(aq) + $5Fe^{3+}(aq)$ + ...$H_2O(l)$

[3]

(b) **(i)** Calculate the number of moles of manganate(VII) used in the titration.

[1]

(ii) Use the equation in **(a)(ii)** and your answer to **(b)(i)** to calculate the number of moles of Fe^{2+} present in the 25.0 cm^3 sample of solution used.

[1]

(iii) Calculate the number of moles of $FeSO_4.xH_2O$ in 3.40 g of the compound.

[1]

(iv) Calculate the relative formula mass of $FeSO_4.xH_2O$.

[1]

(v) The relative formula mass of anhydrous iron(II) sulfate, $FeSO_4$, is 151.8.

Calculate the value of *x* in $FeSO_4.xH_2O$.

[1]

---

2 **9701/23/M/J/14/Q2**

A sample of a hydrated double salt, $Cu(NH_4)_x(SO_4)_2.6H_2O$, was boiled with an excess of sodium hydroxide. Ammonia was given off.

The ammonia produced was absorbed in 40.0 cm^3 of 0.400 mol dm^{-3} hydrochloric acid. The resulting solution required 25 cm^3 of 0.12 mol dm^{-3} sodium hydroxide to neutralise the excess acid.

(a) Write the ionic equation for the reaction between ammonium ions and hydroxide ions.

.... [1]

(b) **(i)** Calculate the amount, in moles, of hydrochloric acid in 40.0 cm^3 of 0.400 mol dm^{-3} solution.

[1]

(ii) Calculate the amount, in moles, of sodium hydroxide needed to neutralise the excess acid. This will be equal to the amount of hydrochloric acid left in excess.

[1]

(iii) Calculate the amount, in moles, of hydrochloric acid that reacted with ammonia.

[1]

(iv) Calculate the amount, in moles, of ammonium ions in the sample of the double salt.

[1]

(v) The sample contained 0.413 g of copper. Use this information and your answer to **(iv)** to calculate the value of *x* in $Cu(NH_4)_x(SO_4)_2.6H_2O$.

[2]

(vi) Calculate the M_r of $Cu(NH_4)_x(SO_4)_2.6H_2O$.

[1]

---

3 **9701/22/O/N/14/Q1(B)(II), (C)(II)**

(a) Explain why there are four different peaks in the mass spectrum of strontium.

....

.... [1]

(b) Calculate the atomic mass, A_r, of this sample of strontium.
Give your answer to **three** significantfigures.

A_r = ... [2]

(c) A compound of barium, **A**, is used in fireworks as an oxidising agent and to produce a green colour.

A has the following percentage composition by mass: Ba, 45.1; C*l*, 23.4; O, 31.5.

Calculate the empirical formula of **A**.

empirical formula of **A** .... [3]

(d) An investigation of the reaction between magnesium oxide and hydrogen chloride gas showed that an intermediate product was formed with the composition by mass Mg, 31.65%; O, 20.84%; H, 1.31% and C*l*, 46.20%.

Calculate the empirical formula of this intermediate compound.

empirical formula .... [2]

4 **9701/21/M/J/15/Q1**

(a) Chemists recognise that atoms are made of three types of particle.

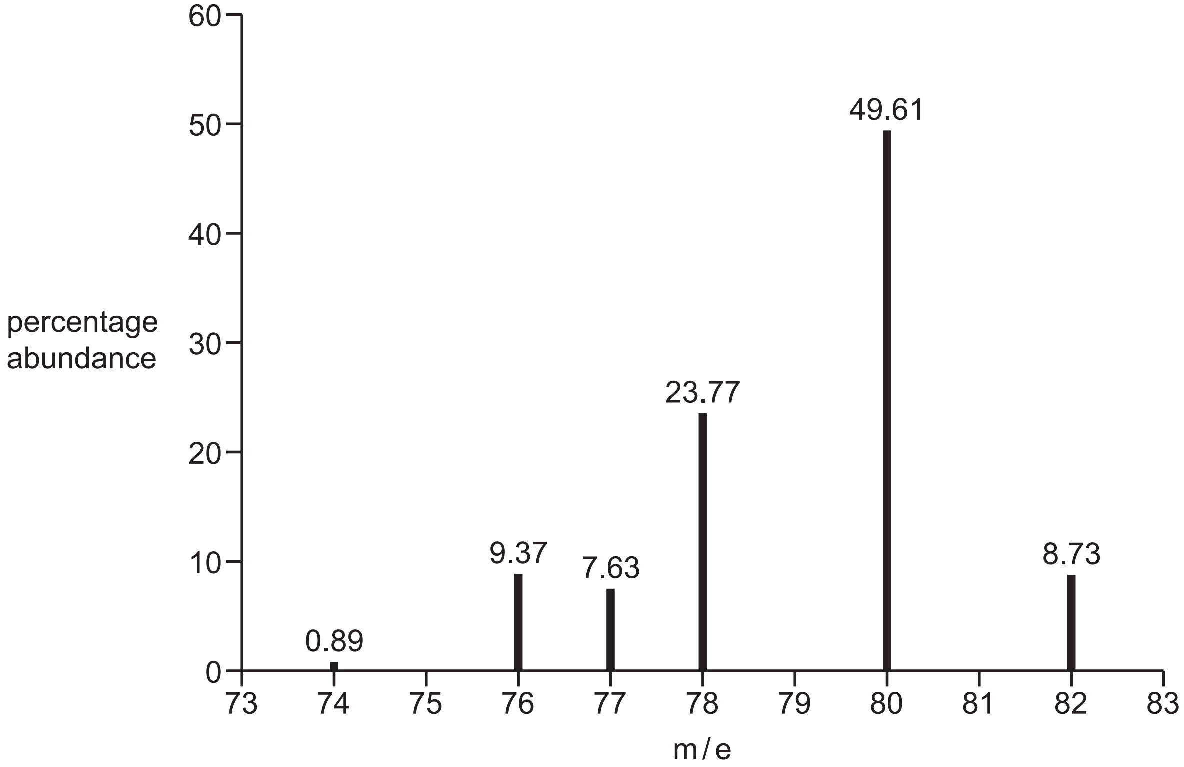
Complete the following table with their names and properties.

name of particle	relative mass	relative charge
		0
	1/1836	

[3]

(b) The relative atomic mass of an element can be determined using data from its mass spectrum.

The mass spectrum of element **X** is shown, with the percentage abundance of each isotope labelled.



(i) Define the terms *relative atomic mass* and *isotope*.

relative atomic mass ....

isotope ....

....

[3]

(ii) Use the data in the mass spectrum to calculate the relative atomic mass, A_r, of **X**. Give your answer to **two** decimal places and suggest the identity of **X**.

A_r of **X** ....

identity of **X** ....

[2]

(c) Antimony, Sb, has been known for about 6000 years. It is present in many ancient forms of bronze, but now its main use is to strengthen lead alloys.

Antimony is produced in a two-stage process from stibnite, a sulphide ore, Sb_2S_3.

The ore is first roasted in oxygen to form the oxide.

$$2Sb_2S_3 + \text{.............} O_2 \longrightarrow Sb_4O_6 + \text{.............}SO_2$$

(i) Balance the above equation.

The oxide is then reduced with carbon.

$$Sb_4O_6 + 3C \longrightarrow 4Sb + 3CO_2$$

(ii) Calculate the volume of carbon dioxide, measured at room temperature and pressure, that would be produced by the processing of 10 moles of Sb_2S_3.

[4]

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5 **9701/22/F/M/16/Q2**

Spathose is an iron ore that contains iron(II) carbonate, $FeCO_3$. The percentage of iron(II) carbonate in spathose can be determined by titration with acidified potassium dichromate(VI) solution using a suitable indicator.

The ionic equation is shown below.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$$

(a) A 5.00 g sample of spathose was reacted with excess concentrated hydrochloric acid and then filtered.
The filtrate was made up to 250 cm^3 in a volumetric flask with distilled water.

A 25.0 cm^3 sample of the standard solution required 27.30 cm^3 of 0.0200 mol dm^{-3} dichromate(VI) solution for complete reaction.

(i) Calculate the amount, in moles, of dichromate(VI) ions used in the titration.

amount = ... mol [1]

(ii) Use your answer to **(i)** to calculate the amount, in moles, of Fe^{2+} present in the 25.0 cm^3 sample.

amount = ... mol [1]

(iii) Use your answer to **(ii)** to calculate the amount, in moles, of Fe^{2+} present in the 250 cm^3 volumetric flask.

amount = ... mol [1]

(iv) Use your answer to **(iii)** to calculate the mass of iron(II) carbonate present in the sample of spathose.

mass = ... g [2]

(v) Calculate the percentage of iron(II) carbonate in the sample of spathose.

percentage of iron(II) carbonate = ... % [1]

(b) Iron ores containing iron(III) compounds can be analysed using a similar method.

A standard solution of an aqueous iron(III) compound is reacted with aqueous tin(II) chloride. Aqueous tin(IV) chloride and aqueous iron(II) chloride are the products of this reaction.

(i) Write an **ionic** equation for this reaction. Do not include state symbols.

.... [2]

(ii) Any excess tin(II) chloride can be removed by reaction with $HgCl_2(aq)$. A white precipitate of Hg_2Cl_2 is produced.

Complete the equation for this reaction.

..........(.....) + ...$HgCl_2(aq) \rightarrow SnCl_4$(.....) + Hg_2Cl_2(.....)

[2]

6 **9701/22/M/J/16/Q3**

Acidified potassium dichromate(VI) can oxidise ethanedioic acid, $H_2C_2O_4$.
The relevant half-equations are shown.

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

$$H_2C_2O_4 \rightarrow 2CO_2 + 2H^+ + 2e^-$$

(a) State the overall equation for the reaction between acidified dichromate(VI) ions and ethanedioic acid.

.... [2]

(b) In an experiment a 0.242 g sample of hydrated ethanedioic acid, $H_2C_2O_4.\boldsymbol{x}H_2O$, was reacted with a 0.0200 mol dm^{-3} solution of acidified potassium dichromate(VI).

32.0 cm^3 of the acidified potassium dichromate(VI) solution was required for complete oxidation of the ethanedioic acid.

(i) Calculate the amount, in moles, of dichromate(VI) ions used to react with the sample of ethanedioic acid.

amount = ... mol [1]

(ii) Calculate the amount, in moles, of ethanedioic acid in the sample.

amount = ... mol [1]

(iii) Calculate the relative molecular mass, M_r, of the hydrated ethanedioic acid.

M_r = ... [1]

(iv) Calculate the value of $\boldsymbol{x}$ in $H_2C_2O_4.\boldsymbol{x}H_2O$.

$\boldsymbol{x}$ = ... [1]

7 9701/23/M/J/16/Q1

An experiment was carried out to determine the percentage of iron in a sample of iron wire.

(a) A 3.35 g piece of the wire was reacted with dilute sulfuric acid, in the absence of air, so that all of the iron atoms were converted to iron(II) ions. The resulting solution was made up to 250 cm^3.

(i) Write a balanced equation for the reaction between the iron in the wire and the sulfuric acid.

.... [1]

A 25.0 cm^3 sample of this solution was acidified and titrated with 0.0250 mol dm^{-3} potassium dichromate(VI). 32.0 cm^3 of the potassium dichromate(VI) solution was required for complete reaction with the iron(II) ions in the sample.

The relevant half-equations are shown.

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

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

(ii) Use the half-equations to write an equation for the reaction between the iron(II) ions and the acidified dichromate(VI) ions.

.... [1]

(iii) Calculate the amount, in moles, of dichromate(VI) ions used in the titration.

amount = ... mol [1]

(iv) Calculate the amount, in moles, of iron(II) ions in the 25.0 cm^3 sample of solution.

amount = ... mol [1]

(v) Calculate the amount, in moles, of iron in the 3.35 g piece of wire.

amount = ... mol [1]

(vi) Calculate the mass of iron in the 3.35 g piece of wire.

mass = ... g [1]

(vii) Calculate the percentage of iron in the iron wire.

percentage = ... % [1]

(b) Some electronegativity values are shown.

element	electronegativity
aluminium	1.5
chlorine	3.0
iron	1.8

(i) Use the data to suggest the nature of the bonding in iron(III) chloride. Explain your answer.

.... [2]

(ii) Suggest an equation for the reaction between iron(III) chloride and water.

.... [1]

8 9701/21/M/J/17/Q1

Combustion data can be used to calculate the empirical formula, molecular formula and relative molecular mass of many organic compounds.

(a) Define the term *relative molecular mass*.

....

.... [2]

(b) **T** is an alcohol, C_xH_yO. A gaseous sample of **T** occupied a volume of 20 cm^3 at 120 °C and 100 kPa.

The sample was completely burned in 200 cm^3 of oxygen (an excess). The final volume, measured under the same conditions as the gaseous sample, was 250 cm^3.

Under these conditions, all water present is vaporised. Removal of the water vapour from the gaseous mixture decreased the volume to 170 cm^3.

Treating the remaining gaseous mixture with concentrated alkali, to absorb carbon dioxide, decreased the volume to 110 cm^3.
The equation for the complete combustion of **T** can be represented as shown.

$$C_xH_yO + zO_2 \rightarrow xCO_2 + \frac{y}{2}H_2O$$

(i) Use the data given to calculate the value of *x*.

x = ... [1]

(ii) Use the data given to calculate the value of *y*.

y = ... [1]

If you were unable to calculate values for *x* and *y* then use *x* = 4 and *y* = 10 for the remaining parts of this question. These are **not** the correct values.

(iii) Complete the equation for the complete combustion of the alcohol, **T**.

.... + ...O_2 $\rightarrow$ ....CO_2 + ...H_2O

[1]

(iv) Give the skeletal formulae for two possible structures of **T**.
Name each alcohol.

....

[2]

(v) Use the general gas equation to calculate the mass of **T** present in the original 20 cm^3 gaseous sample, which was measured at 120 °C and 100 kPa.

Give your answer to **three** significant figures. Show your working.

mass = ... g [3]

9 **9701/22/M/J/17/Q1**

The composition of atoms and ions can be determined from knowledge of atomic number, nucleon number and charge.

(a) Complete the table.

atomic number	nucleon number	number of electrons	number of protons	number of neutrons	symbol
3		2			${}^{6}_{3}Li^{+}$
		23	26	32	

[2]

(b) Boron occurs naturally as a mixture of two stable isotopes, ${}^{10}B$ and ${}^{11}B$. The relative isotopic masses and percentage abundances are shown.

isotope	relative isotopic mass	abundance / %
${}^{10}B$	10.0129	19.78
${}^{11}B$	to be calculated	80.22

(i) Define the term *relative isotopic mass*.

....

.... [2]

(ii) Calculate the relative isotopic mass of ${}^{11}B$.

Give your answer to **six** significant figures. Show your working.

[2]

10 **9701/23/M/J/18/Q1[C(iii)]**

A naturally occurring sample of cerium contains only **four** isotopes. Data for **three** of the isotopes are shown in the table.

isotope	${}^{136}Ce$	${}^{138}Ce$	${}^{140}Ce$	${}^{142}Ce$
relative isotopic mass	135.907	137.906	139.905	to be calculated
percentage abundance	0.185	0.251	88.450	to be calculated

The A_r of the sample is 140.116.

Use these data to calculate the **relative isotopic mass** of the fourth isotope in this sample of cerium.

Give your answer to **three** decimal places.

relative isotopic mass = ... [3]

11 **9701/22/M/J/18/Q2**

Ammonium iron(II) sulfate, $(NH_4)_2Fe(SO_4)_2$, has a relative formula mass, M_r, of 284.

(a) Define the term *relative formula mass*.

....

....

.... [2]

(c) Mohr's salt, $(NH_4)_2Fe(SO_4)_2{\bullet}xH_2O$, is the hydrated form of ammonium iron(II) sulfate.

x represents the number of moles of water in 1 mole of the salt.

A student wanted to determine the value of x. 0.784 g of the hydrated salt was dissolved in water and this solution was acidified.

All of the solution was titrated with 0.0200 mol dm^{-3} potassium manganate(VII). 20.0 cm^3 of this potassium manganate(VII) solution was required for complete reaction with the Fe^{2+} ions.

(i) Use changes in oxidation numbers to balance the equation for the reaction taking place.

$MnO_4^-(aq)$ + ...$Fe^{2+}(aq)$ + ...$H^+(aq)$ → .....$Mn^{2+}(aq)$ + ...$Fe^{3+}(aq)$ + ...$H_2O(l)$ [1]

(ii) State the role of the Fe^{2+} ions in this reaction.

Explain your answer.

....

.... [2]

(iii) Calculate the amount, in moles, of manganate(VII) ions that reacted.

amount = ... mol [1]

(iv) Calculate the amount, in moles, of Fe^{2+} ions in the sample of the salt.

amount = ... mol [1]

(v) Calculate the relative formula mass of $(NH_4)_2Fe(SO_4)_2{\bullet}xH_2O$.

relative formula mass = ... [1]

(vi) Calculate the value of x.

x = ... [1]

---

12 **9701/21/O/N/18/Q1(a,c)**

Iron pyrite, FeS_2, has a yellow colour that makes it look like gold metal. The compound contains the ions Fe^{2+} and S_2^{2-}.

(a) **(i)** Give the full electronic configuration of Fe^{2+}.

$1s^2$ .... [1]

(ii) Calculate the oxidation number of sulfur in the S_2^{2-} ion.
Assume that each sulfur atom in the ion has the same oxidation number.

oxidation number of sulfur in the S_2^{2-} ion = ... [1]

(b) Iron pyrite is often called *fool's gold* because of its appearance. Impure samples of iron pyrite often contain a small amount of gold.

The gold can be obtained from impure iron pyrite. The impure iron pyrite is roasted in oxygen, to produce iron(III) oxide and sulfur dioxide. Gold does not react with oxygen.

(i) The sulfur dioxide produced during roasting would cause environmental consequences if released into the atmosphere.

State and explain **one** of these environmental consequences.

....

.... [2]

(ii) Complete the equation to show the roasting of iron pyrite in oxygen.

$4FeS_2$ + ... → $2Fe_2O_3$ + ...

[2]

(iii) A sample of impure iron pyrite was roasted in oxygen. The composition of the mixture of solid products is shown.

solid product	mass / g
Fe_2O_3	33.18
Au	0.37

Calculate the mass of FeS_2 present in the sample of impure iron pyrite.
Assume that all the FeS_2 was converted to Fe_2O_3 during the roasting process.

(M_r: FeS_2, 120.0; Fe_2O_3, 159.6)

mass of FeS_2 = ... g [2]

(iv) Use your answer to **(iii)** to calculate the percentage by mass of gold in this sample of impure iron pyrite. Assume that gold is the only impurity in this sample of impure iron pyrite.

Give your answer to **two** significant figures.

(If you were unable to calculate an answer to **(iii)**, use 55.00 g as the mass of FeS_2 in this calculation. This is **not** the correct answer.)

percentage by mass of gold = ... % [1]

13 9701/22/O/N/18/Q1(a,c)

The model of the nuclear atom was first proposed by Ernest Rutherford. He developed this model on the basis of results obtained from an experiment using gold metal foil.

(a) Complete the table with information for two of the particles in an atom of ^{197}Au.

particle	relative mass	relative charge	location within atom	total number in an atom of ^{197}Au
electron	0.0005	–1		79
neutron			nucleus	

[4]

(b) A sample of gold found in the earth consists of only one isotope.

(i) Explain what is meant by the term *isotopes*.

....

....

.... [2]

(ii) A different sample of gold contains more than one isotope.

Suggest why this different sample of gold has the same **chemical** properties as the sample found in the earth.

....

.... [1]

---

14 9701/23/M/J/19/Q1(a)

A sample contains three different types of atom: $^{40}_{18}Ar$, $^{40}_{19}K$ and $^{40}_{20}Ca$.

(i) State fully, in terms of the numbers of subatomic particles, what these three atoms have in common.

....

.... [1]

(ii) State fully, in terms of the numbers of **all** subatomic particles, how these three atoms **differ** from each other.

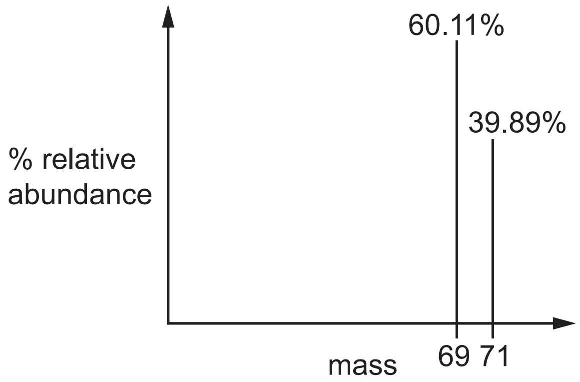
....

.... [1]

15 9701/22/M/J/20/Q1(c)

Gallium is an element in Group 13.

A sample of gallium is analysed using a mass spectrometer. The mass spectrum produced is shown.



Complete the table which describes a gaseous atom of gallium.

isotope	nucleon number	total number of electrons in **lowest** energy level	type of orbital which contains the electron in the **highest** energy level
^{71}Ga			

[3]

16 9701/23/M/J/20/Q2(ciii,d)

(a) (iii) State the electronic configuration of a copper atom.

$1s^2$ ....[1]

(b) A student is provided with a sample of hydrated copper(II) sulfate, $CuSO_4 \cdot xH_2O$, and is asked to determine the value of x.

The student dissolves a sample of the hydrated copper(II) sulfate in water and adds it to an excess of aqueous potassium iodide to make a total volume of 250.0 cm^3 of solution.

$$2CuSO_4 + 4KI \rightarrow 2CuI + I_2 + K_2SO_4$$

The amount of iodine produced during this reaction is found by titrating a sample of this solution with sodium thiosulfate solution.

25.0 cm^3 of the iodine-containing solution requires 20.0 cm^3 of 0.10 $mol\,dm^{-3}$ sodium thiosulfate solution.

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

(i) Calculate the amount, in mol, of copper(II) sulfate present in the original sample of hydrated copper(II) sulfate.

Show your working.

amount of copper(II) sulfate = ... mol [2]

(ii) A total of 7.98 g of $CuSO_4$ is present in 10.68 g of $CuSO_4 \cdot xH_2O$.

Complete each row of the table to calculate the value of x, where x is an integer.

[M_r: $CuSO_4$, 159.6]

amount of $CuSO_4$ in 10.68 g of $CuSO_4 \cdot xH_2O$	.... mol
amount of H_2O in 10.68 g of $CuSO_4 \cdot xH_2O$	.... mol
value of x	x = ...

[3]

17 9701/22/F/M/20/Q3(b)

(b) Glycerol can be used as a starting material in the manufacture of nitroglycerine, $C_3H_5N_3O_9$.

Nitroglycerine decomposes rapidly on heating to form a mixture of gases.

$$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$$

A sample of nitroglycerine decomposes, releasing 1.06 dm^3 of $O_2(g)$ at 850 K and 1.00×10^5 Pa.

(i) Calculate the mass of nitroglycerine that decomposes.

mass of nitroglycerine = ... g [3]

(ii) Calculate the total volume of gas released by this decomposition at 850 K and 1.00×10^5 Pa.

total volume of gas = ... dm^3 [1]

18 9701/21/M/J/20/Q2

(a) The equation shown in **(a)(i)** describes the reaction which occurs when aqueous potassium iodide is added to aqueous copper(II) sulfate. A white precipitate of copper(I) iodide forms in a brown solution of iodine and potassium sulfate.

(i) Balance the equation and include state symbols.

......$CuSO_4$(.....) + ...KI(.....) $\rightarrow$ ......CuI(.....) + ...I_2(.....) + ...K_2SO_4(.....)

[2]

The table gives the oxidation numbers of iodine in the different species in the equation.

iodine-containing species	oxidation number of iodine
KI	–1
CuI	–1
I_2	0

(ii) Deduce the oxidation number of copper in $CuSO_4$ and CuI.

- oxidation number of copper in $CuSO_4$ ....
- oxidation number of copper in CuI ....

[1]

(iii) Describe the type of reaction shown by the equation in **(a)(i)**. Explain your answer in terms of electron transfer.

....

....

.... [2]

(b) In the reaction described in **(a)(i)**, a student uses 17.43 g of $CuSO_4$•**y**H_2O. By further titration of the reaction products the student concludes that the total amount of $CuSO_4$ in the sample is 0.0982 mol.

Use the *Data Booklet* to complete the table to calculate the value of **y**, where **y** is an integer. Show your working.

mass of 0.0982 mol $CuSO_4$	.... g
amount of H_2O in 17.43 g of $CuSO_4$•**y**H_2O	.... mol H_2O
value of **y**	**y** = ...

[4]