# PERIODIC TRENDS WS 1

## SECTION A

**1** Which species represented by the following formulae has the largest radius?

**A** $P^{3-}$ **B** $Cl^-$ **C** Ar **D** $K^+$

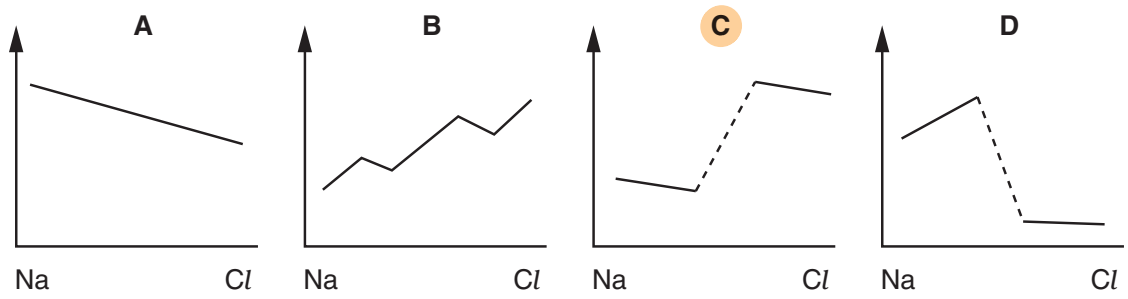
**2** *Use of the Data Booklet is relevant to this question.*

In the gas phase, aluminium and a transition element require the same amount of energy to form one mole of an ion with a 2+ charge.

What is the transition element?

**A** Co
**B** Cr
**C** Cu
**D** Ni

**3** Which diagram represents the change in ionic radius of the elements across the third period (Na to Cl)?



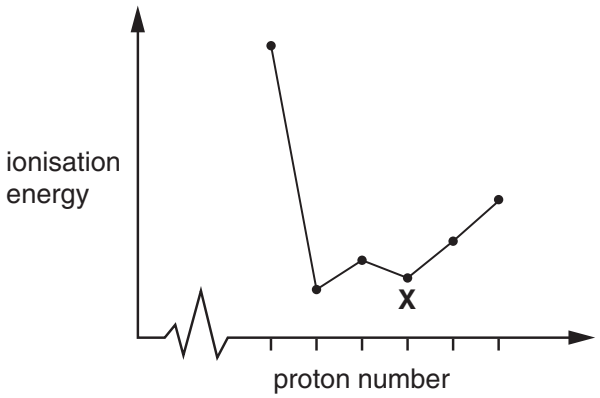

**4** What is the electronic configuration of an element with a **second** ionisation energy higher than that of each of its neighbours in the Periodic Table?

**A** $1s^2 2s^2 2p^6 3s^2$
**B** $1s^2 2s^2 2p^6 3s^2 3p^1$
**C** $1s^2 2s^2 2p^6 3s^2 3p^2$
**D** $1s^2 2s^2 2p^6 3s^2 3p^3$

**5** In which pair is the radius of the second atom greater than that of the first atom?

**A** Na, Mg **B** Sr, Ca **C** P, N **D** Cl, Br

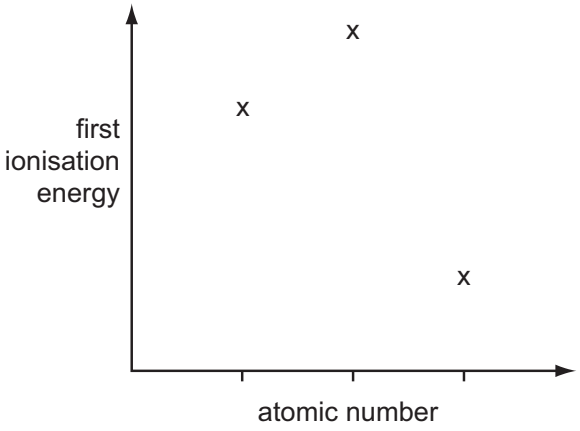
**6** The sketch below shows the variation of first ionisation energy with proton number for six elements of consecutive proton numbers between 1 and 18 (H to Ar).




What is the identity of the element **X**?

**A** Mg **B** Al **C** Si **D** P

**7** Three successive elements in the Periodic Table have first ionisation energies which have the pattern shown in the diagram.




What could be the first element of this sequence?

**A** C **B** N **C** F **D** Na

[W'06 Q2]

**8** Which group of particles is in order of increasing size?

**A** N O F

**B** $N^{3-}$ $O^{2-}$ $F^{-}$

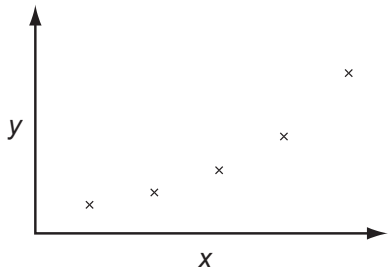
**C** $Na^{+}$ $Mg^{2+}$ $Al^{3+}$

**D** $Na^{+}$ Ne $F^{-}$

[S'09 1 Q12]

**9** *Use of the Data Booklet is relevant to this question.*

The sketch graph shows the variation of one physical or chemical property with another for the Group II elements.




What are the correct labels for the axes?

| | *x*-axis | *y*-axis |
|---|---|---|
| **A** | atomic number | mass number |
| **B** | atomic number | melting point |
| **C** | first ionisation energy | atomic number |
| **D** | first ionisation energy | atomic radius |

[W'07 1 Q14]

**10** Why is the first ionisation energy of phosphorus greater than the first ionisation energy of silicon?

**A** A phosphorus atom has one more proton in its nucleus.

**B** The atomic radius of a phosphorus atom is greater.

**C** The outer electron in a phosphorus atom is more shielded.

**D** The outer electron in a phosphorus atom is paired.

[W'10 1 Q13]

**11** The value of the second ionisation energy of calcium is $1150\,kJ\,mol^{-1}$.

Which equation correctly represents this statement?

**A** $Ca(g) \rightarrow Ca^{2+}(g) + 2e^-$ ; $\Delta H^\ominus = +1150\,kJ\,mol^{-1}$

**B** $Ca^+(g) \rightarrow Ca^{2+}(g) + e^-$ ; $\Delta H^\ominus = +1150\,kJ\,mol^{-1}$

**C** $Ca^+(g) \rightarrow Ca^{2+}(g) + e^-$ ; $\Delta H^\ominus = -1150\,kJ\,mol^{-1}$

**D** $Ca(g) \rightarrow Ca^{2+}(g) + 2e^-$ ; $\Delta H^\ominus = -1150\,kJ\,mol^{-1}$

[S'12 2 Q18]

**12** Sodium and sulfur react together to form sodium sulfide, $Na_2S$.

How do the atomic radius and ionic radius of sodium compare with those of sulfur?

| | atomic radius | ionic radius |
|---|---|---|
| **A** | sodium > sulfur | sodium > sulfur |
| **B** | sodium > sulfur | sodium < sulfur |
| **C** | sodium < sulfur | sodium > sulfur |
| **D** | sodium < sulfur | sodium < sulfur |

[M'1 Q12]

**13** From which particle is the removal of an electron the most difficult?

**A** $Cl^-(g)$ **B** $F^-(g)$ **C** $K^+(g)$ **D** $Na^+(g)$

[W'11 2 Q3]

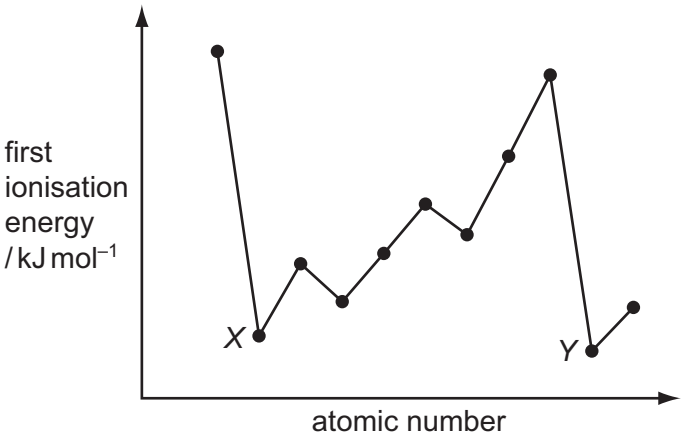
**14** The species Ne, $Na^+$ and $Mg^{2+}$ are isoelectronic. This means that they have the same number of electrons.

In which order do their radii increase?

| | smallest ⟶ | | largest |
|---|---|---|---|
| **A** | Ne | $Na^+$ | $Mg^{2+}$ |
| **B** | Ne | $Mg^{2+}$ | $Na^+$ |
| **C** | $Mg^{2+}$ | Ne | $Na^+$ |
| **D** | $Mg^{2+}$ | $Na^+$ | Ne |

[S'14 3 Q16]

**15** The diagram shows the first ionisation energies of 11 consecutive elements.




Which type of elements are labelled *X* and *Y*?

**A** Group I metals

**B** Group II metals

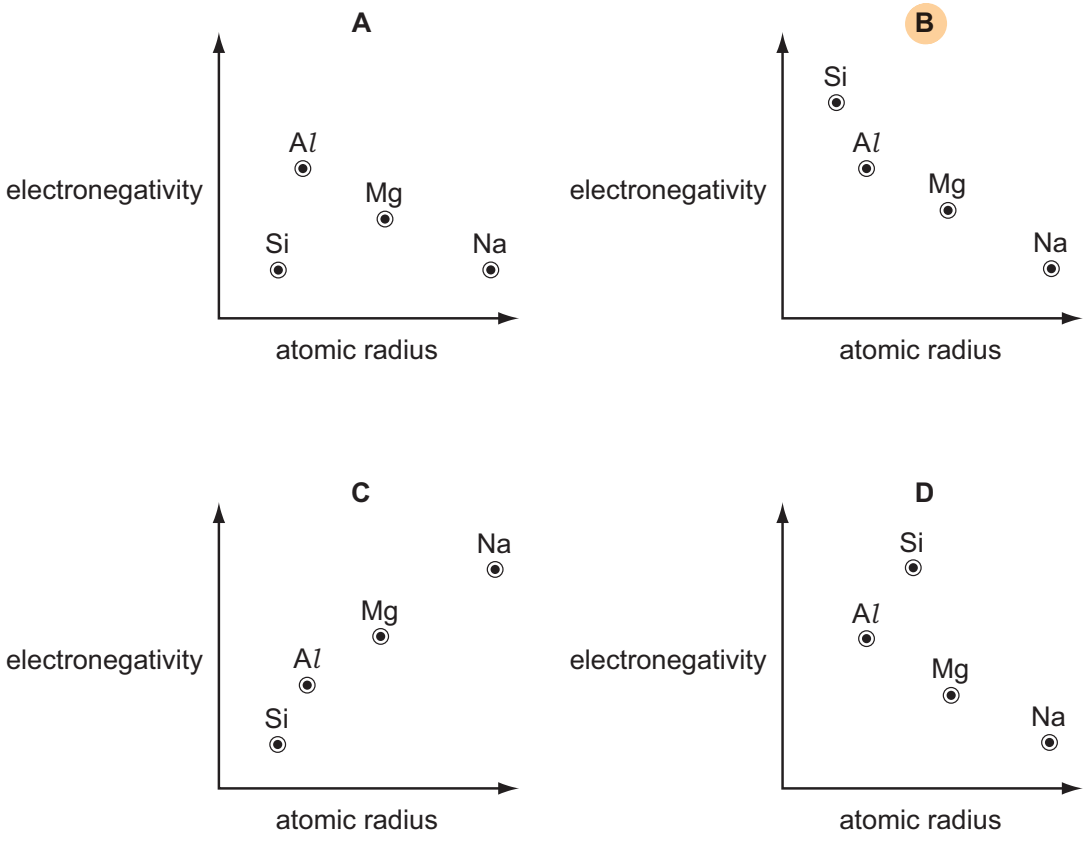
**C** halogens

**D** noble gases

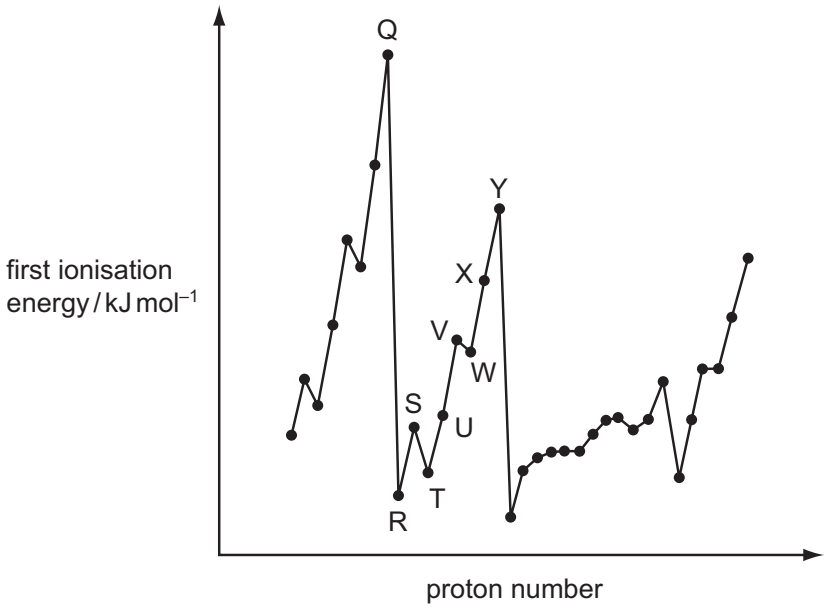
[S'11 12 Q15]

**16** *Use of the Data Booklet is relevant to this question.*

Which graph correctly shows relative electronegativity plotted against relative atomic radius for the elements Na, Mg, Al and Si?



**17** The graph below shows the variation of the first ionisation energy with the number of protons for some elements.




Which statement is correct?

**A** Elements Q and Y are in the same period in the Periodic Table.

**B** The general increase from elements R to Y is due to increasing atomic radius.

**C** The small decrease between elements S and T is due to decreased shielding.

**D** The small decrease between elements V and W is due to repulsion between paired electrons.

[W'13 3 Q18]

**18** Which property **increases** in value going down Group II?

**A** electronegativity

**B** ionic radius

**C** maximum oxidation number

**D** second ionisation energy

[W'13 2 Q14]

**19** Consecutive elements **X**, **Y** and **Z** are in Period 3 of the Periodic Table. Element **Y** has the highest first ionisation energy and the lowest melting point of these three elements.

What are the identities of **X**, **Y** and **Z**?

**A** sodium, magnesium, aluminium

**B** magnesium, aluminium, silicon

**C** aluminium, silicon, phosphorus

**D** silicon, phosphorus, sulfur

[M'16 Q12]

**20** Why is the ionic radius of a chloride ion larger than the ionic radius of a sodium ion?

**A** A chloride ion has one more occupied electron shell than a sodium ion.

**B** Chlorine has a higher proton number than sodium.

**C** Ionic radius increases regularly across the third period.

**D** Sodium is a metal, chlorine is a non-metal.

[W'12 1 Q13]

**21** Why is the ionic radius of a chloride ion larger than the ionic radius of a sodium ion?

**A** A chloride ion has one more occupied electron shell than a sodium ion.

**B** Chlorine has a higher proton number than sodium.

**C** Ionic radius increases regularly across the third period.

**D** Sodium is a metal, chlorine is a non-metal.

[S'12 1 Q13]

**22** Sodium and sulfur react together to form sodium sulfide, $Na_2S$.

How do the atomic radius and ionic radius of sodium compare with those of sulfur?

| | atomic radius | ionic radius |
|---|---|---|
| **A** | sodium < sulfur | sodium > sulfur |
| **B** | sodium < sulfur | sodium < sulfur |
| **C** | sodium > sulfur | sodium > sulfur |
| **D** | sodium > sulfur | sodium < sulfur |

[S'16 2 Q12]

**23** Which element has the **second** smallest atomic radius in its group and the **third** lowest first ionisation energy in its period?

**A** boron

**B** calcium

**C** magnesium

**D** sodium

[S'18 1 Q13]

**24** Element X has a higher first ionisation energy than element Y.

Two students state what they believe is one factor that helps to explain this.

student 1 "X has a higher first ionisation energy than Y because an atom of X has more protons in its nucleus than an atom of Y."

student 2 "X has a higher first ionisation energy than Y because X has a smaller atomic radius than Y."

Only **one** of the two students is correct.

What could X and Y be?

| | X | Y |
|---|---|---|
| **A** | carbon | boron |
| **B** | magnesium | aluminium |
| **C** | oxygen | nitrogen |
| **D** | oxygen | sulfur |

St1 St2 Ans
✓ ✓ X
X X X
X X X
✓ ✓ ✓

[S'18 1 Q10]

## SECTION B

For each of the questions in this section, one or more of the three numbered statements **1** to **3** may be correct.

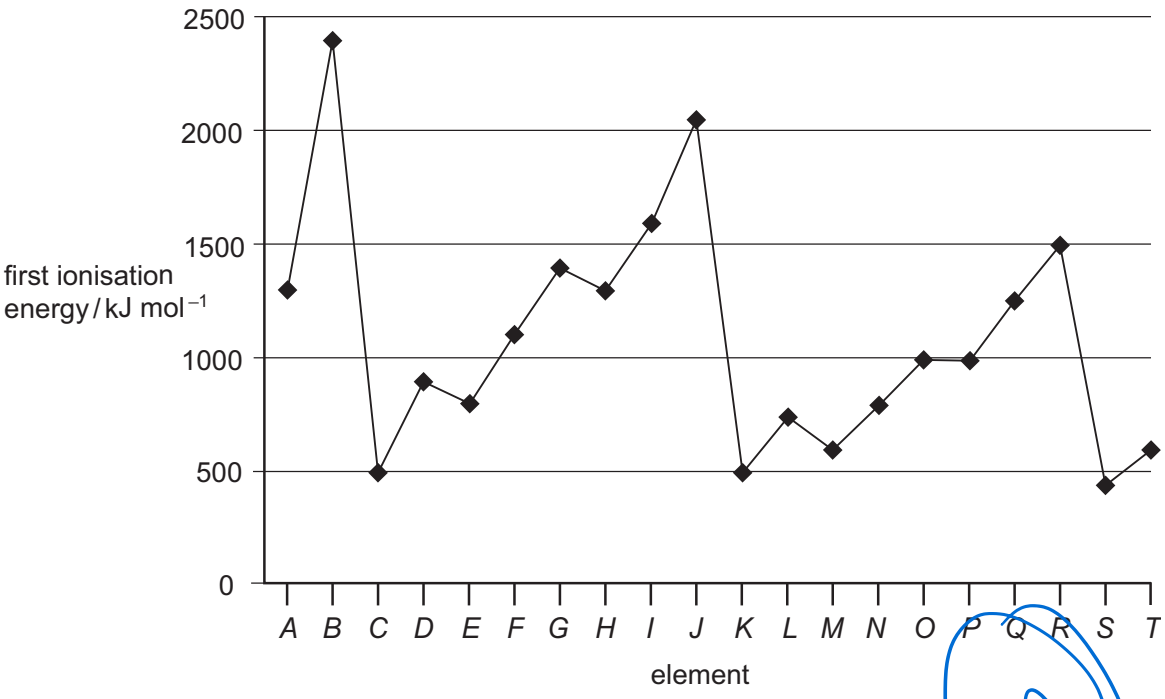
Decide whether each of the statements is or is not correct (you may find it helpful to put a tick against the statements that you consider to be correct).

The responses **A** to **D** should be selected on the basis of

| A | B | C | D |
|---|---|---|---|
| **1, 2** and **3** are correct | **1** and **2** only are correct | **2** and **3** only are correct | **1** only is correct |

No other combination of statements is used as a correct response.

**1** The first ionisation energies of successive elements in the Periodic Table are represented in the graph.




Which of these statements about this graph are correct?

**1** Elements *B*, *J* and *R* are in Group 0 of the Periodic Table.

**2** Atoms of elements *D* and *L* contain 2 electrons in their outer shells.

**3** Atoms of elements *G* and *O* contain half-filled p orbitals.

**2** Which of the following influence the size of the ionisation energy of an atom?

**1** the amount of shielding by the inner electrons

**2** the charge on the nucleus

**3** the distance between the outer electrons and the nucleus

**3** Compound X is made from two elements. One element has the second highest value of first ionisation energy in its group and the other element has the third highest value of first ionisation energy in its group.

Which compounds could be compound X?

**1** calcium chloride

**2** magnesium bromide

**3** potassium sulfide

**4** Why is the first ionisation energy of aluminium less than that of magnesium?

**1** The outer electron in the aluminium atom is more shielded from the nuclear charge.

**2** The outer electron in the aluminium atom is in a higher energy orbital.

**3** The outer electron in the aluminium atom is further from the nucleus.

same as Q3.

**5** Compound X is made from two elements. One element has the second highest value of first ionisation energy in its group and the other element has the third highest value of first ionisation energy in its group.

Which compounds could be compound X?

**1** calcium chloride

**2** magnesium bromide

**3** potassium sulfide

[S'16 3 31]

**6** **X** is an element that has

- its outer electrons in the 4th principal quantum shell,
- a higher 1st ionisation energy than calcium.

What could be the identity of **X**?

**1** bromine

**2** krypton

**3** xenon

# PERIODIC TRENDS WS 2

1 The first six ionisation energies of an element **X** are given below.

| ionisation energy / kJ mol$^{-1}$ | | | | | |
|---|---|---|---|---|---|
| first | second | third | fourth | fifth | sixth |
| 950 | 1800 | 2700 | 4800 | 6000 | 12300 |

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

It is the energy required to remove one electron from each atom in 1 mol of gaseous atoms of an element under standard conditions. [3]

**(b)** Write an equation, with state symbols, for the **second** ionisation energy of element **X**.

$X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-}$ [2]

**(c)** Use the data given above to deduce in which Group of the Periodic Table element **X** is placed. Explain your answer.

Group Group 15.

explanation There is a big jump between the 5$^{th}$ and 6$^{th}$ ionisation energies, indicating there are 5 electrons in the outer most shell. [3]

The first ionisation energies (I.E.) for the elements of Group IV are given below.

| element | C | Si | Ge | Sn | Pb |
|---|---|---|---|---|---|
| 1st I.E. / kJ mol$^{-1}$ | 1090 | 786 | 762 | 707 | 716 |

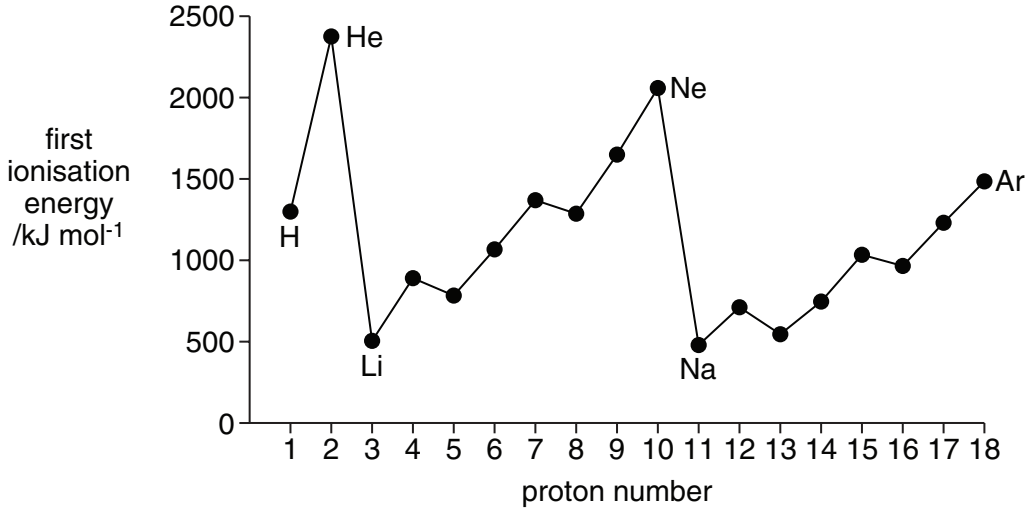
**(d)** Explain the trend shown by these values in terms of the atomic structure of the elements.

There is an increase in size, so the I.E decreases down the group. Number of shells increase down the group. That also leads to more shielding, despite the increase in nuclear charge. [4]

[Total: 12]

2 The Periodic Table we currently use is derived directly from that proposed by Mendeleev in 1869 after he had noticed patterns in the chemical properties of the elements he had studied.

The diagram below shows the first ionisation energies of the first 18 elements of the Periodic Table as we know it today.




**(a)** Give the equation, including state symbols, for the first ionisation energy of fluorine.

$F_{(g)} \rightarrow F^+_{(g)} + e$ [2]

**(b)** Explain why there is a general increase in first ionisation energies from sodium to argon.

Across the period shielding remains the same, radius gets smaller. These factors outweigh the increase in nuclear attraction because of the increase in protons. Hence, ionisation energy increases. [3]

**(c)** **(i)** Explain why the first ionisation energy of aluminium is less than that of magnesium.

The electron in in the p orbital for Al which experiences more shielding from the inner s orbital.

**(ii)** Explain why the first ionisation energy of sulphur is less than that of phosphorus.

Unlike P, S has a paired electrons which repel each other. which makes it easier to remove.

[4]

**3** Magnesium will react on heating with chlorine, or oxygen, or nitrogen to give the chloride, or oxide, or nitride respectively. Each of these compounds is ionic and in them magnesium has the same +2 oxidation state.

**(a)** **(i)** Write an equation, with state symbols, for the **second** ionisation energy of magnesium.

$Mg^{+}_{(g)} \longrightarrow Mg^{2+}_{(g)} + e^{-}$

**(ii)** Use the *Data Booklet* to calculate the enthalpy change that occurs when one mole of gaseous magnesium ions, $Mg^{2+}$, is formed from one mole of gaseous magnesium atoms.

Include a sign in your answer.

$= IE_1 + IE_2$

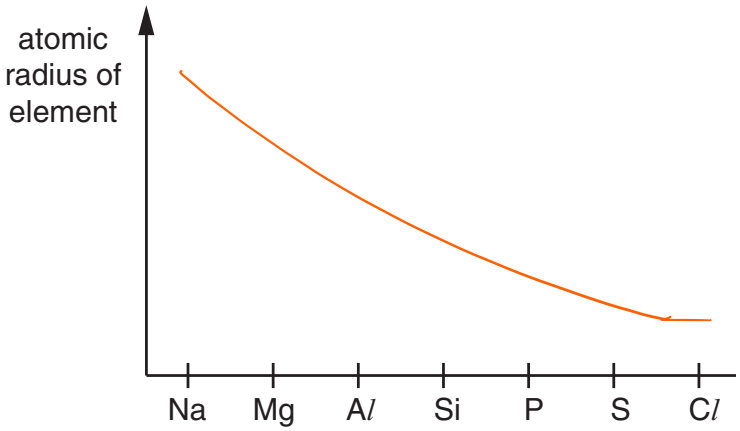
$= 736 + 1450$

$= +2186$

enthalpy change = +2186 kJ $mol^{-1}$

[3]

**4** Elements in the same period of the Periodic Table show trends in physical and chemical properties.
On the grid below, draw a clear sketch to show the variation of the stated property.
Below the grid, briefly explain the variation you have described in your sketch.
You should refer to the important factors that cause the differences in the property you are describing.




explanation Across the period # of protons increases pulling the outer electrons closer in. And as the is no increase in number of shells the radii decreases.

[3]

5 Barium, Ba, was discovered by Davy in 1808. The element gets its name from the Greek 'barys' meaning 'heavy'.

**(a)** The table below compares some properties of barium with caesium.

| element | Cs | Ba |
|---|---|---|
| group | 1 | 2 |
| atomic number | 55 | 56 |
| atomic radius/pm | 531 | 435 |

**(i)** Why do caesium and barium have different atomic numbers?

Because they have different no. of protons [1]

**(ii)** State the block in the Periodic Table in which caesium and barium are found.

S - Block. [1]

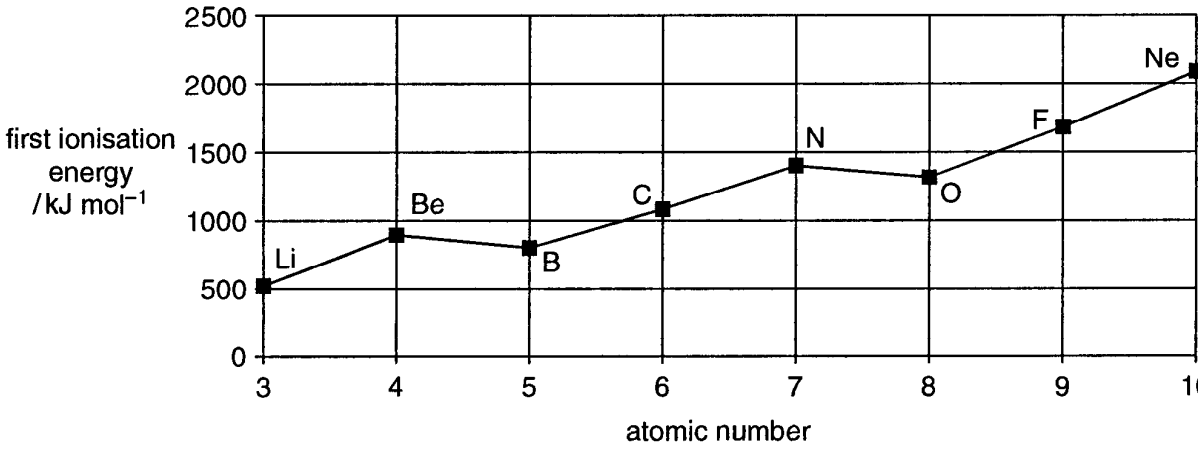
**(iii)** Explain why the atomic radius of barium is **less** than the atomic radius of caesium.

Proton number increases across the group while the shells remain the same. Caesium having greater number of protons pulls on its electrons more, reducing the atomic size. [3]

**(iv)** Predict and explain whether a barium **ion** is *larger*, *smaller* or the *same size* as a barium **atom**.

Smaller as $Ba^{2+}$ loses its outer shell. [2]

6 **(b)** The diagram below shows the variation in the first ionisation energies of elements across Period 2 of the Periodic Table.




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

Energy required to remove one mole of electron from one mole of gaseous atoms under standard conditions. [3]

**(ii)** Explain why the first ionisation energies show a **general** increase across Period 2.

As the orbital level remains the same, there is no change in shielding. However with the increase in number of protons more energy is required to pull away the electrons. [2]

**(iii)** Explain why the first ionisation energy of B is **less** than that of Be.

The outer most electron in 'B' occupies a 2p orbital and experience additional shielding from the 2s orbital. [2]

**(iv)** Estimate a value for the first ionisation energy of the element with atomic number 11. Explain how you made your choice.

First ionisation energy = 470 kJ mol$^{-1}$

Lower than lithium as down the group ionisation energy decreases. [2]

(ii) Decrease in atomic radii due to increase in proton number also contributes to a higher I.E.

7 This question is about the elements in Group II of the Periodic Table, magnesium to barium.

**(a)** Complete the table below to show the electronic configuration of calcium atoms and of strontium ions, $Sr^{2+}$.

| | 1s | 2s | 2p | 3s | 3p | 3d | 4s | 4p | 4d |
|---|---|---|---|---|---|---|---|---|---|
| Ca | 2 | 2 | 6 | 2 | 6 | 0 | 2 | | |
| $Sr^{2+}$ | 2 | 2 | 6 | 2 | 6 | 10 | 2 | 6 | |

[2]

**(b)** Explain the following observations.

**(i)** The atomic radii of Group II elements increase down the Group.

Down the group shells are added increasing the radius.

**(ii)** The strontium ion is smaller than the strontium atom.

when strontium forms a cation it loses its outer shell, hence its decrease in size.

**(iii)** The first ionisation energies of the elements of Group II decrease with increasing proton number.

Despite the increase in nuclear pull, the increase in shielding and radius counter act it and reduce the nuclear pull, decreasing the ionisation energy.

[4]

[S'07 Q3]

8 The alkali metals are a series of six elements in Group I of the Periodic Table. The first ionisation energy of these elements shows a marked trend as the Group is descended.

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

Energy required to remove one mole of electron from one mole of gaseous atoms under standard conditions.

[2]

**(b)** **(i)** State and explain the trend in first ionisation energy as Group I is descended.

Despite the increase in nuclear pull, the increase in shielding and radius counter act it and reduce the nuclear pull, decreasing the ionisation energy.

**(ii)** Suggest how this trend helps to explain the increase in the reactivity of the elements as the Group is descended.

Reactivity increases down group I as its requires less energy to remove electrons.

[3]

**9** Although the actual size of an atom cannot be measured exactly, it is possible to measure the distance between the nuclei of two atoms. For example, the 'covalent radius' of the Cl atom is assumed to be half of the distance between the nuclei in a $Cl_2$ molecule. Similarly, the 'metallic radius' is half of the distance between two metal atoms in the crystal lattice of a metal. These two types of radius are generally known as 'atomic radii'.
The table below contains the resulting atomic radii for the elements of period three of the Periodic Table, Na to Cl.

| element | Na | Mg | Al | Si | P | S | Cl |
|---|---|---|---|---|---|---|---|
| atomic radius / nm | 0.186 | 0.160 | 0.143 | 0.117 | 0.110 | 0.104 | 0.099 |

**(a) (i)** Explain qualitatively this variation in atomic radius.

Nuclear charge increases across the period pulling the outer electrons closer. And so the is no change in shells or shielding the radius decreases across the period.

**(ii)** Suggest why it is not possible to use the same type of measurement for argon, Ar.

Argon doesn't form any bonds.

[4]

**(b) (i)** Use the *Data Booklet* to complete the following table of radii of the cations and anions formed by some of the period three elements.

| radius of cation / nm | | | radius of anion / nm | | |
|---|---|---|---|---|---|
| $Na^+$ | $Mg^{2+}$ | $Al^{3+}$ | $P^{3-}$ | $S^{2-}$ | $Cl^-$ |
| 0.095 | 0.065 | 0.050 | 0.212 | 0.184 | 0.181 |

**(ii)** Explain the differences in size between the cations and the corresponding atoms.

Cations have less electrons than their parent atoms so there is an increase in net nuclear attraction which causes the radius to shrink.

**(iii)** Explain the differences in size between the anions and the corresponding atoms.

There is a decrease in net nuclear attraction as electrons are added to form an anion, resulting in the anion's radius being longer than their parent atom.

[5]

10 The alkali metals are a series of six elements in Group I of the Periodic Table. The first ionisation energy of these elements shows a marked trend as the Group is descended.

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

Energy required to remove one mole of electron from one mole of gaseous atoms under standard conditions. [2]

**(b)** **(i)** State and explain the trend in first ionisation energy as Group I is descended.

Shells increase down the group so the distance increases, added to that, the lower shells cause a sheilding effect reducing the ionisation energies. Even though # protons ↑

**(ii)** Suggest how this trend helps to explain the increase in the reactivity of the elements as the Group is descended.

Outermost electron is easily given off increasing the reactivity of the elements down the group. [3]

**(c)** In a redox reaction, 0.83 g of lithium reacted with water to form 0.50 $dm^3$ of aqueous lithium hydroxide.

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

**(i)** Calculate the amount, in moles, of lithium that reacted.

$n = \frac{0.83}{6.9} = 0.120$ mol.

**(ii)** Calculate the volume of hydrogen produced at room temperature and pressure.

$n_{H_2} = 0.120 / 2 = 0.06$ mol

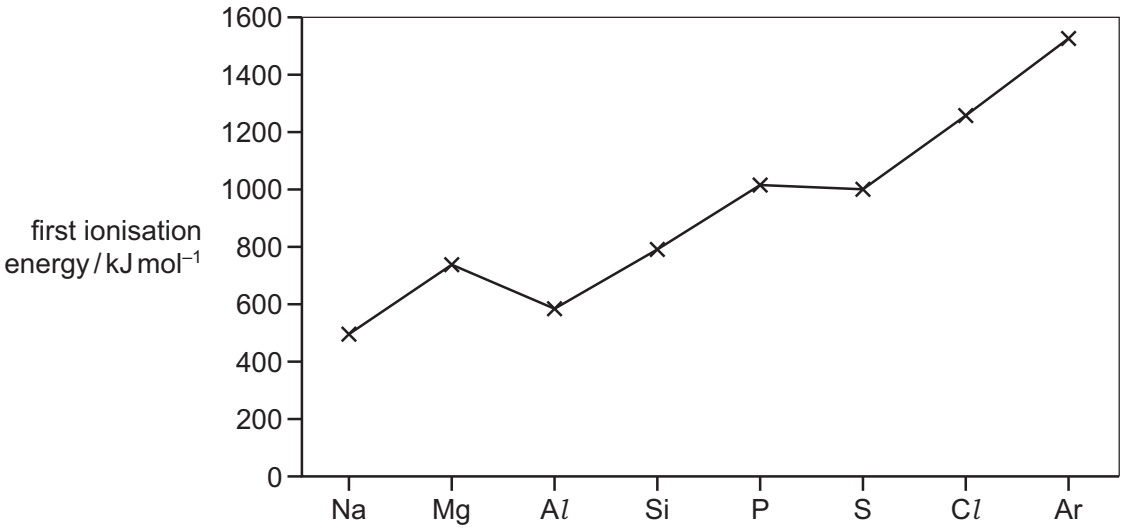
$V_{H_2} = 0.06 \times 24 = 1.44 dm^3$

**(iii)** Calculate the concentration, in mol $dm^{-3}$, of the LiOH(aq) formed.

$C_{LiOH} = \frac{0.12}{0.5} = 0.240$ mol $dm^{-3}$

[5]

**11 (b)** The graph below shows the variation of the first ionisation energies across Period 3.




**(i)** Explain why the first ionisation energy of Ar is greater than that of Cl.

Argon has more protons for the same number of shells, hence the electrons experience more pull.

[1]

**(ii)** Explain why the first ionisation energy of Al is less than that of Mg.

Al electron is in the 3p orbital so it experiences additional shielding from the 3s orbital.

[1]

**(iii)** Explain why the first ionisation energy of S is less than that of P.

Sulphur has a paired electron that adds to the repulsion, hence its easily given off.

[1]

12 **(a)** Complete the full electronic configuration of neon.

$1s^2$ $2s^2\ 2p^6$ [1]

**(b)** **(i)** Explain what is meant by the term *first ionisation energy*.

Energy required to remove one mole of electron from one mole of gaseous atoms under standard conditions. [3]

**(ii)** Explain why the first ionisation energy of neon is greater than that of fluorine.

Neon has more protons than fluorine and its outer electrons experiences greater nuclear attraction. [2]

13 **(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 |

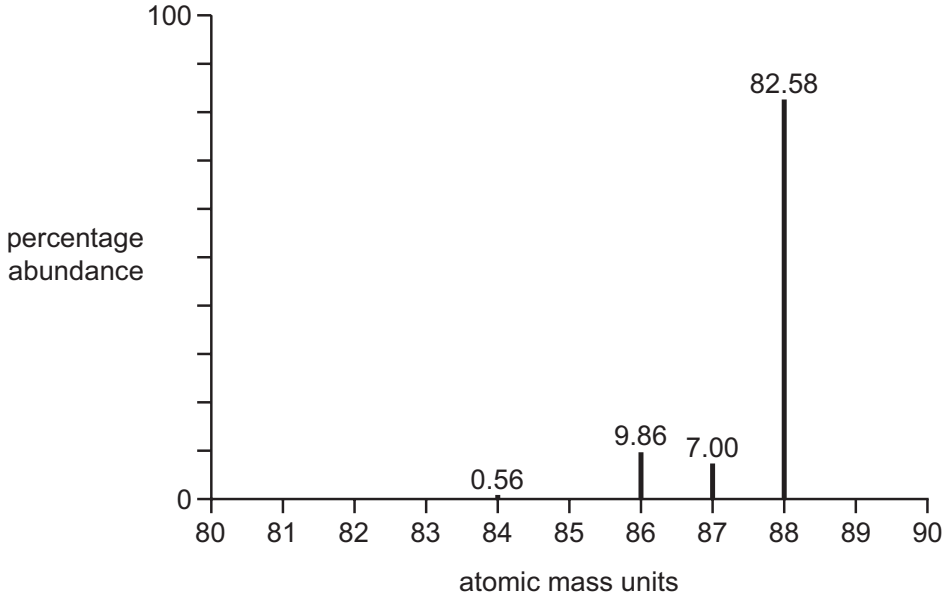
**(i)** Explain why the first ionisation energies decrease down the group.

Despite the increase in nuclear charge
• increase in radius • and in shielding
lower the nuclear attraction on the outer electrons
Hence causing a decrease in IE. [3]

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

Group 2 elements have 2 valence electrons.
So the third ionisation energy is higher because
the $e^-$ is in a lower shell, being closer
to the nucleus and experience lower shielding. [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.

$1s^2\ 2s^2\ 2p^6$ ... $3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^2$ ..... [1]

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

Strontium has 4 naturally occurring isotopes.

..... [1]

**(iii)** Calculate the atomic mass, $A_r$, of this sample of strontium.
Give your answer to **three** significant figures.

$$\frac{0.56(84) + 9.86(86) + 7(87) + 82.58(88)}{100}$$

$A_r$ = 87.7 ..... [2]

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

**(i)** Explain, in terms of electron transfer, what is meant by the term *oxidising agent*.

An oxidising agent is prone to accepting electrons, reducing itself and oxidising the reagents in the reaction. ..... [1]

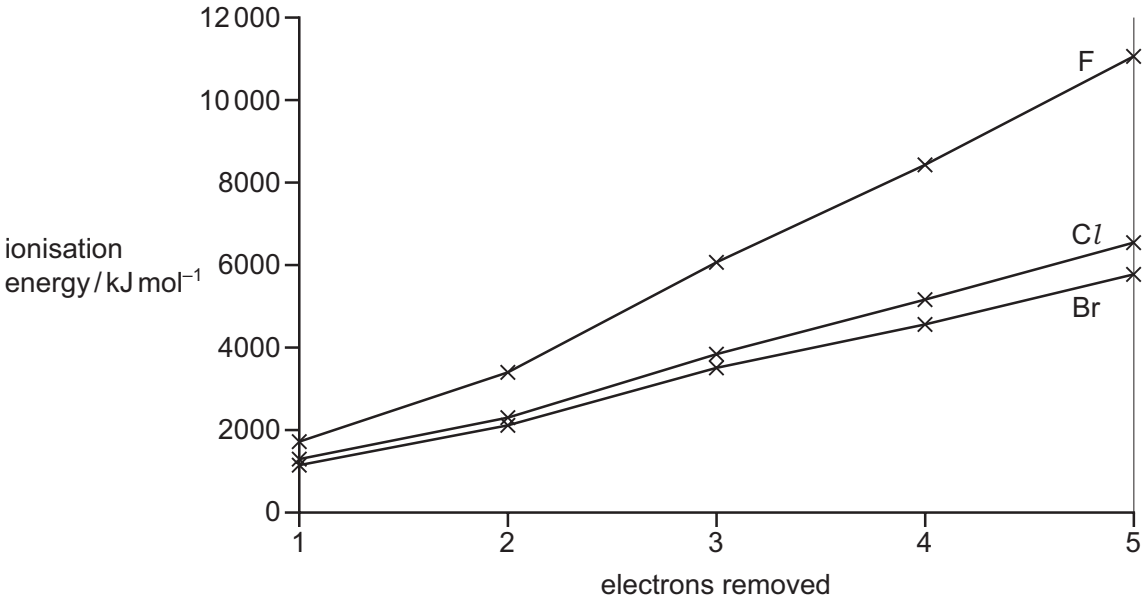
**(ii)** **A** has the following percentage composition by mass: Ba, 45.1; Cl, 23.4; O, 31.5.

Calculate the empirical formula of **A**.

| Ba | Cl | O |
|---|---|---|
| 45.1 / 137 | 23.4 / 35.5 | 31.5 / 16 |
| 0.329 | 0.659 | 1.96 |
| 1 | 2 | 6 |

empirical formula of **A** $BaCl_2O_6$ ..... [3]

**14 (a)** Successive ionisation energies for the elements fluorine, F, to bromine, Br, are shown on the graph.




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

Down the group there is an increase in no. of shells that keep the effective nuclear charge constant through sheilding. The increase in radius also contributes to a low IE required.

[3]

**(ii)** Explain why there is an increase in the successive ionisation energies of fluorine.

There are more protons per electron, increasing the nuclear pull. Also, the orbital for the ionisation energies mentioned is the same.

[2]

15 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** | 6274 | 21 269 | 25 398 | 29 855 |
| **Y** | 7012 | 8496 | 27 107 | 31 671 |
| **Z** | 6542 | 9362 | 11 018 | 33 606 |

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

group number Group 16

explanation big jump between 6th and 7th ionisation energies.

[1]

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

Increases due to increase in nuclear attraction due to the increasing number of protons across the period. [2]

**(iii)** Explain why the **first** ionisation energy of element **Y** is less than that of element **X**.

Y has a pair of electrons in the p orbital which experiences repulsion from the pair.

[2]

**(iv)** Complete the electronic configuration of element **Z**.

1s$^2$ $2s^2 2p^6 3s^2 3p^5$ [1]

16 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 |

sulfur

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

group number Group 17

explanation Big jump b/w 7th and 8th IE. which tells us Y has 7 valence electrons.

[1]

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

- Increase.
- Due to increase in protons the increase nuclear attr.
- Radius↓, sheild remains constant.

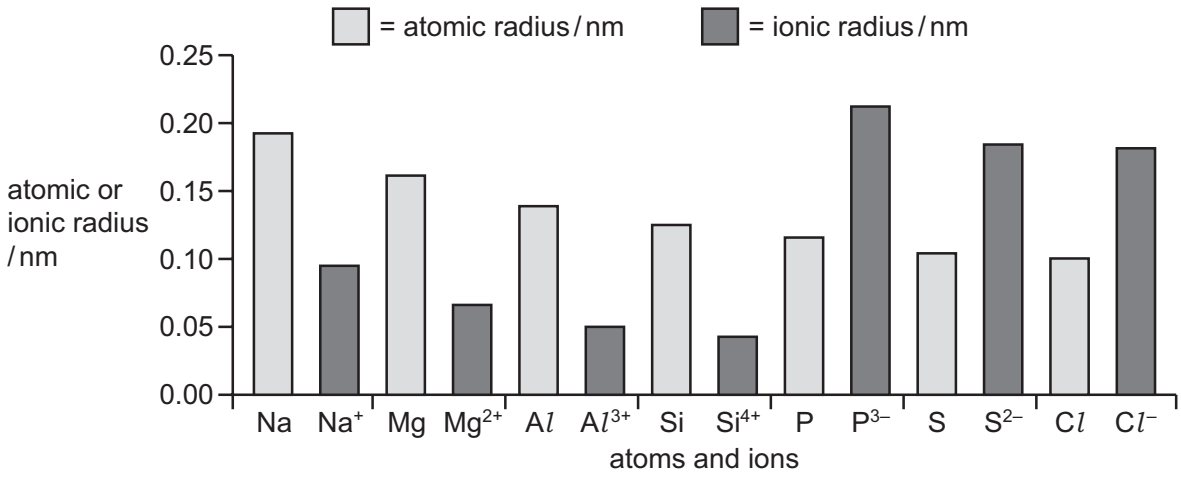
[2]

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

1s$^2$ 2s$^2$2p$^6$ 3s$^2$3p$^4$ [1]

**17** The elements in the third period exhibit periodicity in both their chemical and physical properties.

**(a)** A graph of the atomic and ionic radii across the third period is shown.




**(i)** Explain the decrease in atomic radius across the third period.

There is an increase in nuclear charge across the period. The increase in electrons doesn't change the sheilding or number of shells.

[2]

**(ii)** Explain why, for sodium to silicon, the ionic radii are less than the atomic radii.

These elements form anions and lose their outer shell.

[1]

**(iii)** Explain why, for phosphorus to chlorine, the ionic radii are greater than the atomic radii.

The elements form cations and gain electrons, this increases the repulsion between the outer electrons, expanding the orbital. As there is less net attraction from the nucleus.

[2]

**(b)** The first ionisation energies of the elements across the third period show a general increase.

Aluminium and sulfur do **not** follow this general trend.

**(i)** Explain why aluminium has a lower first ionisation energy than magnesium.

The electron in in the p orbital for Al which experiences more sheilding from the inner s orbital.

[2]

[S'18 1 Q3]

18 The first six successive ionisation energies of an element **D** are shown in Table 4.1 below.

**Table 4.1**

| element | ionisation energy / kJ mol$^{-1}$ | | | | | |
|---|---|---|---|---|---|---|
| | 1st | 2nd | 3rd | 4th | 5th | 6th |
| **D** | 1086 | 2353 | 4621 | 6223 | 37832 | 47278 |

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

It is the energy required to remove one electron from each atom in 1 mol of gaseous atoms of an element under standard conditions. [3]

**(b)** Write an equation, with state symbols, to represent the **third** ionisation energy of element **D**.

$D^{2+}_{(g)} \longrightarrow D^{3+}_{(g)} + e$ [2]

**(c)** Use Table 4.1 to deduce which group of the Periodic Table contains element **D**. Explain your answer.

*group* Group 14.

*explanation* Big jump between 4 and 5 IE. which tells us that D has 4 valence electrons. [3]

19 Sir James Jeans, who was a great populariser of science, once described an atom of carbon as being like six bees buzzing around a space the size of a football stadium.

**(a)** **(i)** Suggest what were represented by the six bees in this description.

six electrons of carbon.

**(ii)** Explain (in terms of an atom of carbon) what stopped the bees from flying away from the space of the football stadium.

The attraction from the positively charged nucleus.

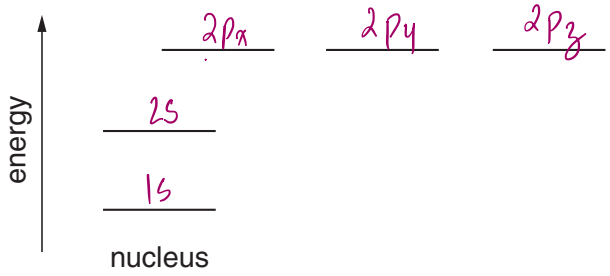
**(iii)** What is missing from Jeans' description when applied to an atom of carbon?

The nucleus.

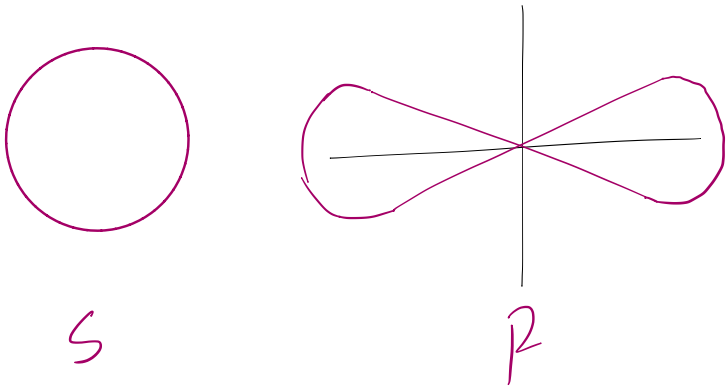
[3]

**(b)** The diagram below represents the energy levels of the orbitals in atoms of the second period, lithium to neon.

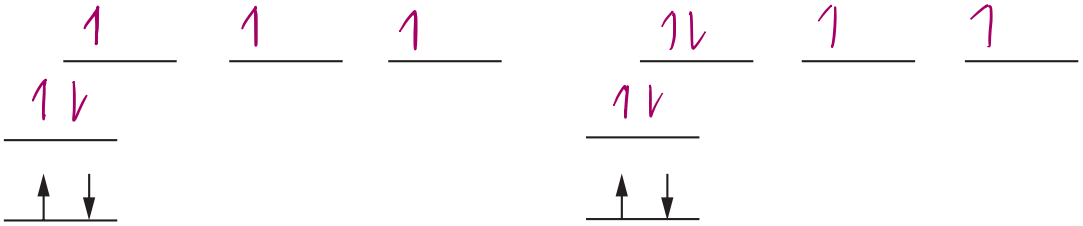
**(i)** Label the energy levels to indicate the principal quantum number and the type of orbital at each energy level.




**(ii)** In the space below, sketch the shapes of the two types of orbital.



**(iii)** Complete the electron configurations of nitrogen and oxygen on the energy level diagrams below, using arrows to represent electrons.



nitrogen oxygen

**(iv)** Explain, with reference to your answer to **(iii)**, the relative values of the first ionisation energies of nitrogen and oxygen. The values are given in the *Data Booklet* and should be quoted in your answer.

IE of N is 1400 and form O is 1310 $KJmol^{-1}$
The IE for oxygen is lesser than that of oxygen as the electron for oxygen experiences repulsion from its pair and is therefore, easier to remove.

[6]

**(c)** **(i)** State the formulae of the negatively charged ions formed by these elements in simple binary compounds (nitrides and oxides).

$N^{3-}$ and $O^{2-}$

**(ii)** Why do nitrogen and oxygen form negative ions, but not positive ions, in simple binary compounds?

The anions have a noble gas configuration.
To form cations a lot of energy will be required to remove the electrons from the outer shells of nitrogen and oxygen.

[2]

[Total : 11]

[S'02 Q1]