Ionisation energy

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First ionisation energy

Definition

The first ionisation energy is the enthalpy change when one mole of electrons is removed from one mole of gaseous atoms forms one mole of gaseous ions with a single positive charge

This is represented by the equation:

$$H_{(g)} \rightarrow H^+_{(g)} + e^- \qquad \Delta H = \text{IONISATION ENTHALPY}$$

Always gaseous

Key idea

Learn it !

The equation for 1st ionisation energy always follow the same pattern.

Learn it!

$X_{(g)} \longrightarrow X^+_{(g)} + e^-$

Always gaseous

$He_{(g)} \longrightarrow He^+_{(g)} + e^-$

$Be_{(g)} \longrightarrow Be^+_{(g)} + e^-$

$Ra_{(g)} \longrightarrow Ra^+_{(g)} + e^-$

It does not matter if the atom does not normally form a +1 ion or is not gaseous

$O_{(g)} \longrightarrow O^+_{(g)} + e^-$

$Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$

Key idea

The second ionisation energy is the enthalpy change when one mole of gaseous ions with a single positive charge forms one mole of gaseous ions with a double positive charge

The equation for 2nd ionisation energy also always follow the same pattern.

Learn it!

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

Always gaseous

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

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

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

Learn it !

Key idea

Factors that effect Ionisation energy

There are three main factors

1.The attraction of the nucleus

(The more protons in the nucleus the greater the attraction)

2. The distance of the electrons from the nucleus

(The bigger the atom the further the outer electrons are from the nucleus and the weaker the attraction to the nucleus)

3. Shielding of the attraction of the nucleus

(An electron in an outer shell is repelled by electrons in complete inner shells, weakening the attraction of the nucleus)

Key idea

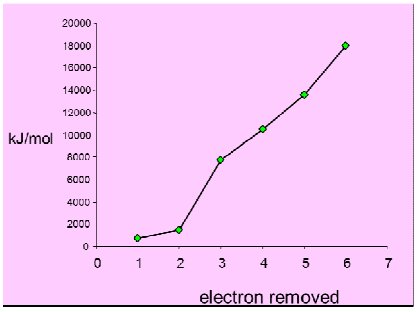
Explaining differences in ionisation energies

Key idea

The second ionisation energy of an element is always bigger than the first ionisation energy

When the first electron is removed the attraction a positive ion is formed.

The ion increases the attraction on the remaining electrons and so the energy required to remove the next electron is larger.

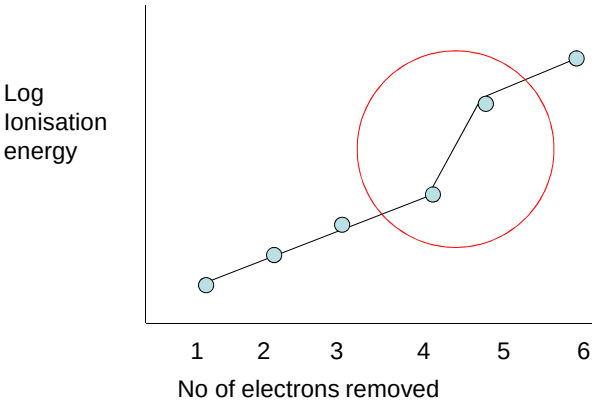


Why the big jump between the 2nd and 3rd?

Ionisation energies and group numbers

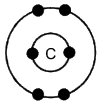
The patterns in successive ionisation energies for an element give us important information about the electronic structure for that element.

Here is the pattern for Carbon



Notice the big jump between 4 and 5.

How can this be related to the electronic structure of carbon ?



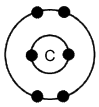
Electronic configuration 2,4

Key idea

Why the big jump ?

The fifth electron is in a inner shell closer to the nucleus and therefore attracted much more strongly by the nucleus than the fourth electron.

It also does not have any shielding by inner complete shells of electron

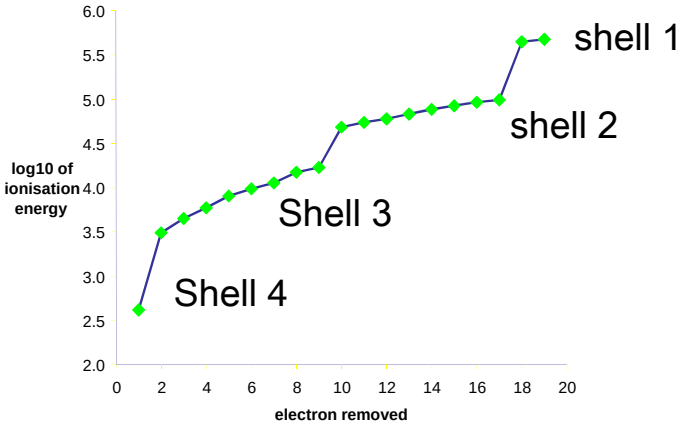


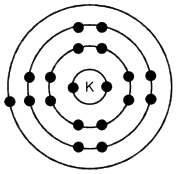
Electronic configuration 2,4

Key idea

Ionisation energies and group numbers

Successive ionisation energies for potassium





Electronic configuration

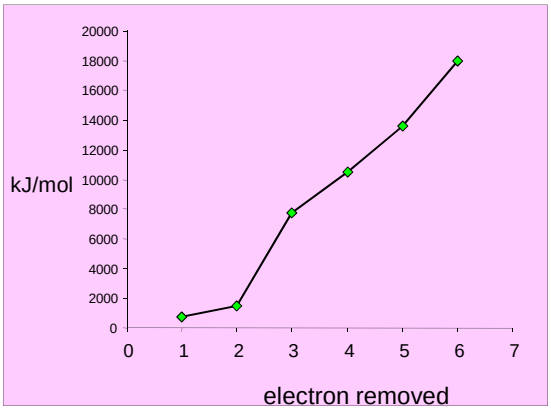
2,8,8,1

The different “jumps” are evidence for the arrangement of electrons in energy levels and sub-levels

Key idea

QUESTION Ionisation energies and group numbers

Which group is this element in?



Group 2

Notice the “jump” in energy needed to remove the 3rd electron

QUESTION Ionisation energies and group numbers

The following table gives the successive ionisation energies of an element in Period 3.

	1	2	3	4	5	6
Ionisation energy kJ mol^{-1}	786	1580	3230	4360	16100	19800

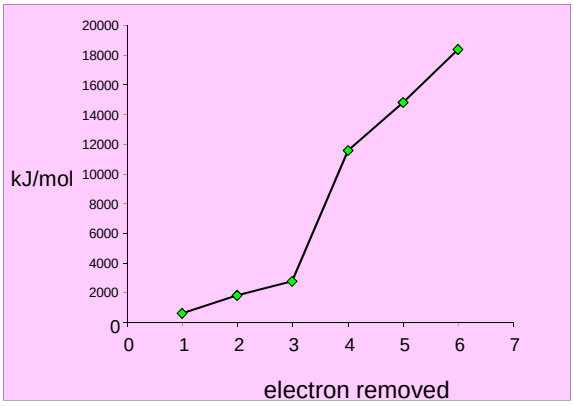
Which period 3 element is this?

Silicon

Notice the "jump" in energy needed to remove the 5th electron

QUESTION Ionisation energies and group numbers

What is the smallest element this could be?



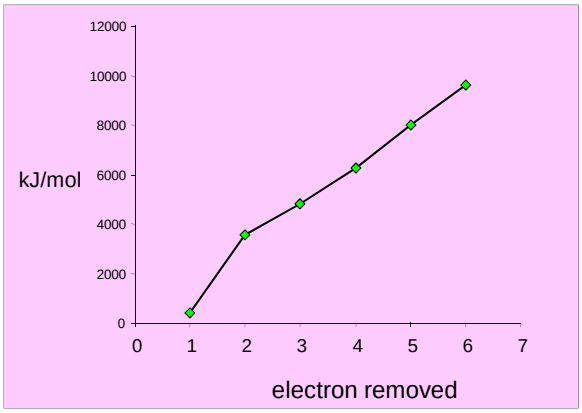
Aluminium

In Group 3 - the "jump" in energy needed to remove the 4th electron.

Can't be Boron as has more than 5 electrons

QUESTION Ionisation energies and group numbers

What is the smallest element this could be?

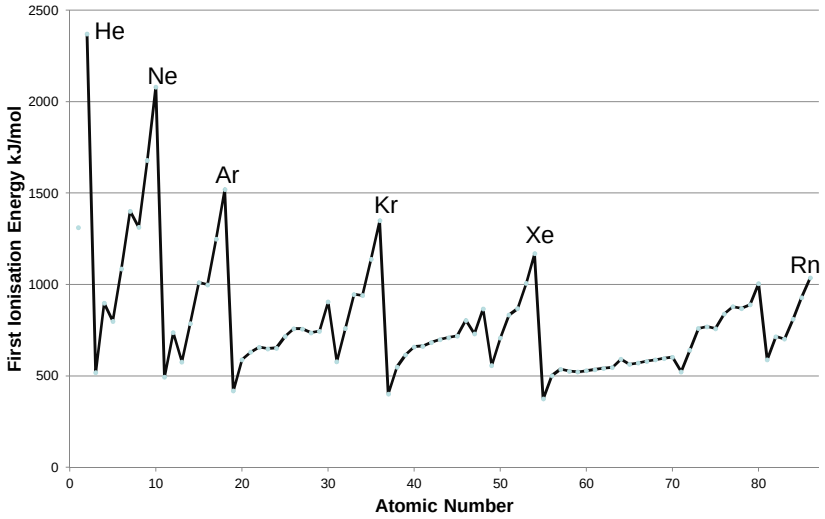


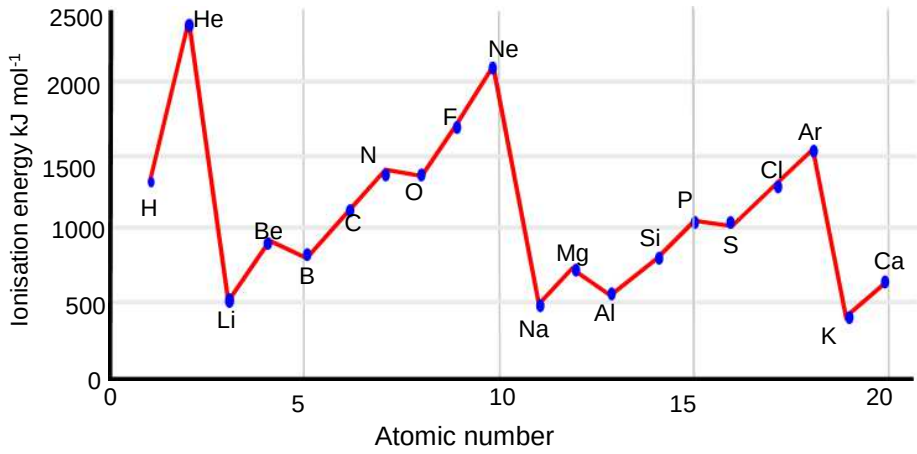
sodium

In Group 1 - the “jump” in energy needed to remove the 2nd electron.

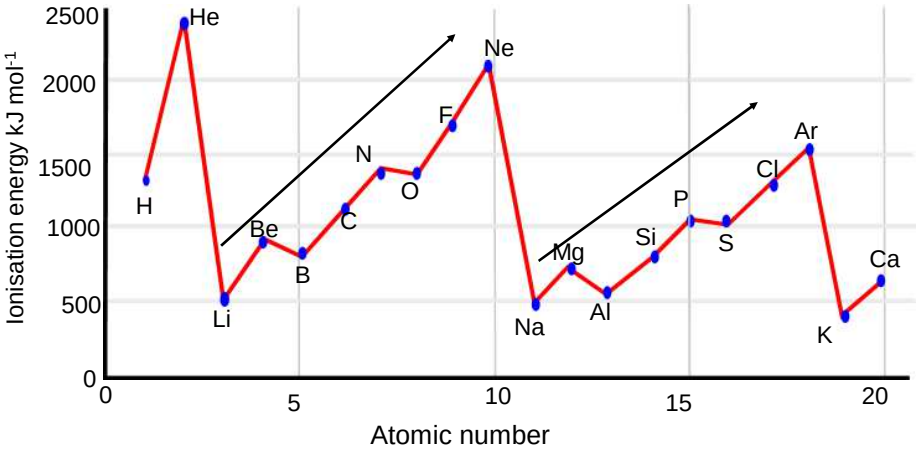
Can’t be lithium as has more than 3 electrons

The first Ionisation energy of the elements

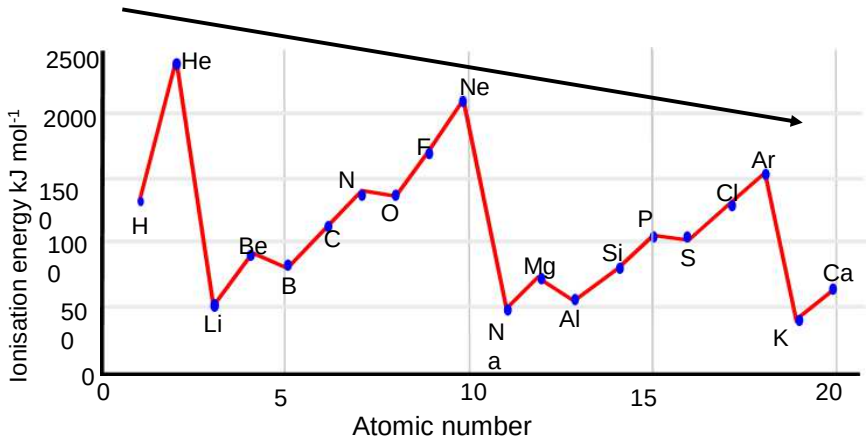




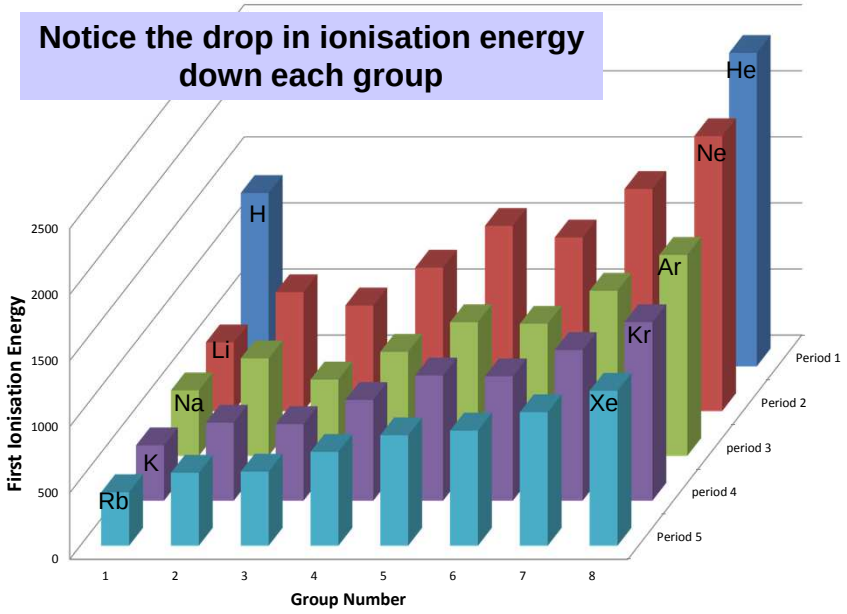
The shape of the graph for periods two and three is similar. A repeating pattern across a period is called periodicity.

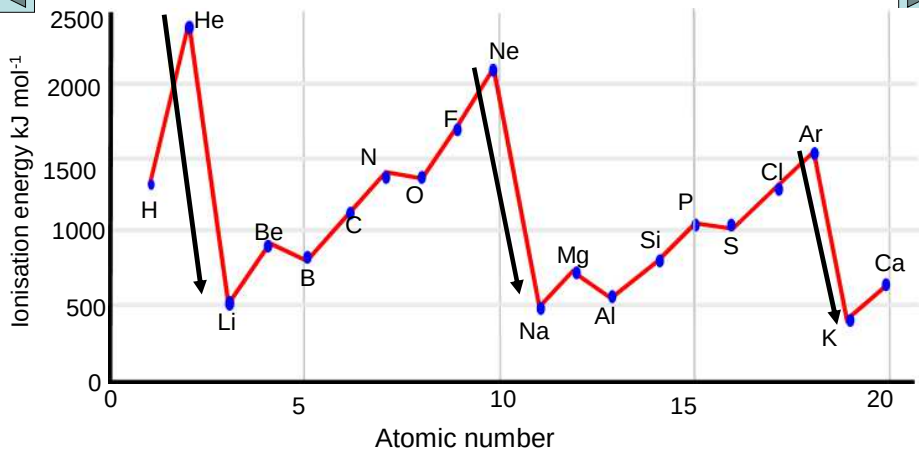


There is a general trend across each period for the ionisation energy to increase. This is due to increasing number of protons in successive elements, which means the electrons are more strongly held to the nucleus.

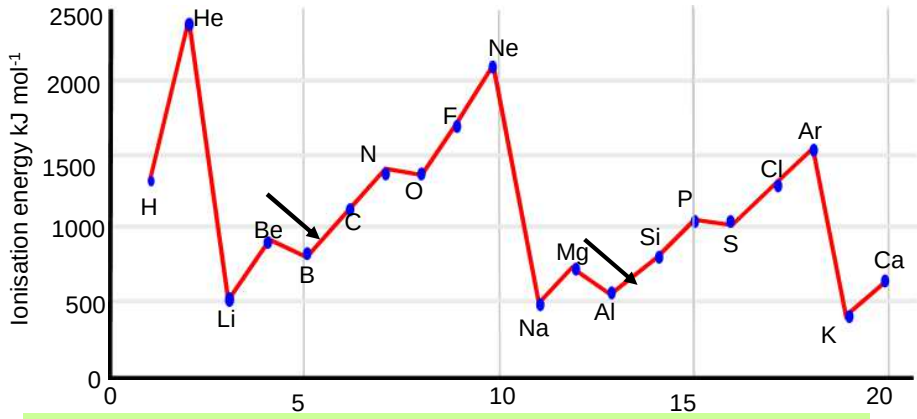


The noble gases are always at the maximum peak for each period, but there is a decrease in ionisation energy down the group. (true of all groups).
This is because as one goes down the group the outer electrons become further from the nucleus and become more shielded from the nuclear pull.





There is a large drop each time between the group 0 elements and the group1 elements.
This is because the element in group 1 will have its outer electron in a new shell further from the nucleus and is more shielded. So the group 1 element is easier to remove and has a lower ionisation energy.



There is a small drop between Be + B , Mg + Al.
B and Al are starting to fill a p sub shell, whereas Be and Mg have their outer electrons in the s sub shell. The electrons in the p subshell are slightly easier to remove because the 2p electrons are higher in energy.

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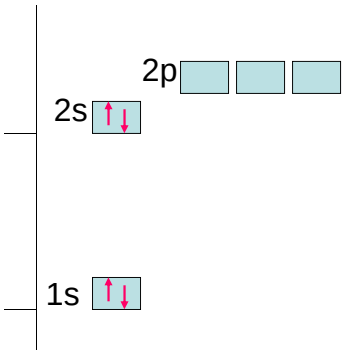
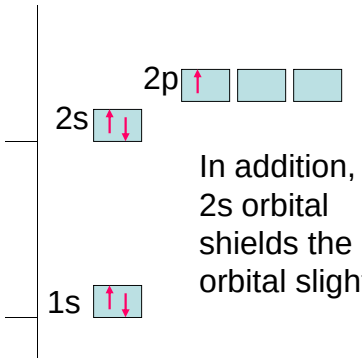
Beryllium

Electronic configuration : $1s^2 2s^2$

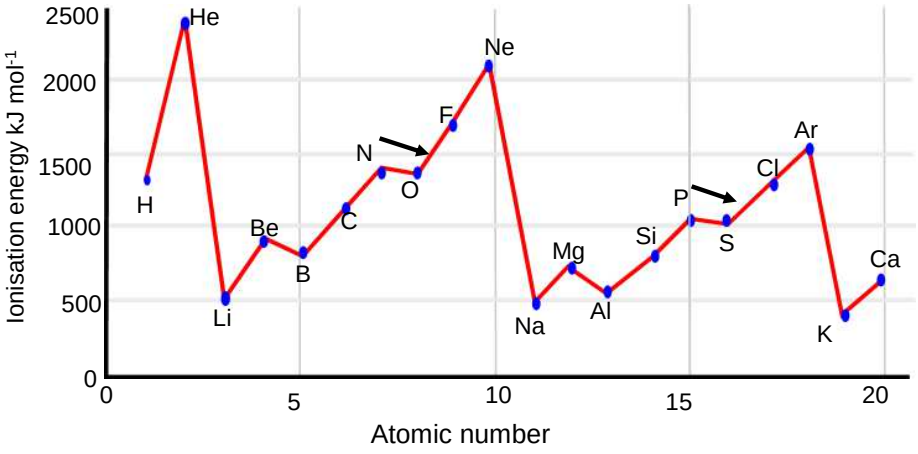


Boron

Electronic configuration : $1s^2 2s^2 2p^1$



In addition, the 2s orbital shields the 2p orbital slightly

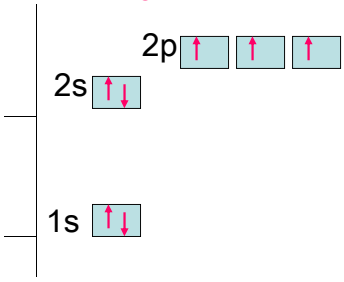


There is a small drop between nitrogen and oxygen and also phosphorous and sulphur.

When the second electron is added to an orbital there is a slight repulsion between the two negatively charged electrons which makes the second electron easier to remove.
For example with oxygen there are 4 p electrons and the 4th is starting to doubly fill the orbitals.

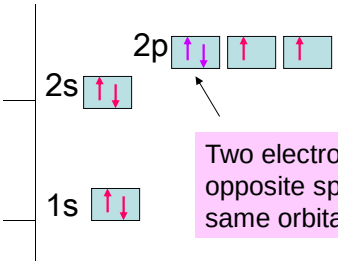
Nitrogen

Electronic configuration : $1s^2\,2s^2\,2p^3$



Oxygen

Electronic configuration : $1s^2 2s^2 2p^4$



Two electrons of opposite spin in same orbital

Problems for discussion

Explain why the ionisation energy of every element is endothermic

Which element has the largest 1st Ionisation energy? Explain why.

Which element has the largest 2nd Ionisation energy? Explain why.

Sketch a graph of the 2nd Ionisation energy for period 3 Na to Ar.

Explain why the ionisation energy of every element is endothermic

Energy is needed to overcome the attractive force between the negatively charged electron and the positively charged nucleus.

What is wrong/insufficient about these answers?

"Heat is needed to pull electron out of the nucleus"

"Energy is needed to break the bond between the electron and the nucleus"

"Energy is needed to remove the electron from the atom"

Which element has the largest 1st Ionisation energy? Explain why

Helium. Its outer electrons are in the first shell closest to the nucleus and has no shielding effects from inner shells. He has a bigger first ionisation energy than H as it has one more proton

What is wrong/insufficient about these answers?

"Helium's outer electron is close to the nucleus with little shielding so hard to remove"

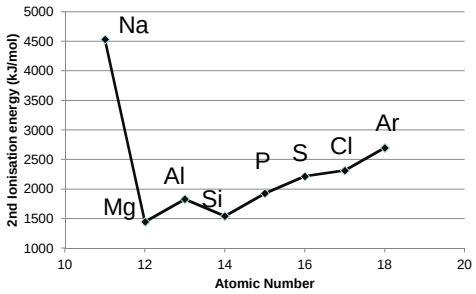
"Helium is a noble gas and has a full shell of electrons. It would lose stability if an electron was removed."

"Helium has a really strong ionic bond between its electron and nucleus because it has no shielding."

Which element has the largest 2nd Ionisation energy? Explain why

Lithium. Its second electron would be removed from the first 1s shell closest to the nucleus and has no shielding effects from inner shells. Li has a bigger second ionisation energy than H and He as it has more protons.

Sketch a graph of the 2nd Ionisation energy for period 3 Na to Ar.



EXAM QUESTIONS

(a) Explain the meaning of the term *first ionisation energy.*

(b) Explain why boron has a lower first ionisation energy than beryllium.

(c) Explain why the first ionisation energy of helium is very large.

(d) Write equations to show the chemical processes which occur when the first and the second ionisation energies of lithium are measured.

(e) Explain why helium has a much higher first ionisation energy than lithium.

(f) Explain why the second ionisation energy of beryllium is greater than the first ionisation energy.

Exam Question

There is a general trend in the values of the first ionisation energies of the elements Na to Ar. The first ionisation energies of the elements Al and S deviate from this trend.

(a) Write an equation, including state symbols, to represent the process for which the energy change is the first ionisation energy of Na. **(2)**

(b) State and explain the general trend in the values of the first ionisation energies of the elements Na to Ar. **(3)**

(c) State how, and explain why, the values of the first ionisation energies of the elements Al and S deviate from the general trend. **(5)**