9 The Periodic Table: chemical periodicity

This topic illustrates the regular patterns in some physical properties of the elements in the Periodic Table.

9.1 Periodicity of physical properties of the elements in the third period ☐

PERIODIC TRENDS

9 The Periodic Table: chemical periodicity

This topic illustrates the regular patterns in chemical and physical properties of the elements in the Periodic Table.

	Learning outcomes Candidates should be able to:
9.1 Periodicity of physical properties of the elements in the third period	a) describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements (see the *Data Booklet*) b) explain qualitatively the variation in atomic radius and ionic radius c) interpret the variation in melting point and electrical conductivity in terms of the presence of simple molecular, giant molecular or metallic bonding in the elements d) explain the variation in first ionisation energy (see the *Data Booklet*) e) explain the strength, high melting point and electrical insulating properties of ceramics in terms of their giant structure; to include magnesium oxide, aluminium oxide and silicon dioxide

PERIODICITY

The outer electronic configuration is a periodic function, it repeats ever so often. Many physical and chemical properties are influenced by the outer shell configuration of an atom and hence also exhibit periodicity, such as:

atomic radius, ionic radius, ionisation energy, electron affinity, electronegativity, electrical conductivity, melting point and boiling point

The first two periods in the periodic table are not typical...

Period 1 (H, He) contains only two elements

Period 2 elements at the top of each group have small sizes and high I.E.values

Period 3 (Na-Ar) is the most suitable period for studying trends

ATOMIC RADIUS

The atomic radius is basically used to describe the size of an atom. Larger the atomic radius, larger the atom.

Atomic radius increases down a group.

This is because, as we go down a group in the periodic table the atoms have increasingly more electron shells.

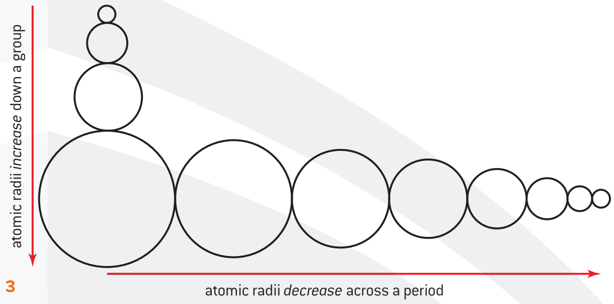
In each new period the outer-shell electrons enter a new energy level and so are located further away from the nucleus.

ATOMIC RADIUS

Atomic radius decreases across a period

The magnitude of the positive charge of the nucleus increases, its "pull" on all of the electrons increases, hence the number of shells remain the same and the electrons are drawn closer to the nucleus.

This results in a contraction of the atomic radius and therefore a decrease in atomic size.



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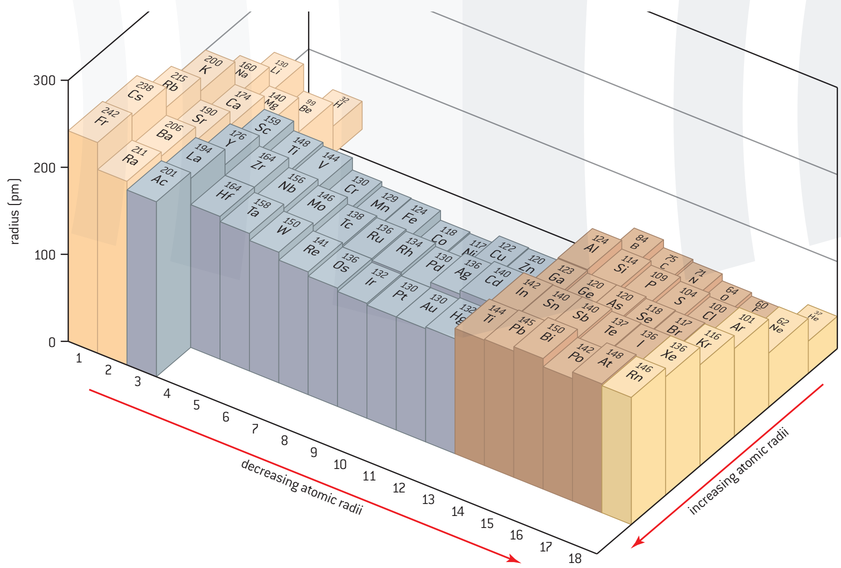
ATOMIC RADIUS



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ATOMIC RADIUS MEASUREMENTS

The radius of an atom is determined by the distance between the center of the nucleus and the boundary of the region where the valence electrons have a probability of being located.

Since the probability of finding an electron in an atom extends up to infinity it is impossible to determine the exact size of the atom.

Hence atomic radius is taken as half the distance between the centre of two adjacent atoms in a molecule or in a closed packed system.

The bond between two adjacent atoms may be covalent, metallic or simply van der Waal's forces.

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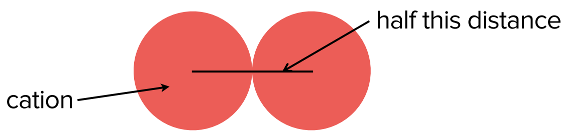
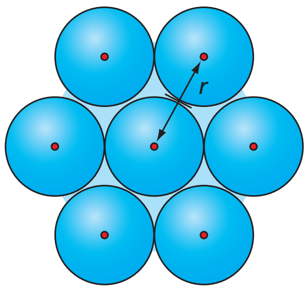
ATOMIC RADIUS

Chemists use X-ray diffraction and other techniques to measure the distance between the nuclei of atoms.

The atomic radius of an atom cannot be defined precisely because it depends on the type of bonding and on the number of bonds.

Metallic radius in metals is half the distance between the inter-nuclear distance of what are effectively ions.





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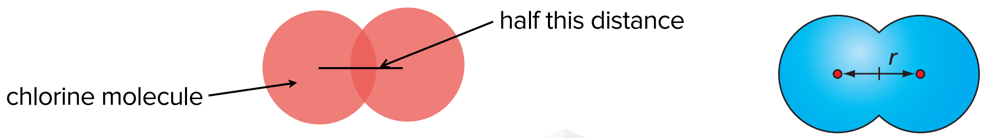
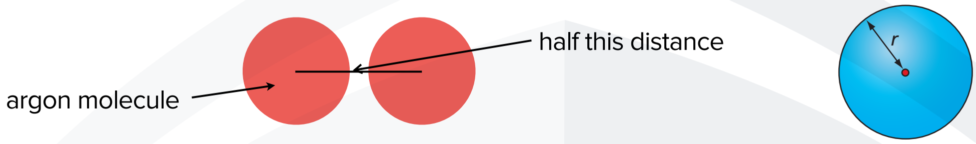
ATOMIC RADIUS MEASUREMENTS

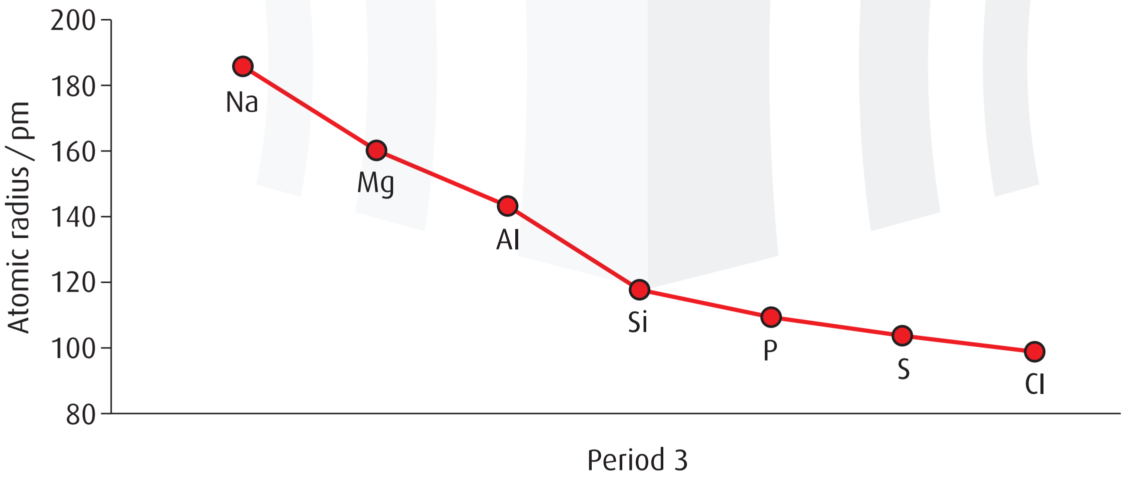
- **Covalent radius** is half the distance between the nuclei of atoms joined by a covalent bond. The values are measured by X-ray or electron diffraction.



- **Van der Waal's** radius is half the distance between two monoatomic molecules (two atoms held by van der Waal's forces)



ATOMIC RADIUS



SKILL CHECK 1

Put the following elements in order of increasing atomic radius. Justify your answers:

A Mg, S, Si

B Mg, K, Al

C Si, Cl, K

SKILL CHECK 2

Explain why the ionic radius of a Group 2 ion is smaller than the atomic radius of the corresponding Group 2 atom.

IONIC RADIUS

Positive ions (cations) are smaller than the parent atom.

The cation has more protons than electrons (an increased nuclear charge).

The excess nuclear charge pulls the remaining electrons closer to the nucleus.

Also, cation formation often results in the loss of all outer-shell electrons, resulting in a significant decrease in radius.

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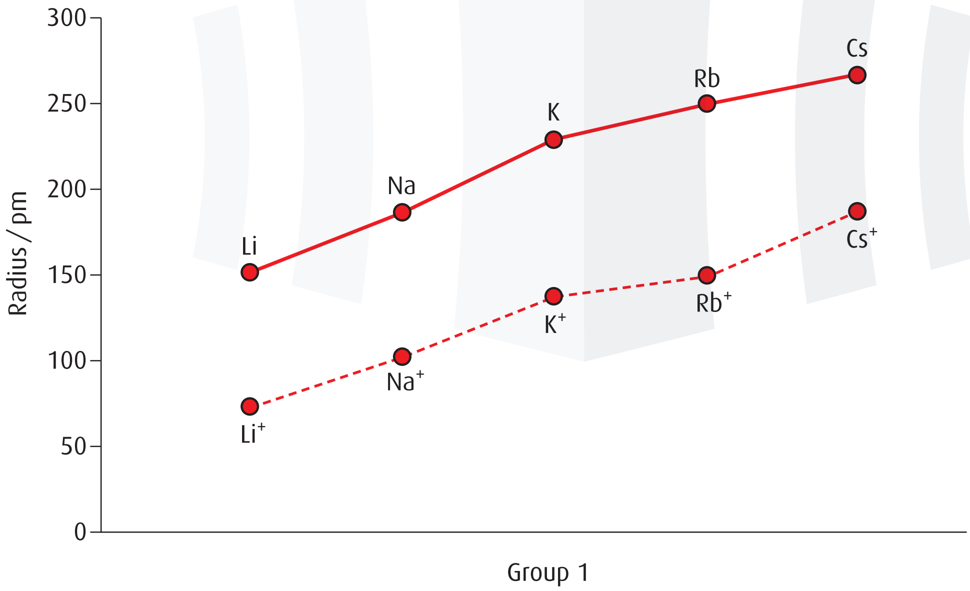
IONIC RADIUS



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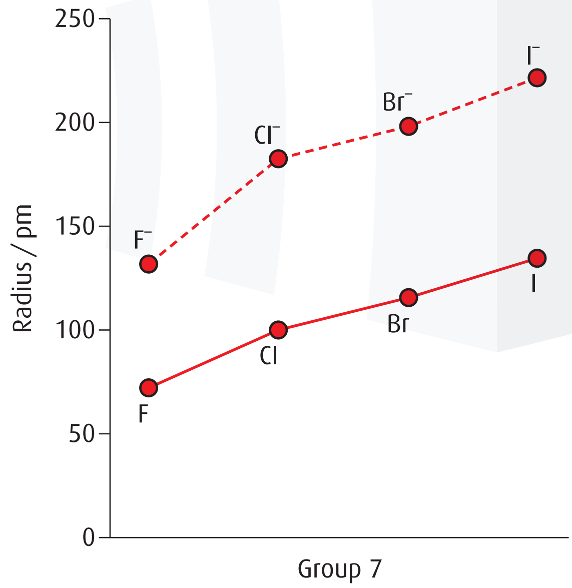
IONIC RADIUS

Negative ions (anions) are larger than the parent atom.

The anion has more electrons than protons.

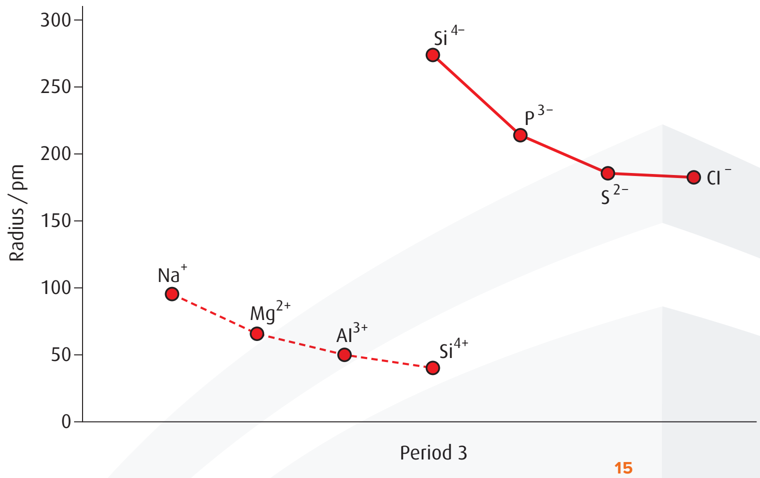
Owing to the excess negative charge, the nuclear "pull" on each individual electron is reduced. The electrons are held less tightly, resulting in a larger anion radius in contrast to the neutral atom.

IONIC RADIUS

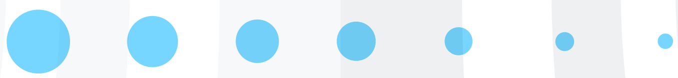


IONIC RADIUS ACROSS A PERIOD

Across a period the radii of ions with the same electron configuration decreases as nuclear charge increases.



IONIC RADIUS IN ISOELECTRONIC PARTICLES



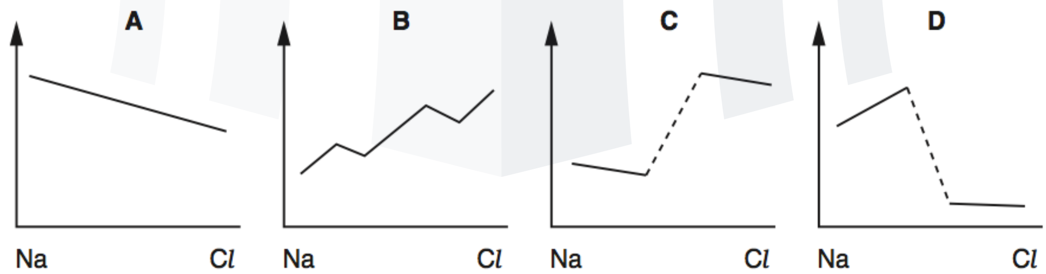
particle	N^{3-}	O^{2-}	F^{-}	Ne	Na^{+}	Mg^{2+}	Al^{3+}
nuclear charge	7	8	9	10	11	12	13
electrons	10	10	10	10	10	10	10

SKILL CHECK 3

Explain why the ionic radii of halides are larger than atomic radii of the corresponding halogens

SKILL CHECK 4

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



SKILL CHECK 5

The species Ar, K^+ and Ca^{2+} are isoelectronic (have the same number of electrons).

In what order do their radii increase?

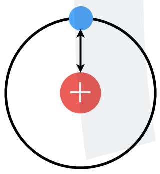
	smallest	→	largest
A	Ar	Ca^{2+}	K^+
B	Ar	K^+	Ca^{2+}
C	Ca^{2+}	K^+	Ar
D	K^+	Ar	Ca^{2+}

WHAT IS IONISATION ENERGY?

Ionisation Energy is the amount of energy needed to remove electrons from atoms.

The ionisation energy of an atom measures how strongly an atom holds its electrons.

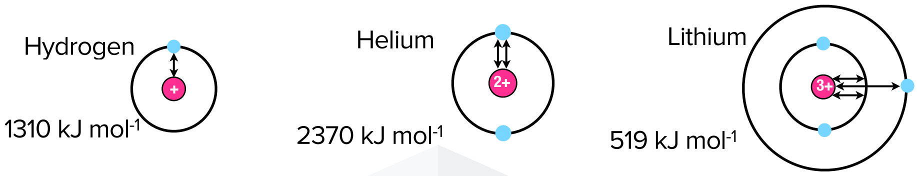
Attraction between the nucleus and an electron



As electrons are negatively charged and protons in the nucleus are positively charged, there will be an attraction between them. The greater the pull of the nucleus, the harder it will be to pull an electron away from an atom.

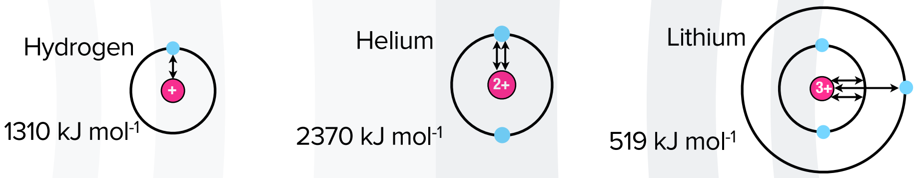
WHAT AFFECTS IONISATION ENERGY?

The value of the 1st Ionisation Energy depends on the electronic structure



Helium has two protons in the nucleus. The nuclear charge is greater so the pull on the outer electrons is larger. More energy will be needed to pull an electron out of the atom.

WHAT AFFECTS IONISATION ENERGY?



Lithium atoms have 3 protons so you would expect the pull on electrons to be greater. However, the 1st I.E. of Li is lower than that of He because

1. Filled inner shells exert a shielding effect - lowers the effective nuclear pull and,

2. Electrons are further away from the nucleus- lowers effective nuclear attraction.

FACTORS INFLUENCING IONISATION ENERGY

Nuclear Charge

As nuclear charge increases the attractive force would increase and hence more energy would be needed to overcome the increasing nuclear attraction.

Atomic Radius

As the atomic radius increases, the outermost electron would be further away from the nucleus experiencing a weaker attractive force. Hence the value decreases with increasing size.

Attraction decreases very rapidly with distance. An electron close to the nucleus will be much more strongly attracted than one further away.

FACTORS INFLUENCING IONISATION ENERGY

Shielding effect/Screening

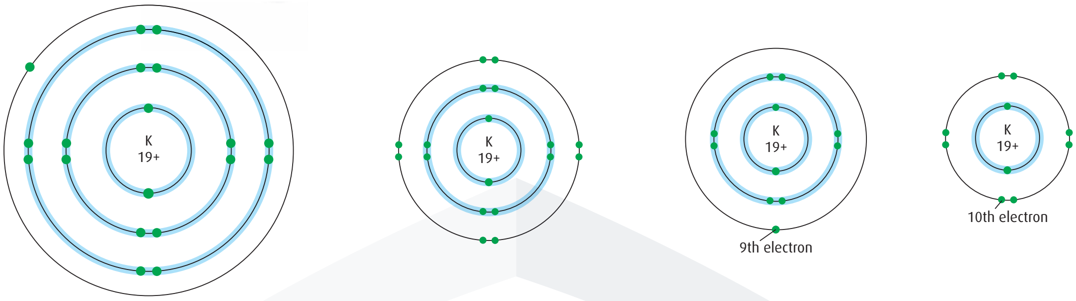
Electrons in full inner shells repel electrons in outer shells. Full inner shells of electrons prevent the full nuclear charge being felt by the outer electrons.

The inner electron orbitals effectively screen or shield the outermost electrons from the nucleus, due to which the ionisation energy decreases.

The lessening of the pull of the nucleus by the inner electrons is known as shielding.

It is due to the shielding effect that the electrons in the outer shell are attracted by the effective nuclear charge which is less than the full charge on the nucleus.

FACTORS INFLUENCING IONISATION ENERGY



FACTORS INFLUENCING IONISATION ENERGY

Nature of sub level

Electrons in the s orbital are closer to the nucleus as compared to the electrons in the p orbital of the same shell.

Hence the s electrons are the more firmly held and require a greater amount of energy to be removed as compared to the p electrons which are slightly further.

Also, the p orbital of the same energy level will experience the shielding effect of the s electrons from the same shell in addition to previous s orbitals.

Note: Only S Filled Orbitals Provide Shielding

FACTORS INFLUENCING IONISATION ENERGY

Stability of Certain Configurations

Two electrons in the same orbital experience a bit of repulsion from each other. This offsets the attraction of the nucleus so that the paired electrons are removed rather more easily.

Completely filled sub levels are more stable than others requiring large amounts of energy for their disruption. This additional stability is associated with half filled sub levels also.

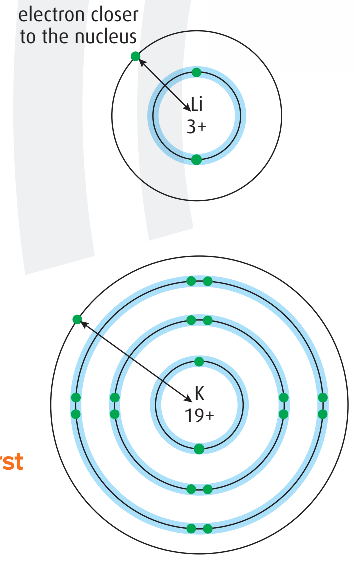
Stability of completely filled sub levels > **Stability of half-filled sub levels** > **Stability of other configurations**

IONISATION ENERGY DOWN A GROUP

Ionisation energy decreases down a group.

Despite an increased nuclear charge the outer electron is easier to remove. This is due to greater distance from the nucleus and increased shielding.

Down the group, as the number of shells increase, the outer electrons are farther away from the nucleus.



Potassium has a lower first ionisation energy than lithium.

IONISATION ENERGY DOWN A GROUP

Also, down the group each successive element contains more electrons in the shells between the nucleus and the outermost electrons which increases the shielding effect.

This increased shielding causes the outermost electrons to be held less tightly to the nucleus.

The increased shielding effect and increased atomic radius/size outweigh the increase in nuclear charge. Therefore ionisation energy decreases.

IONISATION ENERGY ACROSS A PERIOD

General Trend: Ionisation energy increases across a Period

Across a period, the number of protons and the number of electrons increase by one each.

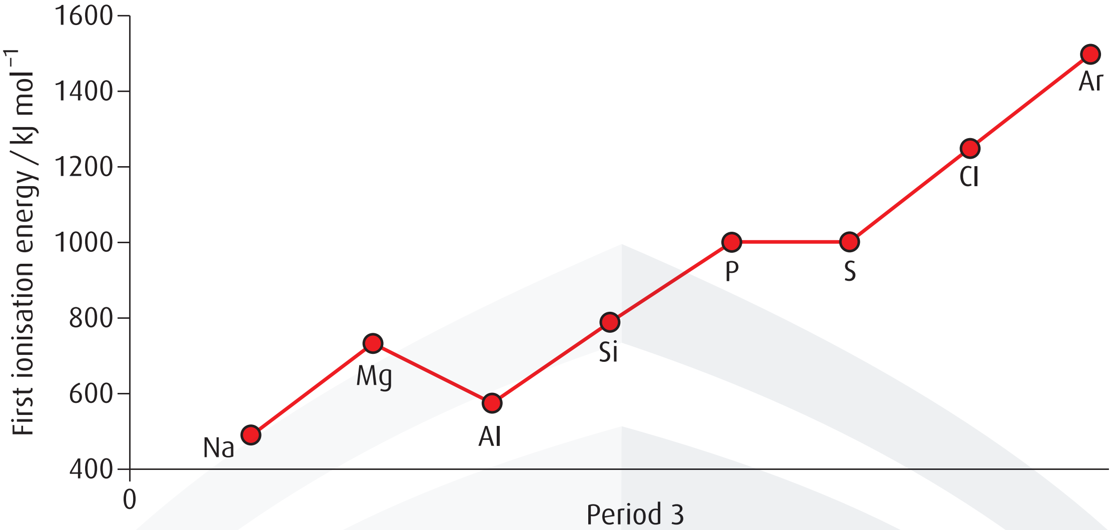
The additional proton increases the nuclear charge.

The additional electron is added to the same outer shell in each of the elements.

A higher nuclear charge more strongly attracts the outer electrons in the same shell, but the electron-shielding effect from inner-level electrons remains the same.

Thus, more energy is required to remove an electron because the attractive force on them is higher.

IONISATION ENERGY ACROSS A PERIOD



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SKILL CHECK 6

For each of the following pairs, state which element has the higher first ionisation energy and explain your answer:

A Mg and Al

B Mg and Ca

C Ne and Na

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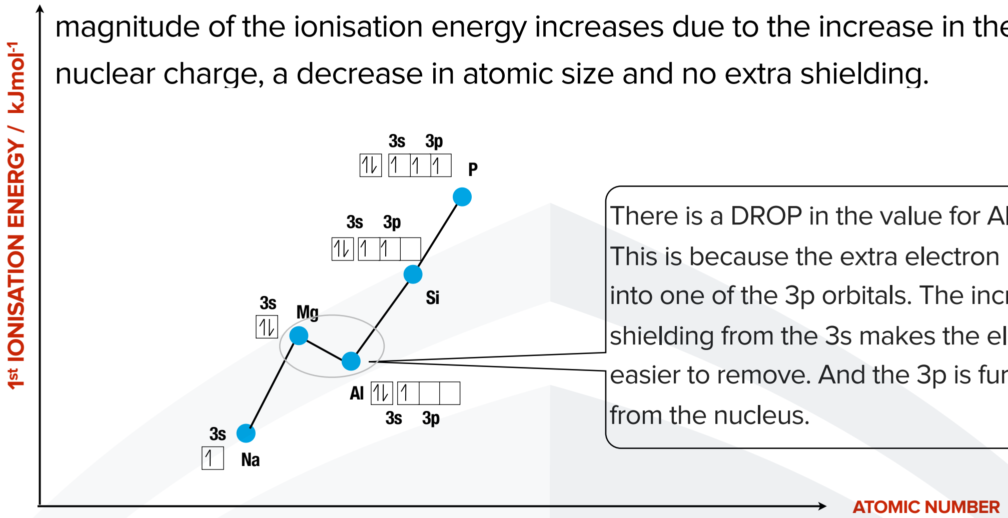
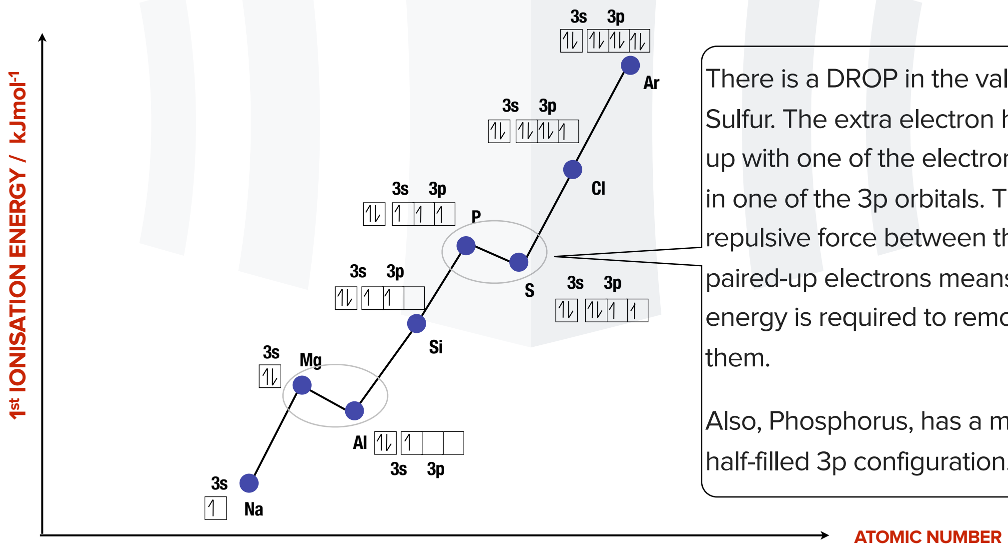
1ST IONISATION ENERGIES OF PERIOD 3

Along a Period as electrons are being added on to the same shell the magnitude of the ionisation energy increases due to the increase in the nuclear charge, a decrease in atomic size and no extra shielding.



There is a DROP in the value for Aluminum. This is because the extra electron has gone into one of the 3p orbitals. The increased shielding from the 3s makes the electron easier to remove. And the 3p is further away from the nucleus.

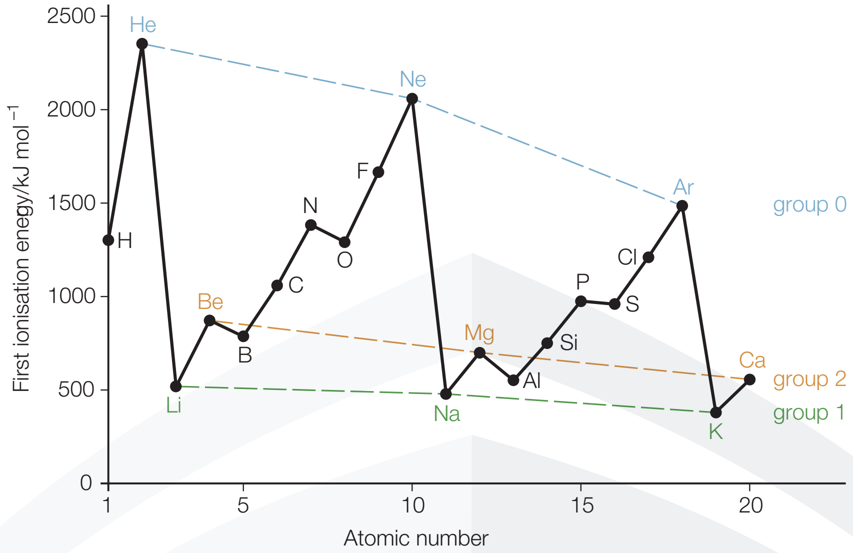
1ST IONISATION ENERGIES OF PERIOD 3



There is a DROP in the value for Sulfur. The extra electron has paired up with one of the electrons already in one of the 3p orbitals. The repulsive force between the two paired-up electrons means that less energy is required to remove one of them.

Also, Phosphorus, has a more stable half-filled 3p configuration.

IONISATION ENERGIES OF ELEMENTS



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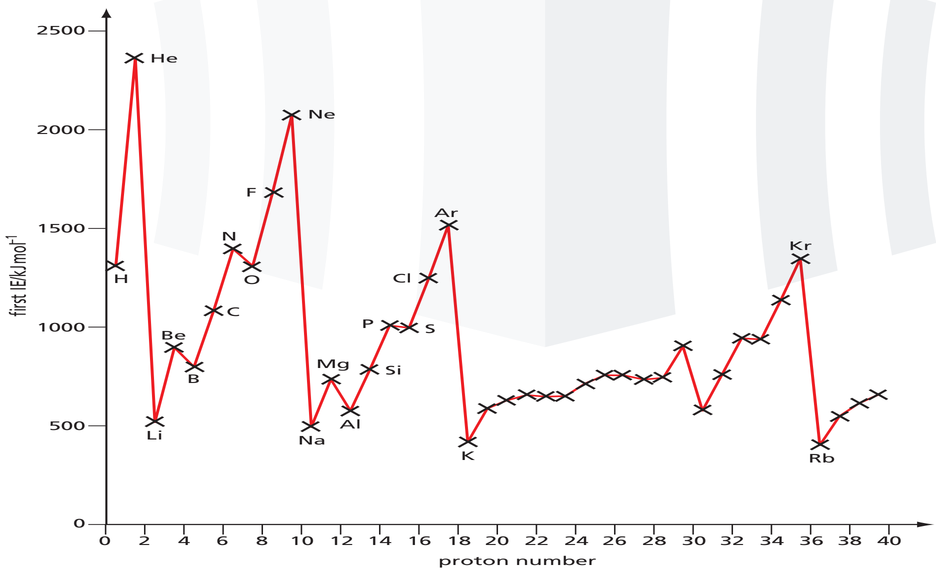
IONISATION ENERGIES OF ELEMENTS



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SKILL CHECK 7

Use of the Data Booklet is relevant to this question.

The elements radon (Rn), francium (Fr) and radium (Ra) have consecutive proton numbers in the Periodic Table.

What is the order of their first ionisation energies?

	least endothermic	→	most endothermic
A	Fr	Ra	Rn
B	Fr	Rn	Ra
C	Ra	Fr	Rn
D	Rn	Ra	Fr

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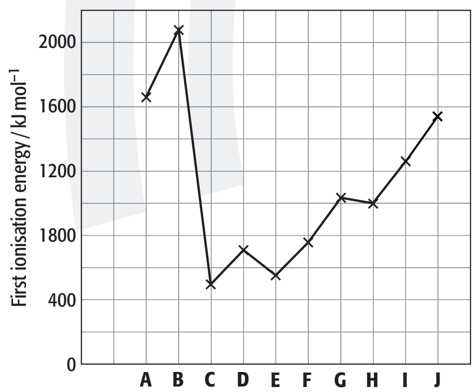
SKILL CHECK 8

The 1st ionisation energies of several elements with consecutive atomic numbers are shown in the graph below. The letters are not the symbols of the elements.



a. Which of the elements A to I belong to Group I in the Periodic Table? Explain your answer.

b. Which of the elements A to I could have the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2$?

c. Explain the rise in 1st ionisation energy between element E and element G.

d. Estimate the 1st ionisation energy of element J.

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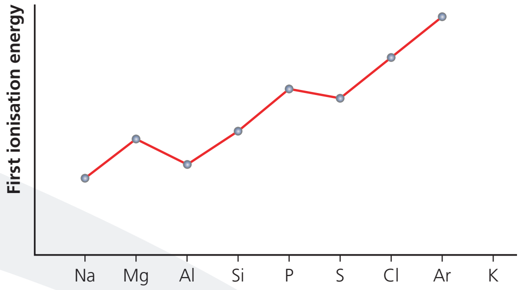
SKILL CHECK 9

The first ionisation of elements sodium to argon is shown below.

A Explain why the general trend from sodium to argon is upwards but why the value for sulfur is less than that for phosphorus.

B Mark on the graph where the value for potassium would be.



C Explain why the value for the second ionisation of sodium is very much larger than that of its first ionisation.

SKILL CHECK 10

Which experience a greater effective nuclear charge: the valence electrons in beryllium or the valence electrons in nitrogen? Why?