

Structure of the Periodic Table

- The periodic table is a list of all known elements arranged in order of increasing atomic number, from 1 to 118.
- In addition to this, the elements are arranged in such a way that atoms with the same number of shells are placed together, and atoms with similar electronic configurations in the outer shell are also placed together. This is achieved as follows:
 - The elements are arranged in rows and columns.
 - Elements with one shell are placed in the first row (i.e. H and He)
 - Elements with two shells are placed in the second row (Li to Ne) and so on.
- A row of elements thus arranged is called a period. The period number, n , is the outer energy level that is occupied by electrons.
- In addition, the elements are aligned vertically (in columns) with other elements in different rows, if they share a similar outer-shell electronic configuration
 - The outer electrons are known as the valence electrons.

- A column of elements thus arranged is called a group

PERIODIC TABLE OF THE ELEMENTS

ALKA LI METALS ALKA LINE EARTH METALS NOBLE GASES

1/I 1/I 2/II 3/III 4/IV 5/V 6/VI 7/VII 0/VIII

1 H 2 He

1 1 1 2 3 4 5 6 7 8 9 10

7 Li Be B C N O F Ne

3 LITHIUM BERYLLIUM BORON CARBON NITROGEN OXYGEN FLUORINE NEON

23 Na Mg Al Si P S Cl Ar

11 SODIUM MAGNESIUM ALUMINIUM SILICON PHOSPHORUS SULPHUR CHLORINE ARGON

39 40 45 48 51 52 55 56 59 59 64 65 70 73 75 79 80 84

19 K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr

20 POTASSIUM CALCIUM SCANDIUM TITANIUM VANADIUM CHROMIUM MANGANESE IRON COBALT NICKEL COPPER ZINC GALLIUM GERMANIUM ARSENIC SELENIUM BROMINE KRYPTON

85 88 89 91 93 96 (99) 101 103 106 108 112 115 119 122 128 127 131

37 Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe

38 STRONTIUM YTTORIUM ZIRCONIUM NI OBIUM MOLYBDENUM TECHNETIUM RUTHENIUM RHODIUM PALLADIUM SILVER CADMIUM INDIUM TIN ANTIMONY TELLURIUM IODINE XENON

55 56 57 72 73 74 75 76 77 78 79 80 81 82 83 84 85 (88) (88)

133 Cs Ba La Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn

137 CAESIUM BARIUM LANTHANUM HAFNIUM TANTALUM TUNGSTEN RHENIUM OSMIUM IRI DIUM PLATINUM GOLD MERCURY THALIUM LEAD BISMUTH POLONIUM ASTATINE RADON

87 88 89 104 105 106

(223) Fr Ra Ac Unq Unp Unh

(226) FRANCIUM RADIUM ACTINIUM UNBIQUADIUM UNBIPENTIUM UNHEXHEMIUM

140 144 (147) 150 152 157 159 162 165 167 169 173 175

58 Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu

59 CERIUM PRASEODYMIUM NEODYMIUM PROMETHIUM SAMARIUM EUROPIUM GADOLINIUM TERBIUM DYSPROSIUM HO LMIUM ERBIUM THULIUM YTTERIUM LUTETIUM

(232) (231) (238) (237) (242) (243) (247) (251) (254) (253) (256) (254) (257)

90 Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

91 THORIUM PROTACTINIUM URANIUM NEPTUNIUM PLUTONIUM AMERICIUM CURIUM BERKELIUM CALIFORNIUM EINSTEINIUM FERMIUM MENDELEVIUM NOBELIUM LAWRENCIUM

LANTHANIDE ELEMENTS

ACTINIDE ELEMENTS

KEY

MASS NUMBER → 4

ELEMENT SYMBOL → He

ATOMIC NUMBER → 2

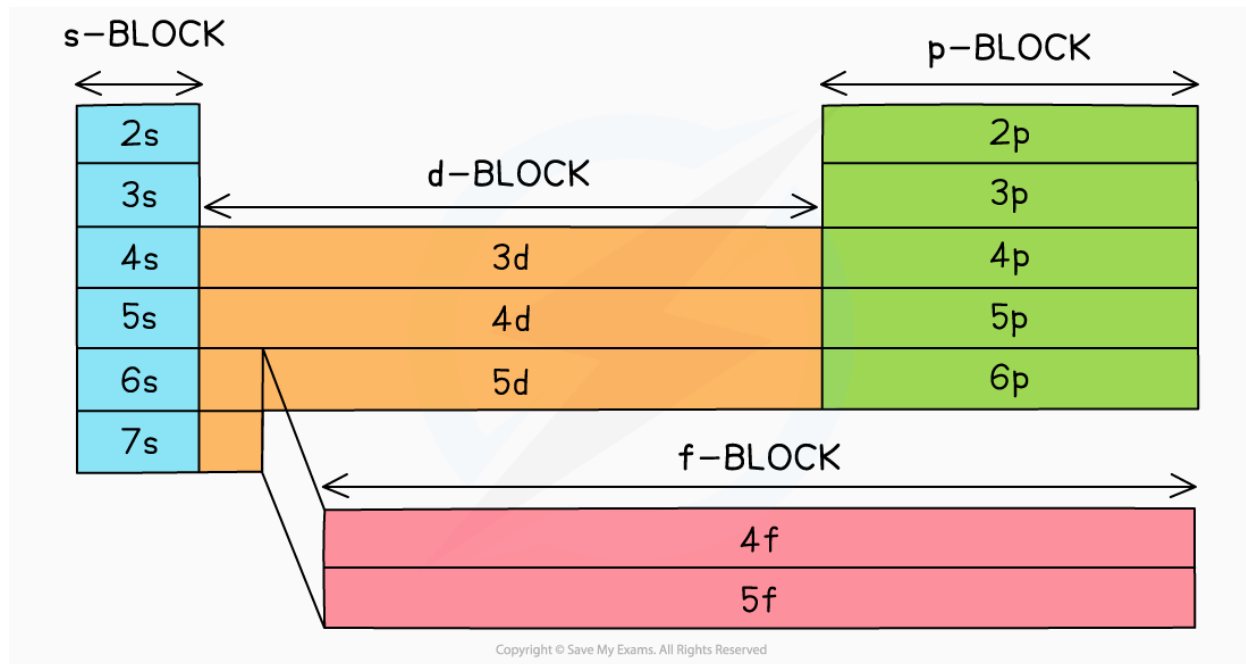
NAME → HELIUM

AT ROOM TEMPERATURE

METALS NON-METALS-SOLID NON-METALS-GAS LIQUIDS METALLOIDS

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- Since the electronic configurations of H and He are unusual, they do not fit comfortably into any group. They are thus allocated a group based on similarities in physical and chemical properties with other members of the group
- He is placed in group 0 on this basis, but hydrogen does not behave like any other element and so is placed in a group of its own



The blocks of the periodic table

- All elements belong to one of four main blocks: the s-block, the p-block, the d-block and the f-block
 - The s-block elements are all those with only s electrons in the outer shell
 - The p-block elements are all those with at least one p-electron in the outer shell
 - The d-block elements are all those with at least one d-electron and at least one s-electron but no f or p electrons in the outer shell (up to 5d)
 - The f-block elements are all those with at least one f-electron and at least one s-electron but no d or p electrons in the outer shell
- The physical and chemical properties of elements in the periodic table show clear patterns related to the position of each element in the table
 - Elements in the same group show similar properties, and properties change gradually as you go across a period
- As atomic number increases, the properties of the elements show trends which repeat themselves in each period of the periodic table
 - These trends are known as periodic trends and the study of these trends is known as periodicity

Deducing Electron Configurations

- The electron configuration of any element can be deduced from its position in the periodic table
- It is like an 'address' that tells you exactly where an element is found
- Using the element germanium as an example to illustrate how it works:

GROUP 4 TELLS YOU THERE ARE
4 VALENCE ELECTRONS AND PERIOD 4
TELLS YOU THEY ARE IN THE 4th SHELL

(FULL INNER SHELLS) + $4s^2 4p^2$

THE SECOND ELEMENT IN THE p-BLOCK
TELLS YOU TWO ELECTRONS IN THE
p-SUBSHELL

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Deducing the electron configuration of germanium

- Germanium is in p block, in group 4 (using the simplify numbering system) and the second element across in period 4
- Group 4 tells you there are four valence electrons and period 4 tells you the valence electrons are in the fourth shell
- The second position in p block tell you that two electrons are in the p subshell
- Similarly, you can deduce the position of an element in the periodic table from its electron configuration:

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

ADD THESE NUMBERS TO
GIVE YOU THE GROUP

$$2 + 5 = 7$$

GROUP 7, THIRD PERIOD,
MEANS IT IS CHLORINE

THIS NUMBER TELLS YOU
THE PERIOD NUMBER WHICH
IS THE OUTER ENERGY LEVEL (n)

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Deducing information from the electron configuration of chlorine

- Test your understanding in the following example:

Worked example

Element Z is in period 4 and group 5 of the periodic table. Which statement is correct?

- Z has 5 occupied energy levels.
- Z can form ions with 3- charge.
- Z is a transition element.

D. Z has 4 valence electrons.

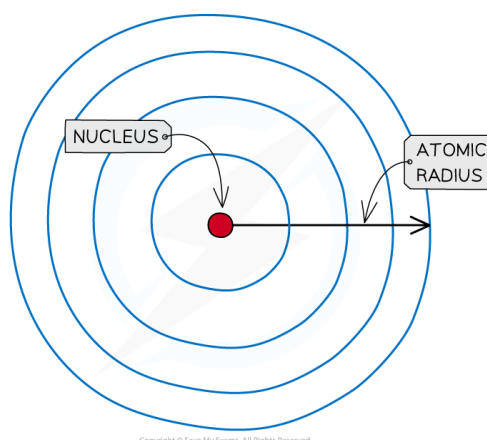
Answer:

- The correct option is B. A group 5 element could form a 3- ion
- It cannot be A because 5 occupied energy levels would place it in period 5
- Transition elements would not be found in group 5, so C is also incorrect
- 4 valence electrons would match an element in group 4, so D must be wrong

Atomic & Ionic Radius

Atomic radius

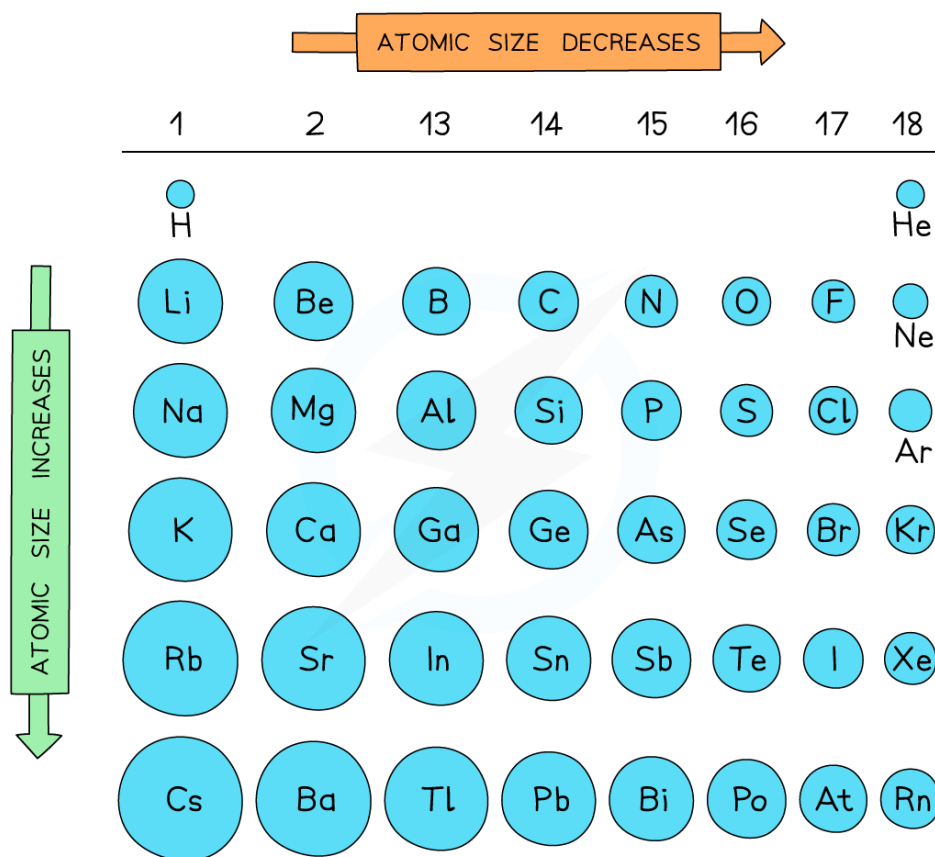
- The atomic radius of an element is a measure of the size of an atom
- It is the distance between the nucleus of an atom and the outermost electron shell
- It can be quite hard to determine exactly where the boundary of an atom lies, so a variety of approaches are taken such as half the mean distance between two adjacent atoms
- This will vary depending on the type of structure and bonding, but it gives a comparative value for atoms



The atomic radius of an atom is the typical distance between the nucleus and the outermost electron shell

- Atomic radii show predictable patterns across the periodic table
 - They generally decrease across each period
 - They generally increase down each group
- These trends can be explained by the electron shell theory
 - Atomic radii decrease as you move across a period as the atomic number increases (increased positive nuclear charge) but at the same time extra electrons are added to the same principal quantum shell
 - The larger the nuclear charge, the greater the pull of the nuclei on the electrons which results in smaller atoms

- Atomic radii increase moving down a group as there is an increased number of shells going down the group
- The electrons in the inner shells repel the electrons in the outermost shells, shielding them from the positive nuclear charge
- This weakens the pull of the nuclei on the electrons resulting in larger atoms



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Trends in the atomic radii across a period and down a group

- The diagram shows that the atomic radius increases sharply between the noble gas at the end of each period and the alkali metal at the beginning of the next period
- This is because the alkali metals at the beginning of the next period have one extra principal quantum shell
 - This increases shielding of the outermost electrons and therefore increases the atomic radius

Ionic radius

- The ionic radius of an element is a measure of the size of an ion
- Ionic radii show predictable patterns
 - The trend down a group is the same as atomic radius - it increases as the number of shells increases

- The trend across a period is not so straightforward as it depends on whether it is positive or negative ions being considered
- Ionic radii increase with increasing negative charge
- Ionic radii decrease with increasing positive charge
- These trends can also be explained by the electron shell theory
 - Ions with negative charges are formed by atoms accepting extra electrons while the nuclear charge remains the same
 - The extra electrons experience repulsion with the other valence electrons which increases the ionic radius
 - The greater the negative charge, the larger the ionic radius
 - Positively charged ions are formed by atoms losing electrons
 - The nuclear charge remains the same but there are now fewer electrons which undergo a greater electrostatic force of attraction towards the nucleus which decreases the ionic radius
 - The greater the positive charge, the smaller the ionic radius

		SIZES OF ATOMS AND THEIR IONS IN PM				
		GROUP 1	GROUP 2	GROUP 13	GROUP 16	GROUP 17
PERIOD 2		Li^+ Li 90 134	Be^{2+} Be 59 90	B^{3+} B 41 82	O^{2-} O 73 126	F^- F 71 119
PERIOD 3		Na^+ Na 116 154	Mg^{2+} Mg 86 130	Al^{3+} Al 68 118	S^{2-} S 102 170	Cl^- Cl 99 167
PERIOD 4		K^+ K 152 196	Ca^{2+} Ca 114 174	Ga^{3+} Ga 76 126	Se^{2-} Se 116 184	Br^- Br 114 182
PERIOD 5		Rb^+ Rb 166 211	Sr^{2+} Sr 132 192	In^{3+} In 94 144	Te^{2-} Te 135 207	I^- I 133 206

Trends in the ionic radii across a period and down a group

Worked example

Which option shows atoms in order of decreasing atomic radius?

- A. $\text{N} > \text{C} > \text{Be} > \text{Mg}$
- B. $\text{Mg} > \text{N} > \text{C} > \text{Be}$
- C. $\text{Be} > \text{C} > \text{N} > \text{Mg}$
- D. $\text{Mg} > \text{Be} > \text{C} > \text{N}$

Answer:

Option D is the correct answer

- First, you need to identify that Be, C and N are all in period 2, but Mg is in period 3, so Mg will have the biggest radius.
- Secondly, atomic radius decreases across the period so Be, C and N decrease in that order as they belong to groups 2, 14 and 15, respectively

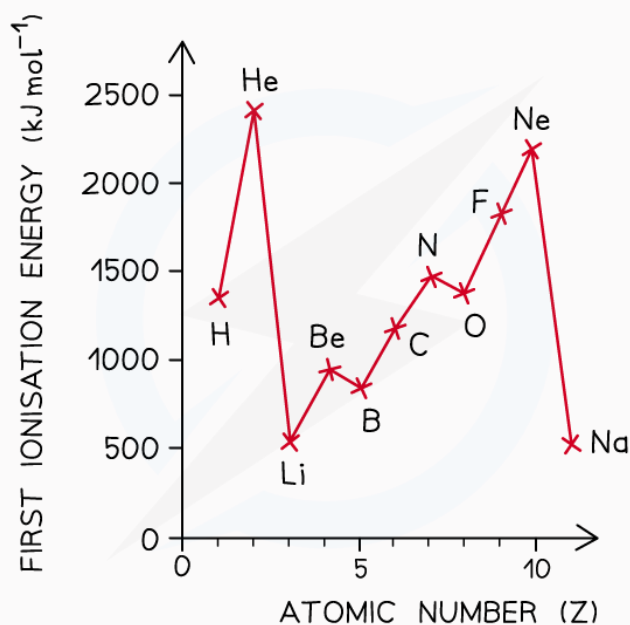
First Ionisation Energy

- The ionisation energy (IE) of an element is the amount of energy required to remove one mole of electrons from one mole of atoms of an element in the gaseous state to form one mole of gaseous ions
- Ionisation energies are measured under standard conditions which are 298 K and 100 kPa
- The units of IE are kilojoules per mole (kJ mol^{-1})
- The first ionisation energy is the energy required to remove the one mole of electrons from one mole the gaseous atoms
 - E.g. the first ionisation energy of calcium:



Ionisation Energies: Trends

- Ionisation energies show periodicity
- As could be expected from their electronic configuration, the group I metals show low IE whereas the noble gases have very high IEs
- The first ionisation energy increases across a period and decreases down a group and is caused by four factors that influence the ionisation energy:
 - Size of the nuclear charge: the nuclear charge increases with increasing atomic number, which means that there are greater attractive forces between the nucleus and outer electrons, so more energy is required to overcome these attractive forces when removing an electron
 - Distance of outer electrons from the nucleus: electrons in shells that are further away from the nucleus are less attracted to the nucleus so the further the outer electron shell is from the nucleus, the lower the ionisation energy
 - Shielding effect of inner electrons: the shielding effect is when the electrons in full inner shells repel electrons in outer shells preventing them to feel the full nuclear charge so the greater the shielding of outer electrons by inner electron shells, the lower the ionisation energy
 - Spin-pair repulsion: paired electrons in the same atomic orbital in a subshell repel each other more than electrons in different atomic orbitals; this makes it easier to remove an electron (which is why the first ionization energy is always the lowest)



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A graph showing the ionisation energies of the elements hydrogen to sodium

Ionisation energy across a period

- The ionisation energy across a period increases due to the following factors:
 - Across a period the nuclear charge increases
 - The distance between the nucleus and outer electron remains reasonably constant
 - The shielding by inner shell electrons remains the same
- There is a rapid decrease in ionisation energy between the last element in one period and the first element in the next period caused by:
 - The increased distance between the nucleus and the outer electrons
 - The increased shielding by inner electrons
 - These two factors outweigh the increased nuclear charge
- There is a slight decrease in 1st I.E. between beryllium and boron as the fifth electron in boron is in the 2p subshell which is further away from the nucleus than the 2s subshell of beryllium
 - Beryllium has a first ionisation energy of 900 kJ mol⁻¹ as its electron configuration is 1s² 2s²
 - Boron has a first ionisation energy of 801 kJ mol⁻¹ as its electron configuration is 1s² 2s² 2p¹
- There is a slight decrease in 1st I.E. between nitrogen and oxygen due to spin-pair repulsion in the 2p subshell of oxygen
 - Nitrogen has a first ionisation energy of 1402 kJ mol⁻¹ as its electron configuration is 1s² 2s² 2p³
 - Oxygen has a first ionisation energy of 1314 kJ mol⁻¹ as its electron configuration is 1s² 2s² 2p⁴

Ionisation energy down a group

- Although going down a group the nuclear charge increases, the ionisation energy down a group decreases and it is due to the following factors:
 - The distance between the nucleus and outer electron increases
 - The shielding by inner shell electrons increases
 - The effective nuclear charge is decreasing as shielding increases

Ionisation Energy Trends across a Period & going down a Group Table

Across a Period: Ionisation Energy Increases	Down a Group: Ionisation Energy Decreases
Increase in nuclear charge	Increase in nuclear charge
Shell number is the same Distance of outer electron to nucleus is the same	Increase in shells Distance of outer electron to nucleus increases Shielding effect increases, therefore, the attraction of valence electrons to the nucleus decreases
Shielding remains reasonably constant	Increased shielding
Decreased atomic/ionic radius	Increased atomic/ionic radius
The outer electron is held more tightly to the nucleus so it gets harder to remove it	The outer electron is held more loosely to the nucleus so it gets easier to remove it

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Successive ionisation energies of an element

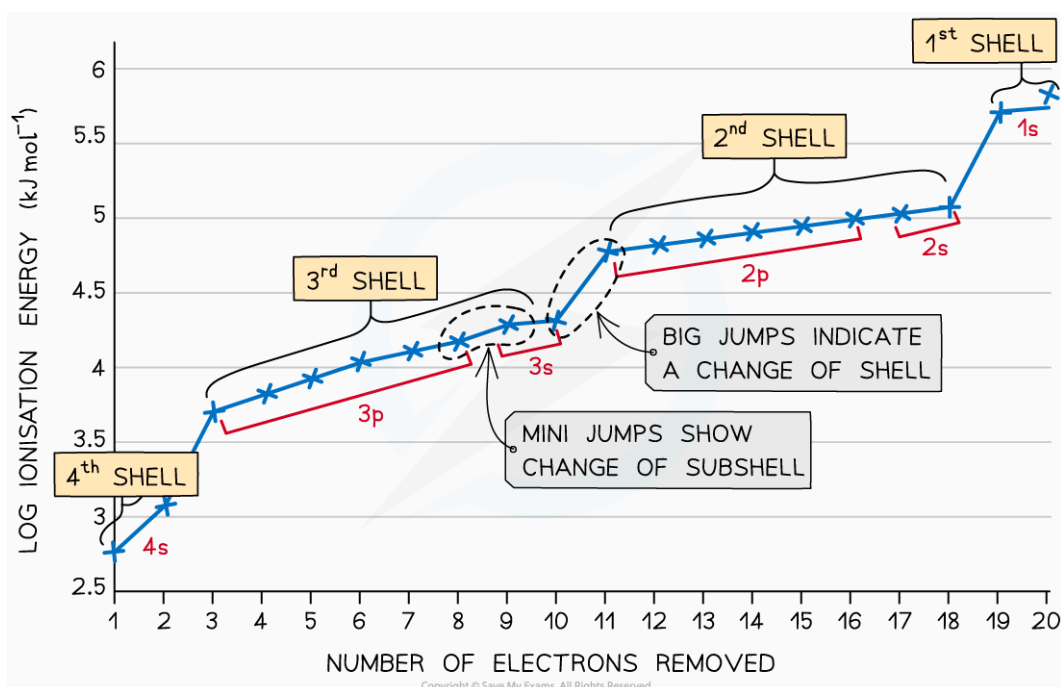
- The successive ionisation energies of an element increase as removing an electron from a positive ion is more difficult than from a neutral atom
- As more electrons are removed the attractive forces increase due to decreasing shielding and an increase in the proton to electron ratio
- The increase in ionisation energy, however, is not constant and is dependent on the atom's electronic configuration
- Taking calcium as an example:

Ionisation Energies of Calcium Table

Electronic Configuration	$1s^2 2s^2 2p^6$ $3s^2 3p^6 4s^2$	$1s^2 2s^2 2p^6$ $3s^2 3p^6 4s^1$	$1s^2 2s^2 2p^6$ $3s^2 3p^6$	$1s^2 2s^2 2p^6$ $3s^2 3p^5$
IE	First	Second	Third	Fourth
IE (kJ mol^{-1})	590	1150	4940	6480

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- The values become very large and difficult to represent meaningfully, so it is more convenient to show the logarithm of the ionisation energies
- This helps us to see significant jumps in I.E.



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Successive ionisation energies for the element calcium

- The first electron removed has a low ionisation energy as it is easily removed from the atom due to the spin-pair repulsion of the electrons in the 4s orbital
- The second electron is a little more difficult to remove than the first electron as you are removing an electron from a positively charged ion
- The third electron is much more difficult to remove than the second one corresponding to the fact that the third electron is in a principal quantum shell which is closer to the nucleus (3p)
- The graph shows there is a large increase in successive ionisation energy as the electrons are being removed from an increasingly positive ion
- The big jumps on the graph show the change of shell and the small jumps are the change of subshell

Exam Tip

Be careful with how you interpret successive ionisation energy graphs as it is common for students to read them the wrong way around and count outer electrons from right to left instead of left to right so they get the jumps in the wrong place. This happens particularly when you are given only a partial successive ionisation energy graph and have to deduce which group the element comes from. It's a good idea if you see an ionisation energy graph in an exam question to label the shells and subshells so you are less likely to make this mistake!

Electron Affinity

- When atoms gain electrons they become negative ions or anions
- Electron affinity (EA) can be thought of as the opposite process of ionisation energy and is defined as
 - The amount of energy released when one mole of electrons is gained by one mole of atoms of an element in the gaseous state to form one mole of gaseous ions
- Electron affinities are measured under standard conditions which are 298 K and 100 kPa
- The units of EA are kilojoules per mole (kJ mol^{-1})
- The first electron affinity is always exothermic
 - E.g. the first electron affinity of chlorine is:

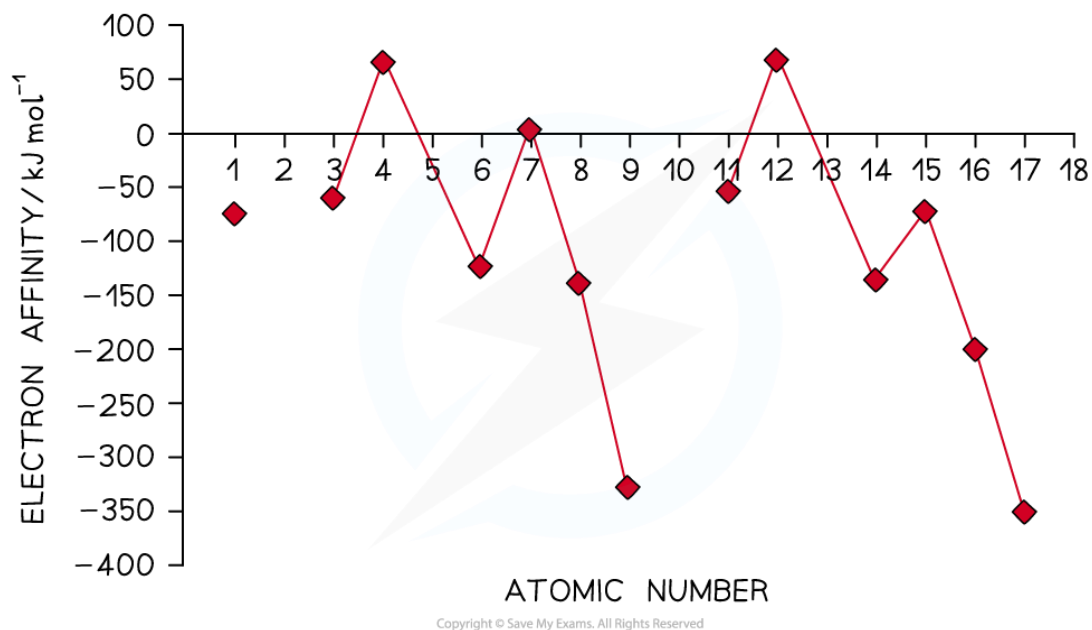


- However, the second electron affinity can be an endothermic process



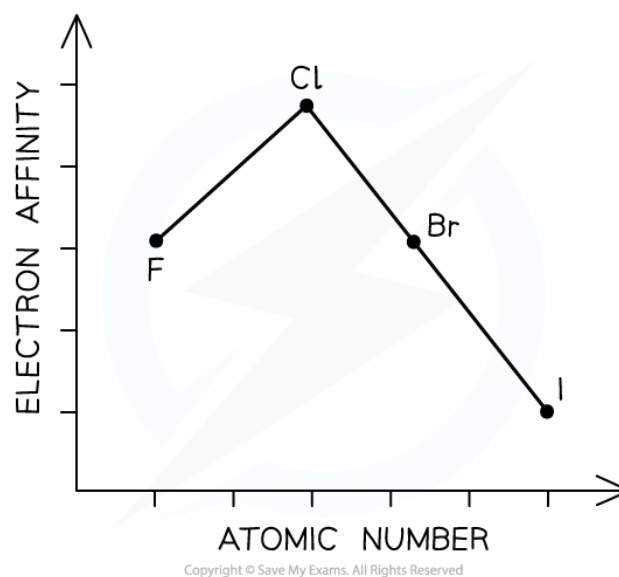
- This is due to the fact that you are overcoming repulsion between the electron and a negative ion, so energy is required making the process endothermic overall

Trends in electron affinity



Electron affinities across a period

- Electron affinities show periodicity
- The pattern is very similar to ionisation energies, except that it is inverted and the minimum points are displaced one element to the right
- As might be expected, the most exothermic electron affinities are for group 17 elements which also have the highest electronegativities
- The strongest pull on electrons correlates with the greater amount of energy released when negative ions are formed
- Noble gases do not form negative ions, so they don't appear in this chart
- The electron affinities reach a peak for group 2 and group 5 elements

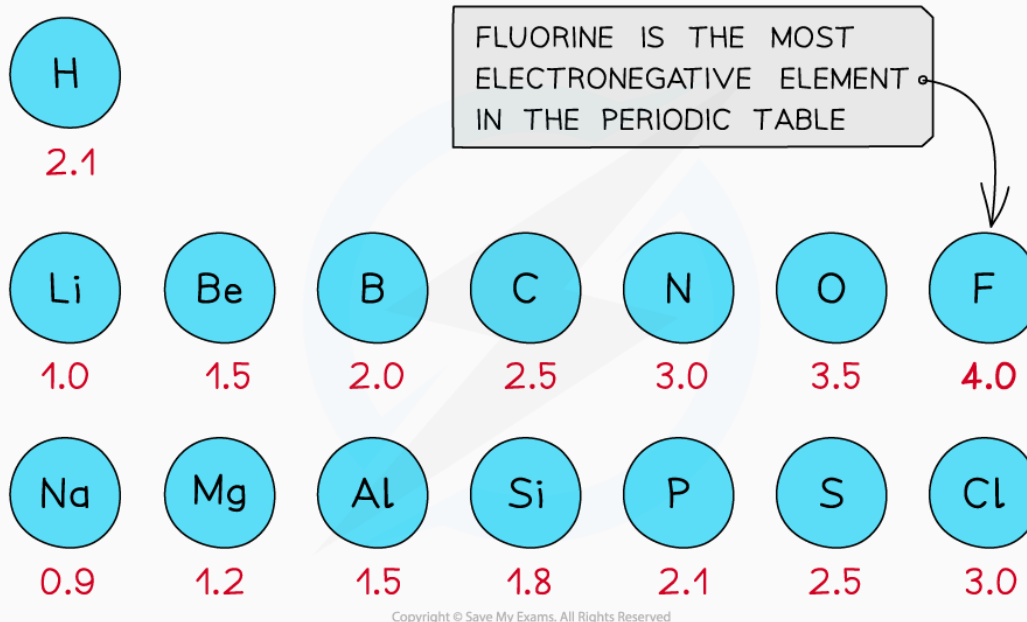


Electron affinities down a group

- Electron affinities generally decrease down a group
- As the atoms become larger the attraction for an additional electron is less, since the effective nuclear charge is reduced due to increased shielding
- Electron affinity become less exothermic going down the group
- An exception to this is fluorine whose electron affinity is smaller than expected
- This is because fluorine is such a small atom and an additional electron in the 2p subshell experiences considerable repulsion with the other valence electrons

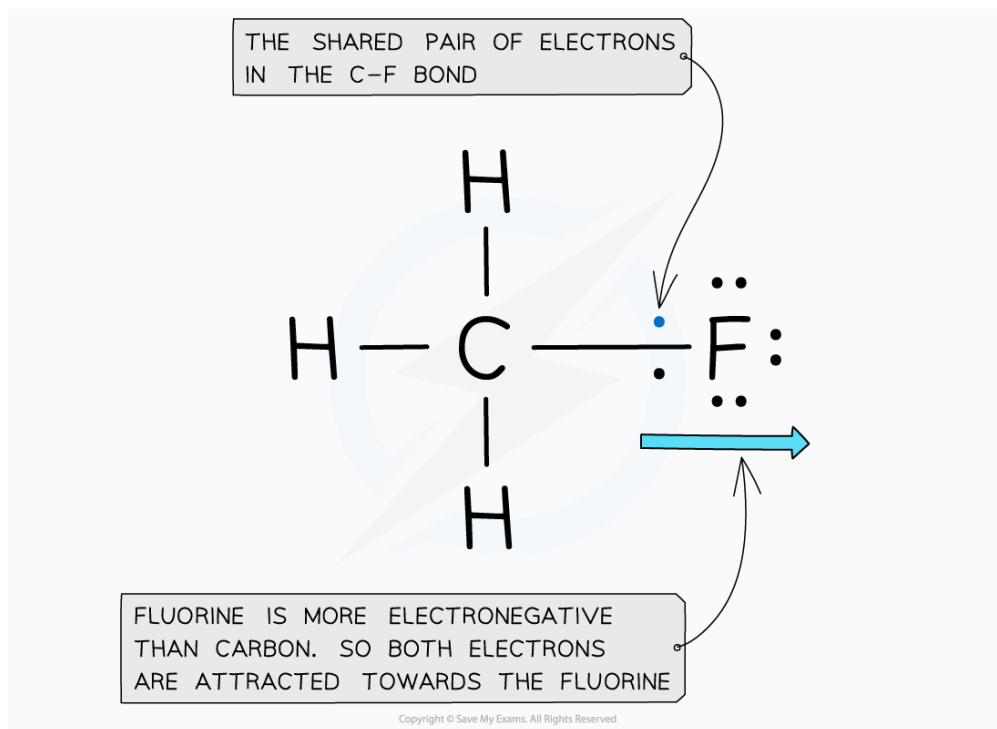
Electronegativity: Definition

- Electronegativity is the ability of an atom to attract a pair of electrons towards itself in a covalent bond
- This phenomenon arises from the positive nucleus's ability to attract the negatively charged electrons, in the outer shells, towards itself
- The Pauling scale is used to assign a value of electronegativity for each atom



First three rows of the periodic table showing electronegativity values

- Fluorine is the most electronegative atom on the Periodic Table, with a value of 4.0 on the Pauling Scale
- It is best at attracting electron density towards itself when covalently bonded to another atom



Electron distribution in the C-F bond of fluoromethane

Electronegativity: Affecting Factors

Nuclear charge

- Attraction exists between the positively charged protons in the nucleus and negatively charged electrons found in the energy levels of an atom
- An increase in the number of protons leads to an increase in nuclear attraction for the electrons in the outer shells
- Therefore, an increased nuclear charge results in an increased electronegativity

	Na	Mg	Al
NUMBER OF PROTONS:	11	12	13
ELECTRONEGATIVITY:	0.9	1.2	1.5

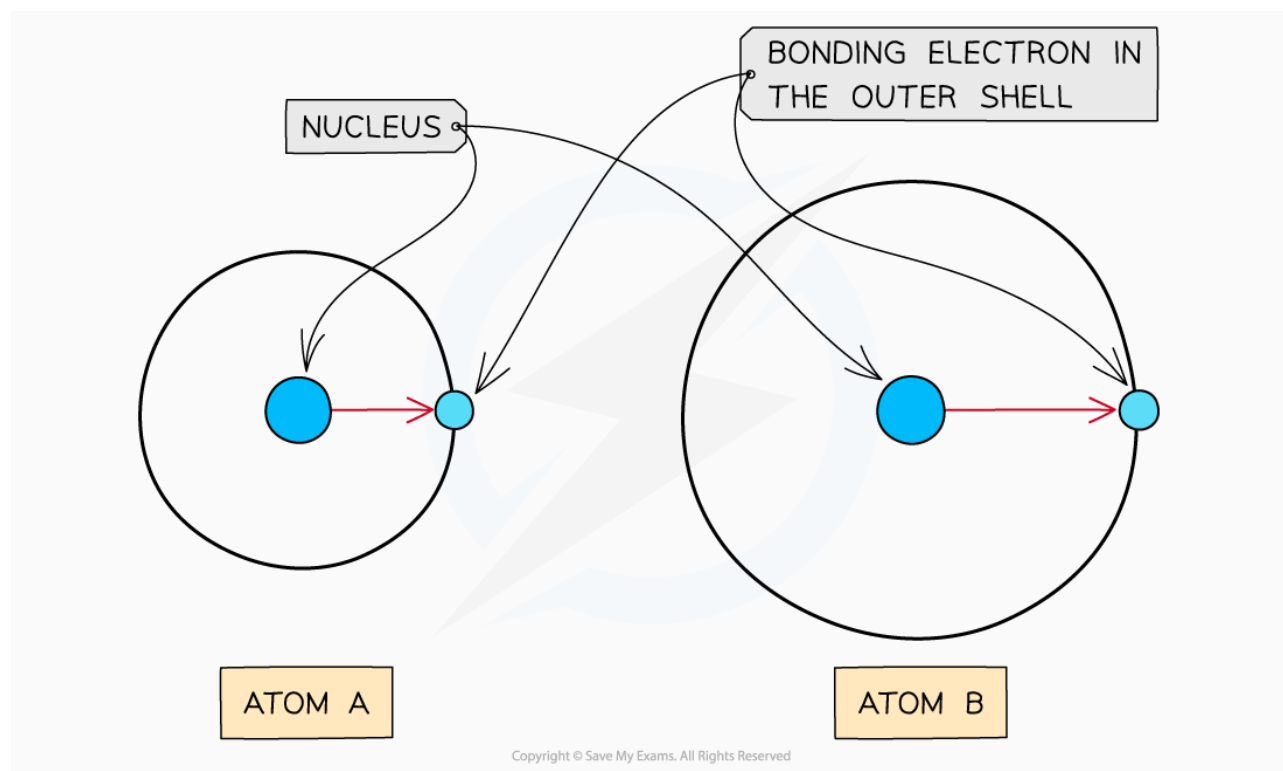
AS THE NUMBER OF PROTONS INCREASE, THE ELECTRONEGATIVITY INCREASES

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As the nuclear charge increases, the electronegativity of an element increases as well

Atomic radius

- The atomic radius is the distance between the nucleus and electrons in the outermost shell
- Electrons closer to the nucleus are more strongly attracted towards its positive nucleus
- Those electrons further away from the nucleus are less strongly attracted towards the nucleus
- Therefore, an increased atomic radius results in a decreased electronegativity



As the atomic radius increases, the nucleus has less of an attraction for the bonding electrons causing atom A to have a higher electronegativity than atom B

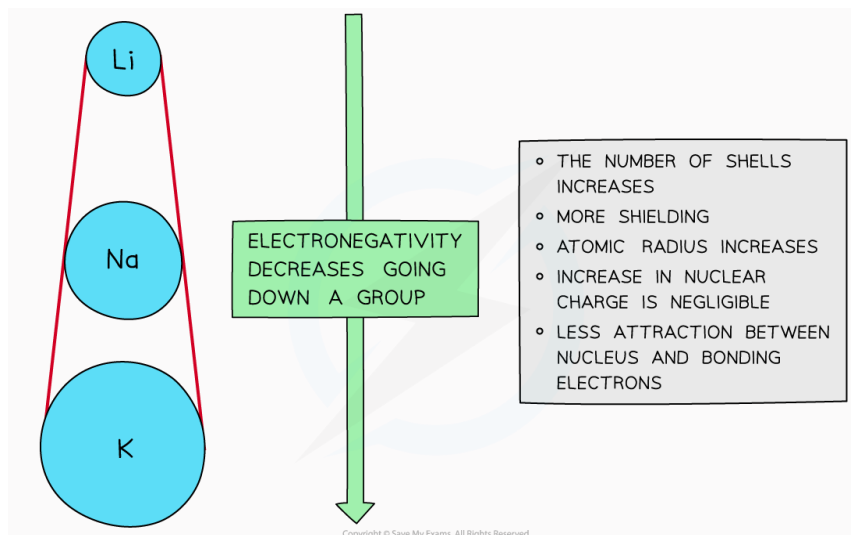
Electronegativity: Trends

- Electronegativity varies across periods and down the groups of the periodic table

Down a group

- There is a decrease in electronegativity going down the group
- The nuclear charge increases as more protons are being added to the nucleus
- However, each element has an extra filled electron shell, which increases shielding
- The addition of the extra shells increases the distance between the nucleus and the outer electrons resulting in larger atomic radii

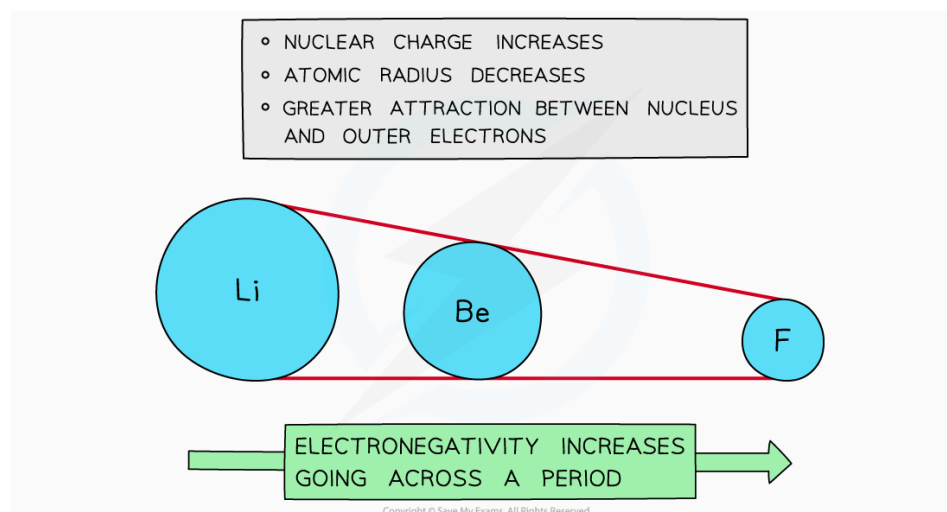
- Overall, there is decrease in attraction between the nucleus and outer bonding electrons
- We say the effective nuclear charge has decreased down the group



Electronegativity decreases going down the groups of the periodic table

Across a period

- Electronegativity increases across a period
- The nuclear charge increases with the addition of protons to the nucleus
- Shielding remains the same across the period as no new shells are being added to the atoms
- The nucleus has an increasingly strong attraction for the bonding pair of electrons of atoms across the period
- This results in smaller atomic radii



Electronegativity increases going across the periods of the periodic table

Trends down a Group & across a Period Table

	Down a Group	Across a Period
Nuclear charge	Increases	Increases
Shielding	Increases	Reasonably constant
Atomic radius	Increases	Decreases
Electronegativity	Decreases	Increases

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Exam Tip

Make sure you learn the definition of electronegativity and can distinguish it from electron affinity as the two are often confused. Electronegativity is about chemical character and only applies to considerations of covalent bonds whereas electron affinity is a thermodynamic value that is measurable and applies to the formation of negative ions. You may come across something called electropositivity - this is a term used to describe the character of elements to form positive ions and is useful when talking about metal atoms and metal ions