

IB Chemistry DP

YOUR NOTES



3. Periodicity

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3.1 The Periodic Table & Periodic Trends

3.1.1 The Periodic Table

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**

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PERIODIC TABLE OF THE ELEMENTS

	ALKALI METALS	ALKALINE EARTH METALS																			NOBLE GASES
1/I	1/I	2/II																			0/VIII
			1																		4
			1																		2
																					4
2/II	3	4																			10
	7	9																			18
	3	4																			18
3/III	11	12																			36
	23	24																			36
	11	12																			18
	19	20																			36
4/IV	39	40																			84
	37	38																			84
	39	40																			36
5/V	85	86																			131
	87	88																			131
	87	88																			54
6/VI	133	134																			131
	133	134																			131
	133	134																			54
7/VII	117	118																			131
	117	118																			131
	117	118																			131

LANTHANIDE ELEMENTS
ACTINIDES ELEMENTS

KEY

MASS NUMBER

ELEMENT SYMBOL

ATOMIC NUMBER

NAME

AT ROOM TEMPERATURE

METALS

NON-METALS - SOLID

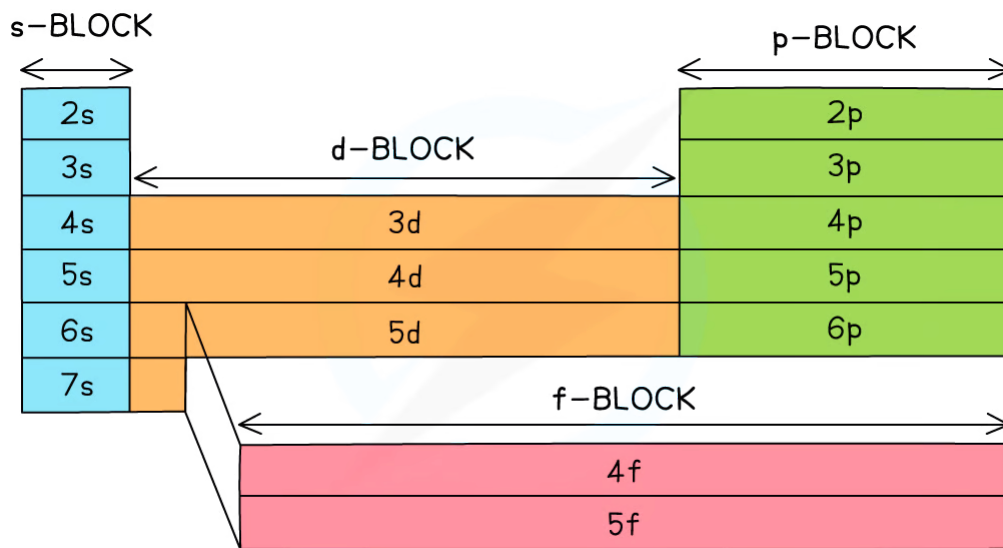
NON-METALS - GAS

LIQUIDS

METALLOID

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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



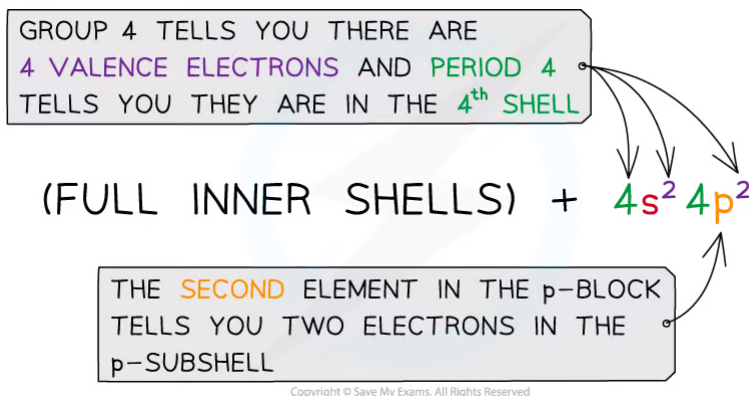
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The blocks of the periodic table

- 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**

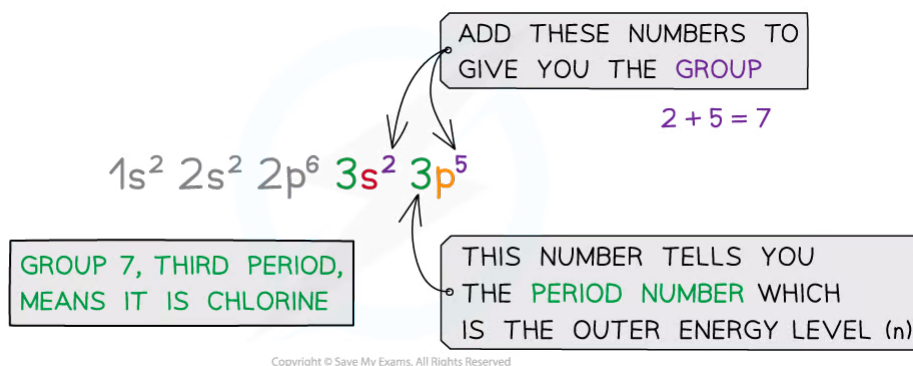
Deducing Electron Configurations

- 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:



Deducing the electron configuration of germanium

- Germanium is in **p block**, in **group 4** (using the simplified 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** tells 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:



Deducing information from the electron configuration of chlorine

- Test your understanding in the following example:

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Worked Example

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

- A. Z has 5 occupied energy levels.
- B. Z can form ions with 3- charge.
- C. 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

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3.1.2 Periodic Trends: Physical - Atomic & Ionic Radius

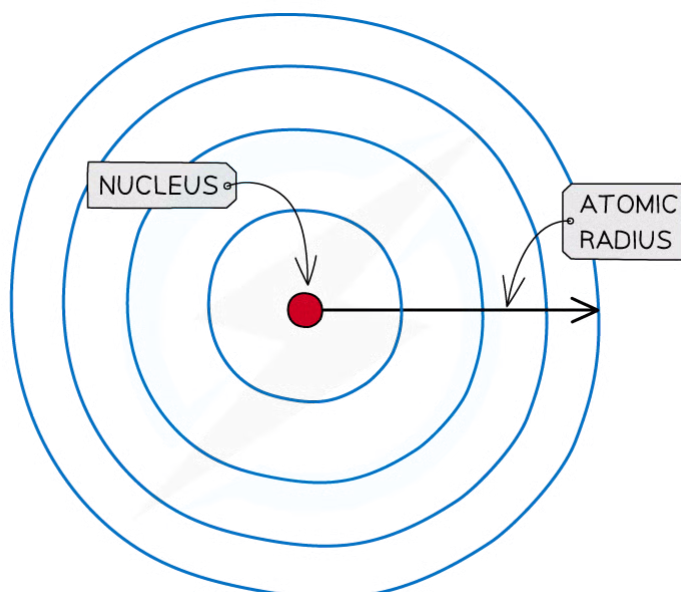
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Atomic & Ionic Radius

Atomic radius

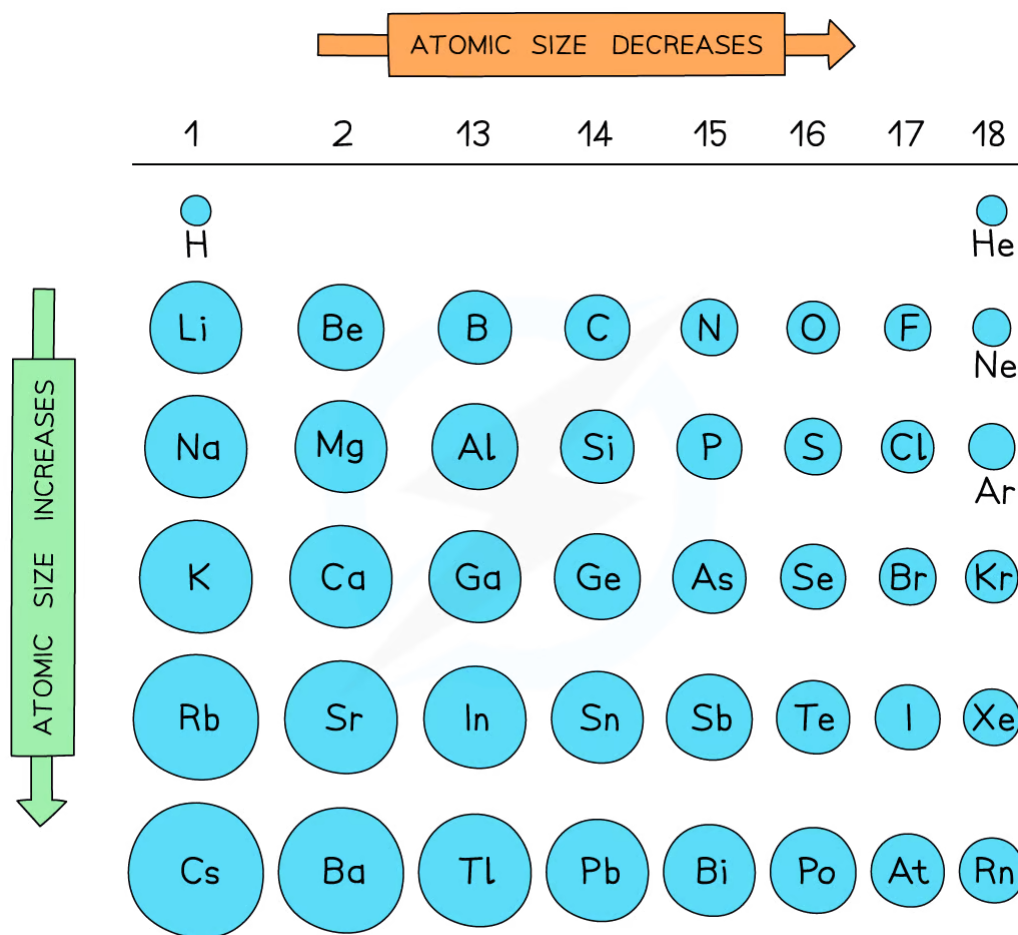
- The **atomic radius** of an element is a measure of the size of an atom
- It is the distance between the **nucleus** of an atom and the outermost electron shell
- It can be quite hard to determine exactly where the boundary of an atom lies, so a variety of approaches are taken such as half the mean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



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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

















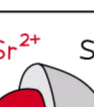
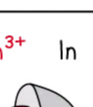
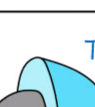
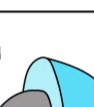
Ionic radius

- 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

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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 O^{2-}  73 126	F F^-  71 119	
PERIOD 3	Na^+ Na  116 154	Mg^{2+} Mg  86 130	Al^{3+} Al  68 118	S S^{2-}  102 170	Cl Cl^-  99 167	
PERIOD 4	K^+ K  152 196	Ca^{2+} Ca  114 174	Ga^{3+} Ga  76 126	Se Se^{2-}  116 184	Br Br^-  114 182	
PERIOD 5	Rb^+ Rb  166 211	Sr^{2+} Sr  132 192	In^{3+} In  94 144	Te Te^{2-}  135 207	I I^-  133 206	

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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. $N > C > Be > Mg$
- B. $Mg > N > C > Be$
- C. $Be > C > N > Mg$
- D. $Mg > Be > C > N$

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

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3.1.3 Periodic Trends: Physical – Ionisation Energy

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:



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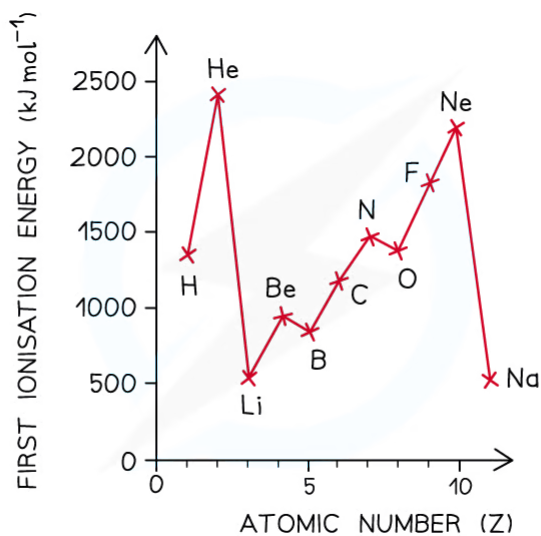


Ionisation Energies: Trends

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- 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

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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

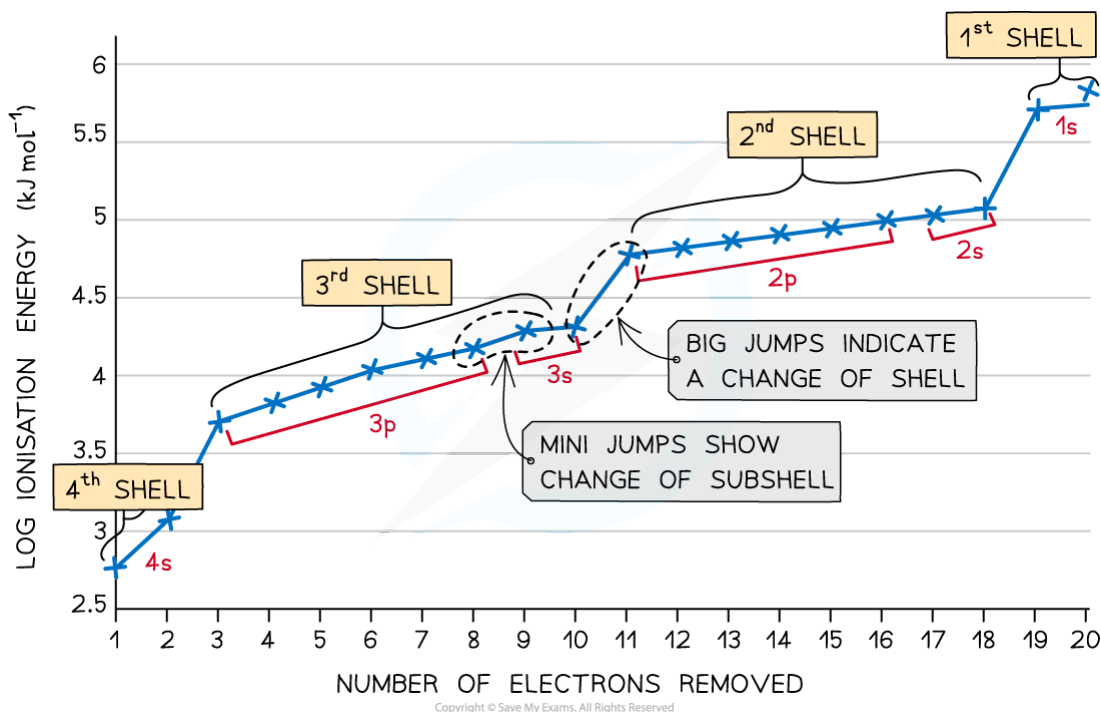
- The **successive** ionisation energies of an element **increase** as removing an electron from a **positive** ion is more difficult than from a **neutral** atom
- As more electrons are removed the attractive forces increase due to **decreasing shielding** and an increase in the proton to electron ratio
- 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!

3.1.4 Periodic Trends: Physical – Electron Affinity

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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:

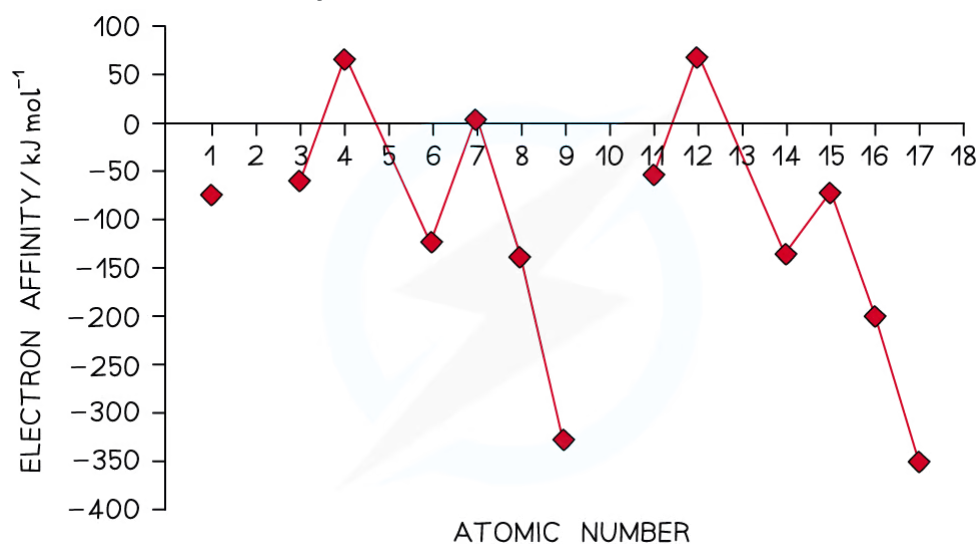


- However, the second electron affinity can be an endothermic process



- 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

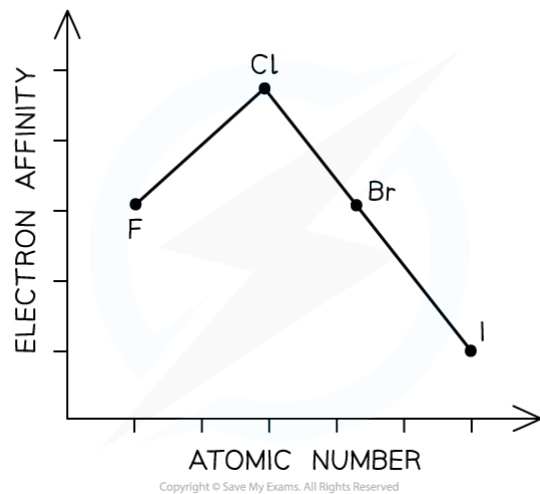
Trends in electron affinity



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

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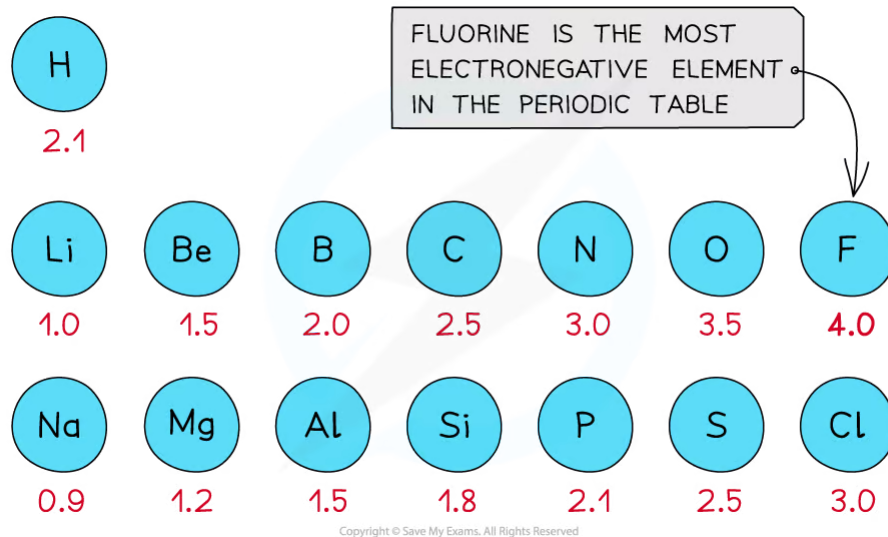
3.1.5 Periodic Trends: Physical – Electronegativity

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Electronegativity: Definition

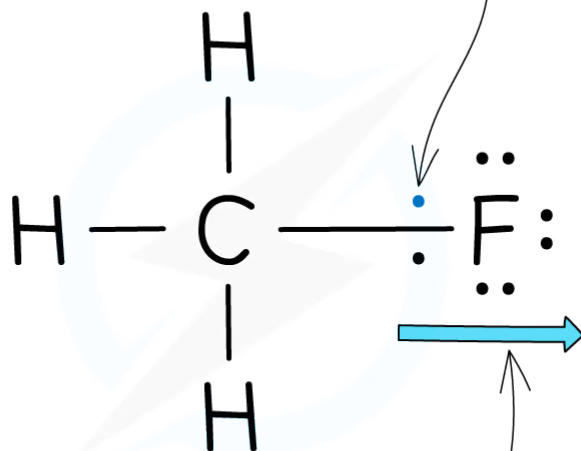
- **Electronegativity** is the ability of an atom to attract a pair of electrons towards itself in a covalent bond
- This phenomenon arises from the **positive** nucleus's ability to attract the **negatively** charged electrons, in the outer shells, towards itself
- The **Pauling scale** is used to assign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

THE SHARED PAIR OF ELECTRONS
IN THE C-F BOND



FLUORINE IS MORE ELECTRONEGATIVE
THAN CARBON. SO BOTH ELECTRONS
ARE ATTRACTED TOWARDS THE FLUORINE

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Electron distribution in the C-F bond of fluoromethane

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Electronegativity: Affecting Factors

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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

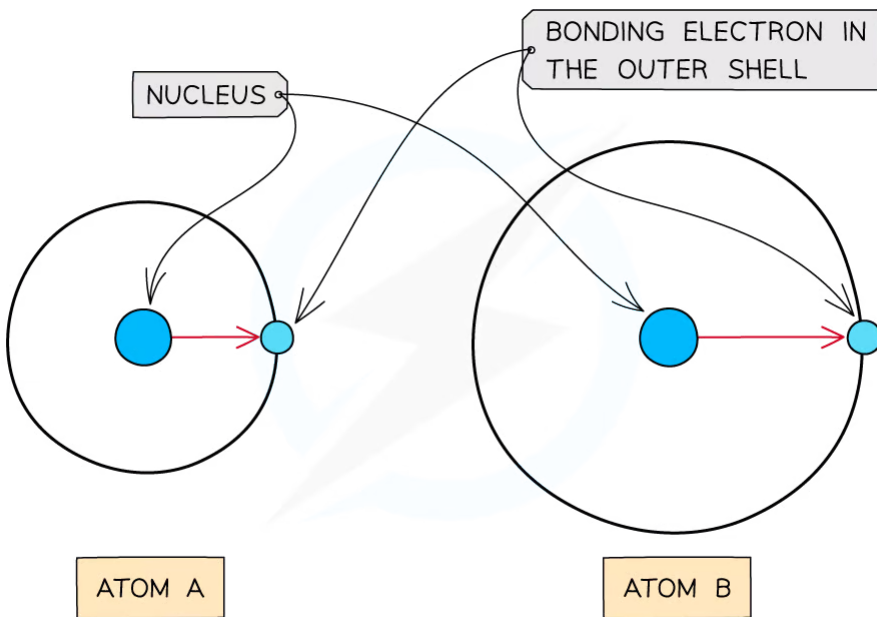
AS THE NUMBER OF PROTONS INCREASE,
THE ELECTRONEGATIVITY INCREASES

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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**



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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

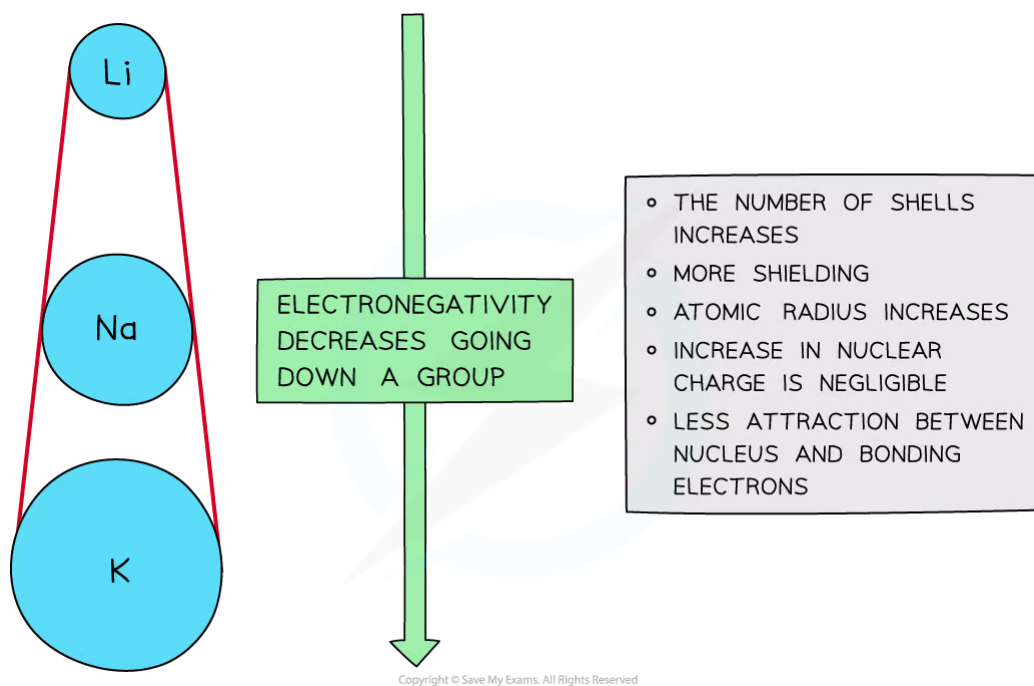
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- **Electronegativity** varies across periods and down the groups of the periodic table

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



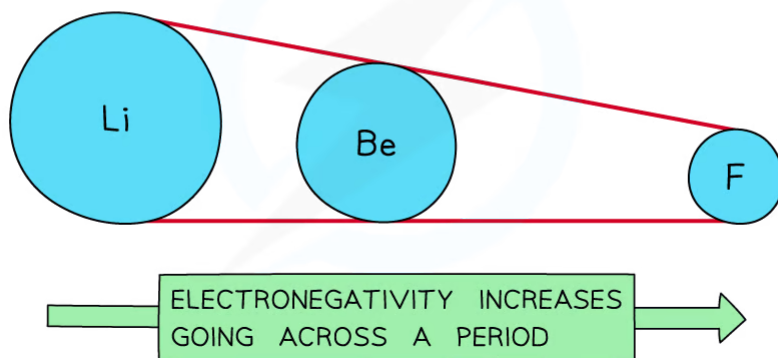
Electronegativity decreases going down the groups of the periodic table

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**



- NUCLEAR CHARGE INCREASES
- ATOMIC RADIUS DECREASES
- GREATER ATTRACTION BETWEEN NUCLEUS AND OUTER ELECTRONS



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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

3.1.6 Periodic Trends: Chemical

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Metallic & Non-metallic

- Periodic trends can also be seen in the typical properties of metals and non-metals which are summarised below:

Property	Metals	Non-metals
Electron arrangement	1–3 (more in Periods 5 & 6) outer shell electrons	4–7 electrons in the outer shell
Bonding	Metallic due to loss of outer shell electrons	Covalent by sharing of outer shell electrons
Electrical conductivity	Good conductors of electricity	Poor conductors of electricity
Type of oxide	Basic oxides (a few are amphoteric)	Acidic oxides (some are neutral)
Reaction with acids	Many react with acids	Do not react with acids
Physical characteristics	Malleable, can be bent and shaped High melting and boiling point	Flaky, brittle Low melting and boiling point

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The typical properties of metals and non-metals

- The typical properties of metals and non-metals can be explained by reference to their trends in atomic radius, ionic radius, ionisation energy, electron affinity and electronegativity
- The low **ionisation energies** and low **electronegativities** of metals can account for the ability of their **valence electrons** to move away from the nucleus
 - This is known as '**delocalisation**' of the electrons
- These properties increase from left to right as you transition from metal to metalloid to non-metal
- The high **electronegativity** and **electron affinity** of non-metals can be related to their tendency to share electrons and form covalent bonds, either with themselves or other non-metal elements
- The similarities in **electronegativities** of the diagonal band of **metalloids** which divides the metals from the non-metals explains the behaviour of **metalloids**

Test yourself on the trends in the periodic table:



Worked Example

Which increase across a period from left to right?

A	Ionic radius	Electronegativity
B	Atomic radius	Ionic radius
C	1st ionization energy	Atomic radius
D	1st ionization energy	Electronegativity

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YOUR NOTES



Answer:

The correct option is **D**

- Across a period atomic radius decreases, electronegativity and ionisation energy increase and ionic radius increases then decreases

3.2 Oxides, Group 1 & Group 17

3.2.1 Periodic Trends: Oxides Across a Period

Oxides

Oxides across a period

- The **acid-base** character of the oxides provides evidence of chemical trends in the periodic table
- The broad trend is that oxides change from **basic** through **amphoteric** to **acidic** across a period
- Aluminium oxide is **amphoteric** which means that it can act both as a base (and react with an acid such as HCl) and an acid (and react with a base such as NaOH)

Acidic & Basic Nature of the Period 3 Oxides

Period 3 oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₂ , SO ₃
Acid/base nature	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic

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- The acidic and basic nature of the Period 3 elements can be explained by looking at their **structure, bonding** and the Period 3 elements' **electronegativity**

Structure, Bonding & Electronegativity of the Period 3 Elements Table

Period 3 oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₂ , SO ₃		
Relative melting point	High	High	Very high	Very high	Low	Low		
Chemical bonding	Ionic	Ionic	Ionic (with some degree of covalent)	Covalent	Covalent	Covalent		
Structure	Giant ionic	Giant ionic	Giant ionic	Giant covalent	Simple molecular	Simple molecular		
Element	Na	Mg	Al	Si	P	S	Cl	O
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0	3.5

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- The difference in electronegativity between oxygen and Na, Mg and Al is the largest

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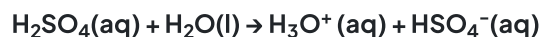
- Electrons will therefore be **transferred** to oxygen when forming oxides giving the oxide an **ionic bond**
- The oxides of Si, P and S will **share** the electrons with the oxygen to form **covalently bonded** oxides
- The oxides of **Na** and **Mg** which show purely **ionic bonding** produce **alkaline** solutions with water as their **oxide** ions (O^{2-}) become **hydroxide** ions (OH^-):



- The oxides of **P** and **S** which show purely **covalent bonding** produce **acidic** solutions with water because when these oxides react with water, they form an acid which donates **H⁺** ions to water
 - Eg. SO_3 reacts with water as follows:



- The H_2SO_4 is an acid which will donate an H^+ to water:



Exam Tip

Only examples of general trends across periods and groups are required, but you should be able to link trends in ionization energy, electron affinity and electronegativity with trends in chemical character such as the nature of the oxides and metallic/ non-metallic behaviour.

YOUR NOTES



Oxides reacting with Water

- The pH changes for the reactions of the oxides with water can be explained by reference to the following equations

Reaction of Oxides with Water

Oxide	Chemical equation	pH
Na_2O	$\text{Na}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{NaOH}(\text{aq})$	14 (strongly alkaline)
MgO	$\text{MgO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Mg}(\text{OH})_2(\text{aq})$	10 (weakly alkaline)
P_4O_{10}	$\text{P}_4\text{O}_{10}(\text{s}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{H}_3\text{PO}_4(\text{aq})$	2 (strongly acidic)
NO_2	$2\text{NO}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HNO}_3(\text{aq}) + \text{HNO}_2(\text{aq})$	1 (strongly acidic)
SO_2 , SO_3	$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$ $\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$	1 (strongly acidic)

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- The pattern here is that:
 - The metallic oxides form hydroxides when they react with water
 - The non-metallic oxides form oxoacids when they react with water



Exam Tip

You should learn how to construct these equations exactly as they are specifically mentioned in the syllabus

YOUR NOTES



Making Predictions

- The position of an element in the periodic table can be used to predict and explain its **metallic** and **non-metallic** behaviour
 - This is illustrated by the bonding of the oxides
 - Metal and non-metal elements generally form **ionic** compounds so the elements **Na** to **Al** have **giant ionic structures**
 - The oxides become more **ionic** as you go **down the group** as the **electronegativity decreases**
 - The oxides become less **ionic** as you go **across a period** as the **electronegativity increases**
 - The oxides of non-metals such as **S**, **N** and **P** form **molecular covalent compounds**
- Sometimes you may be asked to make predictions about oxides that are not specifically mentioned in the syllabus but you should be able deduce their properties if you understand the patterns outlined above, as the following example shows:



Worked Example

Which of these oxides produces the solution with the highest pH when added to water?

- A. CO_2
- B. SO_3
- C. CaO
- D. Na_2O

Answer:

The correct option is **D**.

- CO_2 and SO_3 will produce a pH below 7 as they are non-metal oxides. CaO and Na_2O will produce a pH above 7 as they are metal oxides; however the pH is decreasing as you go across a period, so Na_2O will have a higher pH than CaO

YOUR NOTES





ELEMENT	REACTION
LITHIUM	$2\text{Li}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{LiOH}(aq) + \text{H}_2(g)$ – REACTION SLOWER THAN WITH SODIUM, BUBBLES OF H ₂ GAS – LITHIUM DOESN'T MELT DUE TO ITS HIGH M.P.
SODIUM	$2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$ – BUBBLES OF H ₂ GAS, MELTS INTO A SHINY BALL THAT DASHES AROUND THE SURFACE – NaOH FORMED WHICH PRODUCES A HIGHLY ALKALINE SOLUTION
POTASSIUM	$2\text{K}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{KOH}(aq) + \text{H}_2(g)$ – REACTS MORE VIOLENTLY THAN SODIUM, BUBBLES OF H ₂ GAS, MELTS INTO A SHINY BALL THAT DASHES AROUND THE SURFACE – HOT ENOUGH TO BURN H ₂ WHICH FORMS LILAC FLAME

- Test yourself on this example question:



Worked Example

What would you see when sodium is added to water?

- I. a gas is given off
- II. the temperature of the water increases
- III. a clear, colourless solution is formed

- A. I and II only
- B. I and III only
- C. II and III only
- D. I, II and III

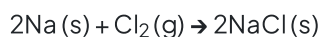
Answer:

The correct option is **D**.

- Bubbles of hydrogen gas are given off. The sodium melts, so that tells you it is an exothermic reaction and the product, sodium hydroxide, is very soluble so a clear, colourless solution would be formed.

Alkali metals with halogens

- All the **alkali metals** react vigorously with the **halogens** in group 17
- The reaction results in an **alkali metal halide salt**



- The reaction becomes increasingly vigorous going down group 1 because
 - The atoms of each element get **larger** going down the group
 - This means that the **ns¹** electron gets **further away** from the nucleus and is **shielded** by more electron shells.
 - The further an electron is from the positive nucleus, the easier it can be lost in reactions

Can you apply what you have just learned? Check out the following example question:



Worked Example

Which pair of elements has the most vigorous reaction?

- A. Cs and I
- B. Li and Cl
- C. Cs and F
- D. Li and F

Answer:

The correct option is **C**.

- You need to choose the lowest element in group 1 and the highest element in group 17 to predict the most vigorous reaction as reactivity increases going down group 1, but decreases going down group 17

YOUR NOTES



3.2.3 Periodic Trends: Group 17 - The Halogens

YOUR NOTES



Halogens

The halogens

- These are the group 17 non-metals that are **poisonous** and include fluorine, chlorine, bromine, iodine and astatine
- Halogens are **diatomic**, meaning they form molecules of **two** atoms
- All halogens have seven electrons in their outer shell
- They form **halide** ions by gaining one more electron to complete their outer shells

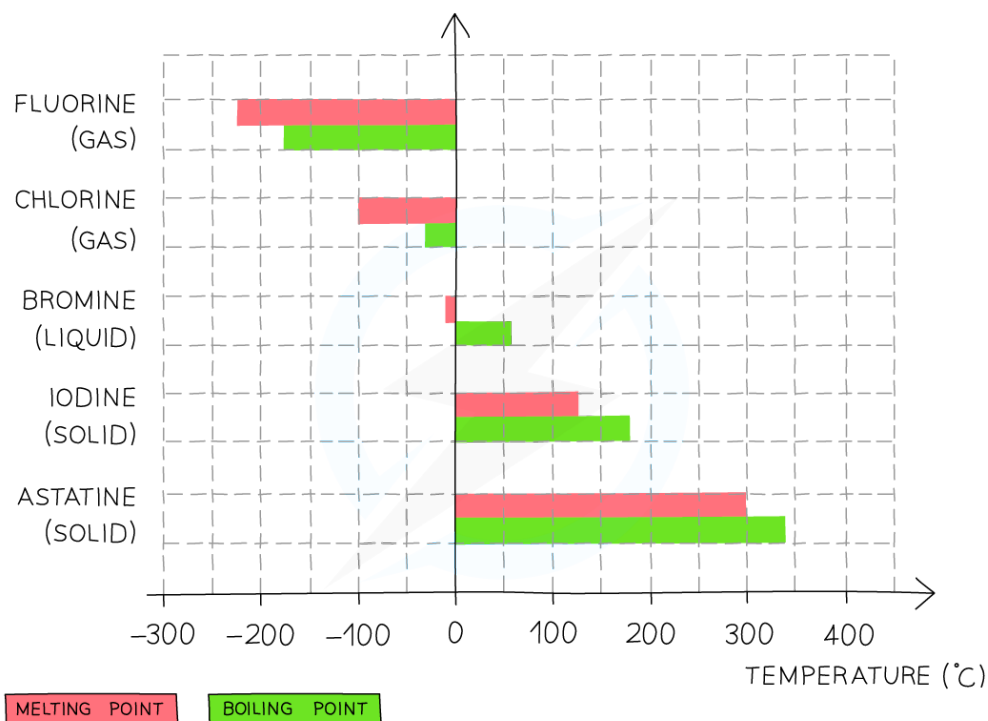
Colours and States at Room Temperature

HALOGEN	PHYSICAL STATE AT ROOM TEMPERATURE	COLOUR	COLOUR IN SOLUTION
FLUORINE	GAS	YELLOW	–
CHLORINE	GAS	PALE GREEN	GREEN-BLUE
BROMINE	LIQUID	RED-BROWN (READILY EVAPORATES TO FORM A BROWN GAS)	ORANGE
IODINE	SOLID	BLACK (SUBLIMES TO FORM A PURPLE GAS)	DARK BROWN

Trends in physical properties of the halogens

Melting point

- The density and melting and boiling points of the halogens **increase** as you go down the group



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Explaining the trend in reactivity in Group VII

- Reactivity of group 17 non-metals **decreases** as you go down the group
- The halogens electron configurations all end in ns^2np^5
- Each outer shell contains seven electrons and when they react, they will need to gain one outer electron to get a full outer shell of electrons
- Going down the group, the **electron affinity** decreases and the **atomic radius** increases
- As you go down group 17, the number of shells of electrons **increases** so **shielding** also increases
- This means that the outer electrons are **further** from the nucleus so there are **weaker** electrostatic forces of attraction that attract the extra electron needed
- The electron is attracted less readily, so the **lower** down the element is in Group 17 the **less reactive** it is

Reaction of the halogens with halide ions in displacement reactions

- A halogen **displacement** occurs when a **more** reactive halogen displaces a **less** reactive halogen from an aqueous solution of its halide
- The reactivity of group 17 non-metals increases as you move up the group
- Out of the 3 halogens, chlorine, bromine and iodine, chlorine is the most reactive and iodine is the least reactive

Aqueous Solution Colour of Halogens



AQUEOUS SOLUTION	COLOUR
CHLORINE	VERY PALE GREEN, BUT USUALLY APPEARS COLOURLESS AS IT IS VERY DILUTE
BROMINE	ORANGE BUT WILL TURN YELLOW WHEN DILUTED
IODINE	BROWN

Halogen displacement reactions

Chlorine and bromine

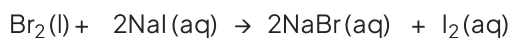
- If you add chlorine solution to colourless potassium bromide solution, the solution becomes orange as bromine is formed
- Chlorine is **above** bromine in group 17 so it is more reactive
- Chlorine will therefore **displace** bromine from an aqueous solution of a metal bromide



potassium bromide + chlorine → potassium chloride + bromine

Bromine and iodine

- Bromine is **above** iodine in group 17 so it is **more** reactive
- Bromine will therefore **displace** iodine from an aqueous solution of a metal iodide



bromine + sodium iodide → sodium bromide + iodine

Test yourself on halogen displacements:



Worked Example

Which of the statements below are correct?

- I. potassium chloride solution will react with fluorine to form chlorine.
- II. sodium chloride solution will react with iodine to form chlorine.
- III. lithium iodide solution will react with bromine to form iodine.

- A. I and II only
- B. I and III only
- C. II and III only
- D. I, II and III

Answer:

The correct option is **B**.

- Fluorine will displace chlorine as it is higher up in the group and bromine will displace iodine for the same reason.
- Iodine is below chlorine so cannot displace chlorine from sodium chloride

YOUR NOTES

