3.1.1 Periodicity

Elements are arranged in increasing atomic number in the periodic table

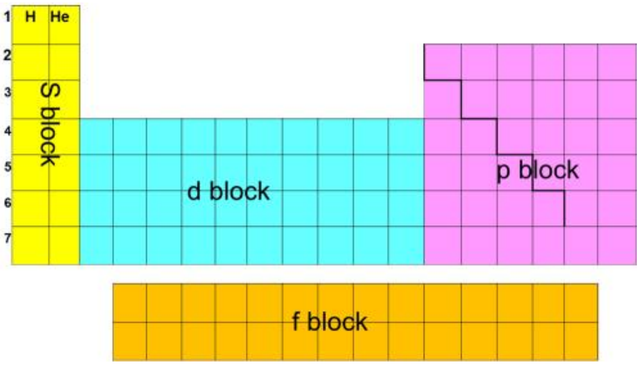
Elements in Groups have similar physical and chemical properties
The atoms of elements in a group have similar outer shell electron configurations, resulting in similar chemical properties;

Elements in periods showing repeating trends in physical and chemical properties

Period 2 = Li, Be, B, C, N, O, F, Ne
Period 3 = Na, Mg, Al, Si, S, Cl, Ar

Classification of elements in s, p, d blocks

Elements are classified as s, p or d block, according to which orbitals the highest energy electrons are in.



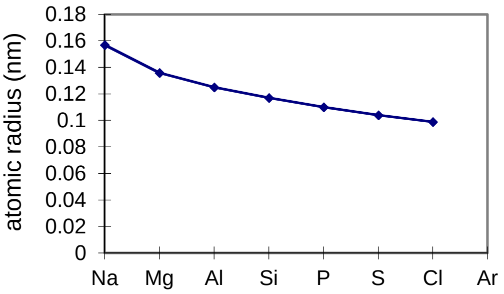
Periodicity *is* a repeating pattern across different periods

Various properties such as atomic radius, melting points, boiling points and ionisation energy display periodicity

Atomic radius

Atomic radii **decrease** as you move from left to right across a period, because the **increased number of protons** create more positive charge attraction for **electrons which are in the same shell** similar shielding.

Exactly the same trend in period 2



Ionisation Energies

Definition :First ionisation energy

The first ionisation energy is Energy needed to remove an electron from **each atom** in **one mole** of **gaseous atoms**

This is represented by the equation: $H_{(g)} \rightarrow H^{+}_{(g)} + e^{-}$

Always gaseous

Remember these definitions very carefully

The equation for 1st ionisation energy always follows the same pattern.
It does not matter if the atom does not normally form a +1 ion or is not gaseous

Factors that affect 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)

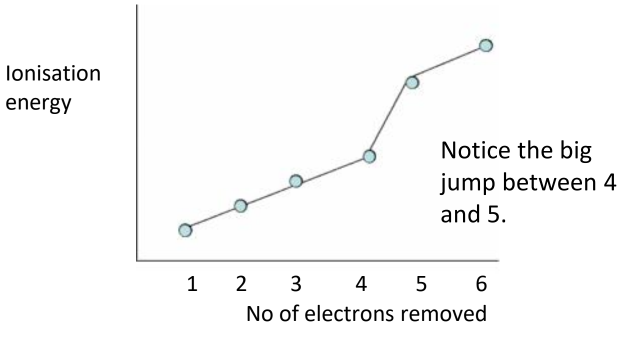
Many questions can be answered by application of these factors

Successive ionisation energies

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

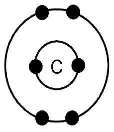
Why are successive ionisation energies always larger?
The second ionisation energy of an element is always bigger than the first ionisation energy. This is because the ion formed, is smaller than the atom and the proton to electron ratio in the 2+ ion is greater than in the 1+ ion. The attraction between nucleus and electron is therefore stronger

How are ionisation energies linked to electronic structure?



Notice the big jump between 4 and 5.

Explanation
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

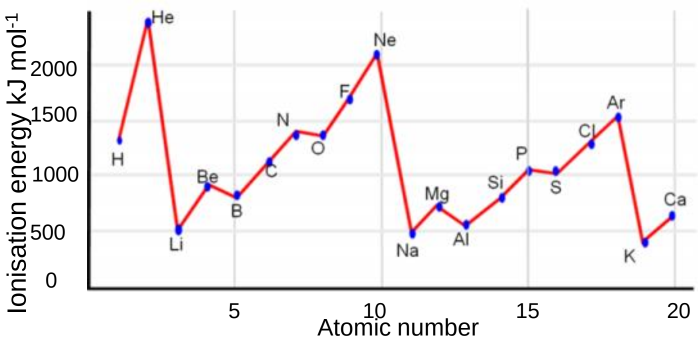


Example: What group must this element be in?

	1	2	3	4	5
Ionisation energy kJ mol^{-1}	590	1150	4940	6480	8120

Here there is a big jump between the 2nd and 3rd ionisations energies which means that this element must be in group 2 of the periodic table as the 3rd electron is removed from an electron shell closer to the nucleus with less shielding and so has a larger ionisation energy

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**.

The pattern in the first ionisation energy gives us useful information about electronic structure

You need to carefully learn the patterns

Q. Why has Helium the largest first ionisation energy?

A. Its first electron is 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

Many questions can be answered by application of the 3 factors that control ionisation energy

Q. Why do first ionisation energies decrease down a group?

A. As one goes down a group, the outer electrons are found in shells further from the nucleus and are more shielded so the attraction of the nucleus becomes smaller

Q. Why is there a general increase in first ionisation energy across a period?

A. As one goes across a period the electrons are being added to the same shell which has the same distance from the nucleus and same shielding effect. The number of protons increases, however, making the effective attraction of the nucleus greater.

Q. Why has Na a much lower first ionisation energy than Neon?

This is because Na will have its outer electron in a 3s shell further from the nucleus and is more shielded. So Na's outer electron is easier to remove and has a lower ionisation energy.

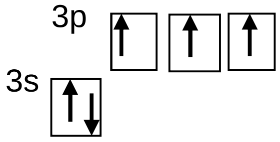
Q. Why is there a small drop from Mg to Al?

Al is starting to fill a 3p sub shell, whereas Mg has its outer electrons in the 3s sub shell. The electrons in the 3p subshell are slightly easier to remove because the 3p electrons are higher in energy and are also slightly shielded by the 3s electrons

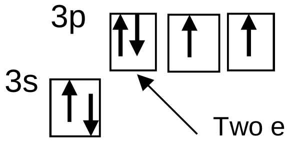
Learn carefully the explanations for these two small drops as they are different to the usual factors

Q. Why is there a small drop from P to S?

With sulphur there are 4 electrons in the 3p sub shell and the 4th is starting to doubly fill the first 3p orbital.
When the second electron is added to a **3p orbital** there is a slight **repulsion** between the two negatively charged electrons which makes the second electron easier to remove.



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



Two electrons of opposite spin in the same orbital

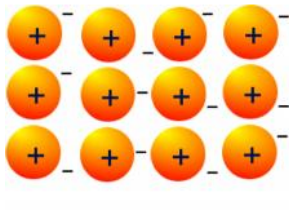
sulphur $1s^2 2s^2 2p^6 3s^2 3p^4$

Metallic bonding

Definition: M**etallic bonding** is the **electrostatic force** of attraction between the **positive metal ions** and the **delocalised electrons**

The three main factors that affect the strength of **metallic bonding** are:

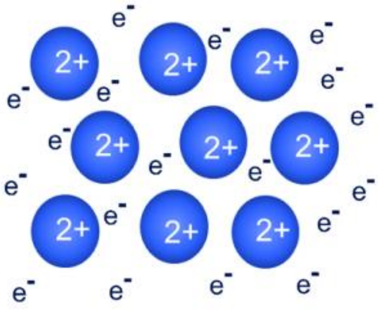
1. Number of protons/ Strength of nuclear attraction.
   The more protons the stronger the bond
2. Number of delocalised electrons per atom (the outer shell electrons are delocalised)
   The more delocalised electrons the stronger the bond
3. Size of ion.
   The smaller the ion, the stronger the bond.



sodium

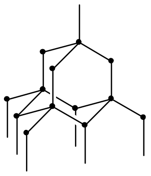
Example

Mg has stronger metallic bonding than Na and hence a higher melting point. The Metallic bonding gets stronger because in Mg there are more electrons in the outer shell that are released to the sea of electrons. The Mg ion is also smaller and has one more proton. There is therefore a stronger electrostatic attraction between the **positive metal ions** and the **delocalised electrons** and **higher** energy is needed to break bonds.



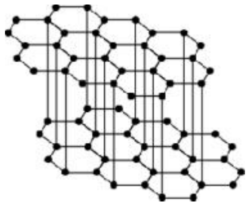
magnesium

Macromolecular: diamond



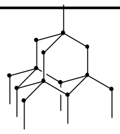
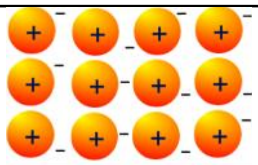
Tetrahedral arrangement of carbon atoms. 4 covalent bonds per atom

Macromolecular: Graphite



Planar arrangement of carbon atoms in layers. 3 covalent bonds per atom in each layer. 4th outer electron per atom is delocalised. Delocalised electrons between layers.

Both these macromolecular structures have very high melting points because of strong covalent forces in the giant structure. It takes a lot of energy to break the **many strong** covalent bonds

Bonding	**Structure**	**Examples**
Covalent : **shared <u>pair</u>** of electrons	**Macromolecular:** giant molecular structures. 	**Diamond** **Graphite** Silicon dioxide Silicon
Metallic: electrostatic force of attraction between the **metal positive ions** and the **delocalised electrons**	 **Giant metallic lattice**	**Magnesium, Sodium** (all metals)

Only use the words **molecules** and **intermolecular forces** when talking about simple molecular substances

Property	**Macromolecular**	**Giant Metallic**
boiling and melting points	high- because of many strong covalent bonds in macromolecular structure. Take a lot of energy to break the many strong bonds	high- strong electrostatic forces between positive ions and sea of delocalised electrons
Solubility in water	insoluble	insoluble
conductivity when solid	diamond and sand: poor, because electrons can't move (localised) graphite: good as free delocalised electrons between layers	good: delocalised electrons can move through structure
conductivity when molten	poor	(good)
general description	solids	shiny metal Malleable as the positive ions in the lattice are all identical. So the planes of ions can slide easily over one another -attractive forces in the lattice are the same whichever ions are adjacent

Melting and boiling points

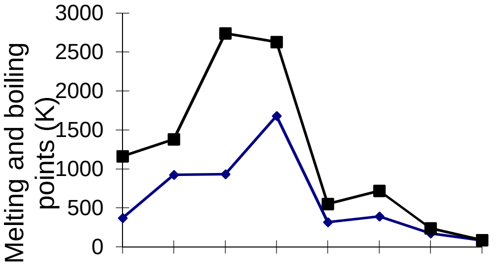
For **Na, Mg, Al- Metallic** bonding : strong bonding – gets stronger the more electrons there are in the outer shell that are released to the sea of electrons. A smaller positive centre also makes the bonding stronger. High energy is needed to break bonds.

Si is Macromolecular: many strong covalent bonds between atoms high energy needed to break covalent bonds– very high mp +bp

$Cl_{2\,(g)}$, $S_{8\,(s)}$, $P_{4\,(S)}$- simple Molecular : weak London forces between molecules, so little energy is needed to break them – low mp+ bp

S_8 has a higher mp than P_4 because it has more electrons (S_8 =128)(P_4=60) so has stronger London forces

Ar is monoatomic weak London forces between atoms



Similar trend in period 2
Li,Be metallic bonding (high mp)
B,C macromolecular (very high mp)
N_2,O_2 molecular (gases! Low mp as small London forces)
Ne monoatomic gas (very low mp)