AQA Chemistry A-level

3.2.1: Periodicity

Detailed Notes

3.2.1.1 - Periodicity

The Periodic Table arranges the known elements according to **proton number**. All the elements along a **period** have the same number of **electron shells**. All the elements down a **group** have the same number of **outer electrons**, this number is indicated by the group number.

Elements are classified into **blocks** within the Periodic Table that show electron configuration:

s-block = groups 1 and 2
p-block = groups 3 to 0
d-block = transition metals
f-block = radioactive elements

These different electron configurations are often **linked to other trends** within the Periodic Table. Periodicity is the study of these trends.

Atomic Radius

Along a **period**, atomic radius **decreases**. This is due to an **increased nuclear charge** for the same number of electron shells. The outer electrons are pulled in closer to the nucleus as the increased charge produces a **greater attraction**. As a result, the atomic radius for that element is reduced.

Down a **group**, atomic radius **increases**. With each increment down a group, an electron shell is added each time. This increases the distance between the outer electrons and the nucleus, **reducing the power of attraction**. More shells also increases electron **shielding** where the inner shells create a 'barrier' that blocks the attractive forces. Therefore the **nuclear attraction is reduced** further and atomic radius increases.

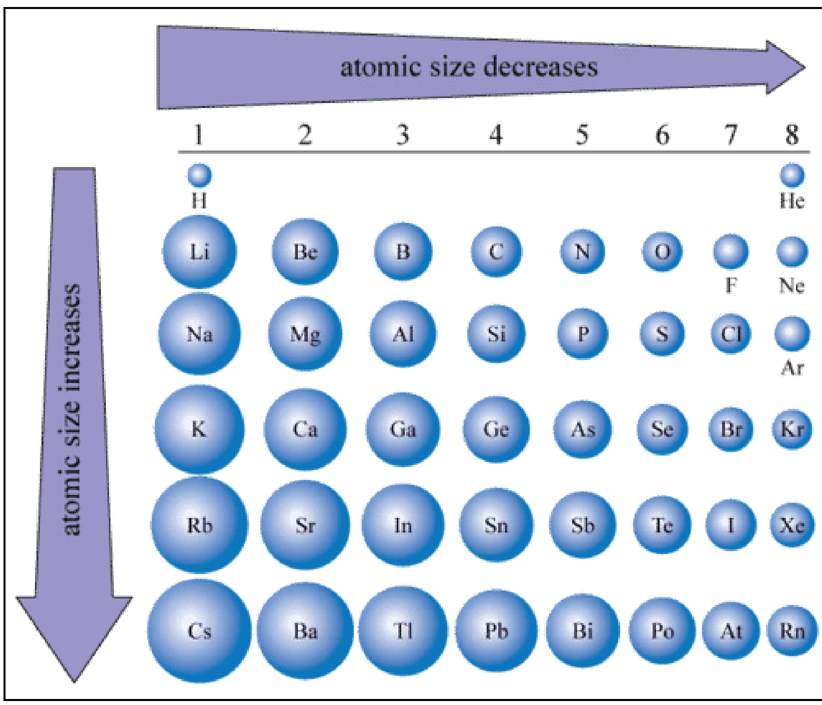
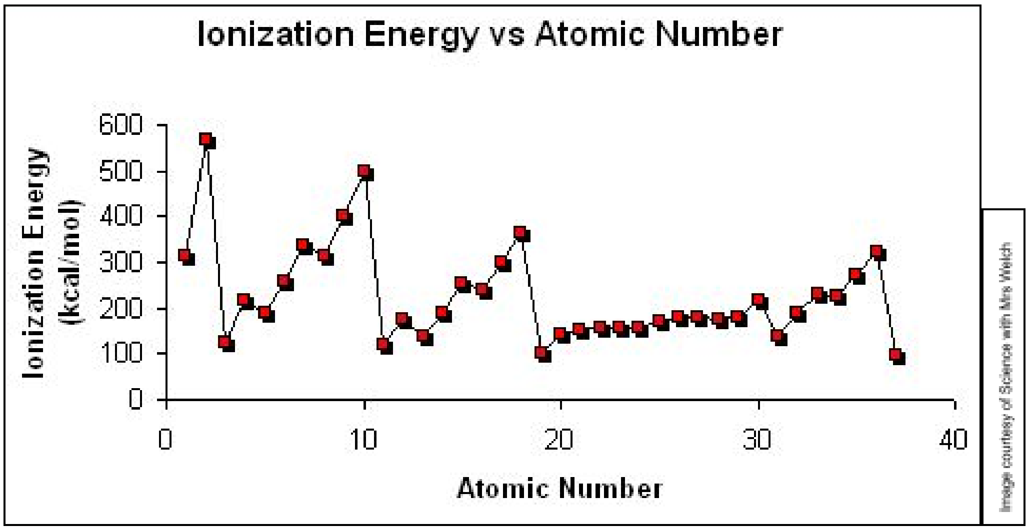


Image courtesy of ck12.org

Ionisation Energy

Along a **period**, ionisation energy **increases**. The decreasing atomic radius and increasing nuclear charge means that the outer electrons are **held more strongly** and therefore **more energy is required** to remove the outer electron and ionise the atom.

Down a **group**, ionisation energy **decreases**. The **nuclear attraction** between the nucleus and outer electrons reduces and increasing amounts of **shielding** means less energy is required to remove the outer electron.



3.2.1.2 - Physical Properties of Period 3

Melting Points

The melting points of the period three elements is linked to the **bond strength and structure**:

Sodium, magnesium and aluminium are all metals with **metallic bonding**. Their melting points increase due to greater **positive charged ions** (Na = +1, Mg = +2, Al = +3). This also means **more electrons are released** as free electrons so the attractive electrostatic forces increase from Na to Al.

Silicon is macromolecular meaning it has a **very strong covalent structure**. These covalent bonds require a lot of energy to break giving it a very high melting point.

Phosphorus, sulphur and chlorine are all **simple covalent molecules** held with weak **van der waals** forces. These intermolecular forces don't require much energy to overcome so these molecule have relatively low, similar melting points.

Argon is a noble gas that exists as **individual atoms** with a **full outer shell of electrons**. This makes the atom **very stable** and the van der waals forces between them very weak. As a result, the melting point of Argon is very low and it exists as a gas at room temperature.

Image courtesy of CapeChemistry