

Periodicity of Some Fundamental Properties

Atomic Radii

Two factors mainly determine atomic radii:

- 1) The attraction of the positively charged nucleus for the electrons – this decreases the atomic radius.
- 2) The 'screening' of the outer electrons from the positive nucleus by the electrons in the inner shells – this increases the atomic radius.

Across a period, each atom can be regarded as built up from the one before by adding an extra proton (+ the requisite number of neutrons) and an extra electron, this added electron going into **the same shell** as the existing electrons. Across a period, factor 1 above outweighs factor 2, so **the atomic radius decreases across a period.**

Down a group, on going from one member to the next, an **extra shell** of electrons is added each time. Down a group factor 2 above outweighs factor 1, so **the atomic radius increases down a group.**

Ionisation Energies

The first ionisation energy of an element is the energy required to remove the outermost electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions. How firmly an outer electron is held is governed by two factors:

- 1) Greater nuclear charge – the electron is held more firmly.
- 2) Further the electron is from the nucleus – the electron is held less firmly.

Across a period, nuclear charge increases and atomic radius decreases. Both these factors make the outer electron more firmly held so, **across a period the ionisation energies increase.**

There are exceptions: there is a **decrease** in ionisation energy on going from Be to B in period 2 and from Mg to Al in period 3. The reason is the unexpectedly high values for Be and Mg.

Reason – consider the electronic configurations of Be, Mg, B and Al:



A full s sub shell has special stability, so to remove an electron from Be or Mg requires removing an electron from the shared pair in a s orbital which requires more energy than removing the lone p electron as in the cases of B and Al.

Down a group, nuclear charge increases. This increases ionisation energy but the electron to be removed is further away from the nucleus each time (one shell more each time) and this tends to decrease ionisation energy (the force of attraction is inversely proportional to the square of the distance). Factor 2 above outweighs factor 1 so, **down a group ionisation energies decrease.**

Electronegativity

Electronegativity is a measure of the attraction of a bonded atom for the pair of electrons in a covalent bond.

This is governed by two factors.

- 1) Increase in nuclear charge – the more readily electrons are attracted so the greater the electronegativity.
- 2) Greater atomic radius – the less readily electrons are attracted so the lesser the electronegativity.

Across a period, nuclear charge increases and atomic radius decreases. Both these factors cause the electrons to be more readily attracted so, **across a period the electronegativities increase.**

Down a group, nuclear charge increases leading to an increase in electronegativity but atomic radius also increases leading to a decrease in electronegativity. Factor 2 above outweighs factor 1 so, **down a group electronegativities decrease.**

Variation in Melting Points, Boiling Points and Densities across Period 3

Ignoring the transition elements, the table below shows the variation in melting points, boiling points and densities of period 3 elements Na to Ar (no, you will not have to remember the table 😊).

Some Properties of Period 3 Elements

| Element | Na | Mg | Al | Si | P(white) | S(α) | Cl | Ar |
|----------------------------|------|------|------|------|----------|---------------|-------|-----------------------|
| Melting Pt/ $^{\circ}$ C | 97.8 | 550 | 660 | 1410 | 44.2 | 113 | -101 | -189 |
| Boiling Pt/ $^{\circ}$ C | 890 | 1110 | 2470 | 2360 | 280 | 445 | -34.7 | -186 |
| Density/ gcm^{-3} | 0.97 | 1.74 | 2.70 | 2.40 | 1.82 | 2.07 | 1.56 | 8.99×10^{-5} |

The melting point of an element indicates the strength of the forces holding the atoms or molecules together.

Na – metal, only 1 mobile electron contributing to the ‘sea of mobile electrons’ which holds the ions together. Thus Na has a relatively low melting point.

Mg – metal, has 2 mobile electrons contributing to the ‘sea of mobile electrons’ thus magnesium has a much higher melting point than sodium.

Al – metal, has 3 electrons in the outer shell but not all three contribute to the ‘sea of mobile electrons’ so the melting point of aluminium is only slightly higher than that of magnesium.

Si – non-metal, uses its 4 outer shell electrons to form an infinite 3-D array of atoms joined by strong, single covalent bonds (compare diamond). Thus the melting point of silicon is very high as melting involves the breaking of many, strong covalent bonds.

P + S – both non-metals, exist as molecules P_4 and S_8 so the only forces holding the molecules together are the weak van der Waals’ forces. Thus both phosphorus and sulfur have relatively low melting points. The melting point of sulfur is higher than that of phosphorus because it is a larger molecule, and as molecular size increases so do van der Waals’ forces.

Cl – non-metal, exists as Cl_2 molecules. Again only weak van der Waals’ forces are holding the small Cl_2 molecules together so a low melting point – a gas at room temperature.

Ar – non-metal, consists of monatomic molecules so again only weak van der Waals’ forces are holding the small Ar molecules together so a very low melting point – again a gas at room temperature.