

## Unit 1 - Key Area 2 – Periodicity Key Area 3 – Structure and Bonding

<p><b>KA 2.1</b></p> <p><i>The first 20 elements in the Periodic Table are categorised according to bonding and structure.</i></p> <p><i>Periodic trends and underlying patterns and principles.</i></p>	Elements are arranged in the Periodic Table in order of increasing atomic number.
	The Periodic Table allows chemists to make accurate predictions of physical properties and chemical behaviour for any element based on its position.
	The first 20 elements in the Periodic Table can be categorised according to bonding and structure:
	<ul style="list-style-type: none"> <li>metallic (Li, Be, Na, Mg, Al, K, Ca)</li> <li>covalent molecular (H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub>, P<sub>4</sub>, S<sub>8</sub> and fullerenes (eg C<sub>60</sub>))</li> <li>covalent network (B, C (diamond, graphite), Si)</li> <li>monatomic (noble gases)</li> </ul>
<p><b>KA 2.2</b></p> <p><i>Covalent radius, ionisation energy, electronegativity and trends in groups and periods, related to atomic structure.</i></p>	The covalent radius is a measure of the size of an atom.
	The trends in covalent radius across periods and down groups can be explained in terms of the number of occupied shells, and the nuclear charge.
	The trends in ionisation energies across periods and down groups can be explained in terms of the atomic size, nuclear charge and the screening effect due to inner shell electrons.
	Atoms of different elements have different attractions for bonding electrons
	Electronegativity is a measure of the attraction an atom involved in a bond has for the electrons of the bond.
	Electronegativity values increase across a period and decrease down a group.
	Electronegativity trends can be rationalised in terms of nuclear charge, covalent radius and the presence of 'screening' inner electrons.
<p><b>KA 3.1</b></p> <p><i>Bonding continuum</i></p> <p><i>Polar covalent bonds and their position on the bonding continuum, dipole formation and notation. <math>\delta^+</math> <math>\delta^-</math>,</i></p> <p><i>eg H <math>\delta^+</math> — Cl <math>\delta^-</math></i></p>	In a covalent bond, atoms share pairs of electrons.
	The covalent bond is a result of two positive nuclei being held together by their common attraction for the shared pair of electrons.
	Polar covalent bonds are formed when the attraction of the atoms for the pair of bonding electrons is different.
	Delta positive ( $\delta^+$ ) and delta negative ( $\delta^-$ ) notation can be used to indicate the partial charges on atoms, which give rise to a dipole.
	Pure covalent bonding and ionic bonding can be considered as being at opposite ends of a bonding continuum with polar covalent bonding lying between these two extremes.
	The larger the difference in electronegativities between bonded atoms, the more polar the bond will be. If the difference is large then the movement of bonding electrons from the element of lower electronegativity to the element of higher electronegativity is complete resulting in the formation of ions.
	Compounds formed between metals and non-metals are often, but not always ionic.

<p><b>KA 3.2</b></p> <p><i>Intermolecular forces are called van der Waals forces.</i></p> <p><i>London dispersion forces, permanent dipole-dipole, hydrogen bonding and the resulting physical properties including solubility.</i></p>	All molecular elements and compounds and monatomic elements condense and freeze at sufficiently low temperatures. For this to occur, some attractive forces must exist between the molecules or discrete atoms.
	Any 'intermolecular' forces acting between molecules are known as van der Waals' forces.
	There are several different types of van der Waals' forces such as London dispersion forces and permanent dipole: permanent dipole interactions, which include hydrogen bonding.
	London dispersion forces are forces of attraction that can operate between all atoms and molecules. These forces are much weaker than all other types of bonding. They are formed as a result of electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules.
	The strength of London dispersion forces is related to the number of electrons within an atom or molecule.
	A molecule is described as polar if it has a permanent dipole.
	The spatial arrangement of polar covalent bonds can result in a molecule being polar.
	Permanent dipole-permanent dipole interactions are additional electrostatic forces of attraction between polar molecules.
	Permanent dipole-permanent dipole interactions are stronger than London dispersion forces for molecules with similar numbers of electrons.
	Bonds consisting of a hydrogen atom bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen are highly polar.
	Hydrogen bonds are electrostatic forces of attraction between molecules which contain these highly polar bonds.
	A hydrogen bond is stronger than other forms of permanent dipole-permanent dipole interaction, but weaker than a covalent bond.
	Melting points, boiling points and viscosity can all be rationalised in terms of the nature and strength of the intermolecular forces which exist between molecules.
	By considering the polarity and number of electrons present in molecules, it is possible to make qualitative predictions of the strength of the intermolecular forces.
	The melting and boiling points of polar substances are higher than the melting and boiling points of non-polar substances with similar numbers of electrons.
	The anomalous boiling points of ammonia, water and hydrogen fluoride are a result of hydrogen bonding. Boiling points, melting points, viscosity and solubility/miscibility in water are properties of substances which are affected by hydrogen bonding.
Hydrogen bonding between molecules in ice results in an expanded structure which causes the density of ice to be less than that of water at low temperatures.	
Ionic compounds and polar compounds tend to be soluble in polar solvents such as water and insoluble in non-polar solvents.	
Non-polar substances tend to be soluble in non-polar solvents and insoluble in polar solvents.	