Definitions and Concepts for Edexcel Chemistry A-level

Topic 2: Bonding and Structure

Ionic bond: Strong electrostatic attraction between two oppositely charged ions. Strength of attraction depends on the relative sizes and charges of ions.

Cation: A positively charged ion, e.g. Na^+.

Anion: A negatively charged ion, e.g. S^{2-}.

Isoelectronic species: Chemical species that have the same number of electrons, e.g. N^{3-}, O^{2-}, F^- ions are isoelectronic - they all have ten electrons. CO and N_2 are isoelectronic molecules - they both have 14 electrons.

Covalent bond: The strong electrostatic attraction between two nuclei and the shared pair of electrons between them. *Polar* covalent bond occurs when there is an asymmetric electron distribution within the covalent bond due to difference in electronegativities.

σ (sigma) **bond**: A bond that results from a direct (end-on) overlap of two orbitals, e.g. a sigma bond in H_2 molecule is formed by overlap of two 1s orbitals. Similarly, a sigma bond in HCl is a result of the end-on overlap of 1s orbital of hydrogen with 3p orbital of chlorine.

π (pi) **bond**: A bond that is formed when two orbitals overlap sideways, e.g. a pi bond in C_2H_4.

Dative covalent bonding: Occurs when one atom donates both electrons in a bond. e.g. in NH_4^+ or H_3O^+ ions. Marked with an arrow.

Shapes of the molecules: Shapes adopted by the molecules so as to minimise the electronic replusions.

Shape of the molecule	Bond angle (°)	Number of bonds made by the central atom	Number of lone pairs on the central atom	Examples
Linear	180	2	0	$BeCl_2$, $Ag(NH_3)_2^+$
Trigonal planar	120	3	0	BF_3, C_2H_4
Tetrahedral	109.5	4	0	CH_4, NH_4^+, $CoCl_4^{2-}$
Trigonal pyramidal	107	3	1	NH_3, H_3O^+
Bent	104.5	2	2	H_2O

Trigonal bipyramidal	120, 90	5	0	PCl_5
Octahedral	90	6	0	SF_6, $Cu(H_2O)_6^{2+}$

Allotropes: Different forms of the same element, e.g. allotropes of carbon are: diamond, graphite, graphene, fullerenes, carbon nanotubes etc.

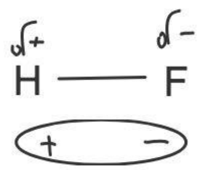
Malleable: A malleable substance can be shaped.

Ductile: A ductile substance can be drawn into a wire.

Intermolecular forces: Forces between the molecules (cf. *bonding*, an *intramolecular* force).

Electronegativity: The ability of atom to attract the bonding electrons in a covalent bond. The most electronegative elements (N,O,F) are small and have a relatively high nuclear charge.

Dipole: Difference in charge between the two atoms of a covalent bond caused by a shift in electron density in the bond due to the electronegativity difference between elements participating in bonding. *Polar molecules* exist as dipoles, e.g.



Metallic bonding: Strong electrostatic attraction between metal ions and the sea of delocalised electrons that surround them.

Delocalised electrons: The electrons that are not contained within a single atom or a covalent bond.

Bond length: Internuclear distance between two covalently bonded atoms.

London forces: Weak intermolecular forces arising due to *fluctuations of electron density* within a nonpolar molecule. These fluctuations may temporarily cause the *asymmetric electron distribution*: the molecule becomes an *instantaneous dipole*. This dipole can *induce a dipole* in another molecule, and so on. The attraction increases with size/shape (points of contact between the molecules) and number of electrons (more fluctuations = more instantaneous/induced dipoles).

Permanent dipole-dipole interactions: Dipole-dipole attractions between polar molecules. Stronger than London forces.

Hydrogen bond: A type of intermolecular force (with some bonding character) between a hydrogen bonded to a more electronegative atom than hydrogen (usually N,O,F) and other atom in a same/different molecule. Directional nature - the bond angle is often 180°. Responsible for anomalous properties of water, e.g. the density of ice < density of water. Ice occupies greater volume than water due to the directional nature of hydrogen bonds within the solid structure.