Chemical Bonding

Electronegativity
- Electronegativity is the ability of an atom to attract or gain electrons
- Fluorine is the most electronegative atom
- Hydrogen is the least electronegative non-metal

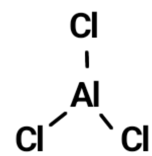
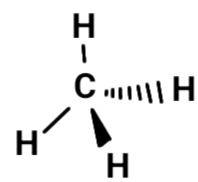
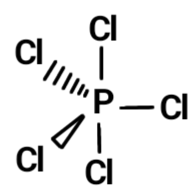
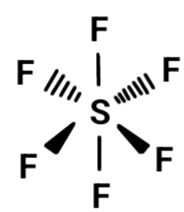
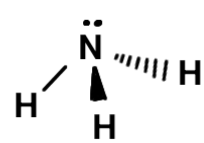
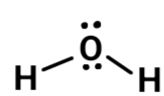
Electronegativity Factors & Trends
- Electronegativity increases left to right across a period
- Electronegativity decreases down the group
- Increased nuclear charge results in increased electronegativity
- Increased atomic radius results in decreased electronegativity
- Increased shielding through increased quantum shells or subshell will decrease electronegativity
- Metals are less electronegative than nonmetals

Electronegativity and Bonding
- A large difference in electronegativities between bonded atoms means the bond is ionic
- Small or no difference means the bond is covalent

Bond Energy
- It is the energy required to break one mole of a covalent bonds in gaseous state
- Bond Energy has units $KJmol^{-1}$
- Bond Energy is always endothermic
- The higher the bond energy, the stronger the covalent bond

Bond Length
- It is the distance between the nuclei of two covalently bonded atoms
- The greater the nuclear attraction, the lesser the bond length
- The lower the bond length, the stronger the covalent bond
- Hence, the lower the bond length, the higher the Bond Energy

Shapes of covalent molecules
- The shapes of covalent molecules are predicted by the Valence Shell Electron Pair Repulsion Theory (VSEPR)
- Electrons are negatively charged and repel each other when close
- The repulsion in bonding pair of electrons a molecule causes it to adopt a shape to minimise the repulsion
- Lone pairs repel each other more than bond pairs
- Lone pair-lone pair **>** lone pair-bond pair **>** bond pair-bond pair repulsion

Bond Pairs	Lone Pairs	Shape Name	Bond Angles	Example
2	0	Linear	180	O=C=O
3	0	Triangular Planar	120	
4	0	Tetrahedral	109.5	
5	0	Triangular Bipyramidal	90 & 120	
6	0	Octahedral	90	
3	1	Triangular Pyramidal	107	
2	2	Bent non-linear	104.5	

Hydrogen Bonding

- Hydrogen Bonding is the strongest form of intermolecular bonding
- It is a type of a permanent dipole-permanent dipole bonding
- It is present in molecules where Hydrogen is covalently bonded to small, highly electronegative atoms; F, O, N only
- Due to the electronegativity difference, the bond becomes highly polarised
- The H becomes so partial positive charged that it can bond with the lone pair of an O or N atom of another molecule
- Hydrogen bonding causes high melting and boiling points, such as in water
- It causes high surface tension in water

Polarity

- Bond Polarity is the charge separation in a covalent molecule to a difference in electronegativities between bonded atoms
- when two covalently bonded atoms have the same electronegativity, the bond is nonpolar
- The lesser electronegativity atom gains a partial positive charge ($\delta+$)
- The higher electronegativity atom gains a partial negative charge ($\delta-$)
- The greater the difference in electronegativity, the more polar the bond becomes

Dipole

- The dipole moment is a measure of how polar a bond is
- It is represented by an arrow pointing towards the partial negative end of dipole

Polarity in Molecules

- The polarity of a molecule is determined by the polarity of each bond and the arrangement of bonds
- Equal and opposite dipoles cancel each other out
- Symmetrical molecules like linear, planar or tetrahedral are non-polar as dipoles are equal and opposite
- Molecules having lone pairs are polar because they distort symmetry

Van der Waals forces

- Intermolecular forces between covalent molecules are known as Van der Waals forces
- There are two types of Van der Waals forces:
  1. temporary/induced dipole - induced dipole forces
  2. Permanent dipole - permanent dipole forces

Instantaneous dipole - Induced dipole forces

- present in non-polar covalent molecules
- when two nonpolar molecules come close together, electrons of the molecules repel each other
- The repulsion causes electrons to be unevenly distributed, inducing a dipole
- the partial positive end of a molecule attracts the partial negative end of another
- this attraction is induced dipole - induced dipole force

- induced dipole - induced dipole forces increase with increasing electrons and greater surface area of molecules to allow more points of contact

Permanent dipole - permanent dipole forces

- polar molecules have permanent dipoles
- the molecules always have partial positive and partial negative ends
- the forces between two polar molecules are permanent dipole-permanent dipole forces
- the partial positive end of one molecule attracts the partial negative end of another

- Permanent dipole-permanent dipole forces are stronger than induced dipole-induced dipole forces in molecules having same number of electrons
- Hence, polar molecules have higher melting and boiling points

Coordinate/Dative bonding

- A dative bond is formed when one atom provides both electrons for a covalent bond
- The sharing is not mutual
- It is a represented by an arrow pointing away from the lone pair of electrons that form the bond
- Al_2Cl_6 is a dimer formed by dative bonding, a chlorine atom of one $AlCl_3$ gives two electrons to the Aluminium atom of another

Incomplete and Expanded Octet

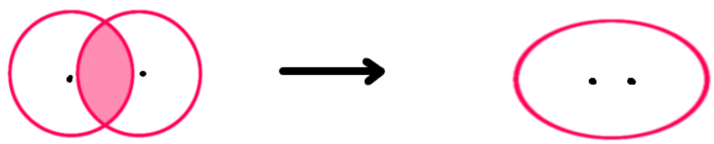
- Some species can have more or less than eight electrons in their outer shell
- $GaCl_3$, $AlCl_3$, $BeCl_2$, BF_3,BCl_3 are examples of molecules having incomplete octet
- PCl_5, SO_2, SO_3, SF_6, SeF_6 are examples of expanded octet

Sigma and Pi bonding

- It is the overlapping of two half filled atomic orbitals
- The greater the atomic orbital overlap, the stronger the bond

Sigma bonds

- They are formed by the end to end overlapping of atomic orbitals
- Both s and p orbitals overlap this way
- the pair of electrons is found between the two nuclei
- the force between the electrons and nuclei bonds the atoms together
- All single covalent bonds are sigma bonds



s-s orbital overlap

s-p orbital overlap

p-p orbital overlap

Pi bonds

- are formed from the sideways overlap of adjacent p orbitals
- double covalent bonds contain 1 sigma and 1 pi bond
- Triple covalent bonds contain 1 sigma and 2 pi bonds