Chemical Bonding and Structure

Ionic bonding

It is the electrostatic attraction between 2 oppositely charge ions.

Metals loses electrons and forms a cation

$Mg \longrightarrow Mg^{2+} + 2e^-$

Non-metals gains electrons and forms an anion

$2e^- + O \longrightarrow O^{2-}$

hydroxide	OH^-	CO_3^{2-}
	NO_3^-	HCO_3^-
	SO_4^{2-}	CH_3COO^-
	HSO_4^-	PO_4^{3-}

These electrostatic forces are strong, so it is difficult to break apart the lattice structure.

Ionic compounds high melting and boiling points due to strong electrostatic forces between oppositely charged ions

- Volatility - very low volatility because of strong electrostatic forces.
- Electrical conductivity - conducts electricity in molten state as ions are free to move (not in solid state)
- Solubility in water - often soluble in water. Breaking of hydrogen bond often pay back energy for required to break apart the ionic structure.

- Electrical conductivity in ionic solutions - Aqueous solutions of ionic substances conduct electricity. (ions can move freely)

- Solubility in non-polar solvents
Ionic substances are not soluble in non-polar solvents as a lot of energy is required to break apart the ionic lattice.

- Covalent bonding
Single covalent bonds - covalent bonding occurs when atoms share electrons, and a covalent bond is the electrostatic attraction between a shared pair of electrons and the nuclei of the atoms that are bonded.

Lewis structure - all the valence electrons are shown

$$:\ddot{\underset{..}{F}}:\ddot{\underset{..}{F}}: \quad \text{or} \quad |\overline{\underline{F}} - \overline{\underline{F}}| \quad \text{or} \quad :\ddot{\underset{..}{F}} - \ddot{\underset{..}{F}}:$$

A double bond results from the sharing of two pairs of electrons, and a triple bond arises from sharing of 3 pairs of electrons.

- Covalent bond is the electrostatic interaction between the positively charged nuclei of both atoms and the shared pair of electrons.

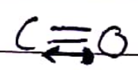
single bond < double bond < triple bond

Electronegativity - higher electronegative more closer the electrons in the bond. Eg:- $^{\delta+}H \cdots : - F^{\delta-}$

Atoms with similar electronegativities will form covalent bonds. Atoms with widely different electronegativities will form ionic bonds. The difference in electronegativity can be taken as a guide to how ionic or how covalent the bond between two atoms is likely to be.

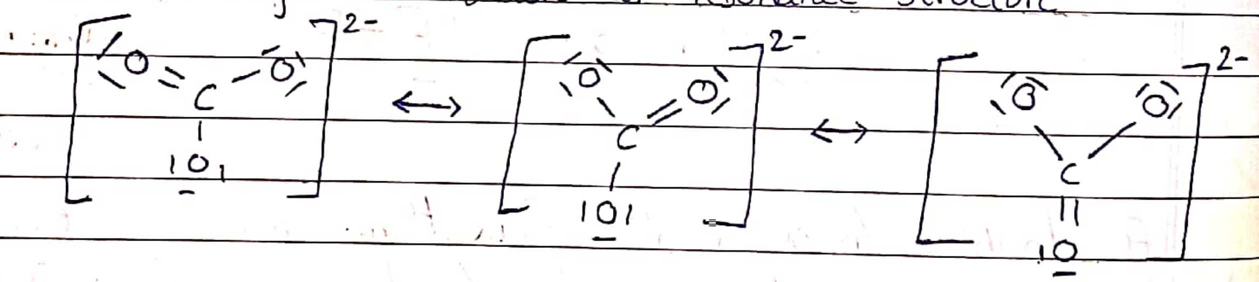
Difference more than 1.7 forms ionic bonds.

- Coordinate covenlent bonds



- Resonance structures

When a compound can have more than one possible structure they are known as resonance structure



Drawing lewis structure

- Add up the total number of valence electrons of all the atoms
- Divide by two to get the total number of valence electron pairs
- Arrange that each atom has 4 lines.

- Valence shell electron pair repulsion theory (VSEPR)

Pairs of electrons (electron domains) in the valence (outer) shell of an atom repel each other and will therefore take up positions in space to minimise these repulsion - to be as far as possible.

order of repulsion

lone - lone pair > lone pair - bonding pair > bonding - bonding pair

Total electron domains	Bonding Pairs	Lone pairs	Basic Shape	Actual Shape	Eg
2	2	0	linear	linear (180°)	CO_2, NCH, NO_2^+
3	3	0	trigonal planar	trigonal planar (120°)	BF_3, SO_3, NO_3^-
3	2	1	trigonal planar	bent, V-shaped (117°)	SO_2, O_3, NO_2^-
4	4	0	tetrahedral	tetrahedral (109.5°)	CCl_4, NH_4^+, SO_4^{2-}
4	3	1	tetrahedral	trigonal pyramidal (107°)	NH_3, PCl_3, H_3O^+
4	2	2	tetrahedral	bent, V-shaped (107.5°)	H_2O, ClF_2^+, I_3^+
5	5	0	trigonal (90°, 120°, 180°)	trigonal bipyramidal	PF_5, XeO_3F_2
5	4	1	bipyramidal	see-saw (102°)	SF_4, XeO_2F_2
5	3	2		T-shape	BrF_3, $XeOF_2$
5	2	3		linear (180°)	I_3^-, XeF_2
6	6	0		octahedral (90°, 180°)	SF_6, PF_6^-
6	5	1	octahedral	square pyramidal	SF_5^-, XeF_5^+
6	4	2		square planar	XeF_4, SF_4^{2-}

Polar Moleoules

electronegativity difference creating electron lying more towards one atom is called a polar bond. These molecule have a overall dipole moment.

Although individual bonds may be polar, a molecule may be non-polar overall if, because of the symmetry of the molecule, the dipole moments of the individual bonds cancel out.

Gaint Covalent Structures

Allotropes of carbon
allotropes are different form of the same element.
- Diamond has a gaint covalent structure. Each carbon atom is joined to four others, in a tetrahedrial array

- Graphite has a gaint covalent structure, but it also has a layer structure. London forces between layers which could easily be broken thus good lubricant, Electrons delocalised between the layers it is a good conductor.

- Silicon dioxide
Each silicon atom is bonded to four oxygen atoms in a tetrahedral array. Each oxygen to two silicon atoms.

Intermolecular forces
Van der Waal's forces is the collective name given to the forces between molecules and includes London forces, dipole-dipole interaction and dipole-induced dipole interactions.

Intermolecular forces are forces ~~within~~ between ~~a~~ molecule.

Intramolebcular forces are forces within a molecule

- London forces are temporary dipole - induced dipole interaction. Electrons will not be symmetrically distributed about the nucleus. This results a ~~k~~temporary dipole in the atom, and induce an opposite dipole in a neighbouring atom. This creates an attractive force.

London forces get stronger as the relative molecular mass increases.

- Polar molecules - because of their polarity of the molecules, there are also intermolecular forces created known as permanent dipole - permanent dipole which is usually just dipole - dipole attractions

- Hydrogen bonding - hydrogen bonding occurs between molecules when a very electronegative atom (N, O, F) is joined to a hydrogen
  (N, O, F who possesses a lone pair of electron.

- Melting or boiling
  Only intermolecular forces are broken when covalent molecular substance are melted or boiled - covalent bonds are not broken.

London < permanent dipole-dipole < hydrogen bonding.

Generally a substance will dissolve in a solvent if the intermolecular forces in the solute and solvent are similar.

- Metallic bonding

Delocalised electrons as electrons doesn't belong to any one metal atom but, rather, are able to move throughout the structure.

Positively charged metal ions in lattice surrounded by sea of delocalised electrons.

Metals are malleable/ductile because of non-directionality of the bonding. The metals can ~~be~~ also be good conductors as free electrons (delocalised electrons). Further when the two layers slide over each other, the bonding in the resulting structure is exactly the same.

Formal charge is the charge that an atom in a molecule would ~~be~~ have if we assume that the electrons in a covalent bond are equally shared between atoms that are bonded.

FC = (number of valence electrons in the uncombined atom) $-\frac{1}{2}$(number of bonding electrons) − (number of non-bonded electrons)

- Sigma and Pi bonds

- Sigma bonds results from the axial (head-on) overlap of atomic orbitals. The electron distribution in a sigma bond lies mostly along the axis joining the two nuclei.

- A pi bond is formed by the sideways overlap of parallel p orbitals. The electron density in the pi bond lies above and below the inter~~mo~~nuclear axis.

- a single bond consists of a sigma bond
- a double bond consists of a sigma and a pi bond
- a triple bond consists of a sigma bond and 2 pi bonds.

- Delocalisation is the sharing of a pair of electrons between three or more atoms.

Bezene has a pi delocalised ring of electrons that extends all around the ~~sig~~ ring of carbon atoms.

- Higher energy is required to break O_2 bond than O_3 as O_2 has bond order of 2 and O_3 has 1.5.

Single bond	1
double bond	2
triple bond	3

Depletion of ozone by CFCs.

$CCl_2F_2 \xrightarrow{UV} \cdot CClF_2 + Cl\cdot$

$O_3 \rightarrow O\cdot + O_2$

$\cdot Cl + O_3 \longrightarrow ClO\cdot + O_2$

$ClO\cdot + O\cdot \longrightarrow O_2 + Cl\cdot$

$$\cancel{\cdot Cl} + O_3 \longrightarrow \cancel{ClO\cdot} + O_2$$
$$\cancel{ClO\cdot} + O\cdot \longrightarrow O_2 + \cancel{Cl\cdot}$$
$$O_3 + O\cdot \longrightarrow 2O_2$$

$N_2O + \cdot O \longrightarrow 2NO$

$\cdot NO_2 + \cdot O \longrightarrow \cdot NO + O_2$

$$\cdot \cancel{NO} + O_3 \longrightarrow \cdot \cancel{NO_2} + O_2$$
$$\cdot \cancel{NO_2} + O\cdot \longrightarrow O_2 + \cdot \cancel{NO}$$
$$O_3 + O\cdot \longrightarrow 2O_2$$

Homolytic fission - the covalent bond breaks so that one electron goes back to each atom making up the bond.

Free radicals are atoms or groups of atoms with unpaired electrons.

- Hybridisation is the mixing of atomic orbitals in a particular atom to produce a new set of orbitals (the same number originally) that have characteristics of the original orbitals and are better arranged in space for covalent bonding.