BONDING

Three main types :
1. Ionic
2. Metallic
3. Covalent
   i) Polar Covalent
   ii) Non-Polar Covalent

Important Terminology :
1. Electronegativity
2. Polarity
   i) Dipoles
   ii) Partial Charges
   iii) Lone pairs

Electronegativity :
- The electronegativity of an element is the relative ability of each of its' atoms to attract the electrons of a covalent bond towards itself

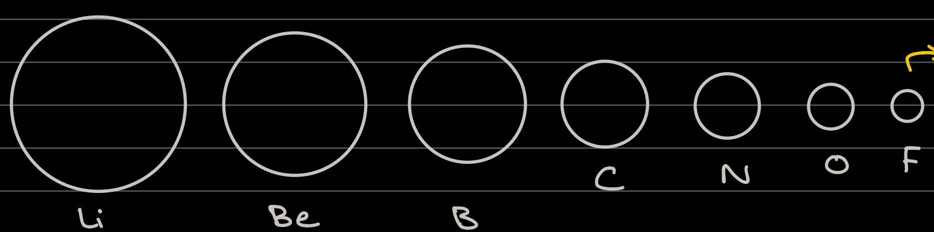
- It is measured on the Pauling Scale, where 0 is the least electronegative and 4 is the most electronegative

Some quick facts :
- Fluorine is the most electronegative atom
- Non-metals in general are more electronegative (smaller atomic radius)
- Metals in general are less electronegative (greater atomic radius)

Factors that determine the electronegativity of an element :
1. Atomic Radius
2. Nuclear Charge

- Atomic radius decreases across a period, as new electrons enter the same energy level and hence experience roughly the same shielding effect but a stronger effective nuclear force due to more protons in the nucleus

- For example, period 2 :



→ F has the most protons, hence the greatest nuclear charge and is also the smallest and therefore, has the greatest charge density, the deciding factor which determines electronegativity

→ Atomic radius decreases across the period

Example: A C–F bond

$\overset{\delta+}{\underset{6p}{C}} \xrightarrow{\quad} \overset{\delta-}{\underset{9p}{F}}$

- Since F has a higher charge density, a greater nuclear charge, and a smaller atomic radius, it pulls the electrons involved in the covalent bond towards itself

Therefore, electrons exist closer to the F nucleus than the C nucleus

- As a result, F develops a partial negative charge (δ^-) and C develops a partial positive charge ($\delta+$)

Note: $\mapsto$ represents the direction in which the pair of electrons in the covalent bond are attracted / pulled towards

Separation of charge refers to the attraction of electrons towards one end of a molecule, leading to the formation of a dipole

- If a dipole exists in a molecule, then that molecule is polar and accordingly, if there is no dipole, the molecule is non-polar

Pauling Scale values of some common elements:

Fluorine	4.0	Lithium	1.0
Oxygen	3.5	Carbon	2.5
Chlorine	3.0	Hydrogen	2.1
Nitrogen	3.0	Caesium	0.7

- Generally speaking, if the difference in electronegativity is greater than 2.0, then the bond can be considered ionic

- Similarly, for covalent bonds, the difference in electronegativity is usually less than 1.0

Note: Although individual bonds may be polar, there may still be no overall charge on the molecule, as the shape of the molecule might be such that the dipoles cancel out
↳ However, it is also possible that the shape might be such that dipoles are reinforced

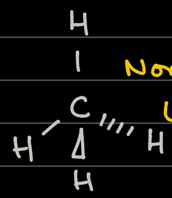
Some common molecules, their dipoles, and their polarity:

H_2: H — H
dipoles cancel out, hence non-polar

HF: δ^+ H — F δ^- → more EN
dipole towards F, hence polar

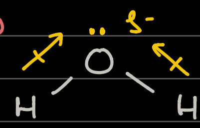
CO_2: O = C = O
dipoles cancel out, hence non-polar

CH_4



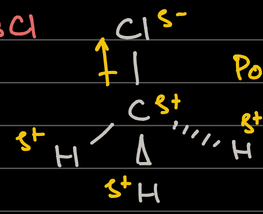
Non-polar (All C-H bonds are non-polar)

H_2O



Polar

CH_3Cl



Polar

Note: Lone pairs and non-bonding electrons do affect the polarity of a molecule
↳ They not only act as areas of negative charge to add to the δ⁻ nature of some atoms, but they can also influence the bond angles and the symmetry of a molecule, leading to dipoles not cancelling each other out and the molecule being polar

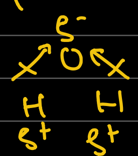
Continuation of ionic properties

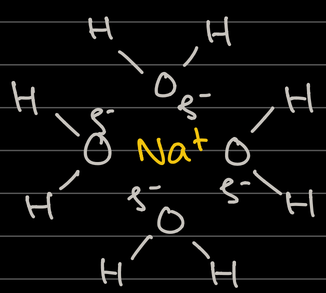
- Electrical conductivity increases as the concentration of the electrolyte increases, as there are more ions to carry charge

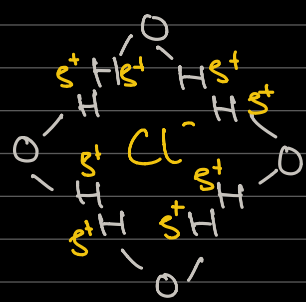
- An electrolyte containing ions with a higher charge (ie. Al^{3+}) conducts better than electrolyte containing lesser charged ions (ie Na^+)
    ↳ as long as the concentrations are the same

Solubility

- Ionic compounds are soluble in polar solvents, ie water

- water  is polar, and it can surround the ions forming ion-dipole forces of attraction
    ↳ These ions are then said to be hydrated
        ↳ This process releases energy





Ion-Dipole Attractions

METALLIC BONDS

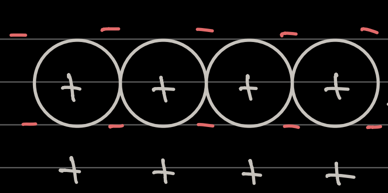
- The "electron sea" model of metallic bonding proposes that all the metal atoms in the sample contribute their valence electrons to form an "electron sea" that is delocalized throughout the metal

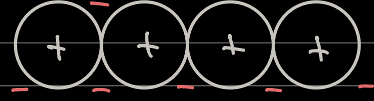
- The nuclei, with their core (inner shell) electrons are submerged within the sea of electrons in an orderly manner.

- The valence electrons are shared amongst all the atoms in the substance and the piece of metal is held together by mutual attractions of the metal cations for the highly delocalised mobile electrons

Structure

- Giant metallic lattice

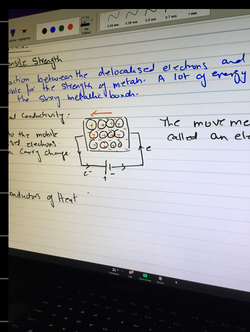




PROPERTIES :

① High tensile strength
- The attraction between the delocalized electrons and the cations is responsible for the strength of metals
  ↳ A lot of energy has to be put in to break the strong metallic bond

② Electrical Conductivity
  ↳ is due to the mobile delocalized electrons
- The movement of electrons is called an electrical current.



③ Good conductors of heat
- The delocalised electrons in the metal conduct heat away from a human hand much faster than the localised electrons in covalent bonds in wood
  ↳ Thus, wood does not seem cool to touch
- Heat energy → kinetic energy of movement of electron

④ High melting + boiling point
- Melting points are moderately high because the electrostatic forces of attraction between the delocalised electrons and the cations have to broken, which requires a large amount of energy

- Boiling a metal requires even more energy as all forces of attraction have to be completely broken

$Na(s)$	$\longrightarrow$	$Na(l)$	$\longrightarrow$	$Na(g)$
solid		molten		gas
regular		less regular		random gaseous atoms

⑤ Malleability / Ductility
- Metals are malleable and ductile due to the fact that the cations are arranged in a regular manner surrounded by the sea of electrons.

- The regular layers can slide over each other without shattering the metal