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**<u>*Important Note*</u>**

For this chapter on Chemical Bonding and Structure, it is important to link how the bonding on a micro scale can translate to physical properties on a macro scale. There are three physical properties which are important to take note of, as well as their respective requirements:

<u>Melting/Boiling Point</u>
Type and strength of bonds
(strong ionic bonds between oppositely charged ions, strong covalent bonds between atoms in the molecule, strong metallic bonds between cations and 'sea' of delocalised electrons, weak IMF between molecules)
Energy to overcome/break bonds
Melting and Boiling point

<u>Electrical Conductivity</u>
Presence of mobile charge carriers (free-moving ions, delocalised electrons, neutral molecules) to conduct electricity

<u>Solubility</u>
“Like dissolves Like” Compounds that carry a charge can interact well with polar molecules like water which are partially charged as well.
Compounds that do not carry charge likewise interact well with non-polar molecules which also do not carry charge.

There are three types of bonding: Ionic, Covalent and Metallic. One important thing to note is that all 3 bonds are electrostatic forces of attraction. The only difference for the different bonds is what this bond is between. However in answers the phrase EFOA is not required and only the actual bond is required. (Eg. Ionic bond, Covalent Bond, Metallic Bond between ...)

**Ionic Bonds** are formed between a positively charged metal cation and a negatively charged non-metal anion.
Metals (Group I, II, III) tend to lose valence electrons, achieving an octet configuration and becoming positively charged cations.
Non-metals (Group V, VI, VII) tend to gain valence electrons, achieving an octet configuration and becoming negatively charged anions.
Ionic Bonds are hence the strong electrostatic forces of attraction between the cations and the anions.
Note that an ionic bond occurs so long as there is a cation and anion, meaning that it does not necessarily have to occur between the atom that lost an electron and the atom that gained that electron. Electron transfer only allows atoms to reach their stable octet configuration, and become ionised.

For dot-and-cross diagrams, do not draw the electron shells. Remember to use a different symbol for each element and remember to draw the square brackets with the charges.
Draw the bond observed in sodium chloride NaCl here:

Giant Ionic Compounds:
- are made up of ions arranged in a giant ionic lattice, held together by ionic bonds which are strong and require a large amount of energy to break. Hence, these compounds have high melting and boiling points.
- can conduct electricity in molten/aqueous state as there are free ions to act as mobile charge carriers to conduct electricity.
- cannot conduct electricity in solid state as the ions are arranged in a giant ionic lattice and there are no free ions to act as mobile charge carriers to conduct electricity.
- can be dissolved in polar substances (both can carry charge and like dissolves like).

**<u>Covalent Bonds</u>** are formed when two atoms share valence electrons from their valence shells. Covalent Bonds are hence the strong electrostatic forces of attraction between the positively charged nucleus and the shared valence electrons.

For dot-and-cross diagrams, do not draw the electron shells. Remember to use a different symbol for each element and alternate the symbols if there is more than one electron pair. Draw the bond observed in water H2O here:

Simple Covalent Molecules:
- are made up of simple and definite molecules which contain strong intramolecular covalent bonds that are hard to break. These molecules are held together by weak intermolecular forces of attraction that require a small amount of energy to overcome. Hence these compounds have low melting and boiling points.
- are made up of neutral molecules, thus there are no particles to act as mobile charge carriers to conduct electricity.
- can dissolve in organic non-polar solvents as most of these molecules do not have charges.

Giant Molecular/Macromolecular Substances:
- are made up of atoms held together in a lattice by strong covalent bonds that require a large amount of energy to break. Hence these compounds have high melting and boiling points.
- are made up of neutral atoms, and all electrons are involved in bonding. Thus there are no mobile charged particles to carry charges and conduct electricity. [Exception of graphite]
- are insoluble in any type of solvent as they are neither organic nor polar.

**Metallic Bonds** are formed between the positively charged metal cations and a 'sea' of delocalised electrons.
Each metal atom's valence electron can break free from the atom, making it a cation, and wander free throughout the entire metal. This produces an array of metal cations embedded in a 'sea' of freely moving delocalised electrons.
Metallic Bonds are hence the electrostatic forces of attraction between the positively charged metal cations and a 'sea' of delocalised electrons.

Metals do not normally have dot-and-cross diagrams. For diagrams of a giant metallic structure, draw the orderly, larger metal cations in a sea of smaller, delocalised electrons. Make sure that there is the right ratio of electrons to metal ions.
Draw the structure of magnesium Mg here:

Giant Metallic Substances:
- are made up of atoms held together in a lattice by strong metallic bonds that require a large amount of energy to break. Hence these compounds have high melting and boiling points.
- consist of positive ions surrounded by a 'sea' of delocalised electrons. These delocalised electrons can act as mobile charge carriers to conduct electricity.
- are insoluble in any type of solvent as they are neither organic nor polar. However they can react with water (but reacting =/= dissolving)

Link to trend of melting points of metals across period:
Across the period, metals have increasing melting points. This is because each metal cation provides more delocalised electrons per cation across the period. (For instance, each sodium cation provides one electron, each magnesium cation provides two electrons and each aluminium cation provides three electrons.) Hence, across the period, because of the increasing numbers of delocalised electrons provided per metal cation, the metallic bond between the metal cation and the sea of delocalised electrons increases across the period. Thus, more energy is required to break this bond, and the melting point of metals increases across the period.

**Dative Bonding/Coordinate Bonding** is a covalent bond where the shared pair of electrons is provided by only one of the bonded atoms, where one atom is the donor and the other is the acceptor.
The donor has to have at least one pair of unshared electrons in its valence shell.
The acceptor has to have at least one vacant orbital in its valence shell.
Examples: Hydronium Ion ($H_3O^+$), Nitrogen atom in Ammonia, Carbon Monoxide

For dot-and-cross diagrams, make sure that the shared electron pair uses the same symbols.
For structural formula, ensure that an arrow points from the donor to the acceptor instead of a line.
Draw the bond observed in ammonia $NH_3$ here:

**Polar Covalent Bonds** occur when the shared pair of electrons are pulled nearer to one atom in the covalent bond, forming a partial negative and partial positive charge.
(Electronegativity is the tendency of an atom to attract valence electrons)
If two atoms of equal electronegativity bond together, the shared pair of electrons would be halfway between the atoms. This is a pure covalent bond and the atoms are normally the same.
If B is slightly more electronegative than A, B has more electron density than A and becomes slightly negative. A is therefore slightly positive. This is a polar covalent bond.
If B is a lot more electronegative than A, the electron pair is dragged right over to B's end of the bond and A has lost control of its electron while B has complete control over both electrons. A therefore becomes an anion while B becomes a cation. An ionic bond has been formed.
The difference in the two atoms' electronegativity thus indicates the type of bond formed.

Polar molecules need to have a net dipole moment. Even if a molecule has polar bonds in it, if there is no net dipole moment, the molecule is not polar.
Example: HCl has a net dipole moment and is polar but $CF_4$ does not and hence is not polar.
Only $H_2O$ needs to be remembered as a polar molecule.

**Case-Study: Diamond vs Graphite**

Diamond and graphite are both giant covalent molecules and are allotropes of carbon (different forms of the same element). However, their atoms are arranged differently and hence they have different properties.

| | Diamond | Graphite |
|---|---|---|
| Hardness | Hardest natural substance<br>Drill tips for drilling equipment<br>Glass cutters | Soft<br>Solid lubricant reduces engine friction<br>Pencil Lead |
| | C atoms strongly and covalently bonded in tetrahedral units, making structure hard and strong. | C atoms arranged in parallel layers. Atoms are covalently bonded within each layer but weak IMF between layers allow them to slide against one another, making graphite soft and slippery. |
| Electrical Conductivity | Does not conduct electricity<br>Each C atom bonds with 4 others, all 4 valence electrons involved in covalent bonding with other C atoms, no mobile charge carriers. | Conducts electricity<br>Each C atom bonds with 3 others, only 3 valence electrons are involved in bonding and there is one delocalised electron that acts as a mobile charge carrier to conduct electricity. |