

Topic 4a – Bonding

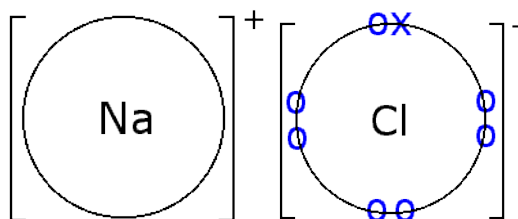
Revision Notes

1) Introduction

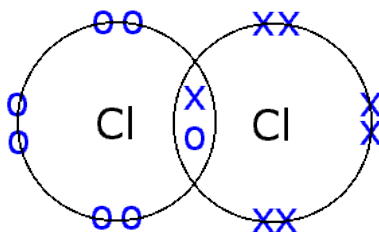
- Atoms form bonds to get a full outer shell of electrons
- Elements in Group 0 (He, Ne, Ar, Kr, Xe, Rn) do not normally undergo chemical reactions as they have full outer shells and, so, do not need to form bonds to other elements
- There are three types of bonding: ionic, covalent and metallic
- Crystals (solids) have one of four types of structure: giant ionic, simple molecular, giant covalent and giant metallic. Giant structures have millions of atoms or ions held together in 3 dimensions. Simple molecular substances consist of a countable number of atoms e.g. H_2O , $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
- Giant covalent is sometimes called giant molecular or macromolecular
- To melt a substance the forces holding the particles together need to be broken
- To conduct electricity there must something charged that can move (ions or electrons).

2) Types of Bonding

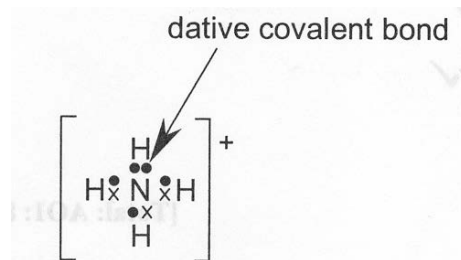
- Ionic bonding – metals transfer electrons to non-metals producing positive metal ions and negative non-metal ions. An ionic bond is an electrostatic attraction between oppositely charged ions. Dot-cross diagrams show outer electrons only e.g. NaCl



- Covalent bonding – A covalent bond is a shared pair of electrons. Only non-metals can get a full shell by sharing electrons. The bond is the attraction of the shared electrons for the two nuclei. Dot-cross diagrams show outer electrons only e.g. Cl_2



- In dative covalent bonds, one atom donates a lone pair of electrons to form the bond. The atom accepting the electrons has an empty orbital
- Examples of species that donate lone pairs include NH_3 , H_2O , Cl^-
- Examples of species that accept lone pairs include H^+ and BF_3

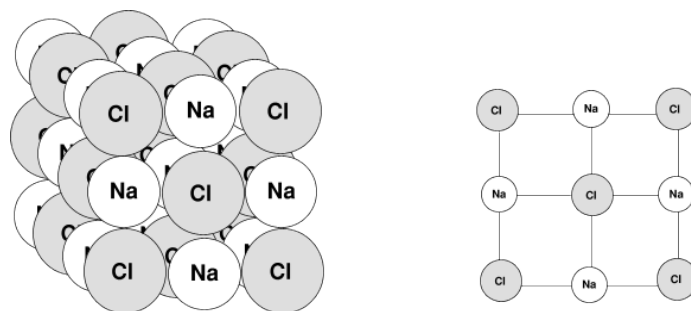


- Metallic bonding – metals lose their outer shell electrons to produce a lattice of positive metal ions surrounded by delocalised (free) electrons.

3) Types of Structure

a) **Giant ionic lattices e.g. sodium chloride**

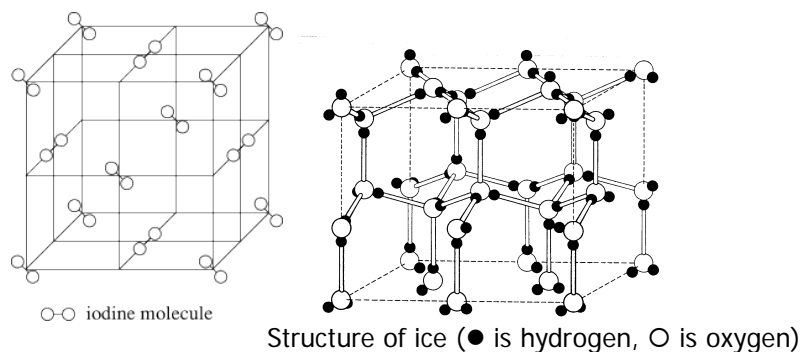
- Lattice of oppositely charged ions.
- High melting and boiling points (strong forces of attraction between **ions** need to be broken).
- Do not conduct when solid (ions not free to move).
- Conduct when molten or dissolved in water (ions then free to move).
- Are brittle (layers of ions cannot slide over each other without repelling)



Note – should be Na^+ and Cl^-

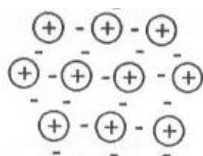
b) **Molecular lattices e.g. iodine and ice**

- Consists of molecules held together by weak intermolecular forces (see section 5 below)
- Low melting and boiling points (weak forces of attraction between **molecules** are easily broken)
- Do not conduct (no free electrons)



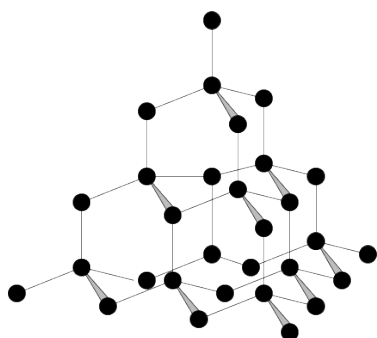
c) Giant metallic lattices e.g. magnesium

- Lattice of metal ions surrounded by delocalised electrons
- High melting and boiling points usually (strong forces of attraction between metal ions and free electrons need to be broken)
- Conduct when solid as delocalised electrons can move
- Are malleable (bendy) and ductile (can be drawn into wires) because the delocalised electrons allow the layers of ions to slide over each other without repelling

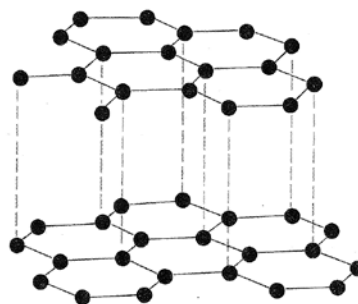


d) Giant covalent lattices e.g. diamond, graphite

- Lattice of non-metal atoms joined by strong covalent bonds
- Very high melting and boiling points usually (many strong covalent bonds to be broken)
- Diamond doesn't conduct (no free electrons)
- In graphite each C forms covalent bonds to 3 other C's within the layers. There are weak forces between the layers allowing them to slide over each other (hence use in pencils and as a lubricant). Fourth outer shell electron is delocalised and can move between the layers allowing graphite to conduct electricity



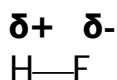
Diamond



graphite

4) Electronegativity and bond polarity

- Electronegativity is the ability of an atom to attract the electrons in a covalent bond.
- When there is a big difference in electronegativity between the atoms at either end of a covalent bond the electrons will be pulled towards the more electronegative atom (shared unequally). This is a polar covalent bond (the molecule has a permanent dipole)
- For example, fluorine is more electronegative than hydrogen so the H-F bond is polar



- Polar molecules have permanent dipoles that don't cancel out (e.g. H₂O) because the dipoles are at an angle
- Non-polar molecules either have no dipoles (e.g. Cl₂) or dipoles that cancel out (e.g. CO₂) because the dipoles are at 180°

5) Intermolecular Forces

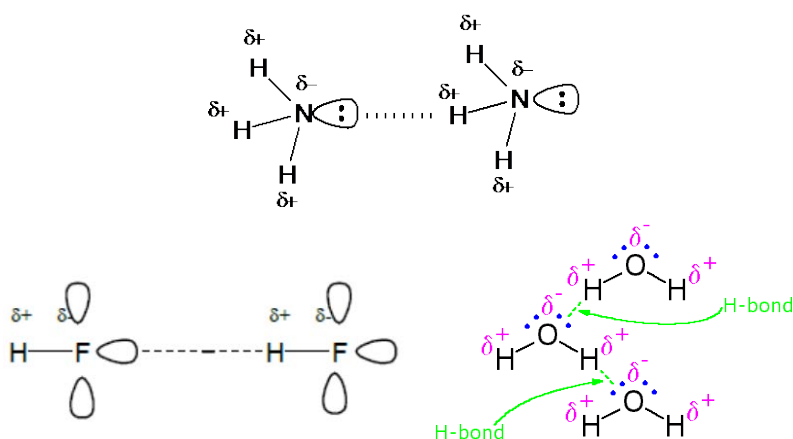
- **Intermolecular forces (IMF) only occur in simple molecular substances. Mentioning molecules or IMF in questions about ionic, metallic or giant covalent substances is a chemical error (CE) and will lose all marks for that question**
- There are 2 types of intermolecular force

a) Van der Waal's

- Arise from temporary dipole (uneven distribution of electrons) in one molecule that induces dipole in another molecule
- The more electrons, the stronger the van der Waal's forces between molecules
- Van der Waal's forces occur in all simple molecular substances

b) Hydrogen bonds

- Only occur when hydrogen is bonded to one of the 3 most electronegative elements: N, O and F. Other elements are not electronegative enough
- There is a large difference in electronegativity between H and N/O/F. This makes H δ⁺ and N/O/F δ⁻
- δ⁺ H is strongly attracted to lone pair on N/O/F in another molecule
- Diagram must show lone pairs, dipoles and H-bond shown by dotted line (see example for NH₃ below. (Note that N in NH₃ has 1 lone pair, O in H₂O has 2 lone pairs and F in HF has 3 lone pairs)



- Water, ammonia and HF have higher melting and boiling points than expected due to the strength of the hydrogen bonds that have to be broken
- Ice is less dense than water because, in ice, the molecules are held further apart so there is more space in the structure

6) Shapes of Molecules

The following procedure allows the shape of a molecule to be worked out:

- Draw a dot-cross diagram
- Count number of electron pairs round the central atom
- **Pairs of electrons repel each other and get as far apart as possible**
- Lone pairs repel more than bonding pairs so bonds are pushed closer together e.g. 107° in ammonia compared with the tetrahedral bond angle of 109.5° in methane

Number of pairs (explanation)	Examples	Name of shape	Bond angle
2 bonding pairs (repel equally)	BeCl ₂	Linear	180°
3 bonding pairs (repel equally)	BF ₃	Trigonal planar	120°
4 bonding pairs (repel equally)	CH ₄ , NH ₄ ⁺ , PCI ₄ ⁺	Tetrahedral	109.5°
5 bonding pairs (repel equally)	PCI ₅	Bipyramidal	90° and 120°
6 bonding pairs (repel equally)	SF ₆ , PCI ₆ ⁻	Octahedral	90°
3 bonding, 1 lone (lp repels more than bps)	NH ₃	Pyramidal	107°
2 bonding, 2 lone (lps repels more than bps)	H ₂ O, NH ₂ ⁻	V-shaped	104.5°
4 bonding, 2 lone (lps repels more than bps)	XeF ₄	Square Planar	90°

This can be done mathematically e.g. PCI₄⁺

Outer shell electrons (P)	5
Bonds formed	4
Adjust for charge (lost e ⁻)	<u>-1</u>
Total for P	8
Pairs for P	4
Lone pairs (pairs-bonds)	0

∴ tetrahedral, bond angle 109.5°

Topic 4b – Periodicity Revision Notes

1) Blocks in the Periodic Table

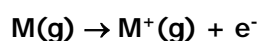
An element can be assigned to the s, p or d block by working out which sub-level its outermost electron is in:

- Na is $1s^2 2s^2 2p^6 3s^1$ so is in the s-block
- O is $1s^2 2s^2 2p^4$ so is in the p-block
- Sc is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ so is in the d-block

2) Ionisation Energies

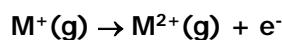
a) First ionisation energy

- Evidence that electrons are arranged in shells or energy levels can be obtained by measuring ionisation energies
- The first ionisation energy of an element is the energy needed to remove one mole of electrons from one mole of gaseous atoms i.e.



b) Successive Ionisation Energies of an element

- Second and subsequent ionisation energies of an element can also be measured i.e.



- Second ionisation is greater than first ionisation energy as it is harder to remove an electron from a positive ion than from a neutral atom
- Jumps in ionisation energies occur when going from one energy level (shell) to another. The jump occurs because the new energy level is closer to nucleus and less shielded

c) Trends Down Group 2 (Be-Ba)

- 1st ionisation energy decreases down Group 2 (and all other groups)
- With each successive element, there is an extra electron shell, so the outer electron is further from the nucleus and more shielded. Less attraction between the nucleus and the outer shell
- This is evidence for electrons being arranged in energy levels (shells)

3) Trends Across Period 3 (Na-Ar)

Be able to describe and explain trends in the following:

1st ionisation energy Describe – general increase
Explain – bigger nuclear charge, same shielding
Dip from Mg to Al because Mg is losing 3s electron, Al is losing 3p. 3p is higher in energy, easier to remove.
Dip from P to S because P is 3p³, S is 3p⁴. Mutual repulsion of paired electrons in S make electron easier to remove than in P.

Atomic radius Describe – it decreases
Explain – bigger nuclear charge, same shielding

Electronegativity Describe – it increases
Explain – bigger nuclear charge, same shielding, stronger attraction between nucleus and shared pair of electrons

Melting & boiling points Na, Mg and Al have metallic bonding. Attraction between positive ions and delocalised free electrons is strong so melting points are high. Melting point increases from Na to Mg to Al because metal ion has greater charge and there are more free electrons per ion so metallic bonding is stronger. Si has very high melting point. Giant covalent structure has many strong covalent bonds to be broken. P₄, S₈ and Cl₂ have low melting points. These are simple covalent molecules held together by weak Van der Waal's forces. Van der Waal's forces increase with molecular mass so S₈ has highest melting point, then P₄ then Cl₂. Ar has simple atomic structure. Fewest electrons, weakest Van der Waal's forces between atoms

Periodic trends are repeated across different rows of the Periodic Table. Any trend in properties across Period 3 will also be shown across Period 2 (and period 4, 5 etc)