# Topic 2: Bonding and Structure

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

**2.1** Know that ionic bonding is the strong electrostatic attraction between oppositely charged ions.

Ionic bonding occurs between metals and non-metals. It is the strong electrostatic attraction between oppositely charged ions in all directions, forming a giant ionic lattice.

This is an exam-mark phrase – always state "strong electrostatic attraction between oppositely charged ions".

**2.2** Understand the effects that ionic radius and ionic charge have on the strength of ionic bonding.

Ionic bonding is stronger when:

- Ionic charge is greater (e.g. $Mg^{2+}O^{2-}$ stronger than $Na^{+}Cl^{-}$) – greater electrostatic attraction.
- Ionic radius is smaller (ions are closer together) – greater charge density, stronger attraction.

This explains trends in melting points and lattice energies of ionic compounds.

**2.3** Understand the formation of ions in terms of electron loss or gain.

Metals lose electrons to form positive ions (cations), achieving a stable noble gas electronic configuration.

Non-metals gain electrons to form negative ions (anions), achieving a stable noble gas electronic configuration.

**2.4** Be able to draw electronic configuration diagrams of cations and anions using dot-and-cross diagrams.

Show the transfer of electrons from metal to non-metal. Use dots for one element and crosses for the other. Show the charge on each ion and square brackets around ions. Only outer shell electrons need to be shown (though full diagrams are accepted).

**2.5** Understand reasons for the trends in ionic radii down a group and for a set of isoelectronic ions.

Down a group: ionic radii increase because there are more electron shells, so the outermost electrons are further from the nucleus.

Isoelectronic ions (e.g. $N^{3-}$, $O^{2-}$, $F^{-}$, $Na^{+}$, $Mg^{2+}$, $Al^{3+}$): all have the same number of electrons but increasing nuclear charge across the series. More protons pull the electron cloud in more tightly, so ionic radius decreases as nuclear charge increases.

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2.6 Understand that the physical properties of ionic compounds and the migration of ions provide evidence for the existence of ions.

Evidence for ions:

- Ionic compounds conduct electricity when molten or dissolved (ions are free to move and carry charge) but not when solid (ions are fixed in the lattice).
- Electrolysis causes ions to migrate to oppositely charged electrodes – cations to cathode, anions to anode.
- These observations confirm the presence of charged particles (ions).

2.7 Know that a covalent bond is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them.

A covalent bond is the strong electrostatic attraction between the two nuclei and the shared pair of electrons between them. This is the exam-mark definition. Covalent bonds form between non-metal atoms.

2.8 Be able to draw dot-and-cross diagrams to show electrons in covalent substances, including: (i) single, double and triple bonds, (ii) dative covalent (coordinate) bonding, including $Al_2Cl_6$ and ammonium ion.

Dot-and-cross diagrams show shared and lone pairs. Use dots for one atom and crosses for the other.

- Single bond: one shared pair
- Double bond: two shared pairs
- Triple bond: three shared pairs (e.g. $N_2$)

Dative covalent bond: both electrons in the shared pair come from one atom (the donor). Show with an arrow from donor to acceptor. E.g. $NH_4^+$ (N donates to $H^+$), $Al_2Cl_6$ (Cl donates lone pair to electron-deficient Al).

2.9 Understand the relationship between bond lengths and bond strengths for covalent bonds.

Shorter bonds are stronger bonds. As bond order increases (single → double → triple), the bond gets shorter and stronger (higher bond enthalpy). More shared electrons means greater electrostatic attraction between nuclei and shared electrons.

2.10 Understand that the shape of a simple molecule or ion is determined by the repulsion between the electron pairs that surround a central atom.

Electron pair repulsion theory (VSEPR): electron pairs around a central atom repel each other and arrange themselves to be as far apart as possible to minimise repulsion.

Lone pair–lone pair repulsion > lone pair–bonding pair > bonding pair–bonding pair.

Lone pairs are held closer to the nucleus and repel more strongly.

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**2.11** Understand reasons for the shapes of, and bond angles in, simple molecules and ions with up to six outer pairs of electrons.

2 electron pairs: linear, 180° (e.g. $BeCl_2$, $CO_2$)

3 electron pairs: trigonal planar, 120° (e.g. $BCl_3$)

4 electron pairs: tetrahedral, 109.5° (e.g. $CH_4$, $NH_4^+$); with 1 lone pair → pyramidal, 107° (e.g. $NH_3$); with 2 lone pairs → bent, 104.5° (e.g. $H_2O$)

5 electron pairs: trigonal bipyramidal, 90°/120° (e.g. $PCl_5$)

6 electron pairs: octahedral, 90° (e.g. $SF_6$)

Lone pairs reduce bond angles by approximately 2.5° per lone pair.

**2.12** Be able to predict the shapes of, and bond angles in, simple molecules and ions analogous to those specified above using electron-pair repulsion theory.

Method: count the total number of electron pairs (bonding + lone) around the central atom. Determine the base shape from the total pairs, then adjust for lone pairs. Always state the name of the shape and the bond angle. Treat double/triple bonds as single "regions of electron density" for shape purposes.

**2.13** Know that electronegativity is the ability of an atom to attract the bonding electrons in a covalent bond.

Electronegativity is the ability of an atom to attract the bonding electrons in a covalent bond. Electronegativity increases across a period (more protons, same shielding) and decreases down a group (more shielding, greater atomic radius). Fluorine is the most electronegative element.

**2.14** Know that ionic and covalent bonding are the extremes of a continuum of bonding type and that electronegativity differences lead to bond polarity.

Bonding exists on a spectrum: purely covalent (equal sharing, e.g. $Cl_2$) → polar covalent (unequal sharing) → ionic (complete electron transfer). The greater the electronegativity difference, the more ionic the bond character. A small difference gives a polar covalent bond (δ+ and δ−).

**2.15** Understand that molecules with polar bonds may not be polar molecules and be able to predict whether or not a given molecule is likely to be polar.

A molecule is polar if it has an overall dipole moment (dipoles do not cancel). Symmetric molecules (e.g. $CO_2$, $CCl_4$, $BF_3$) have polar bonds but are non-polar overall because the dipoles cancel due to symmetry. Asymmetric molecules (e.g. $H_2O$, $CHCl_3$) are polar because the dipoles do not cancel.

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**2.16** Understand the nature of intermolecular forces: (i) London forces, (ii) permanent dipoles, (iii) hydrogen bonds.

(i) London (dispersion) forces: instantaneous dipole–induced dipole attractions. Present in all molecules. Strength increases with number of electrons (larger Mr) and surface area of contact.

(ii) Permanent dipole–dipole: attractions between molecules with permanent dipoles (δ+ attracted to δ−). Stronger than London forces alone for molecules of similar size.

(iii) Hydrogen bonds: a special, strong type of permanent dipole interaction. Occurs when H is bonded to N, O or F (highly electronegative atoms with lone pairs). The lone pair on N/O/F is attracted to the δ+ H on an adjacent molecule.

**2.17** Understand the interactions in molecules, such as $H_2O$, liquid $NH_3$ and liquid HF, which give rise to hydrogen bonding.

Hydrogen bonding requires:

- A hydrogen atom bonded to N, O, or F (δ+ H)
- A lone pair on N, O, or F on an adjacent molecule

In $H_2O$: each molecule can form up to 4 hydrogen bonds (2 lone pairs + 2 H atoms). In $NH_3$: only 1 lone pair limits H-bonding. In HF: only 1 H and 3 lone pairs.

**2.18** Understand the anomalous properties of water resulting from hydrogen bonding: (i) relatively high melting and boiling temperatures, (ii) density of ice compared to water.

(i) Water has a higher boiling point than expected for its molecular mass because extra energy is needed to break the hydrogen bonds between molecules.

(ii) Ice is less dense than liquid water because in ice, hydrogen bonds hold molecules in an open, regular lattice structure with large spaces. When ice melts, some H-bonds break, molecules move closer together, and density increases.

**2.19** Be able to predict the presence of hydrogen bonding in molecules analogous to those mentioned above.

Look for molecules containing H bonded directly to N, O, or F. If the molecule has this feature AND a lone pair on the N/O/F, hydrogen bonding can occur. Examples: alcohols (O–H), carboxylic acids (O–H), amines (N–H).

**2.20** Understand, in terms of intermolecular forces, physical properties shown by materials.

(i) Alkane boiling points increase with chain length: more electrons → stronger London forces → more energy needed.

(ii) Branching decreases boiling point: less surface area of contact → weaker London forces.

(iii) Alcohols have higher boiling points than alkanes of similar Mr: alcohols can form hydrogen bonds (extra energy needed).

(iv) HF has an anomalously high b.p. compared to HCl due to hydrogen bonding. HCl < HBr < HI shows increasing London forces with more electrons. HF breaks the trend due to H-bonding.

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**2.21** Understand factors that influence the choice of solvents.

(i) Water dissolves many ionic compounds: water molecules hydrate the ions (δ+ H attracted to anions, δ− O attracted to cations), overcoming the lattice energy.

(ii) Water dissolves simple alcohols: the OH group forms hydrogen bonds with water.

(iii) Water is a poor solvent for halogenoalkanes: they are polar but cannot form hydrogen bonds with water, and the energy released from interactions with water cannot overcome the existing intermolecular forces.

(iv) Non-aqueous solvents dissolve non-polar substances: "like dissolves like" – similar intermolecular forces in solute and solvent.

**2.22** Know that metallic bonding is the strong electrostatic attraction between metal ions and the delocalised electrons.

Metallic bonding is the strong electrostatic attraction between positive metal ions (cations) and a sea of delocalised electrons. Metallic bond strength increases with greater charge on the ion and smaller ionic radius (e.g. Al > Mg > Na).

**2.23** Know that giant lattices are present in: (i) ionic solids, (ii) covalently bonded solids such as diamond, graphite and silicon(IV) oxide, (iii) solid metals.

Giant ionic lattice: regular arrangement of alternating positive and negative ions (e.g. NaCl). High m.p. due to strong electrostatic attractions.

Giant covalent lattice: atoms bonded by strong covalent bonds throughout (e.g. diamond, graphite, $SiO_2$). Very high m.p.

Giant metallic lattice: regular arrangement of metal cations surrounded by delocalised electrons.

**2.24** Know that the structure of covalently bonded substances such as iodine and ice is simple molecular.

Simple molecular structures consist of small, discrete molecules held together by weak intermolecular forces (London forces, dipole–dipole, or hydrogen bonds). Within each molecule, atoms are joined by strong covalent bonds. Low melting and boiling points because only weak intermolecular forces need to be overcome (not covalent bonds).

**2.25** Know the different structures formed by carbon atoms, including graphite, diamond and graphene.

Diamond: each C bonded to 4 others tetrahedrally; very hard; does not conduct electricity (no free electrons/ions).

Graphite: each C bonded to 3 others in flat hexagonal layers; layers held by weak London forces so can slide (lubricant); one delocalised electron per C → conducts electricity.

Graphene: a single layer of graphite; strong, flexible, excellent conductor of electricity and heat.

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| 2.26 | Be able to predict the type of structure and bonding present in a substance from numerical data and/or other information.<br>High m.p. + conducts when molten/dissolved → giant ionic<br>Very high m.p. + does not conduct (unless graphite) → giant covalent<br>High m.p. + conducts when solid → giant metallic<br>Low m.p. + does not conduct → simple molecular<br>Always link properties to the type of bonding/forces that must be overcome. |
| 2.27 | Be able to predict the physical properties of a substance in terms of particles present, structure, type of bonding and intermolecular forces.<br>Always state: (1) the type of structure (giant ionic/covalent/metallic or simple molecular), (2) the particles present, (3) the type of bonding or intermolecular forces, (4) whether these are strong or weak, and therefore (5) whether a lot or a little energy is needed to overcome them. This determines melting/boiling points, electrical conductivity, and solubility. |

ABOUT THESE NOTES

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