Chemistry with Chloe

AQA A Level Chemistry — 3.1 Physical Chemistry

3.1.3 Bonding

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3.1.3.1 *Ionic bonding: electrostatic attraction between oppositely charged ions in a lattice. Formulas of compound ions.*

Ionic bonding: strong electrostatic attraction between oppositely charged ions in a giant ionic lattice.

Formed by electron transfer from metal to non-metal.

Compound ions: sulfate SO■²■, hydroxide OH■, nitrate NO■■, carbonate CO■²■, ammonium NH■■.

Predict ionic charge from position in the Periodic Table.

3.1.3.2 *Covalent and dative covalent bonds. Single, multiple and co-ordinate bonds.*

Covalent bond: shared pair of electrons between two non-metal atoms. Strong electrostatic attraction between shared electrons and both nuclei.

Multiple bonds: double = 2 shared pairs; triple = 3 shared pairs.

Dative covalent (co-ordinate) bond: both electrons donated by one atom. Shown as an arrow (e.g. NH■■, Al■Cl■).

3.1.3.3 *Metallic bonding: attraction between delocalised electrons and positive ions in a lattice.*

Metallic bonding: strong electrostatic attraction between positive metal cations and a sea of delocalised electrons.

Strength increases with: more delocalised electrons, smaller ionic radius, higher charge.

Explains trends in melting point and hardness across metals.

3.1.3.4 *Four types of crystal structure: ionic, metallic, macromolecular (giant covalent), molecular. Examples and properties.*

Ionic (e.g. NaCl): high m.p., conducts when molten/dissolved, brittle.

Metallic (e.g. Mg): high m.p., conducts when solid, malleable/ductile.

Giant covalent (e.g. diamond, graphite, SiO■): very high m.p., hard (except graphite). Diamond doesn't conduct; graphite does (delocalised electrons in layers).

Simple molecular (e.g. I■, ice): low m.p./b.p., doesn't conduct. Weak intermolecular forces overcome on melting — NOT covalent bonds.

Always link properties to what must be overcome.

3.1.3.5 *Shapes of simple molecules and ions with up to six electron pairs. VSEPR theory and bond angles.*

VSEPR: electron pairs repel and arrange as far apart as possible.

Repulsion order: lone pair–lone pair > lone pair–bond pair > bond pair–bond pair.

2 pairs: linear, 180° (BeCl■, CO■)

3 pairs: trigonal planar, 120° (BF■); with 1 lp: bent ~118°

4 pairs: tetrahedral, 109.5° (CH■, NH■■); with 1 lp: pyramidal 107° (NH■); with 2 lp: bent 104.5° (H■O)

5 pairs: trigonal bipyramidal, 90°/120° (PCl■)

6 pairs: octahedral, 90° (SF■)

Lone pairs reduce bond angles by ~2.5° each.

3.1.3.6 *Electronegativity and bond polarity. Polar bonds vs polar molecules.*

Electronegativity: the power of an atom to attract the bonding pair of electrons in a covalent bond.

Difference in electronegativity → polar bond ($\delta+$ and $\delta-$).

A molecule is polar only if dipoles do NOT cancel.

Symmetric molecules (e.g. CO■, CCl■) have polar bonds but are non-polar overall.

3.1.3.7 *Intermolecular forces: permanent dipole–dipole, London forces, hydrogen bonding. Effect on melting/boiling points.*

London forces (induced dipole–dipole): present in ALL molecules. Increase with number of electrons (higher Mr). Temporary fluctuations in electron density induce dipoles in neighbours.

Permanent dipole–dipole: between molecules with permanent dipoles.

Hydrogen bonding: H bonded to N, O, or F + lone pair on N/O/F of adjacent molecule. Strongest intermolecular force.

Anomalous properties of water: high b.p.; ice less dense than liquid water (open H-bonded lattice).