

Chemical Bonding and Molecular Structure

why atoms pair up, and the shape the pairing leaves behind

Why bonds form at all

- isolated atoms sit higher in energy than bonded ones: a bond forms whenever it lowers energy
- octet rule: atoms gain, lose, or share electrons to reach eight valence electrons
- the count noble gases already have (hydrogen and helium are content with two)
- the octet rule is a guide, not a law: fails for electron-deficient BF_3
- odd-electron NO , and expanded-octet PF_5 , SF_6

Ionic bond vs covalent bond

Ionic (electrovalent)

- complete transfer of electrons, one atom to another
- forms between a low-ionisation-energy metal and a high-electron-affinity non-metal
- e.g. Na loses an electron, Cl gains it: Na^+Cl^-
- held together by electrostatic attraction, no fixed direction

Covalent

- electrons are shared, not handed over, usually between two non-metals
- a single bond shares one pair, a double bond two pairs, a triple bond three
- e.g. two Cl atoms share one pair in Cl_2
- has a definite direction in space, which is why molecules have a shape at all

the split isn't sharp: most real bonds sit somewhere between the two, which is exactly what bond polarity describes

Lattice enthalpy and ionic bond formation



energy released when gaseous ions pack into a solid ionic lattice (the reverse of lattice enthalpy, which NCERT defines as the energy needed to separate the solid back into gaseous ions)



an ionic bond forms readily when the energy released packing the lattice comfortably outweighs the energy spent ionising the metal and forming the anion; a small, highly charged ion pair packs tighter and releases more energy, so it also has the higher lattice enthalpy

Lewis structures and formal charge

- a Lewis structure shows every valence electron: bonding pairs as a dot pair or line
- lone pairs sit on one atom alone
- count electrons: total valence electrons from every atom, adjusted for the overall charge
- formal charge = valence electrons – nonbonding electrons – half the bonding electrons
- preferred structure: formal charges closest to zero, negative charge on the more electronegative atom
- resonance**: some molecules (ozone, carbonate) aren't described by one Lewis structure alone
- the real molecule blends two or more contributing structures, none the actual one