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Topic 3.1: Intermolecular and Interparticle Forces

Revision Notes · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

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Forces Between Particles, Not Bonds Within Them

Intermolecular forces are attractions between separate particles, one molecule to the next. The covalent bonds inside a molecule are intramolecular, and far stronger: an O-H bond costs 463 kJ/mol to break, against a few tens to pull two water molecules apart.

When dry ice sublimes, the CO_2 molecules leave the solid whole: only attractions between molecules are overcome, and no C=O bond breaks. Boiling water is the same story, its O-H bonds surviving unchanged in the steam.

London Dispersion Forces: Every Substance Has Them

Dispersion forces arise because electrons never sit still. At any instant an electron cloud can bunch slightly to one side, producing a temporary dipole. That fleeting dipole induces a matching dipole in a neighbour, and the two attract.

Needing nothing but electrons in motion, dispersion forces act between all atoms and molecules, polar or not. In a nonpolar substance such as CH_4 or I_2 they are the only intermolecular force present.

Strength follows polarizability, how easily an electron cloud is distorted. A large cloud, spread over many electrons and held loosely, distorts readily; a small, tight cloud barely does. Dispersion strength therefore grows with electron count and size: helium (2 electrons) boils at -269 °C, xenon (54) at -108 °C.

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Strength also follows contact. A long straight chain presents a large contact area to its neighbours; a compact branched molecule presents far less. Pentane and neopentane are both C_5H_{12}, yet pentane boils at 36 °C and neopentane at 10 °C, on shape alone.

Dipole-Dipole Interactions: the Permanent Dipoles

A polar molecule carries a permanent dipole: its electron density is permanently uneven, one end δ+ and the other δ-. Dipole-dipole interactions are the attractions between the δ+ end of one such molecule and the δ- end of the next.

The requirement is strict: polar bonds, in a geometry that stops the bond dipoles cancelling. HCl qualifies, with one polar bond and nothing to cancel it. CO_2 does not: its two polar bonds point in opposite directions and cancel.

A polar substance always has dispersion forces too; the dipole term is added, never swapped in. Because dispersion grows with size while the dipole term does not, compare only molecules of similar size. CH_3F and ethane each contain 18 electrons, yet polar CH_3F boils at -78 °C against ethane's -89 °C, purely from the added dipole term.

Hydrogen Bonding: the Strongest Molecular Case

Hydrogen bonding is a dipole-dipole interaction strong enough to earn its own name. It needs two features together: a donor, hydrogen covalently bonded to nitrogen, oxygen or fluorine, and an acceptor, a lone pair on a nitrogen, oxygen or fluorine of a neighbouring molecule.

It is strong for two compounding reasons. Nitrogen, oxygen and fluorine are highly electronegative, so a hydrogen bonded to one of them is left unusually exposed and strongly δ+. Hydrogen has no inner shells, so its δ+ centre approaches a lone pair very closely, where attraction peaks.

Check both features before claiming it. CH_3F contains fluorine but cannot hydrogen bond with itself: none of its hydrogens is bonded to the fluorine. H_2S fails differently: sulfur is not electronegative enough, so it keeps only ordinary dipole-dipole forces.

Counting donors and acceptors predicts the network. Water offers two donors and two lone pairs, joining up to four hydrogen bonds; ammonia offers three donors but one lone pair, capping its network. Hence the boiling points: H_2O at 100 °C, NH_3 at -33 °C.

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Ion-Dipole Forces and the Induced Dipoles

Ion-dipole forces act between a full ionic charge and the permanent dipole of a polar molecule. Around a dissolved Na^+, the δ- oxygen ends of water cluster in; around Cl^-, the δ+ hydrogen ends do. Strength follows Coulomb's law: it grows with ionic charge and falls as ionic radius rises, so Mg^{2+}, twice the charge in a smaller ion (72 pm against 102 pm), pulls water dipoles far harder.

A charge or dipole can also create the dipole it then attracts. A polar molecule beside a nonpolar one distorts its cloud and induces a temporary dipole: a dipole-induced dipole interaction. An ion does the same more strongly, an ion-induced dipole interaction, as between Mg^{2+} and benzene.

Rankings need a size warning. For similar sized particles the usual order holds: dispersion < dipole-dipole < hydrogen bonding < ion-dipole. But dispersion grows without limit as particles grow, and in I_2 it beats the dipole-dipole forces of small HCl.

Worked Examples

Example 1: Ranking two noble gases. Argon (18 electrons) boils at -186 °C, krypton (36) at -153 °C. Both are monatomic and nonpolar, so only dispersion forces act. Krypton's larger cloud is more polarizable, so its temporary dipoles are stronger and its atoms harder to separate. Electron count alone predicts the ranking; the 33 °C gap confirms it.

Example 2: The same electrons, different forces. H_2S and SiH_4 each contain 18 electrons, their molar masses within 6.1%, so their dispersion forces are similar. SiH_4 is tetrahedral and nonpolar, and stops there. H_2S is bent and polar, so dipole-dipole forces act on top, raising its boiling point to -60 °C against SiH_4's -112 °C.

Example 3: Pricing the forces with ΔHvap. ΔHvap prices the total attraction directly: the energy to pull one mole of particles completely apart. CH_4, dispersion only: 8.2 kJ/mol. H_2O, with its hydrogen bond network: 40.7 kJ/mol. The ratio, 40.7 ÷ 8.2 ≈ 5.0, says separating water costs about five times the energy of separating methane.

Example 4: The hydrogen halide anomaly. Boiling points run HF 20 °C, HCl -85 °C, HBr -67 °C, HI -35 °C. From HCl to HI, electron counts climb 18, 36, 54 and boiling points climb with them: dispersion at work. HF breaks the pattern: with only 10 electrons it should boil lowest, yet it boils 105 °C above HCl, because HF alone hydrogen bonds.

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Traps That Cost Marks in Topic 3.1

Claiming hydrogen bonding from the formula alone. A hydrogen bond needs a hydrogen bonded to N, O or F and a lone pair on N, O or F next door. CH_3F, CH_4 and H_2S each fail one half of that test, and questions plant them because they look close.

Forgetting that polar molecules still have dispersion forces. HCl has dispersion forces and dipole-dipole forces. Writing that a polar substance has dipole-dipole forces instead of dispersion forces breaks every ranking built on it: in a large polar molecule, dispersion is often the bigger term.

Treating the strength ladder as absolute. Dispersion < dipole-dipole < hydrogen bonding holds only for comparable sizes. A large enough electron cloud wins outright: solid I_2, gaseous HCl.

Breaking bonds that never break. Boiling, melting and subliming overcome forces between particles; covalent bonds inside them survive. Dry ice does not sublime because its C=O bonds snap.

Explaining a boiling trend with the bonds inside the molecules. Halogen boiling points rise from F_2 to I_2 because dispersion forces between molecules grow with polarizability, not because bonds inside strengthen: Cl-Cl is 242 kJ/mol, I-I 151, yet I_2 boils 218 °C higher.

Ignoring shape when the formulas are identical. Isomers share molar mass and electron count, so neither can explain a boiling point gap. Pentane beats neopentane by 26.6 °C on contact area alone.

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Key Relationships to Remember

- Intermolecular forces act between particles, intramolecular bonds within them; phase changes overcome only the intermolecular forces (dry ice sublimes with its C=O bonds intact)
- London dispersion forces: present in all atoms and molecules, from temporary dipoles inducing dipoles in neighbours; strength grows with polarizability and surface contact area
- Dipole-dipole interactions: between the permanent dipoles of polar molecules, always in addition to dispersion forces, never instead of them
- Hydrogen bonding: needs H bonded to N, O or F (donor) and a lone pair on N, O or F (acceptor); networks follow donor and acceptor counts (H_2O four per molecule, NH_3 capped by one lone pair)
- Ion-dipole forces: between an ion and a polar molecule; strength rises with ionic charge, falls with ionic radius (Mg^{2+}, 72 pm, binds water far more strongly than Na^+, 102 pm)
- Strength ladder for similar sized particles: dispersion < dipole-dipole < hydrogen bonding < ion-dipole; large electron clouds break the ladder (I_2 solid, HCl gas)
- Boiling point and ΔHvap measure total attraction (H_2O 40.7 kJ/mol, CH_4 8.2 kJ/mol)

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