Intermolecular forces

London forces, permanent dipole interactions and hydrogen bonding, the anomalous properties of water, and how these forces explain boiling temperatures and solubility.

Types of intermolecular force

Intermolecular forces act **between** molecules. They are much weaker than the covalent bonds within molecules.

London forces

London (dispersion) forces act between all molecules and atoms. Electrons move constantly, so at any instant the electron cloud is uneven and the molecule has a temporary (**instantaneous**) dipole. This **induces** a dipole in a neighbouring molecule, and the two attract. The force:

- increases with the **number of electrons** in the molecule, because larger electron clouds are more easily distorted,
- is stronger between molecules with a **greater surface area** of contact (unbranched chains).

Permanent dipole–permanent dipole interactions

Molecules with polar bonds that do not cancel have a permanent dipole. The $\delta+$ end of one molecule attracts the $\delta-$ end of another. These forces act **in addition to** London forces.

Hydrogen bonds

A **hydrogen bond** is a strong permanent dipole attraction between a hydrogen atom bonded to **N, O or F** and a **lone pair** on an N, O or F atom of another molecule. The H carries a high $\delta+$ charge, because N, O and F are very electronegative and the hydrogen has no inner electrons to shield its nucleus.

Examples: H_2O, liquid NH_3, HF, alcohols, carboxylic acids and amines. In a diagram show the lone pair, the H, the partial charges and a dotted line, with the three atoms in a straight line.

> **NOTE**
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> Hydrogen bonding can occur in any molecule with $-OH$ or $-NH$ groups. Do not say a hydrogen bond is a bond to hydrogen; it is an attraction between molecules.

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Anomalous properties of water

- **High melting and boiling temperatures** compared with H_2S and other hydrides in Group 6. Each water molecule can form up to four hydrogen bonds (two through its lone pairs and two through its hydrogens), and extra energy is needed to break them.
- **Ice is less dense than liquid water.** In ice each molecule is held by hydrogen bonds in an open hexagonal lattice, with the molecules further apart than in liquid water. On melting, some hydrogen bonds break, the lattice collapses and the molecules pack more closely.

Boiling temperatures

Boiling a liquid overcomes the intermolecular forces.

Observation	Explanation
Alkane boiling temperatures increase with chain length	More electrons, stronger London forces
Branched alkanes boil lower than straight-chain isomers	Branches reduce the surface area of contact, so there are fewer London forces
Alcohols boil higher than alkanes with a similar number of electrons	Alcohols have hydrogen bonding as well as London forces
HF has a high boiling temperature; then $HCl < HBr < HI$	HF forms hydrogen bonds. From HCl to HI London forces increase with the number of electrons

> **TRY IT**
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> Explain why ethanol (M_r = 46) boils at 78 °C but propane (M_r = 44) boils at −42 °C.

Both have a similar number of electrons and similar London forces. Ethanol also has hydrogen bonds between its $-OH$ groups, which need a lot more energy to overcome, so it has a much higher boiling temperature.

Solubility

Choice of solvent depends on the forces between the solute and the solvent.

- **Ionic compounds** dissolve in water because water molecules are polar and form **ion–dipole attractions** with the ions, a process called **hydration**. The energy released by hydration pays for the breaking of the lattice.
- **Small alcohols** dissolve in water because they can form hydrogen bonds with water molecules.

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- **Halogenoalkanes** are poorly soluble in water. They are polar, but they cannot form hydrogen bonds with water, so the new interactions are not strong enough to replace the hydrogen bonds between the water molecules.
- **Non-aqueous solvents** such as hexane dissolve non-polar substances, because the forces in the solute and solvent are similar (London forces), so mixing releases about as much energy as it takes.

"Like dissolves like." A substance dissolves when the new interactions between solute and solvent are comparable in strength with the ones broken.

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