

London, dipole and hydrogen bond attractions between molecules

Chemistry grades 11-12 (A Level · IB · Class 11-12) · Lesson

Quick answer: Intermolecular forces are the attractions between molecules: London forces act between all particles and grow with the number of electrons, permanent dipole-dipole forces act between polar molecules, and hydrogen bonds act when H bonded to N, O or F meets a lone pair. Stronger forces mean higher boiling points.

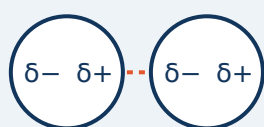
Level: Chemistry grades 11-12 (A Level · IB · Class 11-12) · **Curriculum:** IB S2.2

What are intermolecular forces?

Inside a molecule, atoms are held by strong covalent bonds. Between molecules there are much weaker attractions. When water boils or iodine sublimates, it is these weak forces that break, not the covalent bonds. A typical O-H covalent bond needs about 460 kJ mol^{-1} to break, while the attraction between neighbouring water molecules is about 20 kJ per mole of hydrogen bonds.

Definition

Intermolecular forces are attractions between neighbouring molecules (or atoms, for noble gases). Their strength decides the melting point, the boiling point and the volatility of a molecular substance.



London forces: a momentary dipole in one atom induces a dipole in its neighbour



Permanent dipole-dipole: $\delta+$ end of one molecule attracts $\delta-$ end of the next



Hydrogen bond: H on O, N or F attracted to a lone pair on O, N or F

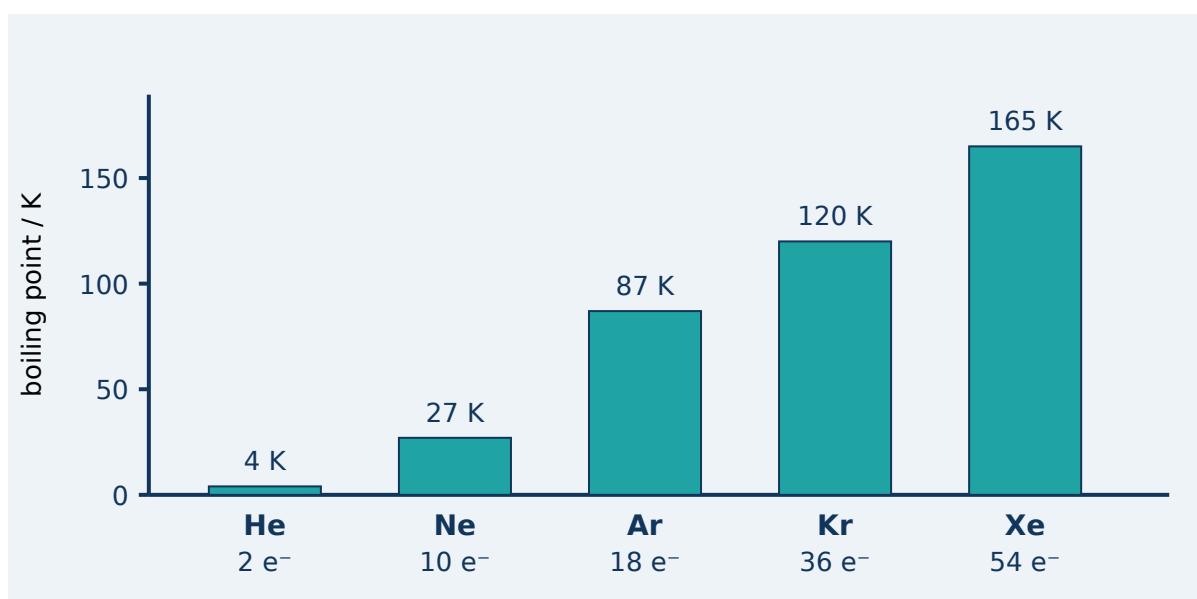
The three kinds of intermolecular force. Red dashes show the attraction.

How do London forces arise?

Electrons are always moving, so at any instant the charge in an atom or molecule is slightly unevenly spread. This **instantaneous dipole** induces an opposite dipole in a neighbour, and the two attract. London (dispersion) forces exist between all particles, even non-polar ones.

Key law

London forces grow stronger as the number of electrons increases, so larger molecules have higher boiling points. For the same number of electrons, long straight chains pack more contact surface than compact, branched molecules, so they also boil higher.



More electrons give stronger London forces and a higher boiling point.

Alkanes show the same pattern: methane boils at -162°C , butane at -1°C and hexane at 69°C . Branching lowers the boiling point: pentane boils at 36°C , but its branched isomer 2,2-dimethylpropane boils at 9.5°C .

What are permanent dipole-dipole forces?

A molecule with a net dipole, such as HCl or propanone, has a permanent $\delta+$ end and $\delta-$ end. Neighbouring molecules line up so that opposite ends attract. These forces come in addition to London forces.

Worked example 1

Fluorine, F_2 , and hydrogen chloride, HCl, each have 18 electrons. F_2 boils at -188°C and HCl at -85°C . Explain the difference.

The number of electrons is equal, so the London forces are similar.

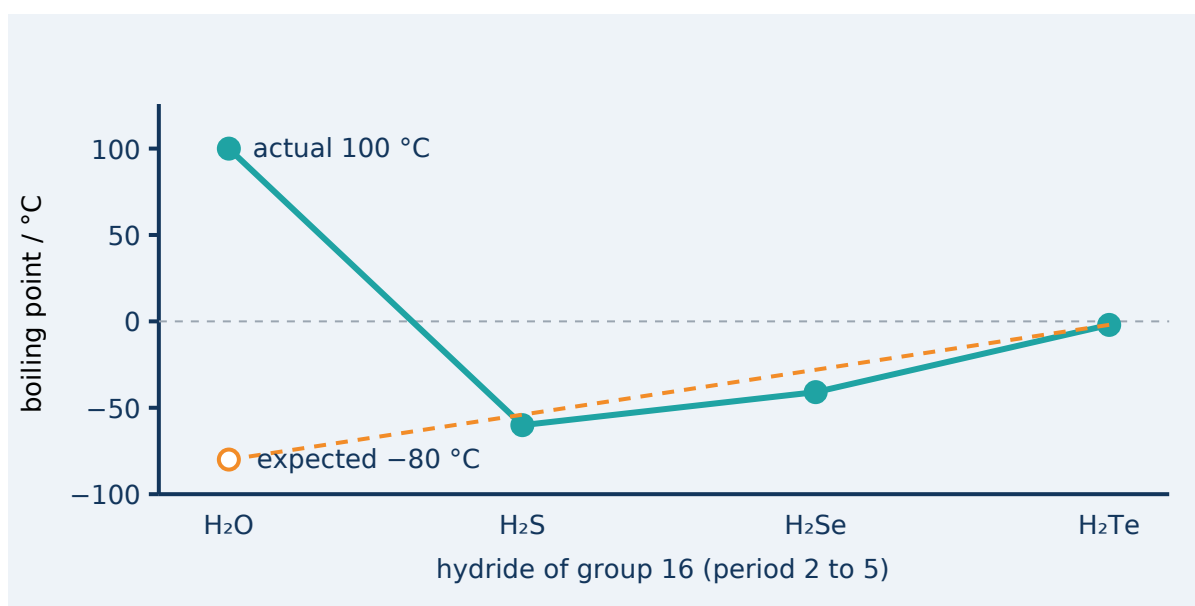
HCl is polar, so it also has permanent dipole-dipole forces, which F_2 lacks. More energy is needed to separate HCl molecules, so **HCl has the higher boiling point.**

What is a hydrogen bond?

Definition

A **hydrogen bond** is a strong intermolecular attraction between a hydrogen atom bonded to N, O or F ($\delta+$) and a lone pair on a different N, O or F atom ($\delta-$). The three atoms lie close to a straight line.

Hydrogen bonds are stronger than ordinary dipole-dipole forces because the small, highly polarised H atom can approach the lone pair very closely. They explain why water (100 °C) boils far above H₂S (−60 °C), as the graph shows.



Water boils about 180 °C above the value expected from the trend of its heavier relatives, because of hydrogen bonding.

Worked example 2

Ethanol (C₂H₅OH) boils at 78 °C and its isomer methoxymethane (CH₃OCH₃) at −24 °C. Explain.

Both have 26 electrons, so London forces are similar.

Ethanol has an O-H group, so molecules form **hydrogen bonds**. Methoxymethane has no H bonded to O, so it has only dipole-dipole and London forces. Ethanol therefore boils far higher.

Common mistake

Boiling water does not break the O-H bonds. Writing that "covalent bonds break when water boils" is wrong: only the intermolecular hydrogen bonds are broken, and the molecules stay as H₂O.

Force	Present in	Relative strength
London forces	all particles	weakest, grows with electrons

Force	Present in	Relative strength
Permanent dipole-dipole	polar molecules such as HCl, propanone	stronger than London for equal size
Hydrogen bonding	molecules with N-H, O-H or F-H	strongest intermolecular force

Lina's tip

To rank boiling points, go through three steps: first check for hydrogen bonding, then for a permanent dipole, and finally compare the number of electrons. Compare molecules with similar numbers of electrons first.

Key points to remember

- London forces: instantaneous dipoles, stronger with more electrons
- Permanent dipole-dipole forces act between polar molecules
- Hydrogen bond: H on N, O or F attracted to a lone pair on N, O or F
- Boiling breaks intermolecular forces, not covalent bonds
- Branching lowers boiling points by reducing surface contact

London, dipole and hydrogen bond attractions between molecules: frequently asked questions

Why does the boiling point increase down a group of noble gases?

Each heavier noble gas has more electrons, so its electron cloud is larger and more easily distorted. London forces become stronger, and more energy is needed to separate the atoms.

What conditions are needed for hydrogen bonding?

A hydrogen atom must be bonded to nitrogen, oxygen or fluorine, and a different molecule must have a lone pair on N, O or F. The H atom then points directly at that lone pair.

Why does branching lower the boiling point?

A branched molecule is more compact, so it touches its neighbours over a smaller surface. The London forces are weaker, and less energy is needed to separate the molecules.

Are intermolecular forces stronger than covalent bonds?

No. Covalent bonds are typically 10 to 100 times stronger. That is why changes of state, which only overcome intermolecular forces, need much less energy than chemical reactions that break bonds.

Why does ice float on water?

In ice each molecule forms four hydrogen bonds in an open hexagonal lattice with empty space. This makes ice about 8 to 9 % less dense than liquid water, so it floats.