A-LEVEL CHEMISTRY • EXERCISE SHEET

3.6 Intermolecular forces, electronegativity and bond properties

Name: ____________ Class: ________ Date: __________ **Total: 31 marks**

KEY WORDS dipole 偶极 • intermolecular forces 分子间作用力 • hydrogen bonding 氢键 • permanent dipole 永久偶极 • induced dipole 诱导偶极

Objective

Build the skills to answer exam questions on **intermolecular forces** — **bond polarity** 键极性 and dipoles, **van der Waals' forces** 范德华力, **hydrogen bonding** 氢键, and the anomalous properties of water.

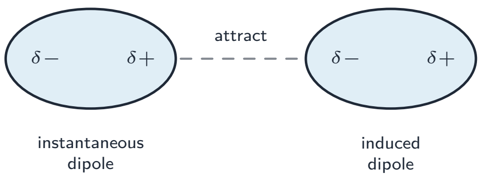
You must be able to:

- explain bond polarity and dipole moments using electronegativity
- describe London (id–id) and permanent dipole (pd–pd) forces, and hydrogen bonding
- use hydrogen bonding to explain the anomalous properties of water

1 Worked examples

Polarity and van der Waals' forces

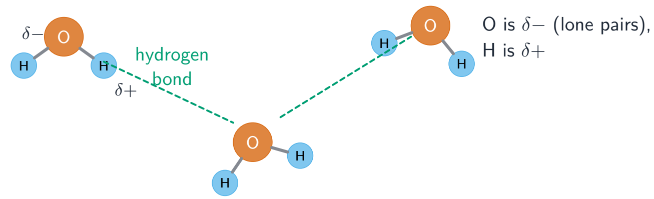
A bond between atoms of different electronegativity has a **dipole** 偶极 ($\delta-$ / $\delta+$). If the dipoles do not cancel, the molecule is **polar**. **Van der Waals' forces** is the general term for all intermolecular forces.



London (id–id) force: a momentary dipole induces one in a neighbour — acts between all molecules, stronger with more electrons

- **London (id–id)** force — acts between all molecules; stronger when there are more electrons.
- **permanent dipole (pd–pd)** force — between always-polar molecules.

Hydrogen bonding and water



Hydrogen bonding: a $\delta+$ H (bonded to N, O or F) is attracted to a lone pair on a $\delta-$ N, O or F

Hydrogen bonding (a strong pd–pd case) explains water's anomalies: high melting/boiling point (many H-bonds to break); high surface tension; and **ice is less dense** than water (H-bonds hold an open structure), so ice floats.

The three forces, in the order you should check for them

Every molecule has **van der Waals' forces** (induced dipole–induced dipole). A polar molecule has those **plus** permanent dipole–dipole. A molecule with H bonded to N, O or F has those **plus** hydrogen bonding. So the forces stack; the question is which one dominates.

Force	Between	Strength
van der Waals' 范德华力	all molecules; arise from instantaneous, fluctuating dipoles inducing dipoles in neighbours	weakest, but grows with the number of electrons and with surface contact
permanent dipole–dipole 永久偶极	polar molecules, from a difference in electronegativity across a bond	intermediate
hydrogen bonding 氢键	a lone pair on N, O or F and an H attached to N, O or F	strongest of the three

A polar bond does not make a polar molecule. CO_2 has two strongly polar bonds, but they are opposite and equal in a linear molecule, so the dipoles cancel and the molecule is non-polar. **Check the shape before deciding.**

Reading a boiling-point trend — the two-part answer the mark scheme wants:

1. Name the force being broken.
2. Explain why it is stronger or weaker here.

So along the alkanes: van der Waals' forces increase with more electrons and greater surface contact, so boiling point rises. A branched isomer boils lower than its straight-chain isomer because branching reduces the surface contact, weakening the van der

Waals' forces — **the covalent bonds are never broken on boiling**, and saying so loses the mark.

Worked example. Explain why H_2O boils at 100 °C but H_2S at −60 °C, even though H_2S has more electrons.

1. Both are polar bent molecules with dipole–dipole forces and van der Waals' forces.
2. In water, H is bonded to the highly electronegative, small oxygen atom, which also has lone pairs, so **hydrogen bonding** occurs; sulfur is not electronegative enough for it.
3. Hydrogen bonds are much stronger than the dipole–dipole and van der Waals' forces in H_2S, so far more energy is needed to separate water molecules — outweighing the larger number of electrons in H_2S.

Ice is less dense than water for the same reason: hydrogen bonds hold the molecules in an **open tetrahedral lattice** with gaps, so the same number of molecules occupies more volume.

2 Practice

2.1 Name the **two** atoms (besides hydrogen) that must be present for hydrogen bonding in water and ammonia. [2]

2.2 State which intermolecular force acts between all molecules. [1]

2.3 Explain why CO_2 is non-polar despite having polar bonds. [2]

2.4 State the three conditions needed for hydrogen bonding. [2]

2.5 Explain why 2-methylpropane boils at a lower temperature than butane. [2]

3 Exam-style questions

3.1 Why does ice float on water? [1]

- **A** ice is warmer
- **B** hydrogen bonds hold ice in an open structure, so it is less dense
- **C** ice has no hydrogen bonds
- **D** ice contains air

3.2 Which has the **strongest** intermolecular forces? [1]

- **A** CH_4
- **B** Cl_2
- **C** H_2O
- **D** O_2

3.3 (a) Explain why water has a relatively high boiling point. [2]

(b) Explain why the boiling points of the noble gases increase down the group. [2]

3.4 Draw and label a diagram to show hydrogen bonding between two water molecules (show the δ charges and the lone pair). [3]

3.5 The table gives the boiling points of four hydrides.

Hydride	$\mathbf{NH_3}$	$\mathbf{PH_3}$	$\mathbf{H_2O}$	$\mathbf{H_2S}$
boiling point / °C	−33	−88	100	−60

(a) Name all the intermolecular forces present in PH_3. [2]

(b) Explain why NH_3 boils at a much higher temperature than PH_3, even though PH_3 has more electrons. [4]

(c) Explain why H_2O boils higher than NH_3, given that both form hydrogen bonds. [2]

(d) State what is broken when H_2O boils, and what is NOT broken. [2]

(e) Explain, in terms of hydrogen bonding, why ice floats on water. [3]

Solutions

2.1 nitrogen, oxygen (and fluorine).

2.2 the London dispersion / instantaneous dipole–induced dipole force.

2.3 The two $C = O$ dipoles are equal in size and point in opposite directions; the molecule is linear, so they cancel and there is no overall dipole.

2.4 A hydrogen atom bonded to N, O or F; and a lone pair on an N, O or F atom of another molecule.

2.5 2-methylpropane is branched, so its molecules have less surface contact with each other; the van der Waals' forces between them are therefore weaker and less energy is needed to separate them.

3.1 B. Hydrogen bonds hold ice in an open lattice, making it less dense than water.

3.2 C. Water has hydrogen bonding, the strongest of the four.

3.3 (a) water molecules are held by hydrogen bonds; much energy is needed to break these, so the boiling point is high.
(b) down the group atoms have more electrons, so stronger London forces, needing more energy to overcome — higher boiling points.

3.4 a dashed hydrogen bond from a $\delta+$ H on one molecule to a lone pair on the $\delta-$ O of the other; both molecules drawn with $\delta+$ on H and $\delta-$ on O; a clear lone pair shown on the accepting oxygen.

3.5 (a) van der Waals' forces; permanent dipole–dipole forces.
(b) NH_3 has hydrogen bonding, because H is bonded to the small, highly electronegative nitrogen, which also carries a lone pair; PH_3 has only van der Waals' and dipole–dipole forces, since phosphorus is not electronegative enough; hydrogen bonds are much stronger, so far more energy is needed to separate ammonia molecules, outweighing PH_3's greater number of electrons.
(c) Oxygen is more electronegative than nitrogen and each water molecule has two hydrogens and two lone pairs, so water forms more hydrogen bonds per molecule, and stronger ones.
(d) The hydrogen bonds and other intermolecular forces between the molecules are broken; the covalent O–H bonds within the molecules are NOT broken.
(e) In ice each molecule is held by hydrogen bonds to four others in a rigid tetrahedral arrangement; this open lattice contains gaps, so the same number of molecules occupies a larger volume; the density is therefore lower than that of liquid water and ice floats.