

Exercise walk-through

# 3.6

## Intermolecular forces, electronegativity and bond properties

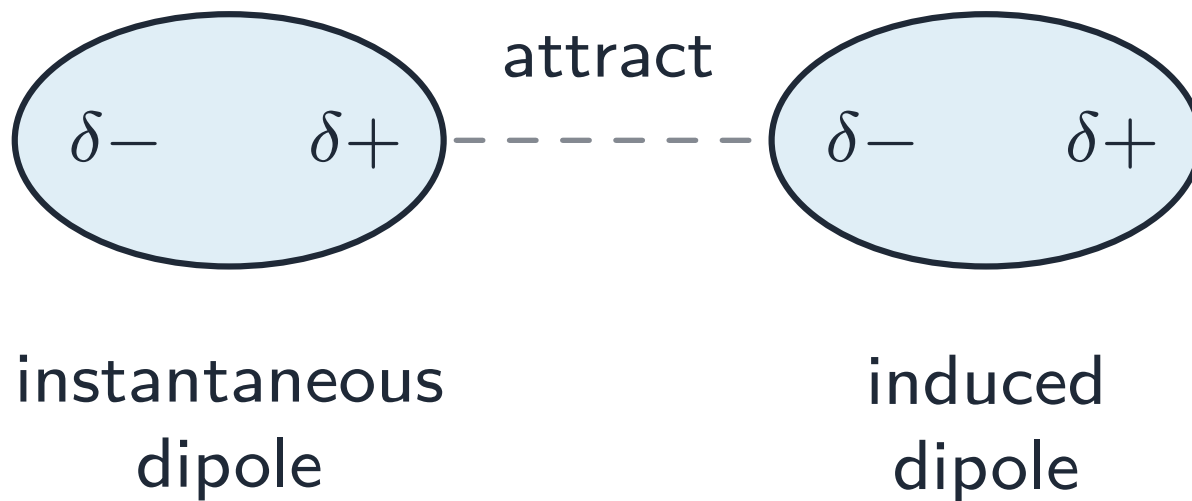
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A-Level Chemistry

# Intermolecular forces

A bond is **polar** if the atoms differ in electronegativity (a **dipole** ).

- **London forces** – between all molecules; stronger with more electrons
- permanent dipole forces – between polar molecules



# Van der Waals' forces, and when a molecule is polar

## Van der Waals' forces

**Van der Waals' forces** is the general name for all intermolecular forces. There are two main types.

## London dispersion

The instantaneous dipole–induced dipole force, or **London dispersion force**. Moving electrons make a brief **instantaneous dipole**, which induces a matching **induced dipole** next door. It acts between **all** molecules.

## Permanent dipole

Between molecules that are always polar, because each carries a **permanent dipole**.

## Dipole moment

If the bond dipoles in a molecule do **not** cancel, the whole molecule has a **dipole moment** and is polar – which is why  $\text{CO}_2$  is not, and  $\text{H}_2\text{O}$  is.

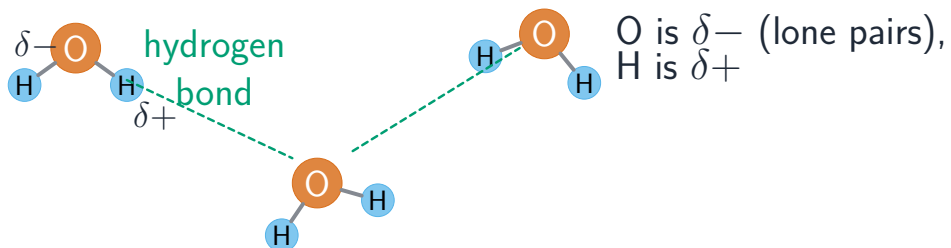
Shape decides polarity: identical bonds arranged symmetrically cancel, however polar each bond is on its own.

## Bond polarity and dipoles

When two atoms with different electronegativity share a bond, the electrons sit closer to the more electronegative atom.

The bond then has a **polarity** : one end is slightly negative ( $\delta-$ ) and the other slightly positive ( $\delta+$ ).

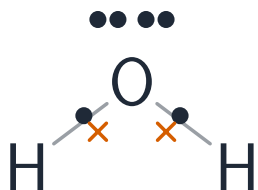
# Hydrogen bonding



**Hydrogen bonding** forms when H is bonded to N, O or F and meets a lone pair on N, O or F.

It gives water a high **boiling point**, high **surface tension**, and makes ice **less dense** (so it floats).

# Dot-and-cross diagrams



H<sub>2</sub>O: 2 bonding pairs  
+ 2 lone pairs



N<sub>2</sub>: triple bond  
(3 shared pairs)

- electron from one atom
- x electron from the other

Covalent dot-and-cross: each shared pair is one electron from each atom. Water has two bonding pairs and two lone pairs; nitrogen shares three pairs (a triple bond)

## Exercises 2.1-2.3

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

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

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**2.3** Explain why  $\text{CO}_2$  is non-polar despite having polar bonds. [2]

## Exercises 2.1-2.3

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

**2.1**

- nitrogen, oxygen (and fluorine)

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

**2.2**

- the London dispersion / instantaneous dipole-induced dipole force

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**2.3** Explain why  $\text{CO}_2$  is non-polar despite having polar bonds. [2]

**2.3**

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

## Exercises 2.4·2.5

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

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

## Exercises 2.4·2.5

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

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

**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

## Exercises 3.1-3.2

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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

☐

A  $\text{CH}_4$

B  $\text{Cl}_2$

C  $\text{H}_2\text{O}$

D  $\text{O}_2$

---

3.2 Which has the **strongest** intermolecular forces?

[1]

## Exercises 3.1·3.2

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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

3.1 B

- Hydrogen bonds hold ice in an open lattice, making it less dense than water

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3.2 Which has the **strongest** intermolecular forces?

[1]

☐

A CH<sub>4</sub>

B Cl<sub>2</sub>

C H<sub>2</sub>O

D O<sub>2</sub>

3.2 C

- Water has hydrogen bonding, the strongest of the four

## Exercise 3.3

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

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

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

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

**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

## Exercise 3.4

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

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### 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

## Exercise 3.5 (a)

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**3.5** The table gives the boiling points of four hydrides.

| Hydride                                  | $\text{NH}_3$ | $\text{PH}_3$ | $\text{H}_2\text{O}$ | $\text{H}_2\text{S}$ |
|------------------------------------------|---------------|---------------|----------------------|----------------------|
| boiling<br>point /<br>$^{\circ}\text{C}$ | −33           | −88           | 100                  | −60                  |

(a) Name all the intermolecular forces present in  $\text{PH}_3$ .

[2]

## Exercise 3.5 (a)

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**3.5** The table gives the boiling points of four hydrides.

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| boiling point / $^{\circ}\text{C}$ | -33           | -88           | 100                  | -60                  |

(a) Name all the intermolecular forces present in  $\text{PH}_3$ . [2]

(a)

- van der Waals' forces; permanent dipole–dipole forces

## Exercise 3.5 (b)

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**3.5** The table gives the boiling points of four hydrides.

| Hydride                                  | $\text{NH}_3$ | $\text{PH}_3$ | $\text{H}_2\text{O}$ | $\text{H}_2\text{S}$ |
|------------------------------------------|---------------|---------------|----------------------|----------------------|
| boiling<br>point /<br>$^{\circ}\text{C}$ | -33           | -88           | 100                  | -60                  |

(b) Explain why  $\text{NH}_3$  boils at a much higher temperature than  $\text{PH}_3$ , even though  $\text{PH}_3$  has more electrons. [4]

## Exercise 3.5 (b)

**3.5** The table gives the boiling points of four hydrides.

| Hydride            | NH <sub>3</sub> | PH <sub>3</sub> | H <sub>2</sub> O | H <sub>2</sub> S |
|--------------------|-----------------|-----------------|------------------|------------------|
| boiling point / °C | −33             | −88             | 100              | −60              |

(b) Explain why NH<sub>3</sub> boils at a much higher temperature than PH<sub>3</sub>, even though PH<sub>3</sub> has more electrons. [4]

(b)

- NH<sub>3</sub> has hydrogen bonding, because H is bonded to the small, highly electronegative nitrogen, which also carries a lone pair; PH<sub>3</sub> 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<sub>3</sub>'s greater number of electrons

## Exercise 3.5 (c)

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**3.5** The table gives the boiling points of four hydrides.

| Hydride                            | $\text{NH}_3$ | $\text{PH}_3$ | $\text{H}_2\text{O}$ | $\text{H}_2\text{S}$ |
|------------------------------------|---------------|---------------|----------------------|----------------------|
| boiling point / $^{\circ}\text{C}$ | -33           | -88           | 100                  | -60                  |

(c) Explain why  $\text{H}_2\text{O}$  boils higher than  $\text{NH}_3$ , given that both form hydrogen bonds. [2]

## Exercise 3.5 (c)

**3.5** The table gives the boiling points of four hydrides.

| Hydride            | NH <sub>3</sub> | PH <sub>3</sub> | H <sub>2</sub> O | H <sub>2</sub> S |
|--------------------|-----------------|-----------------|------------------|------------------|
| boiling point / °C | −33             | −88             | 100              | −60              |

(c) Explain why H<sub>2</sub>O boils higher than NH<sub>3</sub>, given that both form hydrogen bonds. [2]

(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

## Exercise 3.5 (d)

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**3.5** The table gives the boiling points of four hydrides.

| Hydride                                  | $\text{NH}_3$ | $\text{PH}_3$ | $\text{H}_2\text{O}$ | $\text{H}_2\text{S}$ |
|------------------------------------------|---------------|---------------|----------------------|----------------------|
| boiling<br>point /<br>$^{\circ}\text{C}$ | −33           | −88           | 100                  | −60                  |

(d) State what is broken when  $\text{H}_2\text{O}$  boils, and what is NOT broken.

[2]

## Exercise 3.5 (d)

**3.5** The table gives the boiling points of four hydrides.

| <b>Hydride</b>     | <b>NH<sub>3</sub></b> | <b>PH<sub>3</sub></b> | <b>H<sub>2</sub>O</b> | <b>H<sub>2</sub>S</b> |
|--------------------|-----------------------|-----------------------|-----------------------|-----------------------|
| boiling point / °C | −33                   | −88                   | 100                   | −60                   |

(d) State what is broken when H<sub>2</sub>O boils, and what is NOT broken. [2]

(d)

- The hydrogen bonds and other intermolecular forces between the molecules are broken; the covalent O–H bonds within the molecules are NOT broken

## Exercise 3.5 (e)

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**3.5** The table gives the boiling points of four hydrides.

| <b>Hydride</b>           | <b>NH<sub>3</sub></b> | <b>PH<sub>3</sub></b> | <b>H<sub>2</sub>O</b> | <b>H<sub>2</sub>S</b> |
|--------------------------|-----------------------|-----------------------|-----------------------|-----------------------|
| boiling<br>point /<br>°C | −33                   | −88                   | 100                   | −60                   |

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

[3]

## Exercise 3.5 (e)

**3.5** The table gives the boiling points of four hydrides.

| Hydride                            | $\text{NH}_3$ | $\text{PH}_3$ | $\text{H}_2\text{O}$ | $\text{H}_2\text{S}$ |
|------------------------------------|---------------|---------------|----------------------|----------------------|
| boiling point / $^{\circ}\text{C}$ | -33           | -88           | 100                  | -60                  |

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

(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