



# Bonding II

## OCR A Chemistry A H432 | Specification-led notes

SPEC → CRACK IT → WORKED EXAMPLE → TEST IT → DONE

EVERY SPECIFICATION POINT. EXPLAINED. PRACTISED. CHECKED OFF.

### SPECIFICATION COVERAGE

| OCR A REF    | EXACT COVERAGE IN THIS PACK                                                                           | TAUGHT / TESTED              |
|--------------|-------------------------------------------------------------------------------------------------------|------------------------------|
| 2.2.2(g)     | Shapes and bond angles with up to six electron pairs; relative repulsion of bond pairs and lone pairs | Sections 1-4 / Tests 1-4     |
| 2.2.2(h)     | Linear, non-linear, trigonal planar, pyramidal, tetrahedral and octahedral shapes                     | Sections 2-4 / Tests 2-4     |
| 2.2.2(i)     | Electronegativity, Pauling values and the trend towards fluorine                                      | Section 5 / Test 5           |
| 2.2.2(j)     | Polar bonds, permanent dipoles, molecular shape, overall dipoles and dipole cancellation              | Sections 5-6 / Tests 5-6     |
| 2.2.2(k)-(l) | Induced dipole-dipole, permanent dipole-dipole and hydrogen bonding                                   | Sections 7-10 / Tests 7-10   |
| 2.2.2(m)     | Hydrogen bonding and the anomalous properties of water                                                | Section 11 / Test 11         |
| 2.2.2(n)-(o) | Simple molecular lattices and melting, boiling, solubility and electrical conductivity                | Sections 12-15 / Tests 12-15 |

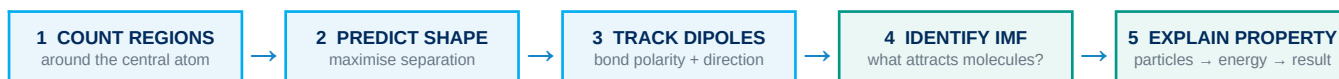
**BOUNDARY** Bonding II completes OCR A 2.2.2(g)-(o). Bonding I is assumed: ionic, covalent, dative and metallic bonding, lattices and structure-property language.

Written for first teaching in Year 12  
Mapped to OCR A Chemistry A H432, specification version 3.1 (May 2026)

# The big picture



**YOU WILL MASTER** How electron-pair repulsion controls shape; how electronegativity creates dipoles; how intermolecular forces control observable properties.



**THE BIG IDEA** Bonding inside a molecule gives its structure. Attractions between molecules decide how easily those molecules separate, move or mix.

## How every teaching section works

| STAGE          | WHAT YOU DO                                                  |
|----------------|--------------------------------------------------------------|
| NOTES          | Learn the required definition, rule, diagram or explanation. |
| WORKED EXAMPLE | Follow a complete question and a visible model answer.       |
| TEST IT        | Answer explicit questions immediately after the teaching.    |
| ANSWERS        | Check every mark point and every model diagram at the end.   |

**EXAM TRAP** A polar bond does not automatically make a polar molecule. Molecular shape decides whether the bond dipoles cancel.



✓ **DONE WHEN** You can move from electron pairs to shape, polarity, intermolecular force and physical property without skipping a link.

# 1. How to determine molecular shape



**SPEC 2.2.2(g)** Predict shapes and bond angles for molecules and ions with up to six electron pairs, including lone pairs.



**CRACK IT** VSEPR means valence-shell electron-pair repulsion. Electron regions repel and arrange as far apart as possible around the central atom.

| TERM            | DEFINITION TO LEARN                                                                                             |
|-----------------|-----------------------------------------------------------------------------------------------------------------|
| central atom    | the atom around which the bonded atoms and electron pairs are arranged                                          |
| bonding pair    | a pair of electrons shared between two covalently bonded atoms                                                  |
| lone pair       | a pair of outer-shell electrons not used in a covalent bond                                                     |
| electron region | one direction of electron density around the central atom; a single, double or triple bond counts as one region |
| molecular shape | the arrangement of atoms around the central atom; lone pairs affect the shape but are not named as positions    |
| bond angle      | the angle between two bonds that meet at the central atom                                                       |

## THE ELECTRON-COUNTING METHOD

**USE THIS METHOD FOR** Common molecules and ions in which H or a halogen is singly bonded to the central atom.

**TOTAL ELECTRON PAIRS** (central-atom group number + one electron for each attached H or halogen + one for each negative charge - one for each positive charge)  $\div$  2

| STEP                  | WHAT TO DO                                                                                 |
|-----------------------|--------------------------------------------------------------------------------------------|
| 1 central atom        | identify the atom around which the shape is being predicted                                |
| 2 group number        | write the number of outer-shell electrons on the central atom                              |
| 3 attached atoms      | add one electron for each attached H or halogen atom                                       |
| 4 charge              | add one electron for each negative charge; subtract one for each positive charge           |
| 5 divide by 2         | divide the electron total by 2 to obtain the total number of electron pairs                |
| 6 bond and lone pairs | bond pairs = attached atoms; lone pairs = total pairs - bond pairs; then predict the shape |

| TOTAL PAIRS | PARENT ELECTRON-PAIR ARRANGEMENT | IDEAL ANGLES |
|-------------|----------------------------------|--------------|
| 2           | linear                           | 180°         |
| 3           | trigonal planar                  | 120°         |
| 4           | tetrahedral                      | 109.5°       |
| 5           | trigonal bipyramidal             | 90° and 120° |
| 6           | octahedral                       | 90°          |

**REPULSION ORDER** lone pair-lone pair > lone pair-bond pair > bond pair-bond pair. Lone pairs repel more strongly and compress bond angles.

**MULTIPLE-BOND CHECK** Do not apply the shortcut blindly to a molecule such as CO<sub>2</sub>. Draw or inspect its structure, then count each double or triple bond as one electron region.



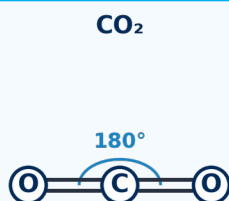
**✓ DONE WHEN** You can calculate the total electron pairs, separate bonding pairs from lone pairs and match the parent arrangement.



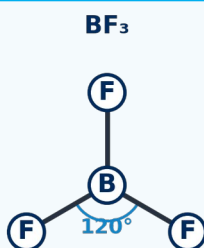
## 2. Shapes with bonding pairs only



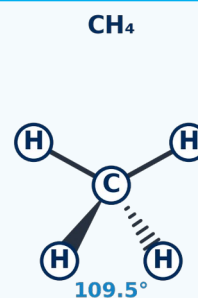
**SPEC 2.2.2(h)** Explain linear, trigonal planar and tetrahedral shapes using electron-pair repulsion.



2 regions • linear



3 regions • trigonal planar



4 regions • tetrahedral

*Every bond is the same sensible length, stops at its atom label and represents one electron region.*

| REGIONS | LONE PAIRS | SHAPE           | ANGLE  | EXAMPLE                                        |
|---------|------------|-----------------|--------|------------------------------------------------|
| 2       | 0          | linear          | 180°   | CO <sub>2</sub>                                |
| 3       | 0          | trigonal planar | 120°   | BF <sub>3</sub>                                |
| 4       | 0          | tetrahedral     | 109.5° | CH <sub>4</sub> , NH <sub>4</sub> <sup>+</sup> |

**WORKED EXAMPLE | Question:** Why is CO<sub>2</sub> linear even though each C=O bond is a double bond?

**COUNT** Each double bond is one region of electron density, so carbon has two regions.

**REPEL** The two regions reach maximum separation by pointing in opposite directions.

**ANSWER** CO<sub>2</sub> is linear with a bond angle of 180°.



**TEST IT 2 [7 marks]**

(a) State the shape and bond angle of BF<sub>3</sub>. [2]

(b) Draw a 3-D diagram of CH<sub>4</sub> using wedge-and-dash bonds. Name its shape and label its bond angle. [4]

(c) Explain why a multiple bond is counted as one electron region in a shape prediction. [1]

DRAWING SPACE FOR TEST IT 2



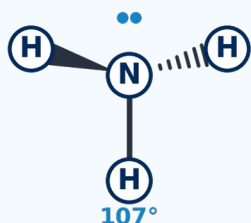
**✓ DONE WHEN** You can move from two, three or four electron regions to the correct bond-pair-only shape.

### 3. Lone pairs change bond angles

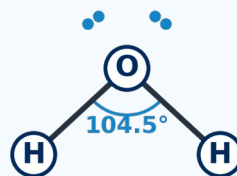


**SPEC 2.2.2(g)-(h)** Use the greater repulsion of lone pairs to explain pyramidal and non-linear shapes and their smaller bond angles.

**REPULSION ORDER** lone pair-lone pair > lone pair-bond pair > bond pair-bond pair. Lone pairs are attracted by only one nucleus, so their electron clouds spread out more and repel more strongly.



**3 bond pairs + 1 lone pair • pyramidal**



**2 bond pairs + 2 lone pairs • non-linear**

*Shape names describe atom positions. Lone pairs control the shape but are not included in its name.*

| MOLECULE         | ELECTRON REGIONS            | ATOM ARRANGEMENT                 | SHAPE             | ANGLE  |
|------------------|-----------------------------|----------------------------------|-------------------|--------|
| NH <sub>3</sub>  | 3 bond pairs + 1 lone pair  | tetrahedral electron arrangement | pyramidal         | 107°   |
| H <sub>2</sub> O | 2 bond pairs + 2 lone pairs | tetrahedral electron arrangement | non-linear (bent) | 104.5° |

**WORKED EXAMPLE | Question: Why is the H-O-H angle smaller than the H-N-H angle?**

**COMPARE** H<sub>2</sub>O has two lone pairs; NH<sub>3</sub> has one lone pair.

**REPULSION** Lone pairs repel more strongly than bonding pairs.

**RESULT** Two lone pairs compress the O-H bonds more, giving 104.5° rather than 107°.



**TEST IT 3 [8 marks]**

- (a) Draw NH<sub>3</sub> in 3-D, show its lone pair, name its shape and label its bond angle. [4]  
 (b) Explain fully why the H-O-H angle in H<sub>2</sub>O is 104.5°, smaller than 109.5°. [4]

DRAWING SPACE FOR TEST IT 3



**✓ DONE WHEN** You use lone-pair repulsion to justify both the shape name and the reduced angle.

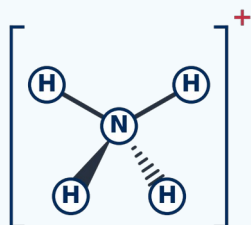
## 4. Molecules and ions with up to six regions



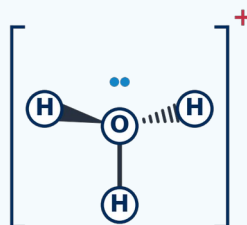
**SPEC 2.2.2(g)-(h)** Predict 3-D shapes of molecules and ions with up to six electron pairs around the central atom.



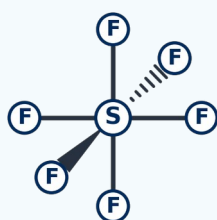
**CRACK IT** For an ion, account for the charge when counting electrons, then draw square brackets and the overall charge. Name the shape from the positions of atoms, not lone pairs.



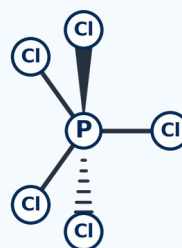
tetrahedral •  $109.5^\circ$



pyramidal • angle not set for OCR recall



octahedral •  $90^\circ$



trigonal bipyramidal •  $90^\circ$  and  $120^\circ$

*PCl<sub>5</sub> is included as a useful five-region reference. OCR's named shape list in 2.2.2(h) focuses on the six shapes used elsewhere on this page and in Sections 2-3.*

| SPECIES                  | AROUND CENTRAL ATOM        | SHAPE                | ANGLE                        |
|--------------------------|----------------------------|----------------------|------------------------------|
| $\text{NH}_4^+$          | 4 bond pairs, 0 lone pairs | tetrahedral          | $109.5^\circ$                |
| $\text{H}_3\text{O}^+$   | 3 bond pairs, 1 lone pair  | pyramidal            | not specified for OCR recall |
| $\text{SF}_6$            | 6 bond pairs, 0 lone pairs | octahedral           | $90^\circ$                   |
| $\text{PCl}_5$ reference | 5 bond pairs, 0 lone pairs | trigonal bipyramidal | $90^\circ$ and $120^\circ$   |

### WORKED EXAMPLE | Question: Predict the shape and bond angle of $\text{NH}_4^+$ .

**COUNT** The central N has four N-H bonding regions and no lone pair in  $\text{NH}_4^+$ .

**REPEL** Four bonding regions arrange tetrahedrally.

**ANSWER**  $\text{NH}_4^+$  is tetrahedral with a bond angle of  $109.5^\circ$ ; show brackets and a + charge.



### TEST IT 4 [10 marks]

- State the number of bond pairs and lone pairs around O in  $\text{H}_3\text{O}^+$ , then name its shape. [3]
- State the shape and smallest bond angle of  $\text{SF}_6$ . [2]
- Draw a 3-D diagram of  $\text{NH}_4^+$ , including brackets and charge, and label its bond angle. [5]

DRAWING SPACE FOR TEST IT 4



**✓ DONE WHEN** You handle charge, brackets, lone pairs and six-region shapes without losing the central-atom count.

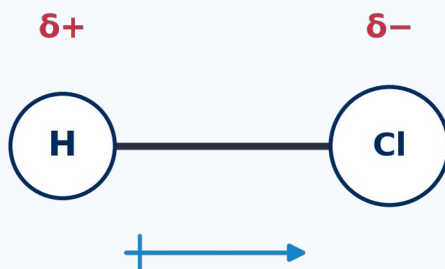
## 5. Electronegativity and polar bonds



**SPEC 2.2.2(i)-(j)** Interpret Pauling electronegativity values and explain polar bonds and permanent dipoles.

| TERM                     | OCR-READY DEFINITION                                                                                                            |
|--------------------------|---------------------------------------------------------------------------------------------------------------------------------|
| <b>electronegativity</b> | the ability of an atom to attract the bonding electrons in a covalent bond                                                      |
| <b>polar bond</b>        | a covalent bond in which the bonding electrons are shared unequally because the bonded atoms have different electronegativities |
| <b>permanent dipole</b>  | a permanent separation of partial charge, shown by $\delta+$ and $\delta-$                                                      |

**PERIODIC TREND** Electronegativity increases across a period and up a group, towards fluorine. Fluorine is the most electronegative element. A greater electronegativity difference usually gives a more polar bond.



**bonding electrons pulled toward the more electronegative Cl atom**

*$\delta$  means partial charge, not a full ionic charge. The crossed tail marks the  $\delta+$  end; the arrowhead points towards  $\delta-$ . The dipole symbol is drawn beside and parallel to the bond.*

| ATOM                    | H   | C   | N   | O   | F   | Cl  |
|-------------------------|-----|-----|-----|-----|-----|-----|
| Pauling value (approx.) | 2.2 | 2.5 | 3.0 | 3.5 | 4.0 | 3.0 |

**WORKED EXAMPLE | Question: Which bond is more polar: C-H or O-H?**

**CALCULATE**  $\Delta\text{EN}(\text{C-H}) = 2.5 - 2.2 = 0.3$ ;  $\Delta\text{EN}(\text{O-H}) = 3.5 - 2.2 = 1.3$ .

**COMPARE** The O-H bond has the larger electronegativity difference.

**ANSWER** O-H is more polar; O is  $\delta-$  and H is  $\delta+$ .



**TEST IT 5 [8 marks]**

- Define electronegativity. [2]
- Use the values above to rank C-H, N-H and O-H from least polar to most polar. Show the differences. [3]
- Add  $\delta+$  and  $\delta-$  to H-Cl and explain why the bond is polar. [3]



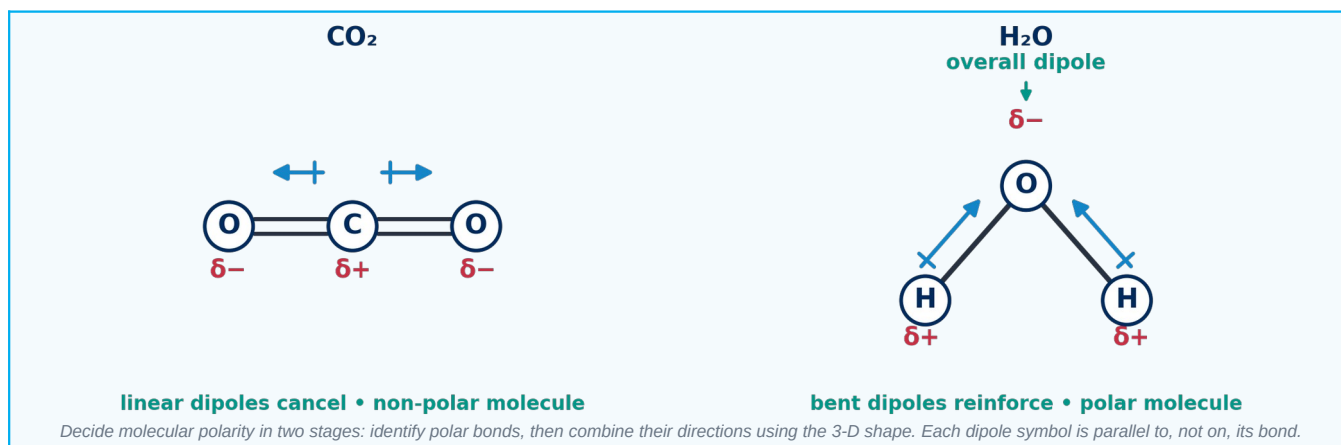
**✓ DONE WHEN** You connect electronegativity difference to unequal sharing and correct partial charges.

## 6. Bond polarity is not molecular polarity



**SPEC 2.2.2(j)** Explain a polar molecule and overall dipole using permanent bond dipoles and molecular shape.

**DEFINITION TO LEARN** A polar molecule has an overall dipole because it contains polar bonds whose dipoles do not cancel due to their direction in the molecular shape.



| MOLECULE         | POLAR BONDS?          | DO DIPOLES CANCEL?                                         | OVERALL   |
|------------------|-----------------------|------------------------------------------------------------|-----------|
| HCl              | yes, H-Cl             | no: there is one bond dipole                               | polar     |
| CO <sub>2</sub>  | yes, C=O              | yes: equal and opposite in a linear molecule               | non-polar |
| BF <sub>3</sub>  | yes, B-F              | yes: symmetrical trigonal planar arrangement               | non-polar |
| CH <sub>4</sub>  | C-H only weakly polar | yes: symmetrical tetrahedral arrangement                   | non-polar |
| NH <sub>3</sub>  | yes, N-H              | no: pyramidal arrangement                                  | polar     |
| H <sub>2</sub> O | yes, O-H              | no: non-linear arrangement                                 | polar     |
| SF <sub>6</sub>  | yes, S-F              | yes: six equal dipoles cancel in an octahedral arrangement | non-polar |

### WORKED EXAMPLE | Question: Why is H<sub>2</sub>O polar but CO<sub>2</sub> non-polar?

**BONDS** Both molecules contain polar bonds.

**SHAPE** CO<sub>2</sub> is linear, so its equal C=O dipoles act in opposite directions and cancel. H<sub>2</sub>O is bent, so its O-H dipoles do not cancel.

**ANSWER** CO<sub>2</sub> has no overall dipole; H<sub>2</sub>O has an overall dipole and is polar.



#### TEST IT 6 [10 marks]

- (a) State whether BF<sub>3</sub>, NH<sub>3</sub>, CH<sub>4</sub> and HCl are polar molecules. Justify each answer using bonds and shape. [8]  
 (b) Draw the bond-dipole directions in CO<sub>2</sub> and H<sub>2</sub>O and state whether each molecule has an overall dipole. [2]

DRAWING SPACE FOR TEST IT 6



**✓ DONE WHEN** You never decide whole-molecule polarity from one bond without checking the 3-D arrangement.

## 7. Intramolecular bonds and intermolecular forces



**SPEC 2.2.2(k)-(o)** Use intermolecular forces to explain simple molecular structures and their physical properties.

| TERM                              | DEFINITION TO LEARN                                                 | WHAT HAPPENS ON BOILING?                          |
|-----------------------------------|---------------------------------------------------------------------|---------------------------------------------------|
| <b>intramolecular bond</b>        | a bond within a particle, such as a covalent bond within a molecule | normally remains intact                           |
| <b>intermolecular force</b>       | an electrostatic attraction between separate molecules              | is overcome so molecules separate                 |
| <b>simple molecular substance</b> | a substance made from small, discrete covalently bonded molecules   | molecules move apart; atoms stay bonded           |
| <b>volatility</b>                 | the tendency of a substance to vaporise                             | high volatility usually means a low boiling point |

**THE NON-NEGOTIABLE DISTINCTION** Covalent bonds hold atoms together inside each molecule. Intermolecular forces attract one molecule to another. Melting or boiling a simple molecular substance overcomes intermolecular forces, not covalent bonds.

| INTERACTION                    | WHERE IT ACTS                                         | KEY CAUSE                                                   |
|--------------------------------|-------------------------------------------------------|-------------------------------------------------------------|
| <b>induced dipole-dipole</b>   | between all atoms and molecules                       | instantaneous and induced dipoles                           |
| <b>permanent dipole-dipole</b> | in addition, between polar molecules                  | attraction between permanent $\delta^+$ and $\delta^-$ ends |
| <b>hydrogen bonding</b>        | between suitable molecules containing N, O or F and H | lone pair attracted to H in -NH, -OH or HF                  |

**OCR TERMINOLOGY** OCR may use van der Waals' forces for permanent dipole-dipole and induced dipole-dipole interactions. Induced dipole-dipole interactions may also be called London (dispersion) forces. These are alternative names, not extra force types. Hydrogen bonding is a particularly strong intermolecular attraction but is still much weaker than a covalent bond.

### WORKED EXAMPLE | Question: What is overcome when iodine boils?

**PARTICLES** Iodine contains discrete  $I_2$  molecules.

**ATTRACTION** The molecules attract through induced dipole-dipole forces.

**CHANGE** Boiling supplies enough energy to overcome those intermolecular forces.

**ANSWER**  $I_2$  molecules separate; the strong I-I covalent bond inside each molecule remains intact.



#### TEST IT 7 [8 marks]

- Define an intermolecular force and an intramolecular bond. [2]
- For solid iodine, identify the intramolecular bond and the intermolecular force. [2]
- Explain why saying 'the covalent bonds break when iodine melts' is incorrect. [2]
- State the relationship between boiling point and volatility. [2]



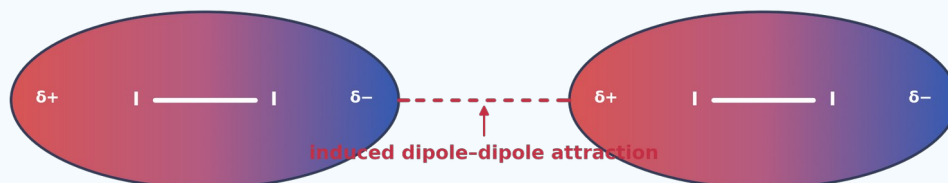
**✓ DONE WHEN** You name exactly what is inside a molecule and what acts between molecules.

## 8. Induced dipole-dipole interactions



**SPEC 2.2.2(k)** Explain induced dipole-dipole interactions, also called London (dispersion) forces.

**DEFINITION TO LEARN** An induced dipole-dipole interaction is an attraction between an instantaneous dipole in one particle and the dipole it induces in a neighbouring particle.



**instantaneous uneven electron density creates opposite temporary dipoles**

*Electron density is always moving. The labelled dipoles are temporary, but fresh dipoles continually form.*

**FORMATION SEQUENCE** 1 Random electron movement makes an instantaneous uneven distribution. → 2 This dipole repels or attracts electrons in a neighbour. → 3 An opposite dipole is induced. → 4 Opposite partial charges attract.

| FACTOR                                 | WHY THE INTERACTION BECOMES STRONGER                                                    |
|----------------------------------------|-----------------------------------------------------------------------------------------|
| more electrons / larger electron cloud | the cloud is more easily distorted (more polarisable), so larger temporary dipoles form |
| greater contact between molecules      | more of the electron clouds can interact at once                                        |

**WORKED EXAMPLE | Question: Why does  $I_2$  have a higher boiling point than  $Br_2$ ?**

**COMPARE**  $I_2$  has more electrons and a larger, more polarisable electron cloud.

**ATTRACTION** Stronger instantaneous and induced dipoles form, so induced dipole-dipole attractions are stronger.

**ENERGY** More energy is required to separate  $I_2$  molecules.

**ANSWER**  $I_2$  has the higher boiling point and lower volatility.



**TEST IT 8 [8 marks]**

- Explain fully why  $I_2$  has a higher boiling point than  $Br_2$ . [4]
- State the intermolecular force present in Ar atoms and in  $CH_4$  molecules. [2]
- Explain why Kr has a higher boiling point than Ne. [2]



**DONE WHEN** You can build the full temporary dipole → induced dipole → attraction explanation.

## 9. Permanent dipole-dipole interactions



**SPEC 2.2.2(k)** Explain attractions between the permanent dipoles of polar molecules.

**DEFINITION TO LEARN** A permanent dipole-dipole interaction is the electrostatic attraction between the permanent  $\delta+$  end of one polar molecule and the permanent  $\delta-$  end of a neighbouring polar molecule.



neighbouring molecules align so  $\delta+$  is attracted to  $\delta-$

*The dotted line is an intermolecular attraction. The H-Cl lines are covalent bonds.*

| MOLECULE TYPE      | EXAMPLES                             | INTERACTIONS PRESENT                                                         |
|--------------------|--------------------------------------|------------------------------------------------------------------------------|
| non-polar molecule | $\text{Cl}_2$ ; linear $\text{CO}_2$ | induced dipole-dipole interactions                                           |
| polar molecule     | $\text{HCl}$ ; bent $\text{SO}_2$    | induced dipole-dipole interactions plus permanent dipole-dipole interactions |

**WHY  $\text{SO}_2$  IS POLAR**  $\text{SO}_2$  is bent, so its polar S-O bond dipoles do not cancel. It therefore has an overall permanent dipole.



**EXAMINER TIP** When comparing boiling points, first compare electron-cloud size. Permanent dipole-dipole attraction does not automatically outweigh a very large difference in induced-dipole strength.

**WORKED EXAMPLE | Question: Why is the boiling point of  $\text{HCl}$  higher than that of  $\text{F}_2$ , which has a similar number of electrons?**

**SIMILAR SIZE**  $\text{HCl}$  and  $\text{F}_2$  have similar numbers of electrons, so their induced interactions are a fair starting comparison.

**POLARITY**  $\text{HCl}$  is polar;  $\text{F}_2$  is non-polar.

**ATTRACTION**  $\text{HCl}$  therefore also has permanent dipole-dipole attractions.

**ANSWER** More energy is needed to separate  $\text{HCl}$  molecules, so  $\text{HCl}$  has the higher boiling point.



**TEST IT 9 [7 marks]**

- State all intermolecular interactions present in  $\text{HCl}$ ,  $\text{Cl}_2$  and  $\text{SO}_2$ . [3]
- Draw two  $\text{HCl}$  molecules aligned correctly and label the permanent dipole-dipole attraction. [3]
- Explain why  $\text{CO}_2$  has no permanent dipole-dipole attractions even though  $\text{C=O}$  bonds are polar. [1]

DRAWING SPACE FOR TEST IT 9



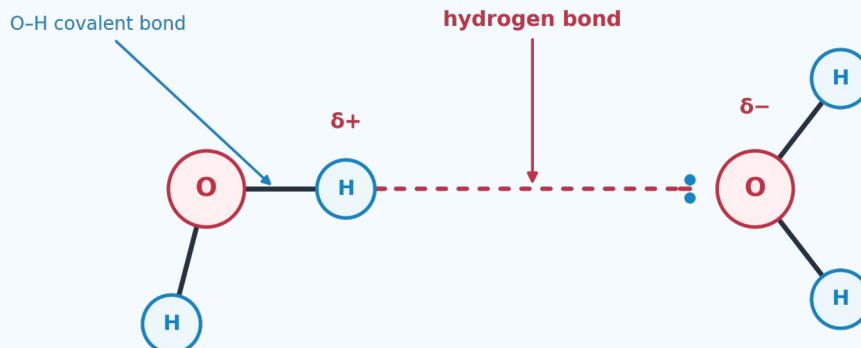
**DONE WHEN** You remember that polar molecules still have induced dipole-dipole interactions as well.

# 10. Hydrogen bonding



**SPEC 2.2.2(i)** Describe hydrogen bonding between molecules containing N, O or F and H in -NH, -OH or HF, including the role of lone pairs.

**DEFINITION TO LEARN** A hydrogen bond is an intermolecular attraction between a lone pair on an N, O or F atom in one molecule and a  $\delta^+$  H atom covalently bonded to N, O or F in another molecule.



**The lone pair, dashed hydrogen bond, H and the O-H covalent bond lie on one straight line.**

*Hydrogen bonds are dashed and intermolecular. The lone pair, dashed line,  $\delta^+$  H and the H-O covalent bond must lie on one straight line.*

| CHECK                  | QUESTION TO ASK                                                                                   |
|------------------------|---------------------------------------------------------------------------------------------------|
| hydrogen-bond donor    | Is there H directly bonded to N, O or F?                                                          |
| hydrogen-bond acceptor | Is there a lone pair on N, O or F?                                                                |
| correct diagram        | Do the lone pair, dashed line, $\delta^+$ H and its H-N/O/F covalent bond form one straight line? |

**CAN HYDROGEN-BOND BETWEEN ITS OWN MOLECULES?** Yes:  $\text{H}_2\text{O}$ ,  $\text{NH}_3$  and HF. No:  $\text{CH}_4$  and HCl. An O or N lone pair alone is not enough; a suitable N-H, O-H or H-F bond is also required.

## WORKED EXAMPLE | Question: Why is the boiling point of $\text{H}_2\text{O}$ much higher than that of $\text{H}_2\text{S}$ ?

**STRUCTURE** Water contains O-H bonds and lone pairs on O, so water molecules form hydrogen bonds.

**COMPARE**  $\text{H}_2\text{S}$  does not form hydrogen bonds between its molecules.

**ENERGY** More energy is required to overcome the stronger attractions between  $\text{H}_2\text{O}$  molecules.

**ANSWER**  $\text{H}_2\text{O}$  has the much higher boiling point and lower volatility.



### TEST IT 10 [9 marks]

- From  $\text{NH}_3$ , HF,  $\text{CH}_4$  and HCl, identify every substance that hydrogen-bonds between its own molecules. [2]
- Draw two  $\text{NH}_3$  molecules and one hydrogen bond. Show the lone pair and partial charges, and keep lone pair ... H-N collinear. [4]
- Explain why  $\text{H}_2\text{O}$  has a higher boiling point than  $\text{H}_2\text{S}$ . [3]

DRAWING SPACE FOR TEST IT 10

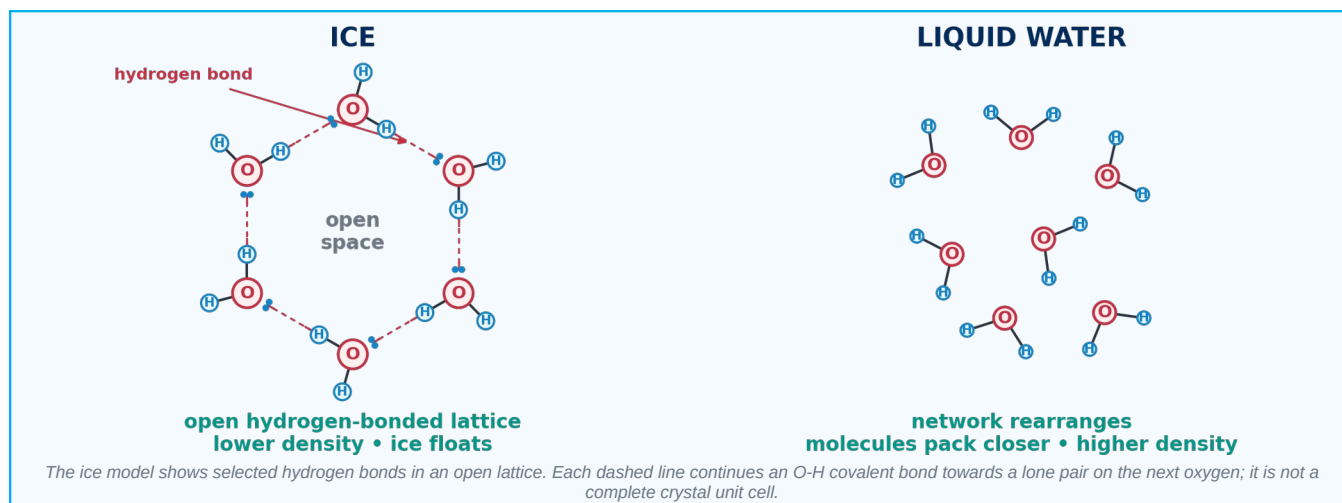


**✓ DONE WHEN** Your lone pair, dashed hydrogen bond,  $\delta^+$  H and H-N/O/F covalent bond form one straight line.

# 11. Why water behaves unusually



**SPEC 2.2.2(m)** Explain the relatively high melting and boiling points of water and why ice is less dense than liquid water.



| ANOMALY                                    | HYDROGEN-BOND EXPLANATION                                                                                                                    |
|--------------------------------------------|----------------------------------------------------------------------------------------------------------------------------------------------|
| relatively high melting and boiling points | many hydrogen bonds act between water molecules; substantial energy is needed to overcome enough of them for molecules to move or separate   |
| ice is less dense than liquid water        | hydrogen bonds hold water molecules in an ordered open lattice; on melting, part of the network collapses and molecules pack closer together |

**DENSITY LINK** For the same mass of water, the open ice lattice occupies a larger volume. Since density = mass ÷ volume, the density is lower, so ice floats.

## WORKED EXAMPLE | Question: Why does ice float on liquid water?

**STRUCTURE** Hydrogen bonds hold molecules in an open, ordered lattice in ice.

**SPACING** The molecules are, on average, farther apart than in liquid water.

**DENSITY** The same mass occupies a greater volume, so ice has a lower density.

**ANSWER** The less-dense ice floats on the denser liquid.



### TEST IT 11 [7 marks]

- (a) Explain why water has a relatively high boiling point for such a small molecule. [4]  
 (b) Explain why ice is less dense than liquid water. Include the role of hydrogen bonds and particle spacing. [3]



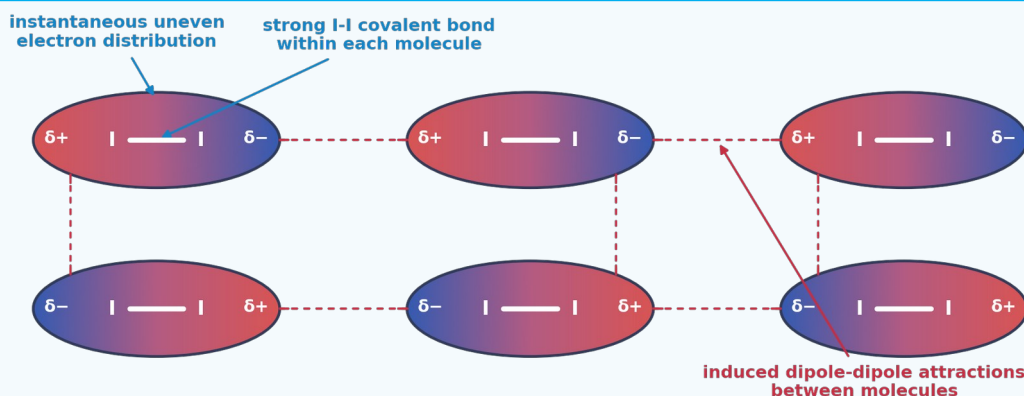
**✓ DONE WHEN** You can link hydrogen bonding to both energy changes and the open structure of ice.

## 12. Simple molecular lattices and state changes



**SPEC 2.2.2(n)-(o)** Explain simple molecular lattices and their melting, boiling and electrical properties.

**DEFINITION TO LEARN** A simple molecular lattice contains discrete covalently bonded molecules held together in the solid by intermolecular forces.



This freezes one instant of the changing electron distribution. On melting, the dashed induced dipole-dipole attractions are overcome; each I-I covalent bond remains intact.

| PROPERTY                | PARTICLES → ATTRACTION / MOBILITY → PROPERTY                                                                                               |
|-------------------------|--------------------------------------------------------------------------------------------------------------------------------------------|
| melting and boiling     | less energy is usually needed to overcome intermolecular forces than giant-lattice bonds, so melting and boiling points are relatively low |
| state change            | molecules separate or move past one another; the covalent bonds inside the molecules remain                                                |
| electrical conductivity | neutral molecules provide no mobile ions or delocalised electrons, so the pure substance does not conduct                                  |

### WORKED EXAMPLE | Question: Why does iodine melt at a much lower temperature than diamond?

**IODINE**  $I_2$  is simple molecular. Melting overcomes intermolecular forces between  $I_2$  molecules.

**DIAMOND** Diamond is a giant covalent lattice. Melting would require many strong C-C covalent bonds to be broken.

**ANSWER** Far more energy is required for diamond, so diamond has the much higher melting point.



### TEST IT 12 [8 marks]

- Explain exactly what happens to particles and attractions when iodine melts. [3]
- Explain why iodine does not conduct electricity as a solid or as a liquid. [2]
- Explain why diamond has a much higher melting point than iodine. [3]



**✓ DONE WHEN** You never claim that covalent bonds break when a simple molecular substance changes state.

## 13. Solubility, conductivity and volatility



**SPEC 2.2.2(o)** Explain the physical properties of simple molecular substances using structure and intermolecular attractions.

| PROPERTY                       | EXPLANATION ROUTE                                                                                                                                                                                                          |
|--------------------------------|----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|
| <b>solubility</b>              | a solute dissolves when new solute-solvent attractions compensate for attractions that must be disrupted; polar substances often dissolve in polar solvents, and non-polar substances often dissolve in non-polar solvents |
| <b>electrical conductivity</b> | simple molecular substances contain neutral molecules, not mobile charged particles; pure molecular solids and liquids therefore do not conduct                                                                            |
| <b>volatility</b>              | weaker intermolecular attractions require less energy to overcome, giving a lower boiling point and greater volatility                                                                                                     |

**SOLUBILITY LANGUAGE** 'Like dissolves like' is a useful first prediction, not an absolute law. State the attractions that can form with the solvent. A molecule with an -OH or -NH group can hydrogen-bond with water, but a large non-polar region may reduce solubility.

**WORKED EXAMPLE | Question: Why is iodine much more soluble in cyclohexane than in water?**

**POLARITY**  $I_2$  and cyclohexane are non-polar; water is polar and strongly hydrogen-bonded.

**NEW ATTRACTIONS**  $I_2$  can form favourable induced dipole-dipole attractions with cyclohexane.

**ANSWER** Those interactions compensate more effectively than interactions between  $I_2$  and water, so iodine is more soluble in cyclohexane.

**WORKED EXAMPLE | Question: Why does methanol mix completely with water?**

**FUNCTIONAL GROUP** Methanol contains an O-H group and oxygen lone pairs.

**ATTRACTION** Methanol and water form hydrogen bonds with one another.

**ANSWER** The strong solute-solvent attractions make methanol miscible with water.

**EXAM TRAP** A molecular substance can conduct after dissolving only if it reacts or ionises to make mobile ions. Do not say every aqueous molecular solution conducts.

**TEST IT 13 [8 marks]**

- (a) Explain why iodine is more soluble in cyclohexane than in water. [3]  
 (b) Explain why methanol is soluble in water. [2]  
 (c) Explain why pure liquid hexane does not conduct electricity. [2]  
 (d) State how volatility changes as intermolecular forces become stronger. [1]



**✓ DONE WHEN** You justify solubility, conductivity and volatility with particles and attractions rather than slogans alone.



# 15. Mixed OCR A mastery



**SPEC SYNTHESIS 2.2.2(g)-(o)** Combine shape, bond polarity, overall polarity, intermolecular forces and physical properties.

**WORKED EXAMPLE | Question:**  $\text{BF}_3$  contains polar B-F bonds. Predict its shape, bond angle and molecular polarity.

**COUNT** B has three bonding regions and no lone pairs.

**SHAPE** Three regions repel equally into a trigonal planar arrangement with  $120^\circ$  angles.

**DIPOLES** The three equal B-F dipoles are symmetrically arranged and cancel.

**ANSWER**  $\text{BF}_3$  is trigonal planar,  $120^\circ$ , and non-polar overall.

**EXAM TRAP** Do not use 'symmetrical' by itself. State which bond dipoles are equal, how they are arranged and that their vector effects cancel.

## TEST IT 15 [18 marks]

- (a)  $\text{SF}_6$  has six S-F bonds and no lone pairs on S. State its shape and smallest bond angle, then explain why the molecule is non-polar overall. [4]
- (b)  $\text{CO}_2$  and  $\text{H}_2\text{O}$  both contain polar bonds. Explain why only  $\text{H}_2\text{O}$  has an overall dipole. [4]
- (c) Explain why  $\text{I}_2$  has a higher boiling point than  $\text{Cl}_2$ . [4]
- (d) Explain why ice is less dense than liquid water. [3]
- (e) Explain why a pure simple molecular liquid does not conduct electricity. [3]



**EXAMINER TIP** For each explanation, circle the property, underline the particles, then check you have written attraction → energy or mobility → conclusion.

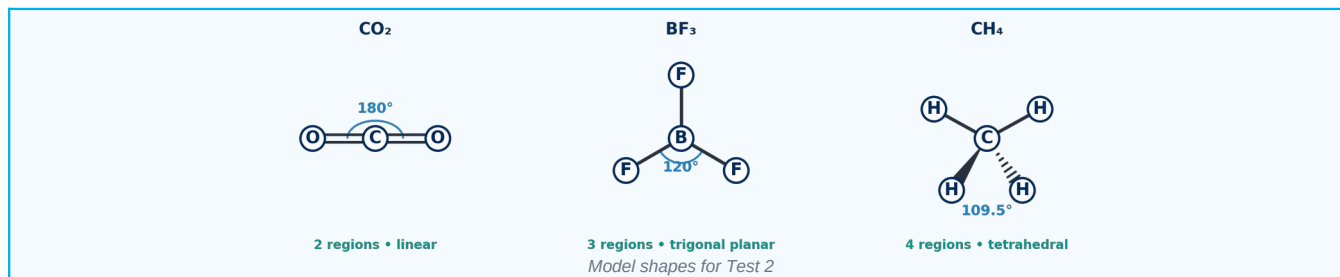
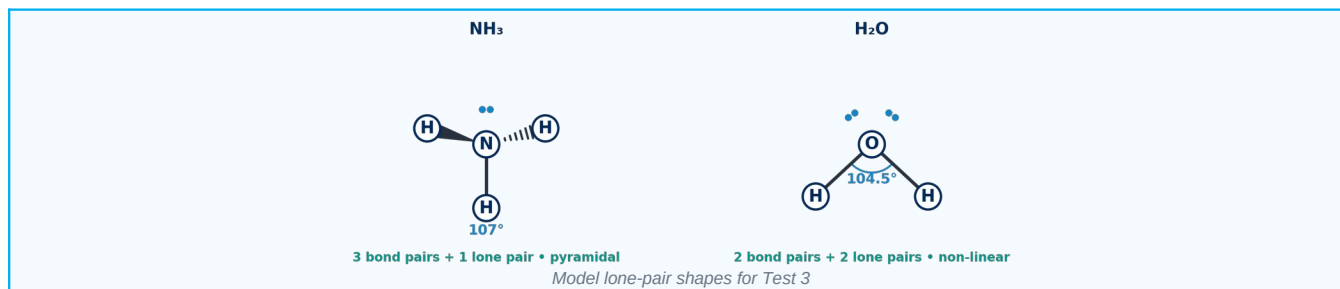
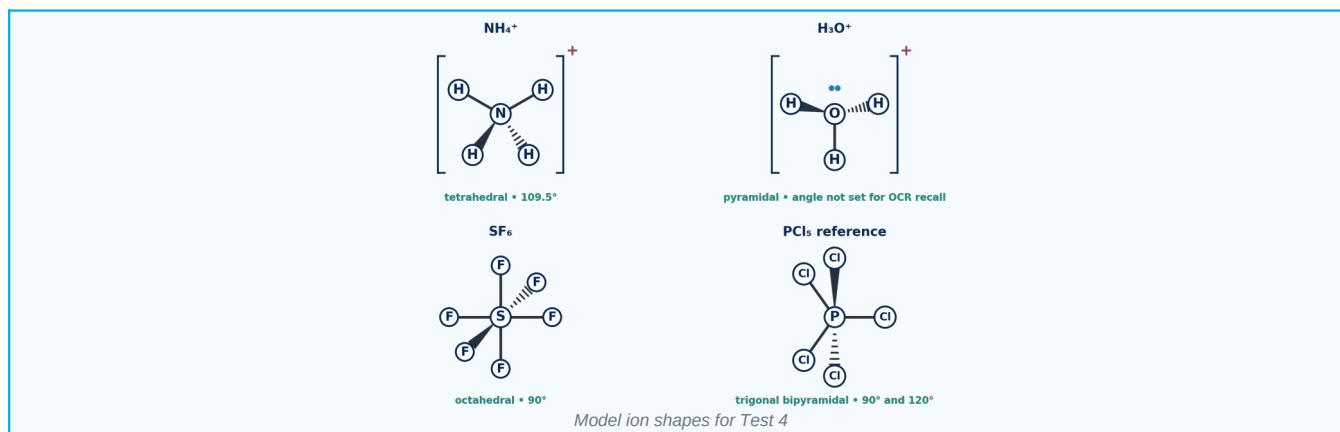


**✓ DONE WHEN** You can combine the whole Bonding II chain in an unfamiliar OCR-style context.

# TEST IT answers

**MARKING ROUTINE** Award one mark for each distinct correct point or correctly labelled feature. For drawing questions, compare shape, bond directions, lone pairs, brackets, charges and angle labels.

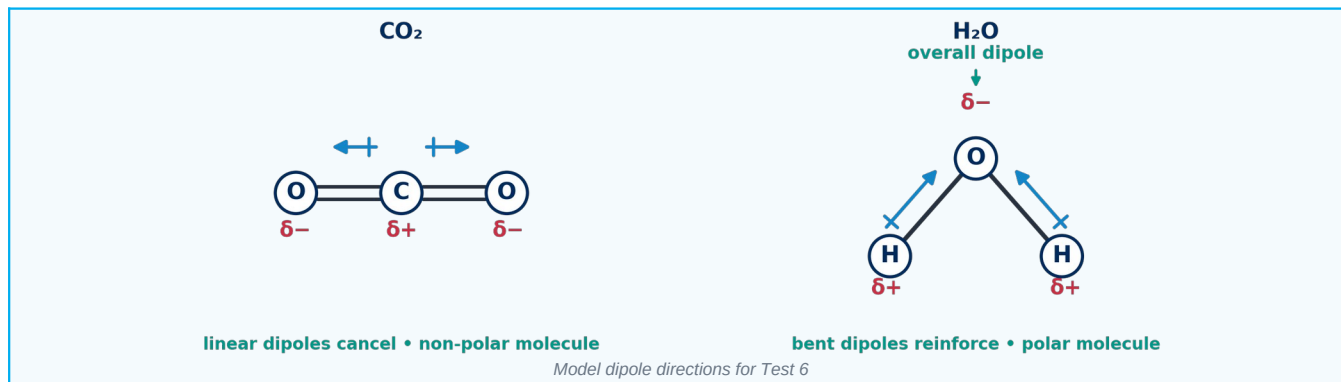
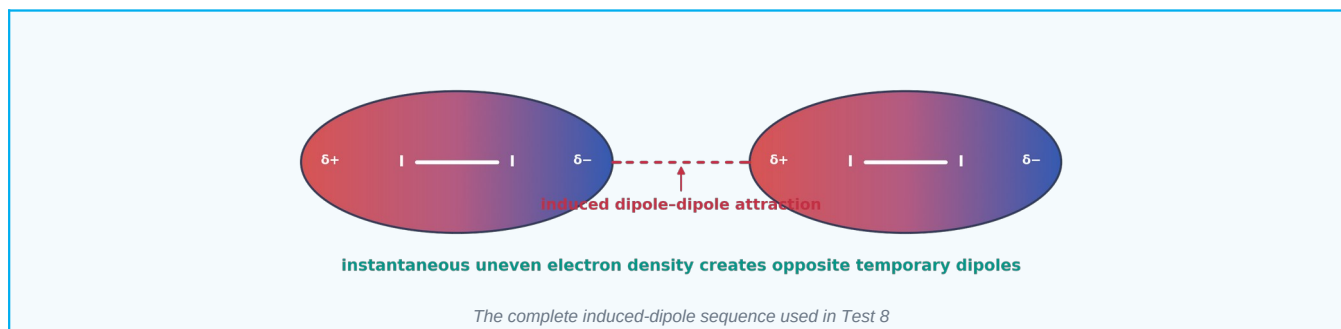
| TEST | COMPLETE ANSWER                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                 | MARKS |
|------|-------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-------|
| 1    | (a) Bonding pair: pair of electrons shared between two covalently bonded atoms. Lone pair: pair of outer-shell electrons not used in a covalent bond. (b) lone pair-lone pair > lone pair-bond pair > bond pair-bond pair. (c) Be is in Group 2, so it contributes 2 electrons; two attached Cl atoms add 2 more. Total = 4 electrons; $4 \div 2 = 2$ electron pairs around Be. (d) The two electron regions repel and reach maximum separation by pointing in opposite directions, so $\text{BeCl}_2$ is linear with a $180^\circ$ bond angle. | 8     |
| 2    | (a) $\text{BF}_3$ : trigonal planar, $120^\circ$ . (b) $\text{CH}_4$ : tetrahedral, $109.5^\circ$ , with two ordinary bonds, one solid wedge and one dashed wedge. (c) A multiple bond occupies one direction/region of electron density around the central atom.                                                                                                                                                                                                                                                                               | 7     |
| 3    | (a) $\text{NH}_3$ has three N-H bonds and one lone pair around N; pyramidal; $107^\circ$ . (b) $\text{H}_2\text{O}$ has two lone pairs. Lone pairs repel more strongly than bond pairs, so the two lone pairs compress the O-H bonds from the tetrahedral $109.5^\circ$ to $104.5^\circ$ .                                                                                                                                                                                                                                                      | 8     |
| 4    | (a) $\text{H}_3\text{O}^+$ : three bond pairs and one lone pair; pyramidal. (b) $\text{SF}_6$ : octahedral; $90^\circ$ . (c) $\text{NH}_4^+$ : bracketed tetrahedral ion, four N-H bonds, no lone pair on N, overall + charge, $109.5^\circ$ .                                                                                                                                                                                                                                                                                                  | 10    |



**SELF-CHECK** Every atom label is clear, every bond stops at its atom, and lone pairs are shown on the central atom where required.

# TEST IT answers

| TEST | COMPLETE ANSWER                                                                                                                                                                                                                                                                                                                                                                                                                                          | MARKS |
|------|----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-------|
| 5    | (a) Electronegativity is the ability of an atom to attract the bonding electrons in a covalent bond. (b) C-H: 0.3; N-H: 0.8; O-H: 1.3, so C-H < N-H < O-H in polarity. (c) H is $\delta+$ and Cl is $\delta-$ . Cl is more electronegative, so it attracts the shared bonding electrons more strongly.                                                                                                                                                   | 8     |
| 6    | (a) $\text{BF}_3$ non-polar: three equal B-F dipoles cancel in a trigonal planar shape. $\text{NH}_3$ polar: N-H dipoles do not cancel in a pyramidal shape. $\text{CH}_4$ non-polar: equal C-H dipoles cancel tetrahedrally. HCl polar: one polar bond gives an overall dipole. (b) $\text{CO}_2$ dipoles point from C towards each O and cancel; $\text{H}_2\text{O}$ dipoles point towards O and reinforce.                                           | 10    |
| 7    | (a) Intermolecular force: electrostatic attraction between separate molecules. Intramolecular bond: bond within a particle. (b) Strong I-I covalent bond within $\text{I}_2$ ; induced dipole-dipole forces between $\text{I}_2$ molecules. (c) Melting only lets intact $\text{I}_2$ molecules move past one another by overcoming intermolecular forces. (d) Higher boiling point means lower volatility; lower boiling point means higher volatility. | 8     |
| 8    | (a) $\text{I}_2$ has more electrons and a larger, more polarisable cloud than $\text{Br}_2$ ; larger temporary dipoles form; induced dipole-dipole forces are stronger; more energy is needed, so boiling point is higher. (b) Induced dipole-dipole interactions in both Ar and $\text{CH}_4$ . (c) Kr has more electrons and a more polarisable cloud than Ne, so its induced dipole-dipole interactions are stronger and its boiling point is higher. | 8     |



**SELF-CHECK** A polarity answer must mention both the polar bonds and whether their directional dipoles cancel in the molecular shape.

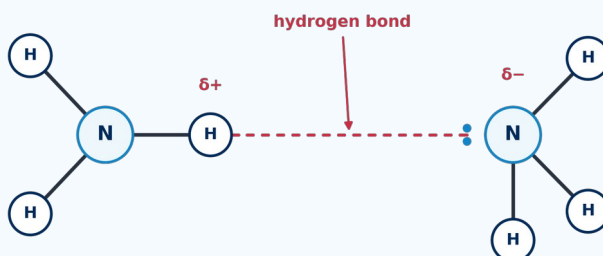
# TEST IT answers

| TEST | COMPLETE ANSWER                                                                                                                                                                                                                                                                                                                                                                                                                                        | MARKS |
|------|--------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-------|
| 9    | (a) HCl: induced plus permanent dipole-dipole. Cl <sub>2</sub> : induced only. SO <sub>2</sub> : induced plus permanent dipole-dipole because SO <sub>2</sub> is bent and polar. (b) Two HCl molecules must align Cl $\delta^-$ ... H $\delta^+$ ; the dotted line is the permanent dipole-dipole attraction. (c) CO <sub>2</sub> is linear, so its equal, opposite C=O dipoles cancel; it has no overall dipole.                                      | 7     |
| 10   | (a) NH <sub>3</sub> and HF hydrogen-bond between their own molecules; CH <sub>4</sub> and HCl do not. (b) Show the lone pair, dashed attraction, $\delta^+$ H and its N-H covalent bond on one straight line between the two NH <sub>3</sub> molecules. (c) H <sub>2</sub> O forms hydrogen bonds; H <sub>2</sub> S does not. More energy is needed to overcome attractions between water molecules, so H <sub>2</sub> O has the higher boiling point. | 9     |
| 11   | (a) Water molecules form many hydrogen bonds. Substantial energy is needed to overcome enough of these intermolecular attractions, so water has a relatively high boiling point. (b) Hydrogen bonds hold molecules in an open ordered lattice in ice. Molecules are farther apart than in liquid water, so the same mass occupies more volume and ice has lower density.                                                                               | 7     |



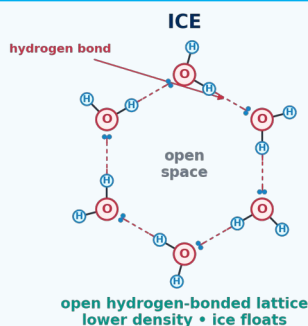
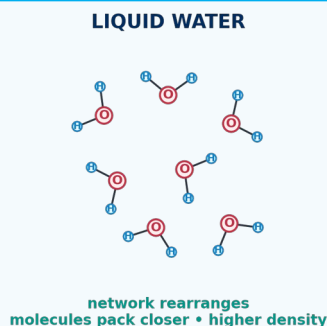
neighbouring molecules align so  $\delta^+$  is attracted to  $\delta^-$

Model alignment for Test 9



The lone pair, dashed hydrogen bond, H and the N-H covalent bond lie on one straight line.

Model NH<sub>3</sub> hydrogen-bond diagram for Test 10

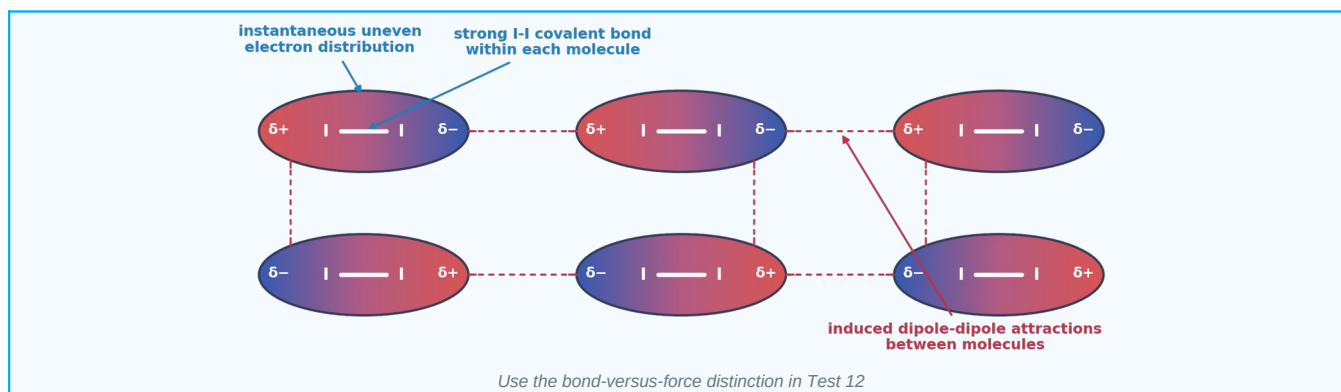


Model particle explanation for Test 11

**SELF-CHECK** Hydrogen bonding is an intermolecular attraction. The lone pair, dashed line,  $\delta^+$  H and the N-H or O-H covalent bond must be collinear.

## TEST IT answers

| TEST | COMPLETE ANSWER                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                                | MARKS |
|------|------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|-------|
| 12   | (a) $I_2$ molecules gain energy; induced dipole-dipole forces between molecules are overcome; intact molecules move past one another. (b) Solid and liquid iodine contain neutral molecules and no mobile ions or delocalised electrons. (c) Diamond is a giant covalent lattice; many strong C-C covalent bonds must be broken, requiring much more energy than overcoming intermolecular forces in iodine.                                                                                                                                                                                   | 8     |
| 13   | (a) $I_2$ and cyclohexane are non-polar and form favourable induced dipole-dipole attractions; water is polar and strongly hydrogen-bonded, so $I_2$ is more soluble in cyclohexane. (b) Methanol forms hydrogen bonds with water through its O-H group and O lone pairs. (c) Hexane contains neutral molecules and has no mobile charged particles. (d) Stronger intermolecular forces give lower volatility.                                                                                                                                                                                 | 8     |
| 14   | (a) $CH_4$ tetrahedral and non-polar; $NH_3$ pyramidal and polar. (b) $CH_4$ : induced dipole-dipole only. $NH_3$ : induced dipole-dipole, permanent dipole-dipole and hydrogen bonding. (c) $NH_3$ molecules form hydrogen bonds; $CH_4$ does not. More energy is required to overcome the stronger attractions between $NH_3$ molecules, so $NH_3$ has the higher boiling point.                                                                                                                                                                                                             | 12    |
| 15   | (a) $SF_6$ is octahedral, $90^\circ$ ; six equal S-F dipoles cancel symmetrically, so it is non-polar. (b) $CO_2$ is linear and equal C=O dipoles cancel; $H_2O$ is bent and O-H dipoles do not cancel. (c) $I_2$ has more electrons and a more polarisable cloud than $Cl_2$ , so induced dipole-dipole forces are stronger; more energy is required; boiling point is higher. (d) Ice has an open hydrogen-bonded lattice, greater molecular spacing and therefore lower density. (e) A pure simple molecular liquid contains neutral molecules and no mobile ions or delocalised electrons. | 18    |



**MARKING TIP** For comparison questions, a complete chain usually contains: named particles → named attraction → stronger or weaker → more or less energy → stated property.



**READY** Reattempt the lowest-scoring TEST IT without looking at the notes, then mark it again.

# Final mastery checklist

- ☐ I can define bonding pair, lone pair and electron-pair repulsion theory.
- ☐ I can count electron regions and treat a multiple bond as one region.
- ☐ I can draw and name linear, trigonal planar, tetrahedral, pyramidal, non-linear and octahedral shapes.
- ☐ I know  $180^\circ$ ,  $120^\circ$ ,  $109.5^\circ$ ,  $107^\circ$ ,  $104.5^\circ$  and  $90^\circ$  in the correct contexts.
- ☐ I can use lone-pair repulsion to explain reduced bond angles.
- ☐ I can define electronegativity and assign  $\delta^+$  and  $\delta^-$  correctly.
- ☐ I can decide molecular polarity by combining bond dipoles with 3-D shape.
- ☐ I can distinguish intramolecular bonds from intermolecular forces.
- ☐ I can explain induced dipole-dipole and permanent dipole-dipole interactions.
- ☐ I can draw and define hydrogen bonding with a collinear lone pair ... H-N/O/F arrangement.
- ☐ I can explain water's high melting/boiling points and why ice is less dense.
- ☐ I can explain melting, boiling, volatility, solubility and conductivity of simple molecular substances.

## Mastery map

| AREA                                     | TESTS       | MARKS      |
|------------------------------------------|-------------|------------|
| Shapes and electron-pair repulsion       | 1-4         | 33         |
| Electronegativity and molecular polarity | 5-6         | 18         |
| Intermolecular-force language and types  | 7-10        | 32         |
| Water and simple molecular properties    | 11-13       | 23         |
| Explanation and mixed mastery            | 14-15       | 30         |
| <b>TOTAL</b>                             | <b>1-15</b> | <b>136</b> |

**OCR A SCOPE CHECK** Covered in full: 2.2.2(g)-(o). All later reaction chemistry is deliberately excluded from this Bonding II pack.



**READY** You can now use shape, polarity and intermolecular forces to explain unfamiliar molecular properties in OCR A questions.