OCR A | Year 12 chemistry

Bonding

Complete notes

Clear explanations, precise definitions and worked examples for the whole topic. Use the contents or PDF bookmarks to find a section. TEST IT references point to the matching questions in the separate topic Questions & answers pack.

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Ionic bonding & formulae

1. Ionic formulae and common ions

SPEC 2.1.2(a)

CRACK IT

An ionic formula must have **no overall charge**. Choose the smallest whole-number ratio of ions that makes the positive and negative charges balance.

For example, one 2+ ion needs two 1- ions. Use the charges to find the ratio, then write the subscripts in the formula.

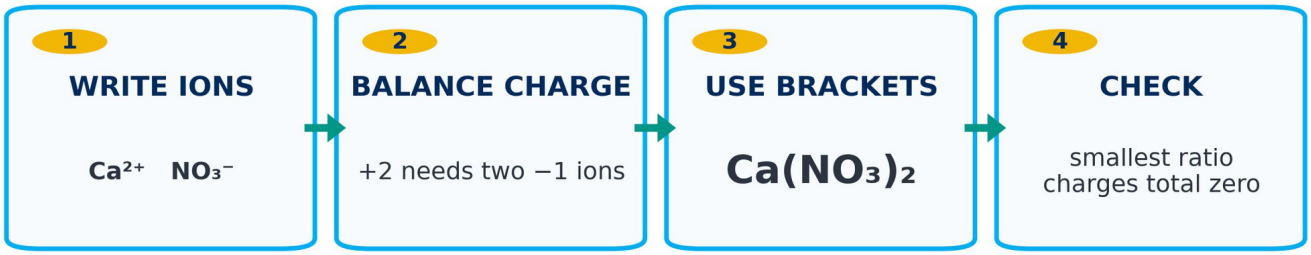
A compound ion stays together as a group. If you need more than one of that ion, put brackets around its formula before adding the subscript. The table below gives common ions to recognise.

DEFINITIONS & KEY TERMS

Ionic formula

The smallest whole-number ratio of ions giving zero overall charge.

Group / ion	Charge or formula
Group 1	1+
Group 2	2+
Group 16	2−
Group 17	1−
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Sulfate	SO_4^{2-}
Hydroxide	OH^-
Ammonium	NH_4^+
Zinc	Zn^{2+}
Silver	Ag^+



The final formula is neutral and uses the smallest whole-number ratio.

WORKED EXAMPLE

Calcium nitrate

Write the formula from Ca^{2+} and NO_3^-.

Worked answer

1. One Ca^{2+} ion has charge +2.

2. Two nitrate ions give charge −2.

3. Keep each nitrate ion intact and use brackets.

$Ca(NO_3)_2$

EXAM TRAP Use brackets only when more than one compound ion is required. Simplify Ca_2O_2 to CaO; never change the internal formula of a compound ion.

TEST IT

Try questions 1, 2 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Every formula is neutral, simplified and uses brackets correctly.

2. Ionic bonding and electron origin

SPEC 2.2.2(a)

CRACK IT

There are two ideas to keep separate. **Electron transfer makes the ions.** The attraction between the resulting positive and negative ions is the ionic bonding.

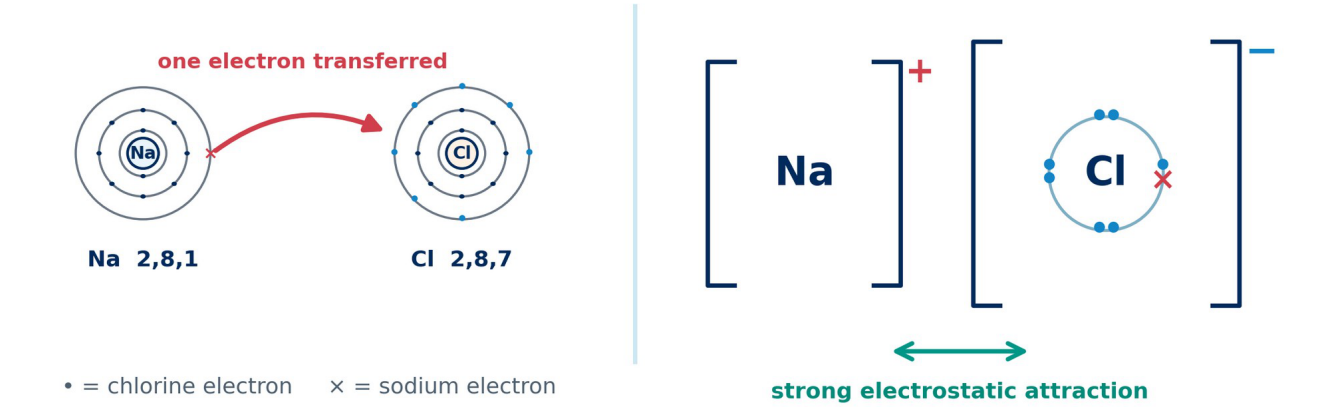
A metal atom loses electrons and becomes a positive ion. A non-metal atom gains electrons and becomes a negative ion. In a giant ionic lattice, each ion attracts surrounding ions of the opposite charge.

In a dot-and-cross diagram, use dots and crosses consistently to show where the electrons came from. Show outer-shell electrons, put each ion in square brackets and write its charge outside. Check both electron totals and the ratio of ions.

DEFINITIONS & KEY TERMS

Ionic bonding

Electrostatic attraction between positive and negative ions.



EXAM TRAP A transfer arrow represents a process, not the ionic bond itself.

TEST IT

Try questions 3, 4, 5, 6, 7 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Your ionic diagram shows electron origin, correct electron totals, brackets and charges.

3. Giant ionic lattices and their properties

SPEC 2.2.2(b-c)

CRACK IT

An ionic solid is a giant repeating arrangement of positive and negative ions. Strong electrostatic attractions act throughout the lattice, so a lot of energy is needed to separate the ions on melting.

For conductivity, ask whether charged particles can move. In the solid, the ions are held in fixed positions. When molten, the ions can move and carry charge.

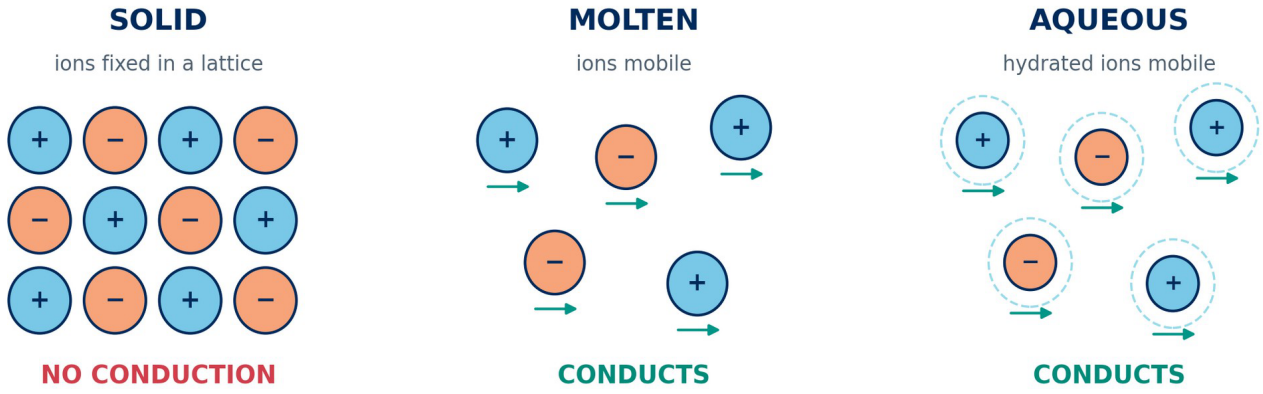
Many ionic compounds dissolve in water, although not all do. If an ionic compound forms an aqueous solution, its mobile ions can carry charge. Do not explain ionic conductivity using moving electrons.

Higher ion charges and smaller ionic radii generally strengthen the electrostatic attraction. For the same charges, smaller ions let the centres of opposite ions get closer together. Stronger attractions need more energy to overcome, so the melting point tends to be higher. Compare similar lattice arrangements and change one factor at a time.

DEFINITIONS & KEY TERMS

Giant ionic lattice

A regular repeating arrangement of oppositely charged ions strongly attracted in all directions.



Property	Explanation
High melting / boiling points	Many strong electrostatic attractions throughout the lattice require much energy to overcome.
Solid does not conduct	Ions are fixed; no mobile charge carriers.
Molten or dissolved sample conducts	Ions are mobile and carry charge.
Solubility	Often soluble in water, but not always: new solvent-ion attractions must compensate for attractions disrupted.

EXAM TRAP Molten ionic compounds conduct through mobile ions, not delocalised electrons. A charged ion does not automatically make every ionic compound soluble.

WORKED EXAMPLE

Comparing ion charges: MgO and NaF

MgO and NaF have similar ion sizes and comparable lattice arrangements. Why does MgO have the higher melting point?

Worked answer

1. Both have giant ionic lattices. MgO contains Mg^{2+} and O^{2-}; NaF contains Na^{+} and F^{-}.
2. The ion sizes are similar, so focus on charge: the ions in MgO have larger charges.
3. The electrostatic attractions between opposite ions are stronger in MgO.
4. More energy is needed to overcome these attractions on melting.

Higher charges → stronger electrostatic attraction → more energy → higher melting point.

WORKED EXAMPLE

Comparing ion sizes: NaCl and KCl

NaCl and KCl have the same ion charges and comparable lattice arrangements. Na^+ is smaller than K^+. Why does NaCl have the higher melting point?

Worked answer

1. Both contain a 1+ metal ion and the same Cl^- ion, so charge does not explain the difference.

2. The smaller Na^+ ion allows the centres of opposite ions to be closer together.

3. The electrostatic attraction between the ions is stronger in NaCl.

4. More energy is needed to overcome these attractions on melting.

Smaller ions, with charge unchanged → stronger electrostatic attraction → higher melting point.

TEST IT

Try questions 8, 9, 10, 11 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You explain every property using the correct particles, attraction and mobility or energy.

Covalent, dative & bond strength

4. Covalent bonds and dot-and-cross diagrams

SPEC 2.2.2(d-e)

CRACK IT

In a covalent bond, two atoms share a pair of electrons. The shared electrons are attracted to both positive nuclei, holding the atoms together.

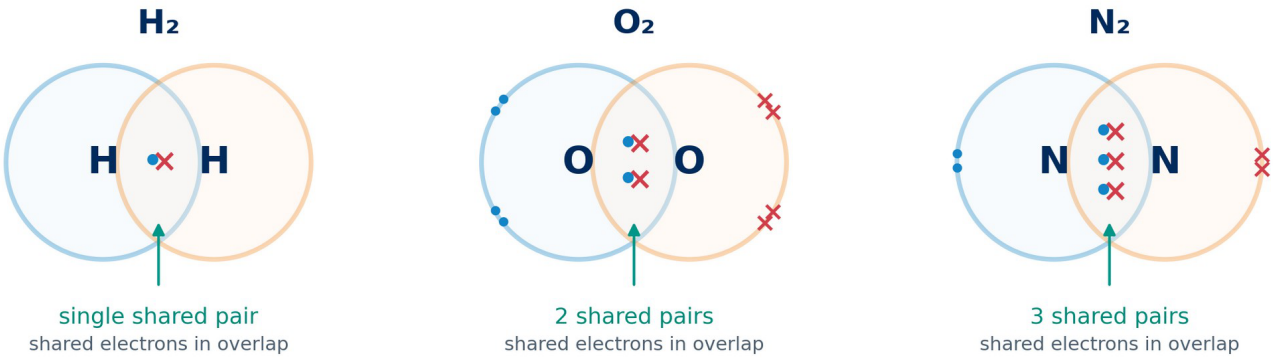
A single bond contains one shared pair, a double bond two and a triple bond three. In a dot-and-cross diagram, put every shared pair in the overlap between the atoms.

A lone pair is not shared in a bond. Show these electrons outside the overlap, and use dots and crosses consistently to show their atom of origin. Count all the outer-shell electrons before finishing.

DEFINITIONS & KEY TERMS

Covalent bond

The strong electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms.



• = electron from left atom × = electron from right atom

Bond	Shared pairs	Line model
Single	1	H–H
Double	2	O=O
Triple	3	N≡N

EXAM TRAP “A shared pair of electrons” alone is incomplete as the full definition: include strong electrostatic attraction to both nuclei.

TEST IT

Try questions 12, 13, 14 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can define a covalent bond precisely and account for every outer-shell electron.

5. Dative bonds: donor, acceptor and arrow

SPEC 2.2.2(e)(iii)

CRACK IT

A covalent bond contains a shared pair of electrons. Usually, each atom supplies one electron. In a **dative covalent bond**, one atom supplies both.

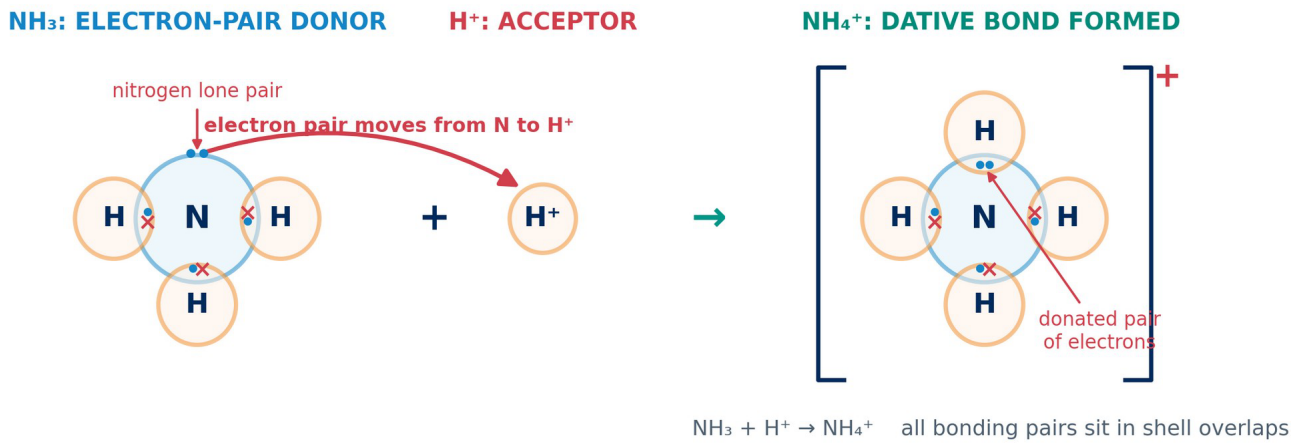
The atom supplying the pair is the donor. The species receiving it is the acceptor. Draw the arrow **from the donor's lone pair to the acceptor**.

For NH_3 reacting with H^+, nitrogen supplies its lone pair to the hydrogen ion. Once NH_4^+ has formed, all four N-H bonds are equivalent. The origin of the electrons does not make one of those bonds different afterwards.

DEFINITIONS & KEY TERMS

Dative covalent bond

A shared pair of electrons in which both electrons were supplied by the same atom.



Role	$NH_3 + H^+ \rightarrow NH_4^+$
Donor	Nitrogen in NH_3 supplies its lone pair.
Acceptor	H^+ accepts the pair.
After formation	All four N–H bonds are equivalent.

EXAM TRAP Start the arrow at the donated electron pair and point towards the acceptor. The charge and brackets belong around the whole product ion.

TEST IT

Try questions 15, 16, 17 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You identify donor and acceptor before drawing the correct arrow and complete product ion.

6. Average bond enthalpy and bond strength

SPEC 2.2.2(f)

CRACK IT

Bond strength tells you how hard it is to separate the bonded atoms. A larger average bond enthalpy means more energy is needed to break the bond, so the bond is stronger.

Use the values supplied to make a comparison. Do not confuse the strength of a covalent bond inside a molecule with the attractions between separate molecules.

At this stage, use the comparison. Formal definitions and bond-enthalpy calculations are developed in Energetics.

Larger average bond enthalpy → stronger covalent bond → more energy needed to break it

WORKED EXAMPLE

Compare two bonds

Bond A has average bond enthalpy 347 kJ mol^{-1}; bond B has 612 kJ mol^{-1}.

State which bond is stronger and explain your choice using the bond enthalpies.

Worked answer

612 is the larger value, so more energy is required to break a mole of B bonds.

Bond B is stronger.

EXAM TRAP Do not reverse the relationship: a large positive bond-breaking energy means a strong bond.

TEST IT

Try questions 18, 19 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You identify the stronger bond and justify it using the energy needed to break it.

Metallic bonding

7. Metallic bonding and mobile electrons

SPEC 3.1.1(d-f) foundation

CRACK IT

A metal contains a giant lattice of positive ions surrounded by delocalised electrons. 'Delocalised' means the electrons are not restricted to one atom or one bond.

The attraction between the positive ions and these electrons holds the metal together. Strong attractions often give metals high melting and boiling points.

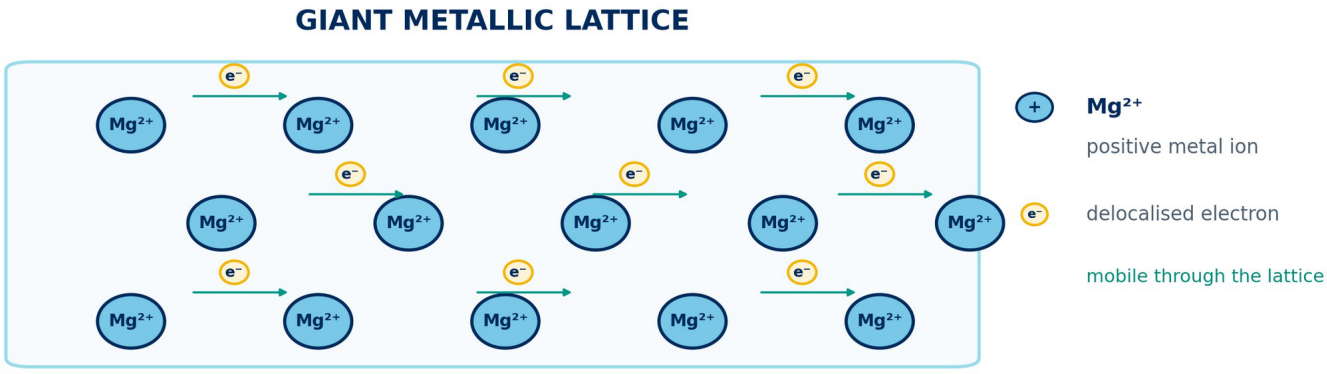
The delocalised electrons can move and carry charge in both solid and liquid metals. The positive ions do not travel through the solid to conduct electricity.

Metals do not normally dissolve as intact metallic lattices in molecular solvents. A metal reacting with a substance is a different process.

DEFINITIONS & KEY TERMS

Metallic bonding

Strong electrostatic attraction between positive metal ions and delocalised electrons.



Strong electrostatic attraction acts between the cations and the delocalised electrons.

EXAM TRAP Name the mobile charge carriers as delocalised electrons. The positive ions are fixed in a solid metal.

TEST IT

Try question 20 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You explain metallic conductivity in both states and link melting to electrostatic attraction.

Structures & properties

8. Diamond and graphite

SPEC 3.1.1(e-f) foundation

CRACK IT

Diamond and graphite both contain giant covalent networks, but each carbon atom bonds differently. That difference explains their different properties.

In diamond, each carbon forms four bonds in a rigid three-dimensional network. There are no mobile electrons, so it does not conduct.

In graphite, each carbon forms three bonds within a layer. One electron per carbon is delocalised and can carry charge. Weak van der Waals attractions between layers allow the layers to slide.

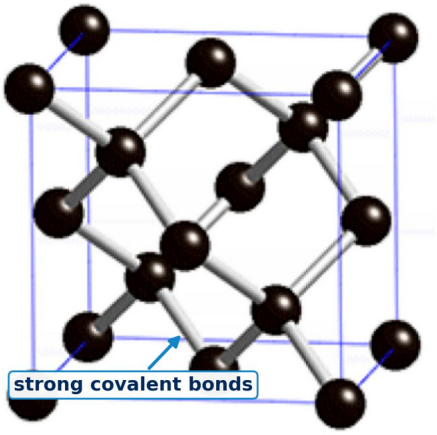
Both structures have high melting points because many strong covalent bonds must be broken. Graphite is not made of separate molecules, even if an older diagram calls the between-layer attractions 'intermolecular'. Silicon and graphene are also giant covalent networks; their fuller comparisons belong to Periodicity.

DEFINITIONS & KEY TERMS

Giant covalent structure

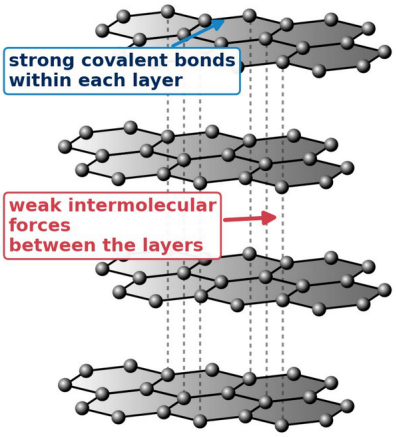
A continuous network of atoms joined by many strong covalent bonds.

DIAMOND



3D giant covalent network • four bonds per carbon

GRAPHITE



layered network • three bonds per carbon • weak intermolecular forces between layers

Feature	Diamond	Graphite
Bonding	Four covalent bonds per carbon	Three covalent bonds per carbon
Structure	Rigid three-dimensional network	Flat layers with weak attractions between them
Melting	Very high: many strong bonds must be broken	Very high: many strong bonds within layers must be broken
Conductivity	No mobile charge carriers	Delocalised electrons move along layers
Mechanical property	Hard, rigid network	Layers can slide; relatively soft

EXAM TRAP Graphite does not conduct because it has mobile ions. Sliding layers do not break the strong covalent bonds within each layer.

TEST IT

Try questions 21, 22 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You explain the similarities and differences using structure, bonds and mobile electrons.

9. Simple molecular lattices and state changes

SPEC 2.2.2(n-o)

CRACK IT

A simple molecular substance contains separate molecules. Each molecule has strong covalent bonds inside it, while neighbouring molecules attract through intermolecular forces.

Melting and boiling overcome attractions between molecules; the molecules usually stay intact. This is why a molecular substance can have a low boiling point despite having strong covalent bonds.

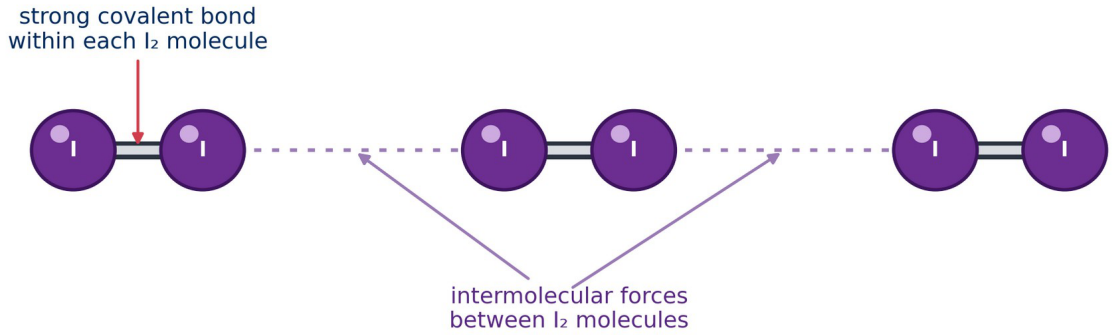
Pure molecular substances normally have no mobile ions or delocalised electrons, so they do not conduct. To explain solubility, compare the attractions disrupted with the new attractions formed with the solvent.

DEFINITIONS & KEY TERMS

Simple molecular solid

Discrete covalently bonded molecules held together by intermolecular attractions.

SOLID IODINE: DISCRETE I_2 MOLECULES



On melting, intermolecular forces are overcome; the I-I covalent bonds remain intact.

EXAM TRAP Do not say iodine melts because its I–I covalent bonds break. Detailed intermolecular force types follow in Bonding II.

TEST IT

Try question 23 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You name which attractions change during melting and explain the lack of charge carriers.

10. Identify structures from physical properties

SPEC 2.2.2(b-c,n-o); 3.1.1(d-f)

CRACK IT

Start a structure question by identifying the particles: ions, atoms or molecules. Then name the attraction between them. This stops you using the right explanation for the wrong structure.

1. Choose the structure type from the evidence: giant ionic, metallic, giant covalent or simple molecular.

2. For melting or boiling, name the attractions that must be overcome and connect their strength to the energy needed.

3. For conductivity, identify the charged particles and say whether they can move.

4. Use all the evidence together. A solid conductor could be a metal or graphite, so one property alone may not identify it.

Structure	Particles	Main attraction	Conductivity	Melting / boiling
Giant ionic	Positive and negative ions	Electrostatic attraction between ions	Molten / aqueous only	High
Metallic	Cations and delocalised electrons	Ions attracted to delocalised electrons	Solid and liquid	Often high
Giant covalent	Atoms	Strong covalent bonds throughout a network	Usually none; graphite is an exception	Very high
Simple molecular	Molecules	Intermolecular forces	None in pure substance	Usually lower

WORKED EXAMPLE

An unknown solid

High melting point; does not conduct as a solid; conducts when molten.

Deduce the likely structure type of the solid and explain how its properties support your answer.

Worked answer

Charged particles exist, but they become mobile only after the solid lattice breaks down.

The evidence supports a giant ionic structure.

EXAM TRAP Use all the data. No conductivity alone cannot distinguish a molecular solid from diamond.

TEST IT

Try questions 24, 25 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You justify an identification using more than just a structure label.

11. Build complete property explanations

SPEC 2.2.2(a-f,n-o); 3.1.1(d-f)

CRACK IT

Build an explanation as a linked chain. Each sentence should explain why the next part follows.

1. Name the structure and the particles present.

2. Identify the attraction between those particles.

3. For a melting-point question, link attraction strength to the energy needed to separate the particles.

4. For a conductivity question, name the charge carrier and say whether it can move.

Use specific terms such as ions, delocalised electrons or intermolecular forces. A statement that a substance 'has strong bonds' does not identify the attraction.

Structure → particles → attraction → energy or mobility → property

WORKED EXAMPLE

Sodium chloride versus iodine

Explain why sodium chloride has a much higher melting point than iodine. Refer to their structures and the attractions overcome on melting.

Worked answer

1. NaCl is giant ionic; iodine is simple molecular.

2. Strong electrostatic attractions act between ions throughout the NaCl lattice.

3. I_2 molecules are attracted by weaker intermolecular forces.

4. Much more energy is needed to overcome the ionic attractions.

NaCl has the higher melting point.

EXAM TRAP Avoid “ionic molecules”, “covalent bonds melt” and “all ionic compounds dissolve”. Use the named particles and their attraction.

TEST IT

Try question 26 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can compare structures without skipping a link in the explanation.

Shapes & bond angles

12. Predict shape by counting electron regions

SPEC 2.2.2(g)

CRACK IT

The shape comes from electron regions repelling one another around the central atom. They arrange as far apart as possible. Count **bonding regions and lone pairs** before naming the shape.

1. Find the central atom's outer-shell electron number: for example, 5 for N or 6 for O.

2. For a suitable species with single bonds to H or halogens, add one electron for each attached atom.

3. Add electrons for a negative charge and subtract them for a positive charge.

4. Divide by two to get the total pairs. Subtract the bonding pairs to find the lone pairs.

5. Choose the parent electron arrangement. Then name the molecular shape using the positions of the atoms.

The shortcut has limits. With multiple bonds or a supplied electron diagram, count regions directly: a single, double or triple bond is one bonding region. Lone pairs count as regions too.

DEFINITIONS & KEY TERMS

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: a single, double or triple bond counts as one region.

Molecular shape

The arrangement of atoms around the central atom.

Bond angle

The angle between two bonds meeting at the central atom.

Total pairs = (central valence electrons + attached H or halogens + negative charge − positive charge) ÷ 2

Total regions	Parent arrangement	Ideal angles
2	Linear	180°
3	Trigonal planar	120°
4	Tetrahedral	109.5°
5	Trigonal bipyramidal	90°, 120° and 180°
6	Octahedral	90° and 180°

WORKED EXAMPLE

Count the pairs in ammonia

Predict the shape of NH_3.

State its bond angle and explain your prediction in terms of bonding pairs and lone pairs.

Worked answer

1. Nitrogen contributes five outer electrons: (5 + 3)/2 = 4 pairs.

2. Three pairs form N–H bonds, leaving one lone pair.

3. Four regions have a tetrahedral parent arrangement; the lone pair compresses the bonds.

Pyramidal, with a 107° H–N–H angle.

EXAM TRAP The shortcut uses the number of outer-shell electrons, not modern group number 15 or 16. For multiple bonds, inspect the structure and count each multiple bond as one region.

TEST IT

Try question 27 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You distinguish total electron regions, bonding regions and lone pairs and minimise their repulsion.

13. Linear, trigonal planar and tetrahedral shapes

SPEC 2.2.2(g-h)

CRACK IT

When every electron region is a bond, the arrangement of the regions gives the molecular shape directly: two gives linear, three gives trigonal planar and four gives tetrahedral.

The bond angles are 180°, 120° and 109.5° respectively. Count a multiple bond as one region when deciding the shape.

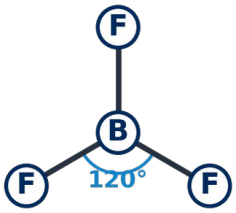
A tetrahedral diagram needs depth. Use ordinary lines for bonds in the paper, a solid wedge towards you and a hashed wedge away. Keep equivalent bond lengths consistent and finish each bond at its atom label.

CO_2



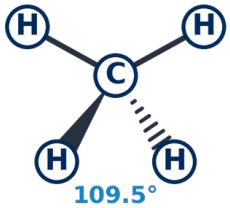
2 regions • linear

BF_3



3 regions • trigonal planar

CH_4



4 regions • tetrahedral

Regions	Lone pairs	Shape	Angle	Examples
2	0	Linear	180°	CO_2, $BeCl_2$
3	0	Trigonal planar	120°	BF_3
4	0	Tetrahedral	109.5°	CH_4, NH_4^+

WORKED EXAMPLE

Why is CO2 linear?

Each C=O bond in CO_2 is a double bond. Predict the molecular shape and O–C–O bond angle, and explain them using electron-pair repulsion.

Worked answer

Each double bond occupies one direction of electron density. The two regions reach maximum separation on opposite sides of carbon.

Linear, 180°.

EXAM TRAP A tetrahedral molecule is three-dimensional, not four flat bonds at right angles.

TEST IT

Try question 28 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can name, draw and give the angles for the three bonding-only arrangements.

14. Lone pairs compress bond angles

SPEC 2.2.2(g-h)

CRACK IT

A lone pair takes up more space near the central atom than a bonding pair. It is attracted to one nucleus rather than shared between two nuclei. It repels nearby bonding regions more strongly and pushes them closer together.

CH_4, NH_3 and H_2O each have four electron regions around the central atom. Replacing bonds with lone pairs changes the positions of the atoms and therefore the molecular shape.

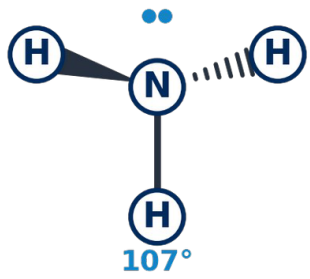
- CH_4: four bonds, no lone pairs; tetrahedral, 109.5°.
- NH_3: three bonds, one lone pair; pyramidal, about 107°.
- H_2O: two bonds, two lone pairs; bent/non-linear, about 104.5°.

Subtracting about 2.5° for each lone pair is a memory aid for this familiar comparison. It is not a universal formula for every molecule or parent arrangement.

Terminology: pyramidal and **trigonal pyramidal** are two names for this required shape. NH_3 is the common example, with a bond angle of about 107°.

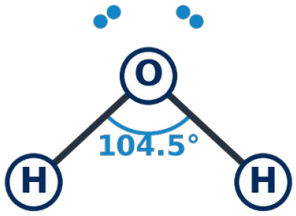
Lone pair–lone pair > lone pair–bond pair > bond pair–bond pair

NH_3



3 bond pairs + 1 lone pair • pyramidal

H_2O



2 bond pairs + 2 lone pairs • non-linear

Molecule	Regions	Shape	Angle
NH_3	3 bonding + 1 lone	Pyramidal (trigonal pyramidal)	107°
H_2O	2 bonding + 2 lone	Non-linear (bent)	104.5°

WORKED EXAMPLE

Why is water's angle smaller?

State the H–O–H bond angle in water and the H–N–H bond angle in ammonia. Explain why the angle in water is smaller.

Worked answer

1. Water has two lone pairs; ammonia has one.
2. Lone pairs repel more strongly than bonding pairs.
3. The O–H bonds are pushed closer together.

104.5° in H_2O compared with 107° in NH_3.

EXAM TRAP The familiar 109.5° → 107° → 104.5° sequence is not a universal “subtract 2.5° per lone pair” rule. Use known examples or supplied data for other species.

TEST IT

Try question 29 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You name the molecular shape from atom positions and explain the angle using lone-pair repulsion.

15. Molecular ions and up to six regions

SPEC 2.2.2(g-h)

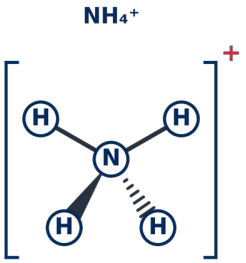
CRACK IT

Treat a molecular ion with the same method as a neutral molecule. Include the charge when counting electrons, then separate the bonding regions from the lone pairs.

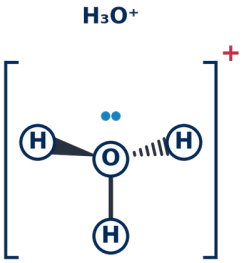
Four regions give NH_4^+ a tetrahedral shape because all four are bonds. In H_3O^+, one region is a lone pair, so the atom arrangement is pyramidal.

The named shapes include linear, non-linear, trigonal planar, **pyramidal (also called trigonal pyramidal)**, tetrahedral and octahedral.

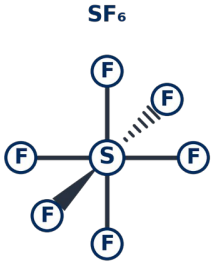
Trigonal bipyramidal, as in PCl_5, is a different shape and is not in that named list. You must still apply electron-pair repulsion to molecules and ions with up to six electron pairs around the central atom.



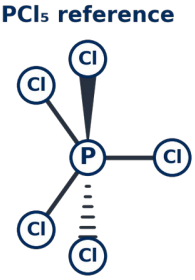
tetrahedral • 109.5°



pyramidal • angle not set for OCR recall



octahedral • 90°



trigonal bipyramidal • 90° and 120°

Species	Bonding + lone regions	Shape	Angles
NH_4^+	4 + 0	Tetrahedral	109.5°
H_3O^+	3 + 1	Pyramidal	No fixed recall angle stated here
SF_6	6 + 0	Octahedral	90° adjacent, 180° opposite
PCl_5	5 + 0	Trigonal bipyramidal	120° equatorial; 90° axial-equatorial; 180° axial-axial

WORKED EXAMPLE

Account for the charge in NH4+

Predict the shape of the ammonium ion.

State its bond angle and explain your prediction in terms of electron pairs.

Worked answer

1. Electron total = 5 + 4 − 1 = 8; divide by two to get four pairs.

2. All four pairs are bonding pairs, so they repel equally.

3. Draw the ion in square brackets with an overall + charge.

Tetrahedral, 109.5°.

EXAM TRAP Do not give H_3O^+ an angle merely by copying NH_3. An ion's charge belongs around the whole species.

TEST IT

Try question 30 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You handle charge, brackets and up to six electron regions in a three-dimensional drawing.

Electronegativity & polarity

16. Electronegativity and polar bonds

SPEC 2.2.2(i-j)

CRACK IT

Electronegativity tells you how strongly an atom attracts the shared pair in a covalent bond. Unequal attraction makes the bond polar: the more electronegative atom is δ− and its partner is δ+.

Electronegativity generally increases across a period and up a group. Fluorine is the most electronegative element.

Calculate a difference by subtracting the smaller value from the larger one. Always use the numbers supplied in that question: some practice questions use H = 2.1, so they need not give the same numerical difference as a different notes example.

A partial charge is not an ionic charge. The atoms are still sharing the bonding pair.

DEFINITIONS & KEY TERMS

Electronegativity

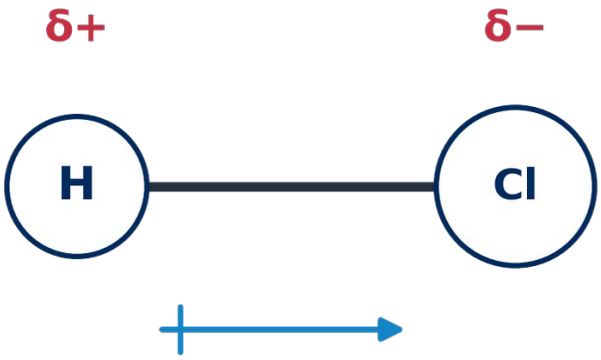
The ability of an atom to attract the bonding electrons in a covalent bond.

Polar bond

A covalent bond with an unequal distribution of bonding electron density because the atoms have different electronegativities.

Permanent dipole

A permanent separation of partial charge, represented by δ+ and δ−.



bonding electrons pulled toward the more electronegative Cl atom

Atom	H	C	N	O	F	Cl
Approximate Pauling value for these notes	2.2	2.5	3.0	3.5	4.0	3.0

WORKED EXAMPLE

Compare C–H and O–H

The Pauling electronegativities are C = 2.5, H = 2.2 and O = 3.5. Calculate the electronegativity difference for C–H and O–H. State which bond is more polar and identify the δ+ and δ− atoms in that bond.

Worked answer

1. ΔEN(C–H) = 2.5 − 2.2 = 0.3.

2. ΔEN(O–H) = 3.5 − 2.2 = 1.3.

3. The larger difference gives more unequal sharing.

O–H is more polar; O is δ− and H is δ+.

EXAM TRAP δ is a partial charge, not a full ionic charge. Bond character changes continuously; no single numerical electronegativity threshold is an absolute law.

TEST IT

Try questions 31, 32, 33 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You define electronegativity precisely, calculate differences and orient partial charges and dipole arrows correctly.

17. Bond polarity and molecular polarity

SPEC 2.2.2(j)

CRACK IT

A molecule can contain polar bonds and still be non-polar overall. The answer depends on whether the bond dipoles cancel in its three-dimensional shape.

1. Identify which bonds are polar and mark their dipole directions.

2. Work out the molecular shape, including the effect of lone pairs.

3. Check whether the bond dipoles balance in all directions.

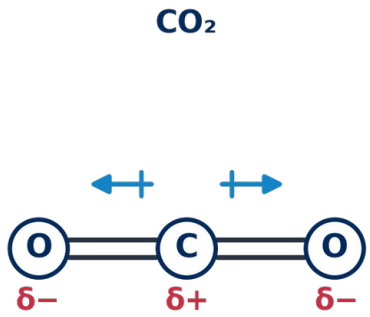
4. If they cancel completely, the molecule has no permanent dipole. If they do not, the molecule is polar.

CO_2 is linear, so its equal C=O dipoles cancel. H_2O is bent, so its O-H dipoles do not cancel. Draw the shape before deciding.

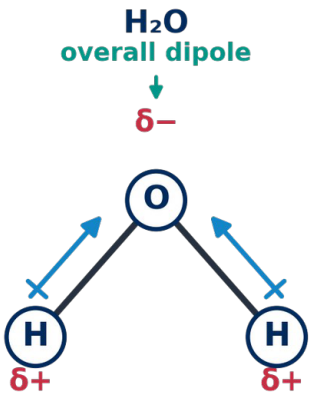
DEFINITIONS & KEY TERMS

Polar molecule

A molecule with an overall dipole because its bond dipoles do not cancel in its molecular shape.



linear dipoles cancel • non-polar molecule



bent dipoles reinforce • polar molecule

Molecule	Shape / dipole arrangement	Overall polarity
HCl	One bond dipole	Polar
CO_2	Linear, equal opposite dipoles	Non-polar
BF_3	Symmetrical trigonal planar	Non-polar
CH_4	Symmetrical tetrahedral	Non-polar
NH_3	Pyramidal; dipoles do not cancel	Polar
H_2O	Bent; dipoles do not cancel	Polar
SF_6	Symmetrical octahedral	Non-polar

WORKED EXAMPLE

H2O versus CO2

H_2O and CO_2 both contain polar covalent bonds. State which molecule is polar overall and explain the difference using molecular shape and bond dipoles.

Worked answer

CO_2 is linear, so the equal C=O dipoles point in opposite directions and cancel. H_2O is bent, so its O–H dipoles do not cancel.

CO_2 has no overall dipole; H_2O is polar.

EXAM TRAP “Symmetrical” alone is not a complete explanation. State which equal dipoles cancel and how they are arranged.

TEST IT

Try questions 34, 35 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You decide overall polarity from bond dipoles and geometry rather than the presence of one polar bond.

Intermolecular forces

18. Bonds within molecules and forces between them

SPEC 2.2.2(k-o)

CRACK IT

Keep the location of an attraction clear. Covalent bonds hold atoms together **within** a molecule. Intermolecular forces attract one molecule to another.

Start with induced dipole-dipole attractions, which act between all atoms and molecules. Polar molecules also have permanent dipole-dipole attractions. Suitable molecules can form hydrogen bonds as well.

List all the forces present before comparing substances. Naming hydrogen bonding does not mean the other applicable forces disappear.

London dispersion is another name for induced dipole-dipole interactions. Van der Waals is also used as a broader term for induced and permanent dipole-dipole interactions; the specific name makes your meaning clearer.

DEFINITIONS & KEY TERMS

Intramolecular bond

A bond within a particle, such as an O–H covalent bond within water.

Intermolecular force

An electrostatic attraction between separate molecules.

Volatility

The tendency to vaporise; higher volatility usually accompanies a lower boiling point.

Interaction	Where it acts	Cause
Induced dipole-dipole (London dispersion)	All atoms and molecules	Instantaneous dipoles induce neighbouring dipoles.
Permanent dipole-dipole	In addition, between polar molecules	Opposite permanent partial charges attract.
Hydrogen bonding	H directly bonded to N/O/F and a neighbouring N/O/F lone pair	A lone pair attracts a strongly δ+ hydrogen on another molecule.

WORKED EXAMPLE

What changes when iodine boils?

When iodine boils, identify the particles that separate and the attractions overcome. State whether the I–I covalent bonds break.

Worked answer

1. Iodine contains discrete I_2 molecules.
2. Induced dipole-dipole attractions act between the molecules.
3. Boiling overcomes those intermolecular attractions.

The molecules separate; I–I covalent bonds remain intact.

EXAM TRAP Do not replace a named intermolecular force with “strong bonds”, which could be mistaken for covalent bonds.

TEST IT

Try questions 36, 37 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You name the interaction correctly and keep it separate from bonds inside each molecule.

19. Induced dipole-dipole interactions

SPEC 2.2.2(k)

CRACK IT

Electrons are always moving. For an instant, they can be distributed unevenly, giving one side of a particle a small negative charge and the other a small positive charge.

1. This instantaneous dipole distorts the electron cloud of a neighbouring particle.

2. That distortion creates an induced dipole in the neighbour.

3. Opposite partial charges attract. New dipoles continually form as the electrons move.

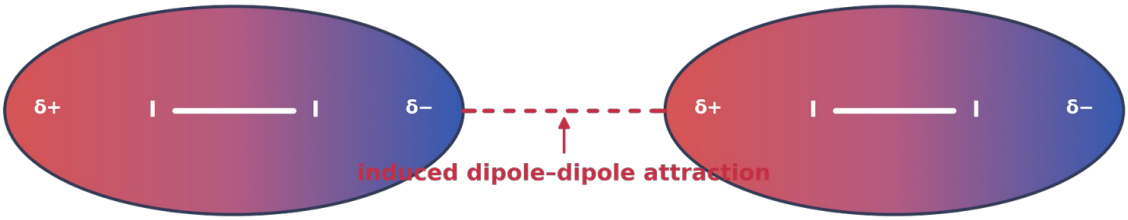
A larger electron cloud is usually more easily distorted, or **more polarisable**. This tends to give stronger van der Waals attractions and a higher boiling point.

For isomers with the same electron count, more branching usually reduces surface contact between molecules. The overall attractions are then weaker, so less energy is needed to separate them.

DEFINITIONS & KEY TERMS

Induced dipole-dipole interaction

Attraction between an instantaneous dipole in one particle and the dipole it induces in a neighbour.



instantaneous uneven electron density creates opposite temporary dipoles

Factor	Effect
More electrons / larger cloud	Usually more polarisable, so larger temporary dipoles and stronger attractions.
Greater molecular contact	More electron-cloud contact can strengthen the overall attraction.

WORKED EXAMPLE

I2 compared with Br2

Why does I_2 have the higher boiling point?

Explain your answer in terms of intermolecular attractions and state which substance is more volatile.

Worked answer

I_2 has more electrons and a larger, more polarisable cloud. Stronger induced dipole-dipole attractions require more energy to overcome.

I_2 has the higher boiling point and lower volatility.

EXAM TRAP All molecules have these interactions, including polar molecules. Explain the electron cloud rather than saying only “the molecule is heavier”.

TEST IT

Try question 38 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You connect electron movement, polarisability, attraction and energy to the boiling-point trend.

20. Permanent dipole-dipole interactions

SPEC 2.2.2(k)

CRACK IT

A polar molecule has a permanent uneven distribution of charge. The δ+ end of one molecule attracts the δ− end of a neighbouring molecule.

These permanent dipole-dipole attractions act **as well as** induced dipole-dipole attractions. For otherwise comparable molecules, the extra attractions can increase the energy needed to separate them.

HCl has an overall dipole because it has one polar bond. SO_2 is bent, so its bond dipoles do not cancel. Use the shape as well as bond polarity when deciding whether this force is present.

Do not rank boiling points from polarity alone. A large non-polar molecule can have stronger overall attractions than a small polar molecule.

DEFINITIONS & KEY TERMS

Permanent dipole-dipole interaction

Electrostatic attraction between the permanent δ+ end of one polar molecule and the permanent δ− end of a neighbour.



neighbouring molecules align so δ+ is attracted to δ−

Molecule	Interactions
Cl_2 or linear CO_2	Induced dipole-dipole
HCl or bent SO_2	Induced plus permanent dipole-dipole

WORKED EXAMPLE

HCl compared with F2

HCl and F_2 have similar numbers of electrons. Explain why HCl has the higher boiling point.

Worked answer

1. Their electron counts provide a useful starting comparison for induced interactions.
2. HCl is polar; F_2 is non-polar.
3. HCl also has permanent dipole-dipole attractions.

More energy is required to separate HCl molecules.

EXAM TRAP Permanent dipoles do not automatically outweigh a much larger electron cloud. Compare the whole set of interactions.

TEST IT

Try question 39 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You draw correctly aligned partial charges and distinguish the dotted attraction from a covalent bond.

Hydrogen bonding & properties

21. Hydrogen bonding and correct diagrams

SPEC 2.2.2(l)

CRACK IT

Hydrogen bonding needs a particular arrangement. Find a δ+ hydrogen **directly bonded to N, O or F**, then a lone pair on N, O or F in a neighbouring molecule.

1. Find a δ+ H atom directly bonded to N, O or F in one molecule.

2. Find a lone pair on N, O or F in a neighbouring molecule.

3. Draw a dotted attraction from that lone pair to the δ+ H. The covalent bond from H to its own atom stays in place.

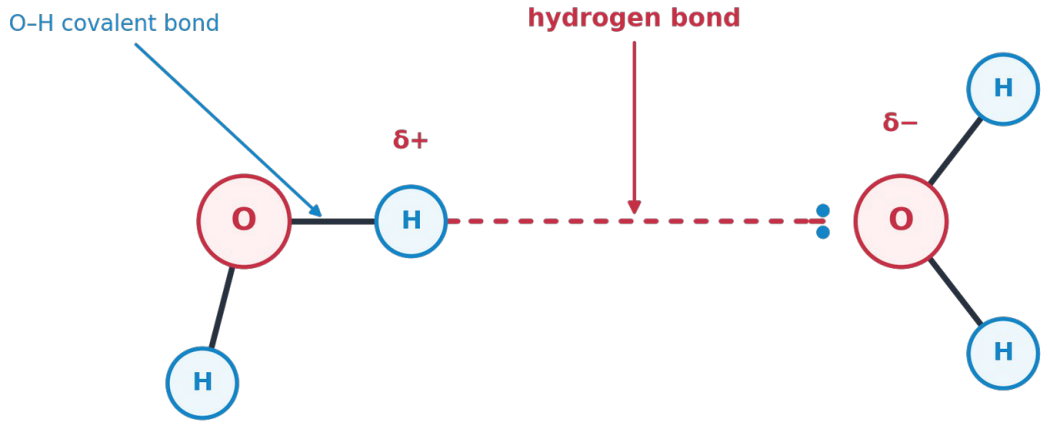
4. Add the appropriate partial charges and keep the hydrogen-bond arrangement clear.

Water, ammonia and HF can hydrogen-bond between their own molecules. An oxygen lone pair can attract a δ+ H atom in water, even when the oxygen-containing molecule has no O-H bond and cannot hydrogen-bond with itself. H bonded only to carbon does not meet the hydrogen-bonding condition here.

DEFINITIONS & KEY TERMS

Hydrogen bond

An intermolecular attraction between a lone pair on N, O or F in one molecule and a δ+ 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.

Check	Requirement
Hydrogen atom	H directly bonded to N, O or F.
Lone pair	An available lone pair on N, O or F.
Drawing	Solid covalent bonds; dotted intermolecular line from the neighbouring N/O/F lone pair to δ+ H; correct δ signs.

WORKED EXAMPLE

Water compared with hydrogen sulfide

Why does H_2O have a much higher boiling point than H_2S?

Worked answer

1. H_2O has O–H bonds and oxygen lone pairs, so its molecules hydrogen-bond.
2. H_2S does not form the same hydrogen-bonding network.
3. More energy is needed to overcome the attractions between water molecules.

H_2O boils higher and is less volatile.

EXAM TRAP Keep the N/O/F lone pair, dotted line, H and its covalently bonded N/O/F approximately collinear. Do not draw a solid bond between molecules.

TEST IT

Try questions 40, 41, 42, 43, 44 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can identify every part of a hydrogen bond and explain its effect on properties.

22. Water’s boiling point and the density of ice

SPEC 2.2.2(m)

CRACK IT

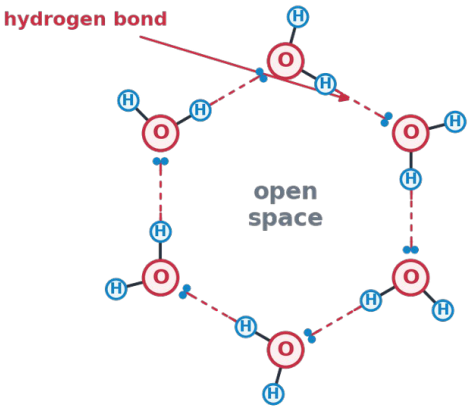
Hydrogen bonds help explain two unusual properties of water. They make water's boiling point higher than the simple trend for comparable hydrides suggests, and they give ice its open structure.

In ice, hydrogen bonds hold water molecules further apart than they are on average in liquid water. The same mass occupies a larger volume, so ice is less dense and floats.

On melting, some hydrogen bonds are overcome and the open arrangement partly collapses. The O-H covalent bonds inside each water molecule remain intact.

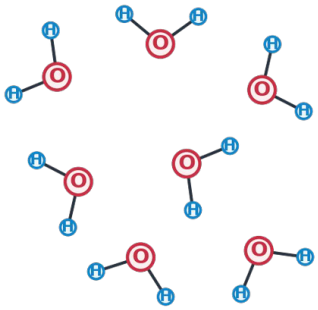
Use the same link for boiling-point comparisons: stronger overall attractions between molecules mean more energy is needed to separate them.

ICE



open hydrogen-bonded lattice
lower density • ice floats

LIQUID WATER



network rearranges
molecules pack closer • higher density

Anomaly	Explanation
Relatively high melting and boiling points	Many hydrogen bonds require a substantial energy input to overcome enough attractions for molecules to move or separate.
Ice is less dense than liquid water	An ordered open lattice holds water molecules farther apart. Part of the network collapses on melting.

WORKED EXAMPLE

Why does ice float?

Explain why solid ice is less dense than liquid water and therefore floats on water.

Worked answer

1. Hydrogen bonds hold water molecules in an open, ordered lattice.
2. For the same mass, ice occupies a larger volume.
3. Density = mass ÷ volume, so ice is less dense.

Ice floats on the denser liquid.

EXAM TRAP O–H covalent bonds remain intact during melting and boiling. The diagram shows selected links, not a complete crystal unit cell.

TEST IT

Try question 45 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You connect hydrogen bonding both to the energy needed for state changes and to the spacing of molecules.

23. Molecular lattices, solubility and volatility

SPEC 2.2.2(n-o)

CRACK IT

For a molecular substance, separate the strong covalent bonds inside the molecules from the attractions between them. Melting and boiling normally overcome the between-molecule attractions.

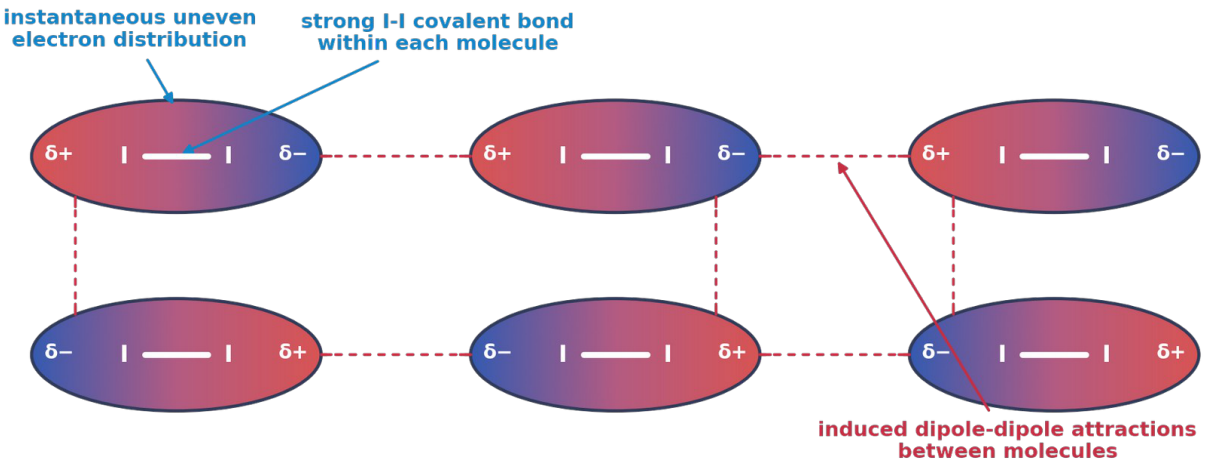
Stronger overall attractions tend to increase boiling point and reduce volatility. Neutral molecules normally provide no mobile ions or delocalised electrons, so the pure molecular substance does not conduct.

For solubility, identify the attractions disrupted and the new ones formed with the solvent. A molecule can accept a hydrogen bond from water without being able to hydrogen-bond to itself. The examples below show how to name the relevant interactions.

DEFINITIONS & KEY TERMS

Simple molecular lattice

Discrete covalently bonded molecules held together in a solid by intermolecular forces.



Property	Explanation
Melting / boiling	Intermolecular forces are overcome; molecules usually stay intact.
Pure electrical conductivity	No mobile ions or delocalised electrons.
Solubility	Depends on new solute-solvent attractions compared with those disrupted.
Volatility	Weaker attractions usually mean a lower boiling point and higher volatility.

WORKED EXAMPLE

Iodine versus diamond

Explain why iodine has a much lower melting point than diamond. Refer to their structures and the attractions that must be overcome.

Worked answer

Iodine is simple molecular, so melting overcomes attractions between I_2 molecules. Diamond is a giant covalent network whose strong C–C bonds would need to be broken.

Much more energy is required for diamond.

WORKED EXAMPLE

Iodine in cyclohexane

Why is iodine much more soluble in cyclohexane than in water?

Worked answer

I_2 and cyclohexane are non-polar. Favourable induced dipole-dipole attractions can form between them, whereas iodine does not replace water's hydrogen-bond attractions as effectively.

Iodine is more soluble in cyclohexane.

WORKED EXAMPLE

Methanol in water

Why does methanol mix completely with water?

Worked answer

Methanol has an O–H group and oxygen lone pairs. Methanol and water form favourable hydrogen bonds with one another.

Methanol is miscible with water.

EXAM TRAP "Like dissolves like" is a prediction, not an absolute law. A molecular solute conducts only if mobile ions are present, for example after reaction or ionisation.

TEST IT

Try question 46 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You justify each physical property with particles and attractions rather than a slogan.

24. Build complete intermolecular-force comparisons

SPEC 2.2.2(k-o)

CRACK IT

To compare boiling points, explain what has to be overcome when the particles separate. Stronger overall intermolecular attractions need more energy, so boiling point tends to be higher.

1. Identify the molecules and all the forces between them.

2. Compare electron count, polarisability, shape and any extra permanent-dipole or hydrogen-bonding attractions.

3. Decide which substance has stronger overall attractions.

4. Link that to more energy needed and the higher boiling point.

A more volatile substance evaporates more readily. Stronger attractions generally reduce volatility at a given temperature. Melting points also depend on how well particles pack in the crystal, so avoid using one simple force ranking for every comparison.

Particles → named attraction → relative strength → energy → property

Weak explanation	Improve it
"NH_3 has stronger bonds"	Name hydrogen bonds between molecules, not N–H covalent bonds.
"I_2 is bigger"	Explain electron count, polarisability, induced dipoles, attraction and energy.
"CO_2 bonds are non-polar"	The C=O bonds are polar; equal opposite dipoles cancel.

WORKED EXAMPLE

Ammonia compared with phosphine

Why does NH_3 have a higher boiling point than PH_3?

Worked answer

NH_3 contains N–H bonds and a nitrogen lone pair, so its molecules hydrogen-bond. PH_3 does not. More energy is needed to separate NH_3 molecules.

NH_3 has a higher boiling point and lower volatility.

EXAM TRAP Molecular polarity is one factor. Larger electron clouds can make a non-polar substance boil higher than a smaller polar one.

TEST IT

Try questions 47, 48 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN Every comparison names the interactions before linking their strength to energy and the observed property.

25. Combine shape, polarity and physical properties

SPEC 2.2.2(g-o)

CRACK IT

A longer question may connect shape, polarity and boiling point. Work through those ideas in that order so one answer supports the next.

1. Count the central-atom electron regions, including lone pairs.

2. Draw and name the three-dimensional molecular shape.

3. Use that shape to decide whether the bond dipoles cancel.

4. Identify all the applicable intermolecular forces, including induced dipoles.

5. Explain the property using energy needed to separate particles, or the mobility of charge carriers.

Keep the reasoning visible. A correct shape or a named force on its own is only part of the explanation.

Electron regions → shape → bond dipoles → overall polarity → intermolecular forces → property

WORKED EXAMPLE

BF3 has polar B–F bonds

BF_3 contains polar B–F bonds. Predict its molecular shape, bond angle and overall polarity, explaining your reasoning.

Worked answer

1. Boron has three bonding regions and no lone pairs.

2. The regions arrange trigonal planar at 120°.

3. Three equal B–F dipoles are symmetrically arranged and cancel.

BF_3 is trigonal planar, 120° and non-polar overall.

EXAM TRAP Do not decide molecular polarity from the name of a bond, or boiling point from polarity alone.

TEST IT

Try question 49 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can link the whole chain without confusing covalent bonds with attractions between molecules.