AQA | 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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Covalent & dative bonds page 7
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Ionic bonding & formulae

1. Ion formation and ionic bonding

SPEC 3.1.3.1

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

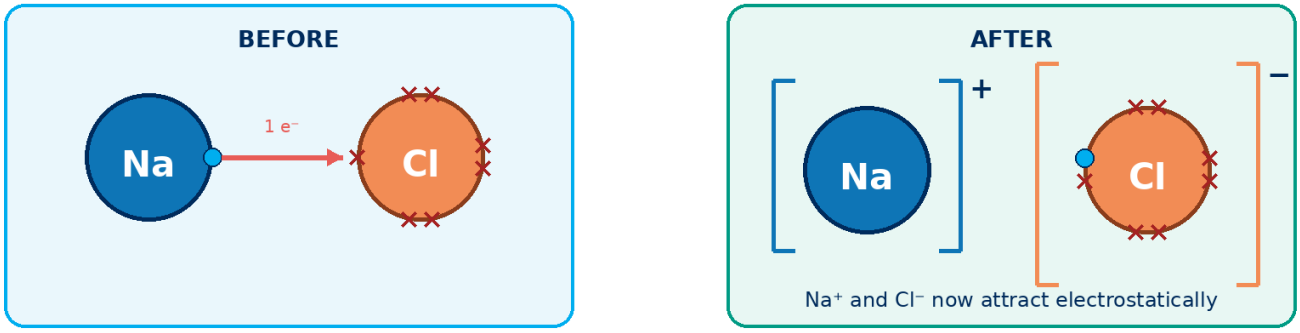
Ion formation

Metals lose electrons to form positive ions; non-metals gain electrons to form negative ions.

Ionic bonding

Strong electrostatic attraction between oppositely charged ions in a lattice.

ION FORMATION: electron transfer



Group	Electron change	Typical ion charge	Example
1	Loses 1	1+	Na^{+}
2	Loses 2	2+	Mg^{2+}
3	Loses 3	3+	Al^{3+}
5	Gains 3	3−	N^{3-}
6	Gains 2	2−	O^{2-}
7	Gains 1	1−	Cl^{-}

EXAM TRAP Electron transfer makes the ions. It is not the definition of the ionic bond.

TEST IT

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

DONE WHEN You can predict common ion charges and define ionic bonding precisely.

2. Ionic formulae and compound ions

SPEC 3.1.3.1

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.

Compound ion	Formula
Sulfate	SO_4^{2-}
Hydroxide	OH^-
Nitrate	NO_3^-
Carbonate	CO_3^{2-}
Ammonium	NH_4^+

WORKED EXAMPLE

Aluminium sulfate

Construct the formula from Al^{3+} and SO_4^{2-}.

Worked answer

1. Two Al^{3+} ions give +6.

2. Three sulfate ions give −6.

3. Keep sulfate intact and use brackets: $Al_2(SO_4)_3$.

4. Check neutrality and the smallest ratio.

$Al_2(SO_4)_3$

Ions	Ratio	Formula
Mg^{2+} and Cl^{-}	1 : 2	$MgCl_2$
Al^{3+} and O^{2-}	2 : 3	Al_2O_3
Ca^{2+} and NO_3^{-}	1 : 2	$Ca(NO_3)_2$
Ca^{2+} and O^{2-}	1 : 1	CaO

EXAM TRAP “Drop and swap” can leave an unsimplified formula. Use brackets for repeated compound ions, never alter their internal formulae.

TEST IT

Try questions 2, 3 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can balance charges, use brackets and simplify the ratio.

3. Giant ionic lattices and conductivity

SPEC 3.1.3.4

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.

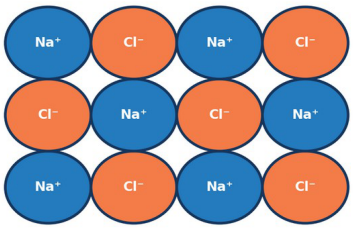
Many ionic compounds dissolve in water, although not all do. If an ionic compound forms an aqueous solution, its mobile ions can carry charge. Do not explain ionic conductivity using moving electrons.

DEFINITIONS & KEY TERMS

Giant ionic lattice

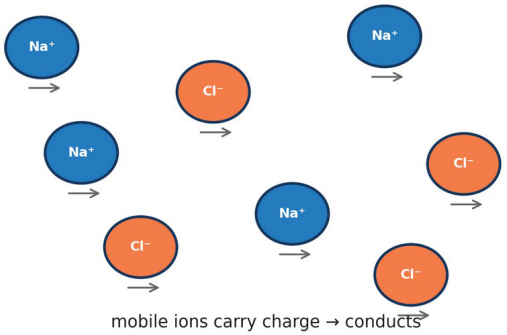
A regular repeating three-dimensional arrangement of oppositely charged ions.

SOLID: ions touching and fixed



no mobile charged particles → does not conduct

MOLTEN: ions separated and mobile



mobile ions carry charge → conducts

State	Ions	Conducts?
Solid	Fixed in lattice positions	No: no mobile charge carriers
Molten	Mobile	Yes: ions carry charge
Aqueous, if dissolved	Hydrated and mobile	Yes: ions carry charge

TEST IT

Try question 4 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can explain solid and molten conductivity by naming the charged particles and their mobility.

4. Explaining ionic melting points

SPEC 3.1.3.4

CRACK IT

A high melting point needs an explanation of **which particles attract** and **why separating them needs energy**. For an ionic compound, those particles are oppositely charged ions.

Higher ion charges and smaller ionic radii generally strengthen the electrostatic attraction. More energy is then needed to overcome the attractions, so the melting point tends to be higher.

Use the actual ions in a comparison. Saying that one substance has 'stronger bonds' without identifying the attraction leaves out the chemistry.

Structure → particles → attraction → energy → property

WORKED EXAMPLE

Why sodium chloride has a high melting point

Explain why sodium chloride has a high melting point. Refer to its structure, the particles present and the attractions between them.

Worked answer

1. NaCl has a giant ionic lattice.

2. Strong electrostatic attractions act between oppositely charged Na^+ and Cl^- ions throughout the lattice.

3. Much energy is needed to overcome these attractions.

Therefore sodium chloride has a high melting point.

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.

EXAM TRAP "Lots of strong bonds" does not name the ions or the electrostatic attraction.

TEST IT

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

DONE WHEN You can explain high melting point using ions, electrostatic attraction and energy.

Covalent & dative bonds

5. Covalent bonds and dot-and-cross diagrams

SPEC 3.1.3.2

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.

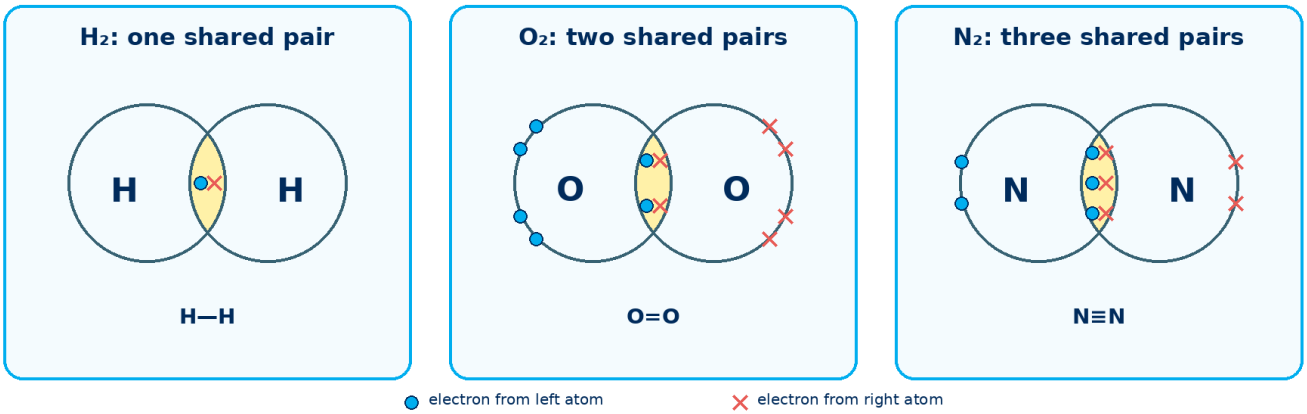
A lone pair is not shared in a bond. Show these electrons outside the overlap, and use dots and crosses consistently to show their atom of origin. Count all the outer-shell electrons before finishing.

DEFINITIONS & KEY TERMS

Single covalent bond

A shared pair of electrons attracted to both positive nuclei.

DOT-AND-CROSS: outer-shell electrons



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

EXAM TRAP Covalent bonds within a molecule remain intact when a molecular substance melts or boils. Forces between molecules are overcome.

TEST IT

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

DONE WHEN You can draw complete outer-shell diagrams and convert shared pairs into bond lines.

6. Dative bonds: donor to acceptor

SPEC 3.1.3.2

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**.

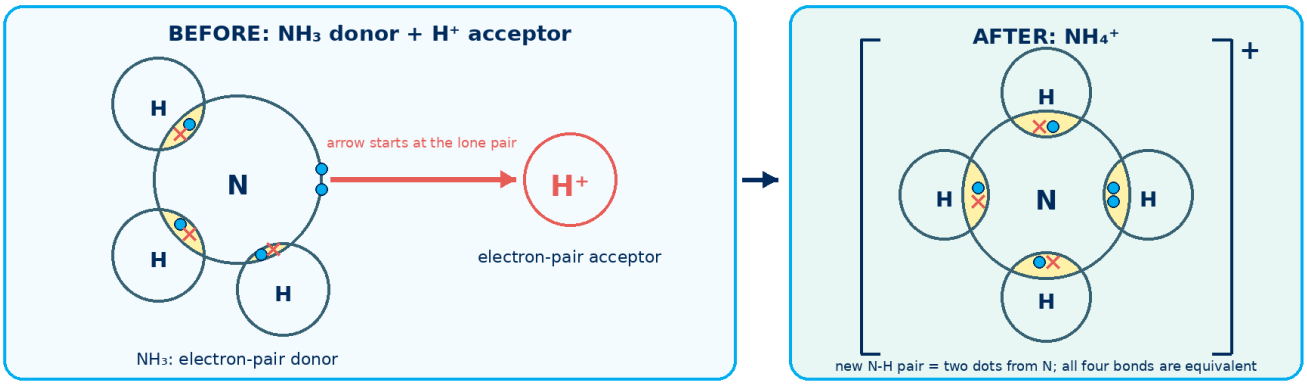
For NH_3 reacting with H^+, nitrogen supplies its lone pair to the hydrogen ion. Once NH_4^+ has formed, all four N-H bonds are equivalent. The origin of the electrons does not make one of those bonds different afterwards.

DEFINITIONS & KEY TERMS

Dative covalent bond

A shared pair of electrons with both electrons supplied by one atom.

DATIVE BOND: a lone pair is donated to H^+



Role	$NH_3 + H^+ \rightarrow NH_4^+$
Donor	Nitrogen in NH_3 has an available lone pair
Acceptor	H^+ accepts the pair

EXAM TRAP The arrow follows the electron pair: donor → acceptor.

TEST IT

Try questions 10, 11, 12 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can identify donor and acceptor and draw an arrow from the lone pair to the acceptor.

Metallic bonding

7. Metallic bonding and mobile electrons

SPEC 3.1.3.3

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

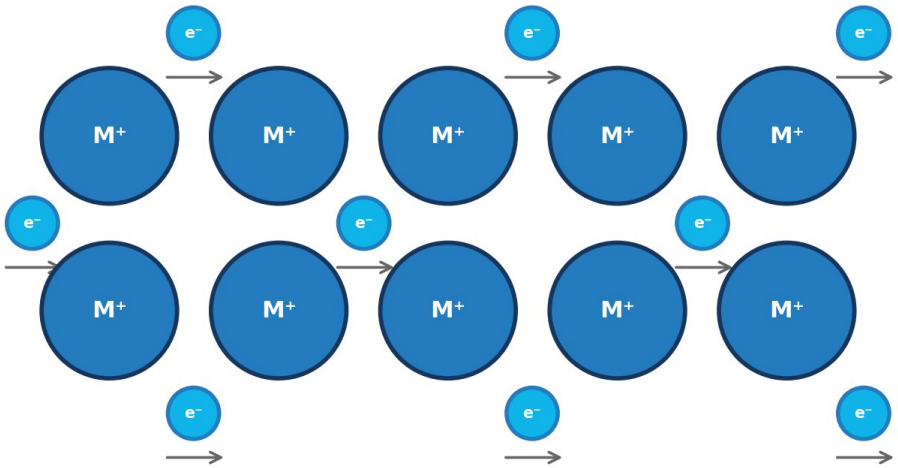
Metallic bonding

Strong electrostatic attraction between positive metal ions and delocalised electrons.

Delocalised electron

An electron not confined to one atom or one bond, able to move through the structure.

GIANT METALLIC LATTICE



positive ions are close together; mobile delocalised electrons carry charge

EXAM TRAP Name the charge carrier as delocalised electrons, rather than merely “electrons” or “free electrons”.

TEST IT

Try question 13 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can define metallic bonding and explain conductivity and melting point.

8. Sodium chloride and magnesium

SPEC 3.1.3.4

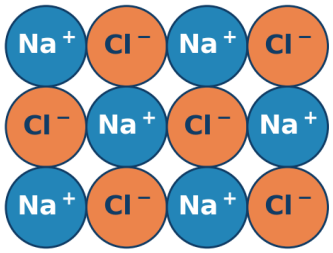
CRACK IT

Sodium chloride and magnesium both have giant structures with strong electrostatic attractions. The difference is **what is attracting what**.

In sodium chloride, positive sodium ions attract negative chloride ions. In magnesium, positive metal ions attract delocalised electrons.

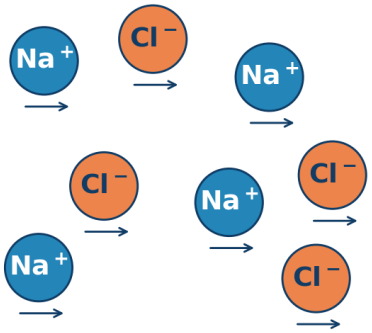
This explains their different conductivity patterns. Magnesium already has mobile electrons in the solid. Solid sodium chloride has ions, but they cannot move through the lattice until it melts or dissolves.

Solid sodium chloride



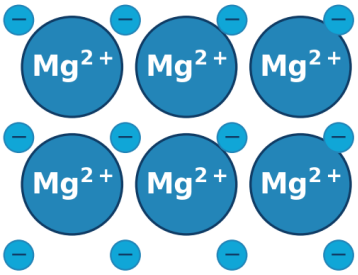
Ions fixed.
No mobile charge carriers.
Does not conduct.

Molten sodium chloride



Mobile ions carry charge.
Conducts.

Magnesium solid or liquid



Mobile delocalised electrons carry charge.
Conducts.

Feature	Sodium chloride	Magnesium
Particles	Na^+ and Cl^- ions	Mg^{2+} ions and delocalised electrons
Solid	Does not conduct: ions fixed	Conducts: delocalised electrons mobile
Molten	Conducts: ions mobile	Conducts: delocalised electrons mobile

TEST IT

Try question 14 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can distinguish mobile ions from mobile delocalised electrons.

Structures & properties

9. Four structure types: particles and attractions

SPEC 3.1.3.4

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	Bond or attraction
Giant ionic	Positive and negative ions	Electrostatic attractions between oppositely charged ions
Metallic	Positive metal ions and delocalised electrons	Electrostatic attraction between positive ions and delocalised electrons
Giant covalent	Atoms	Many strong covalent bonds throughout a network
Simple molecular	Molecules	Forces between molecules

EXAM TRAP Do not jump straight from the structure name to the property. The explanation lies in the particles and attractions.

TEST IT

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

DONE WHEN You can identify particles and attractions in all four crystal types.

10. Identifying structures from physical properties

SPEC 3.1.3.4

CRACK IT

Treat the properties as clues that must fit the same structure. A high melting point suggests strong attractions, but it does not by itself tell you which particles are present.

Conductivity is especially useful. A solid metal conducts through mobile electrons. An ionic compound normally conducts only when its ions can move, such as when molten or dissolved.

Check the full set of observations against the table. Graphite is an important exception to the idea that giant covalent substances never conduct.

Structure	Melting point	Conductivity
Giant ionic	High	Solid: no; molten/aqueous: yes, via ions
Metallic	Often high	Solid and liquid: yes, via delocalised electrons
Giant covalent	Very high	Usually no; graphite is an exception
Simple molecular	Usually low relative to giant structures	No mobile charged particles in the pure substance

TEST IT

Try questions 17, 18, 19, 20 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can justify the most likely structure from property evidence.

11. Diamond: a rigid covalent network

SPEC 3.1.3.4

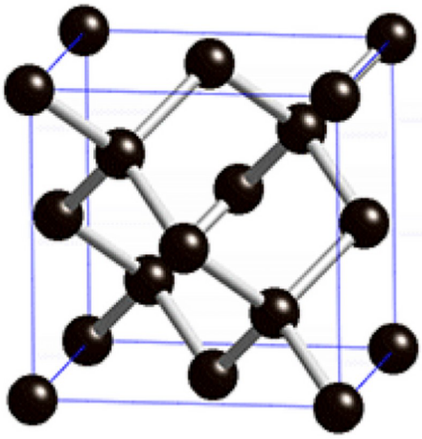
CRACK IT

Diamond is one giant covalent network, rather than a collection of small carbon molecules. Each carbon atom forms four strong covalent bonds in a rigid three-dimensional arrangement.

Melting requires many strong covalent bonds to be broken, so a lot of energy is needed. The rigid network also makes diamond very hard.

All four outer electrons of each carbon are used in bonds. There are no mobile charge carriers, so diamond does not conduct electricity.

Diamond: 3D giant covalent



each C forms 4 strong covalent bonds
no mobile charged particles

EXAM TRAP Diamond is not made of discrete molecules. Do not explain its properties using intermolecular forces.

TEST IT

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

DONE WHEN You can explain diamond’s hardness, very high melting point and lack of conductivity.

12. Graphite: layers and delocalised electrons

SPEC 3.1.3.4

CRACK IT

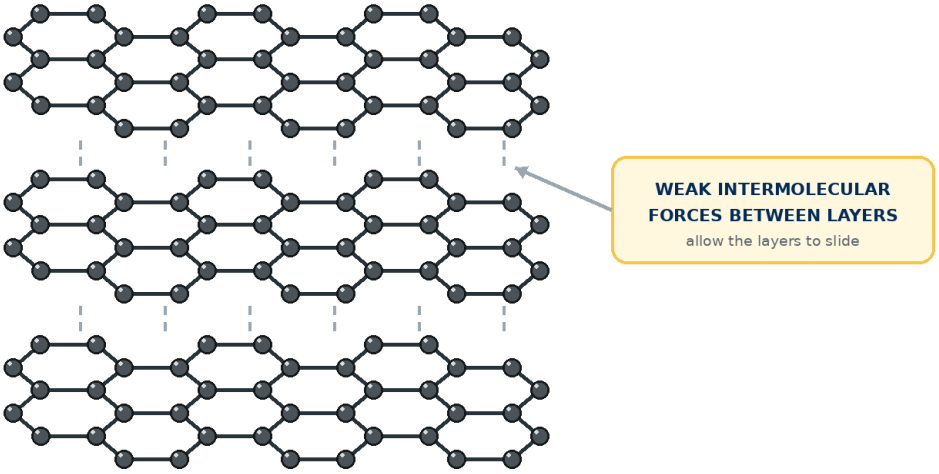
Graphite contains flat layers of carbon atoms. Each carbon makes three strong covalent bonds, leaving one electron per carbon delocalised.

The strong bonds within the layers explain the high melting point. The delocalised electrons can move along the layers and carry charge, so graphite conducts electricity.

Weak van der Waals attractions between the layers allow them to slide over one another. This explains softness; it is a different explanation from the one for melting point.

Graphite is a giant covalent structure. The diagram's forces between layers do not mean that each layer is a collection of small molecules.

Graphite: layered giant covalent



each carbon forms 3 covalent bonds in a flat layer
one electron per carbon is delocalised and mobile within each layer

EXAM TRAP Sliding layers do not require the strong covalent bonds within each layer to break.

TEST IT

Try questions 22, 23 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can compare diamond and graphite using bonding, electron mobility and layers.

13. Molecular crystals: iodine and ice

SPEC 3.1.3.4

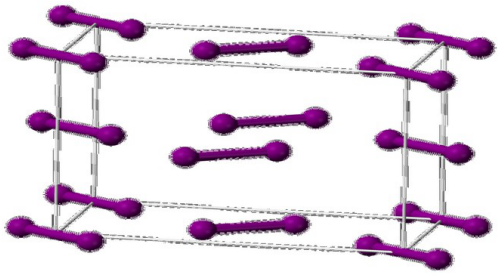
CRACK IT

Iodine and ice are made of separate molecules arranged in a solid. There are covalent bonds **inside** each molecule and attractions **between** neighbouring molecules.

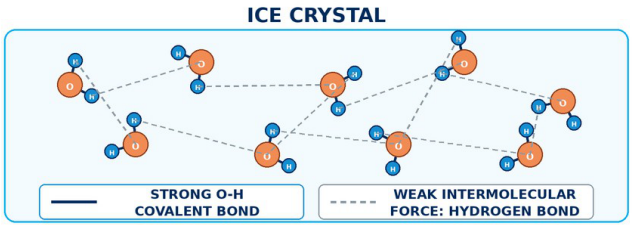
Melting partly overcomes the attractions between molecules. Boiling separates the molecules further. The molecules themselves remain intact.

For iodine, name van der Waals forces between I_2 molecules. For ice, name hydrogen bonds between water molecules. Do not say that melting breaks the I-I or O-H covalent bonds.

Solid iodine: I_2 molecular lattice



strong I—I bonds inside each molecule
weaker forces between molecules



Crystal	Particles	Within molecules	Between molecules
Iodine	I_2 molecules	Strong I–I covalent bonds	van der Waals forces
Ice	H_2O molecules	Strong O–H covalent bonds	Hydrogen bonds

Change	Energy
Melting or boiling	Energy absorbed to overcome attractions
Freezing or condensing	Energy released as attractions form

EXAM TRAP Do not call iodine's covalent bonds weak or claim that melting ice breaks O–H bonds.

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 can identify molecules and the attractions involved in changes of state.

14. Drawing exact numbers of particles

SPEC 3.1.3.4

CRACK IT

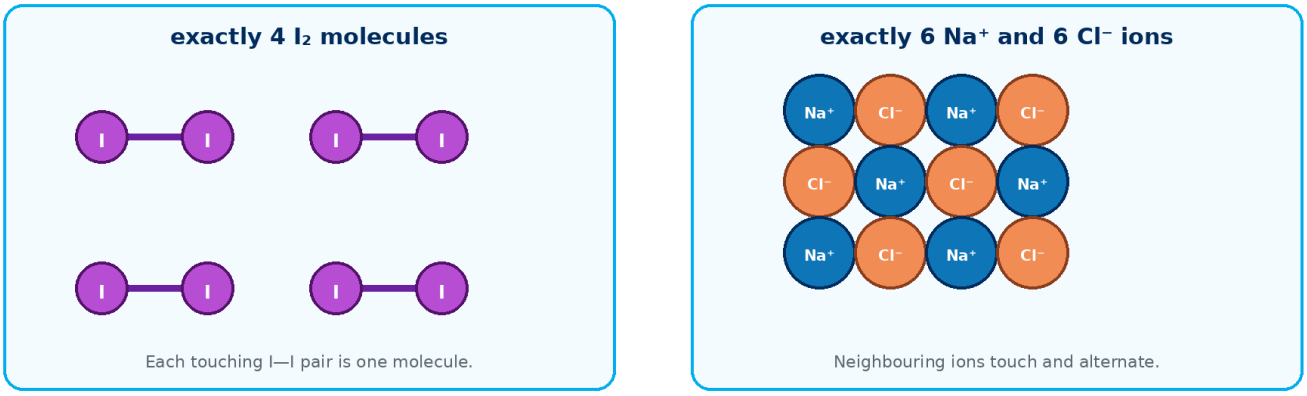
Before drawing, underline what you have been asked to count: atoms, ions or molecules. The number of circles you need depends on that choice.

1. For I_2, draw separate touching pairs. One pair is one molecule, so four molecules need eight atoms.

2. For NaCl, show an alternating arrangement of touching positive and negative ions, with the charges labelled.

3. Count the requested entities again at the end. Check that the overall composition and charge match the substance.

SPECIFIED PARTICLES: count the requested entities



EXAM TRAP Four I_2 molecules contain eight iodine atoms. A correct pattern with the wrong number of particles can lose marks.

TEST IT

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

DONE WHEN Your drawing has the requested entity, exact count and correct charges.

Shapes & bond angles

15. VSEPR: count charge clouds, then minimise repulsion

SPEC 3.1.3.5

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

VSEPR

Valence Shell Electron Pair Repulsion: outer-shell electron pairs repel and arrange as far apart as possible.

Charge cloud

One region of electron density: each bond region and each lone pair counts as one.

Total pairs = (central valence electrons + singly bonded atoms + negative charge − positive charge) ÷ 2

WORKED EXAMPLE

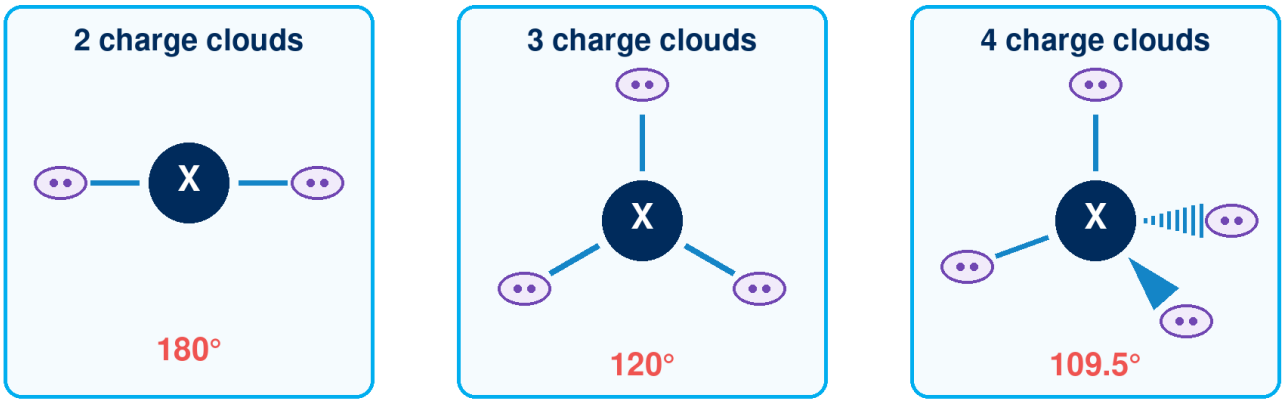
Ammonia versus ammonium

For NH_3 and NH_4^+, calculate the total number of electron pairs around nitrogen, state how many are bonding pairs and lone pairs, and predict the shape.

Worked answer

1. NH_3: (5 + 3)/2 = 4 pairs; 3 bonding and 1 lone.
2. NH_4^+: (5 + 4 − 1)/2 = 4 pairs; all 4 bonding.

NH_3: trigonal pyramidal. NH_4^+: tetrahedral.



EXAM TRAP This shortcut is for suitable singly bonded species. For multiple bonds or supplied electron diagrams, count regions directly: a single, double or triple bond is one bonding 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 can count regions, identify lone pairs and explain minimisation of repulsion.

16. Two and three charge clouds

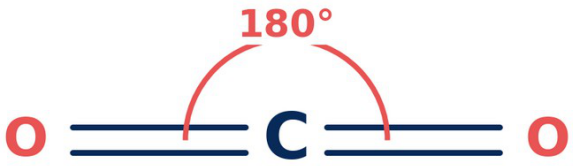
SPEC 3.1.3.5

CRACK IT

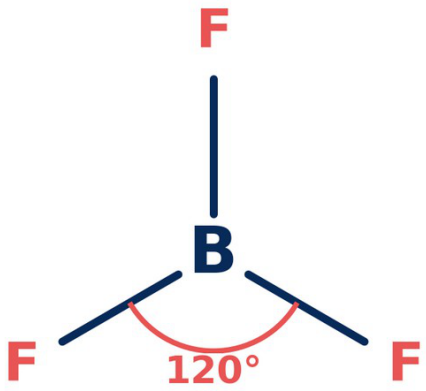
With two electron regions and no lone pairs, the greatest separation is opposite sides of the central atom. The shape is **linear, 180°**.

With three electron regions and no lone pairs, the regions spread into one plane. The shape is **trigonal planar, 120°**.

Count regions, not individual shared pairs. Each double bond in CO_2 counts as one region, so CO_2 has two regions around carbon and is linear.



linear • two charge clouds



trigonal planar • 120°

Species	Bonding regions	Lone pairs	Shape	Angle
CO_2	2	0	Linear	180°
$BeCl_2$	2	0	Linear	180°
BF_3	3	0	Trigonal planar	120°
SO_3	3	0	Trigonal planar	120°

WORKED EXAMPLE

Sulfur trioxide

Deduce SO_3's shape from three bonding regions and no lone pairs.

State the bond angle and explain the arrangement of the bonding regions.

Worked answer

The three charge clouds repel equally into one plane.

Trigonal planar, 120°.

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 recognise two- and three-region shapes without counting double bonds twice.

17. Four bonding regions: tetrahedral shapes

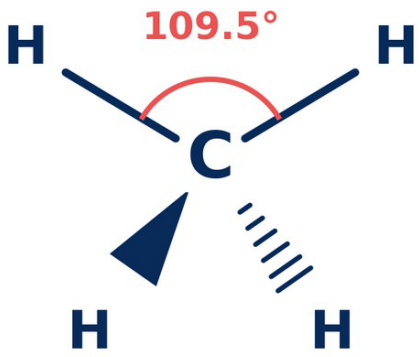
SPEC 3.1.3.5

CRACK IT

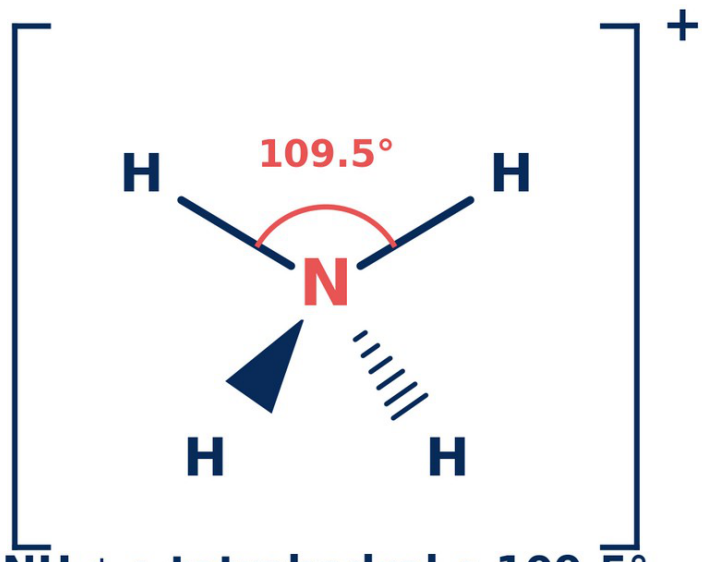
Four bonding regions spread out in three dimensions, giving a **tetrahedral shape with angles of 109.5°**. Four flat bonds at right angles would put the regions too close together.

CH_4, $SiCl_4$, NH_4^+ and BF_4^- all have four bonding regions and no lone pairs around the central atom.

- An ordinary line shows a bond in the plane of the paper.
- A solid wedge shows a bond coming towards you.
- A hashed wedge shows a bond going away from you.
- For an ion, put brackets and the overall charge around the whole species.



tetrahedral • 109.5°



NH_4^+ • tetrahedral • 109.5°

EXAM TRAP Do not draw tetrahedral as four flat bonds at right angles.

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 can draw tetrahedral molecules and ions using valid wedge-and-dash bonds.

18. Lone pairs and compressed bond angles

SPEC 3.1.3.5

CRACK IT

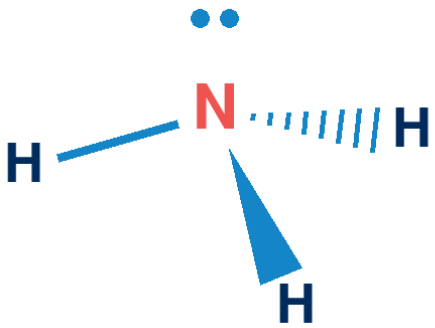
A lone pair takes up more space near the central atom than a bonding pair. It is attracted to one nucleus rather than shared between two nuclei. It repels nearby bonding regions more strongly and pushes them closer together.

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.

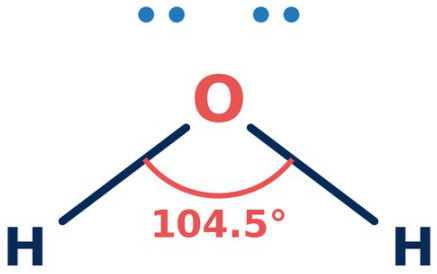
- CH_4: four bonds, no lone pairs; tetrahedral, 109.5°.
- NH_3: three bonds, one lone pair; pyramidal, about 107°.
- H_2O: two bonds, two lone pairs; bent/non-linear, about 104.5°.

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.

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



trigonal pyramidal · 107°



non-linear (bent) • 104.5°

Species	Bonding regions	Lone pairs	Shape	Angle
SO_2	2	1	Bent	Less than 120°
NH_3	3	1	Trigonal pyramidal	107°
H_2O	2	2	Bent	104.5°
H_3O^+	3	1	Trigonal pyramidal	Use supplied data; do not automatically assume 107°

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 can explain the familiar angle sequence and avoid treating 2.5° as a universal rule.

19. Five and six charge clouds

SPEC 3.1.3.5

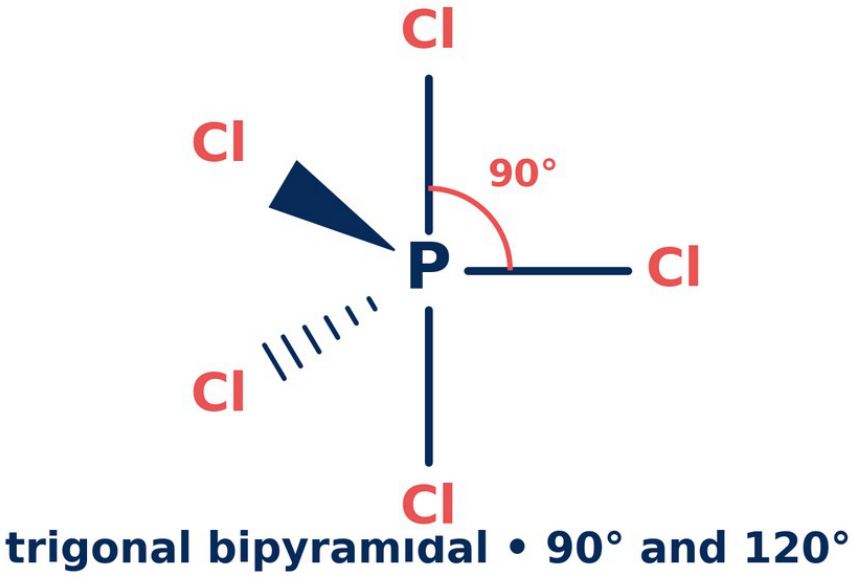
CRACK IT

Five regions arrange in a trigonal bipyramid. Three equatorial positions lie in one plane; two axial positions point above and below it. Six regions arrange octahedrally.

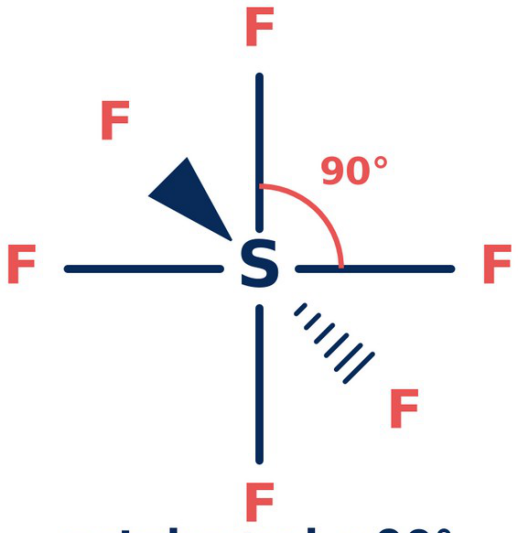
Lone pairs still occupy positions in those arrangements, even though they are left out when naming the molecular shape. That gives the seesaw, T-shaped, linear, square-pyramidal and square-planar shapes in the table.

With five regions, lone pairs take equatorial positions to reduce the number of 90° repulsions. With six regions and two lone pairs, as in XeF_4, the lone pairs sit opposite one another.

Use the table's angles and any data supplied. Do not apply the methane/ammonia/water subtraction shortcut to these shapes.



trigonal bipyramidal • 90° and 120°



octahedral • 90°

Total regions	Bonding + lone	Example	Molecular shape	Angles
5	5 + 0	PCl_5	Trigonal bipyramidal	90°, 120°, 180°
5	4 + 1	SF_4	Seesaw	Distorted below ideal 90°, 120°, 180°
5	3 + 2	ClF_3	T-shaped	Approximately 90°, 180°, slightly compressed
5	2 + 3	XeF_2	Linear	180°
6	6 + 0	SF_6	Octahedral	90°, 180°
6	5 + 1	BrF_5	Square pyramidal	Approximately 90°, 180°
6	4 + 2	XeF_4	Square planar	90°, 180°

EXAM TRAP Five bonds are not a flat pentagon, and six bonds are not a flat hexagon.

TEST IT

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

DONE WHEN You can name and justify the five- and six-region shapes, including lone-pair variants.

20. Deducing unfamiliar molecular shapes

SPEC 3.1.3.5

CRACK IT

An unfamiliar formula does not need a new method. Work out its electron regions, place any lone pairs, then look at where the atoms sit.

1. Count the bonding regions and lone pairs around the central atom.

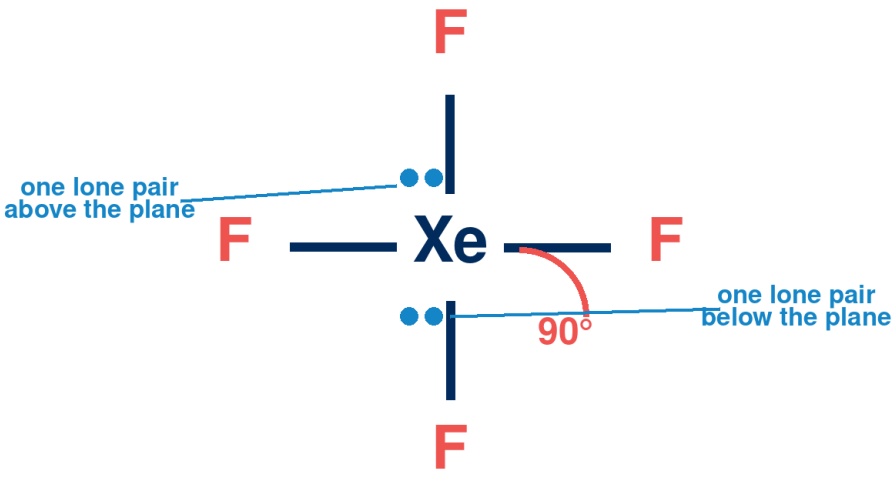
2. Choose the parent electron arrangement from the total number of regions.

3. Place lone pairs where repulsion is minimised.

4. Ignore the lone-pair positions when naming the molecular shape, but keep their effect on the arrangement.

For XeF_4, six regions include four bonds and two opposite lone pairs. The atoms left in the middle plane form a square, so the shape is square planar.

Central atom → bonding regions → lone pairs → parent geometry → molecular shape



square planar • 90° and 180°

Species	Bonding regions	Lone pairs	Reasoning	Shape
NO_2^+	2	0	Opposite regions	Linear
SO_2	2	1	Three-region parent; omit lone-pair position	Bent
XeF_4	4	2	Six-region parent; opposite lone pairs	Square planar

TEST IT

Try questions 33, 34 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can work from supplied bonding regions and lone pairs without guessing.

Electronegativity & polarity

21. Electronegativity and polar bonds

SPEC 3.1.3.6

CRACK IT

In a covalent bond, both atoms attract the shared electrons. Electronegativity describes how strongly an atom attracts that bonding pair.

If one atom attracts the pair more strongly, the electrons are shared unequally. That atom becomes δ− and the other becomes δ+. These are partial charges, not the full charges of ions.

Use the difference between the supplied electronegativity values. A larger difference generally gives a more polar bond, but there is no single numerical cut-off that works for every bond.

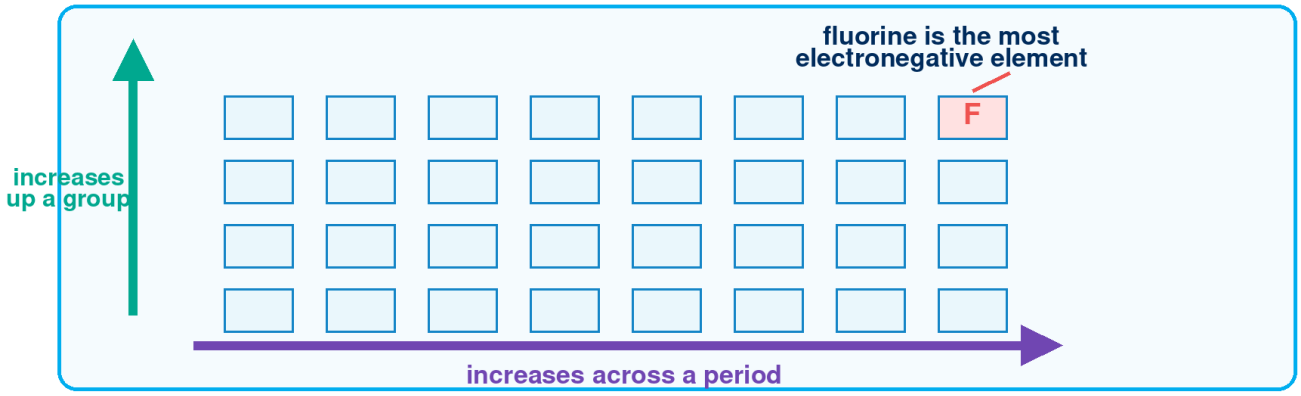
DEFINITIONS & KEY TERMS

Electronegativity

The power of an atom to attract the pair of electrons in a covalent bond.

Polar covalent bond

Unequal sharing of bonding electrons produces δ+ and δ− ends.



Electronegativity difference = |value for atom 1 − value for atom 2|

EXAM TRAP Subtract the electronegativities; do not average them. δ denotes partial, not full ionic charge.

TEST IT

Try questions 35, 36 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can define electronegativity, compare differences and assign partial charges.

22. From bond polarity to molecular polarity

SPEC 3.1.3.6

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.

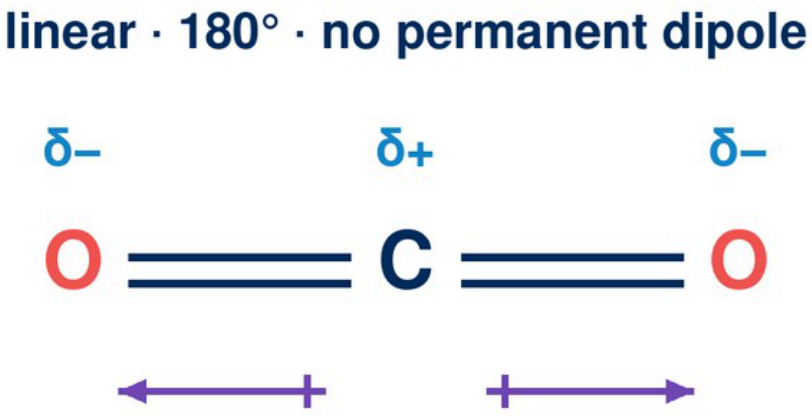
1. Identify which bonds are polar and mark their dipole directions.

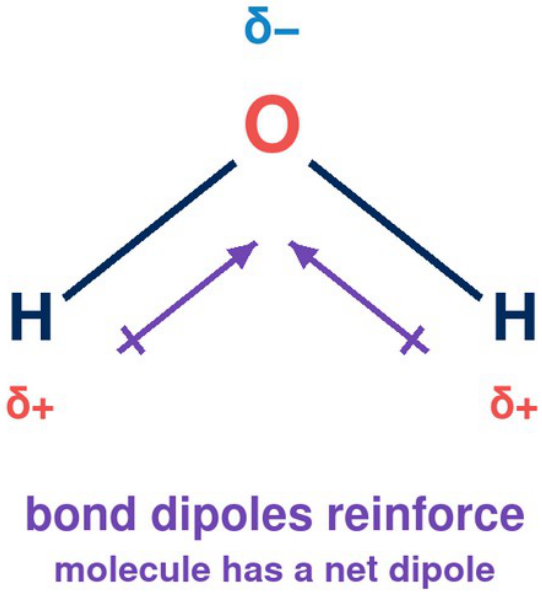
2. Work out the molecular shape, including the effect of lone pairs.

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.





WORKED EXAMPLE

SO3 and SO2

SO_3 and SO_2 both contain polar S–O bonds. Explain why SO_2 is polar overall but SO_3 is not.

Worked answer

SO_3 is symmetrical trigonal planar, so equivalent dipoles cancel. SO_2 is bent, so its dipoles do not cancel.

SO_3 is non-polar; SO_2 has a permanent dipole.

TEST IT

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

DONE WHEN You can distinguish a polar bond from a polar molecule.

23. Symmetry and cancellation of dipoles

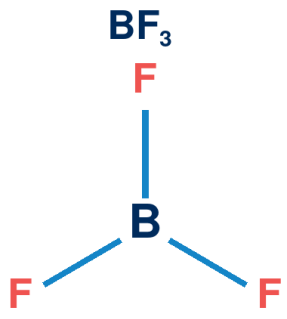
SPEC 3.1.3.6

CRACK IT

To decide whether a molecule is polar, look at the **whole arrangement of bond dipoles**. Their sizes and directions must balance for complete cancellation.

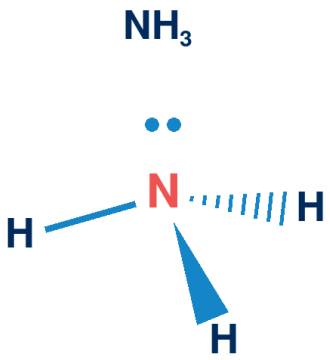
BF_3, CO_2 and CCl_4 have equivalent surrounding atoms arranged so the dipoles cancel. NH_3 and H_2O have shapes in which they do not.

A tetrahedral shape alone does not guarantee cancellation. In CH_2Cl_2 the surrounding atoms are different, so its bond dipoles do not balance and the molecule is polar.



trigonal planar · 120°

equal bond dipoles cancel



trigonal pyramidal · 107°

bond dipoles do not cancel

Molecule	Shape	Permanent dipole?
BF_3	Symmetrical trigonal planar	No: equal dipoles cancel
CO_2	Symmetrical linear	No: equal dipoles cancel
CCl_4	Symmetrical tetrahedral	No: equal dipoles cancel
H_2O	Bent	Yes: dipoles do not cancel
NH_3	Trigonal pyramidal	Yes: dipoles do not cancel
CH_2Cl_2	Tetrahedral with unlike surrounding atoms	Yes: dipoles do not cancel

EXAM TRAP Tetrahedral does not automatically mean non-polar. Cancellation depends on the sizes and directions of all bond dipoles.

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 can justify polarity using shape and whether dipoles cancel.

Intermolecular forces

24. Identifying all intermolecular forces

SPEC 3.1.3.7

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. Use the specific force name when explaining an attraction.

Force	When present	Attraction
Induced dipole–dipole (van der Waals)	Between all atoms and molecules	Opposite instantaneous and induced partial charges
Permanent dipole–dipole	Also between polar molecules	Opposite permanent partial charges
Hydrogen bonding	H bonded to N/O/F and a neighbouring N/O/F lone pair	δ+ H bonded to N/O/F and a lone pair on N/O/F

EXAM TRAP A molecule with oxygen but no O–H/N–H/F–H may form hydrogen bonds to δ+ H atoms in water while being unable to hydrogen-bond to itself.

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 can identify every applicable force rather than naming only the strongest.

25. Induced dipole–dipole (van der Waals) forces

SPEC 3.1.3.7

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.

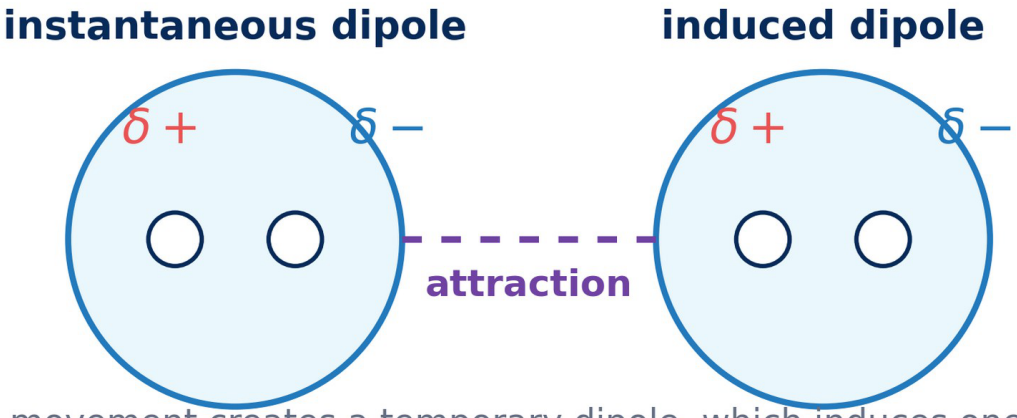
1. This instantaneous dipole distorts the electron cloud of a neighbouring particle.

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.



electron movement creates a temporary dipole, which induces one next door

Factor	Consequence
More electrons / larger electron cloud	Greater polarisability, stronger temporary dipoles
Greater molecular surface contact	More neighbouring electron clouds can interact

WORKED EXAMPLE

Halogen boiling-point trend

Explain $Cl_2 < Br_2 < I_2$ in boiling point.

Worked answer

Electron number and polarisability increase down the group. van der Waals forces strengthen, requiring more energy to separate molecules.

Boiling point increases.

TEST IT

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

DONE WHEN You can explain trends through polarisability or surface contact and energy.

26. Permanent dipole–dipole attractions

SPEC 3.1.3.7

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 force

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



Clδ− is aligned towards Hδ+ on a neighbouring molecule

WORKED EXAMPLE

HCl versus F2

HCl and F_2 have similar relative molecular masses. Explain why HCl has the higher boiling point.

Worked answer

Both have van der Waals forces. HCl is polar and has additional permanent dipole–dipole attractions, increasing the energy needed to separate molecules.

HCl has stronger overall attractions and the higher boiling point.

EXAM TRAP Do not rank boiling points from dipole size alone. A larger non-polar molecule can have stronger overall forces than a smaller polar molecule.

TEST IT

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

DONE WHEN You can explain permanent-dipole effects while also considering van der Waals forces.

Hydrogen bonding & properties

27. Hydrogen bonding: conditions and diagrams

SPEC 3.1.3.7

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.

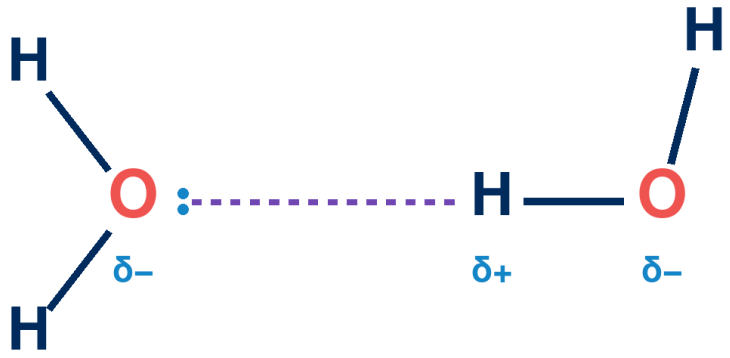
4. Add the appropriate partial charges and keep the hydrogen-bond arrangement clear.

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 attraction between a lone pair on N, O or F and a δ+ H atom covalently bonded to N, O or F in a neighbouring molecule.



O lone pair ··· H—O: one straight line

Substance	Self hydrogen-bonding?	Reason
HF	Yes	H–F bonds and F lone pairs
H_2O	Yes	O–H bonds and O lone pairs
NH_3	Yes	N–H bonds and an N lone pair
CH_3F	No	H is bonded to C, not F
CH_3OCH_3	No	O lone pairs but no H directly bonded to O

WORKED EXAMPLE

Propan-1-ol and methoxyethane

Propan-1-ol and methoxyethane have the same molecular formula. Explain why propan-1-ol has the higher boiling point.

Worked answer

Both have van der Waals and permanent dipole–dipole forces. Propan-1-ol also self hydrogen-bonds through O–H groups; methoxyethane has no H directly bonded to N, O or F.

Stronger overall attractions in propan-1-ol require more energy to overcome.

TEST IT

Try questions 44, 45 in the matching topic Questions & answers pack. Attempt them before checking the worked answers at the back.

DONE WHEN You can identify the required H atom and lone pair and draw the correct intermolecular attraction.

28. Hydrogen bonding, ice and anomalous boiling points

SPEC 3.1.3.7

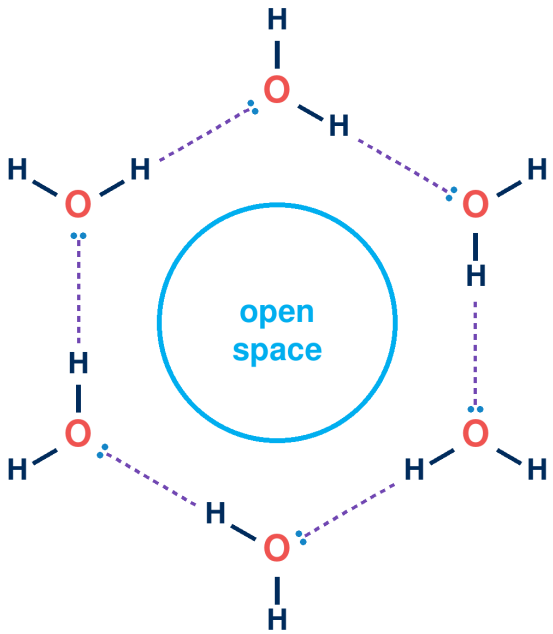
CRACK IT

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.



Comparison	Explanation
NH_3 versus PH_3	NH_3 forms hydrogen bonds; its boiling point is unusually high
HF versus HCl	HF forms hydrogen bonds; more energy is required to separate molecules
H_2O versus other Group 6 hydrides	Water's hydrogen-bonded network raises its boiling point

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 can explain floating ice and anomalous boiling points with a complete force-to-property chain.

29. Boiling, melting and competing intermolecular effects

SPEC 3.1.3.7

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.

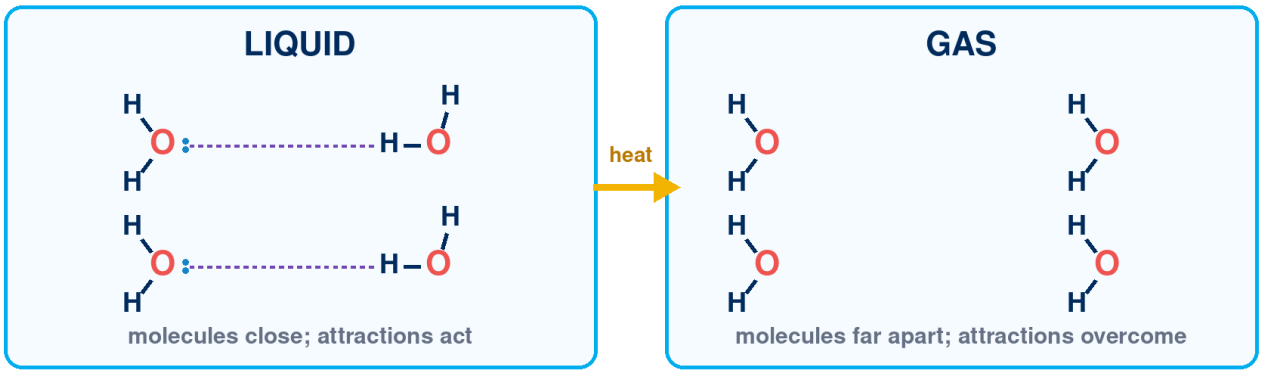
1. Identify the molecules and all the forces between them.

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 → all forces → overall strength → energy required → boiling point

Process	Overcome	Remains intact
Boiling water	Intermolecular hydrogen bonds	O–H covalent bonds
Melting iodine	Intermolecular van der Waals forces	I–I covalent bonds

WORKED EXAMPLE

Ethanol and ethane

Explain why ethanol has a higher boiling point than ethane.

Worked answer

Both have van der Waals forces. Ethanol is polar and also forms hydrogen bonds. Its overall attractions are stronger, so more energy is needed to separate molecules.

Ethanol has the higher boiling point.

EXAM TRAP “Hydrogen bonding beats every other molecule” is not a universal rule. Consider electron count, shape and every force present; melting points also depend on crystal packing.

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 You can evaluate competing effects without claiming that boiling breaks covalent bonds.

30. Solubility: attractions disrupted and formed

SPEC 3.1.3.7

CRACK IT

Dissolving means particles of the solute mix with the solvent. Some attractions within the solute and between solvent particles are disrupted, while new attractions form between solute and solvent.

Name those new attractions. Ethanol can hydrogen-bond with water. Propanone can accept hydrogen bonds from water through its oxygen lone pairs, even though propanone does not hydrogen-bond to itself.

Hexane cannot form hydrogen bonds that adequately replace water-water attractions, so its solubility in water is very low. Iodine mixes more readily with non-polar hexane through favourable van der Waals interactions.

'Like dissolves like' is a useful starting point. An exam explanation needs the actual particles and attractions, rather than the slogan alone.

Pair	Interaction reasoning	Prediction
Ethanol + water	O–H groups and O lone pairs allow hydrogen bonds between the two substances	Mix readily
Hexane + water	Hexane cannot form hydrogen bonds that adequately replace water–water attractions	Very low solubility
Iodine + hexane	Both are non-polar; favourable van der Waals interactions form	Iodine dissolves better than in water
Propanone + water	O lone pairs attract δ+ H atoms in water	Mix readily

EXAM TRAP Oxygen-containing molecules may form hydrogen bonds to δ+ H atoms in water without hydrogen-bonding to themselves. Polar does not automatically mean water-soluble.

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 name attractions disrupted and formed rather than relying on a slogan.