

Chemical bonding

AS

Read online at pastopic.com/cambridge/chemistry-9701/notes/chemical-bonding

Last checked 1 October 2026

Read first: [Atomic structure](#)

© 2026 Pastopic. Free for personal study. Please share the link, not the file.

What you need to know

By the end of this topic you should be able to:

1. **Define electronegativity:** the power of an atom to attract electrons to itself.
2. **Explain what makes electronegativity high or low:** nuclear charge, atomic radius, and shielding by inner shells and sub-shells.
3. **State and explain the trends:** electronegativity increases across a period and decreases down a group.
4. **Use differences in Pauling electronegativity values** (given in the question when you need them) to predict whether two elements form an ionic or a covalent bond.
5. **Define ionic bonding** as the electrostatic attraction between oppositely charged ions (cations and anions), and describe it in sodium chloride, magnesium oxide and calcium fluoride.
6. **Define metallic bonding** as the electrostatic attraction between positive metal ions and delocalised electrons.
7. **Define covalent bonding** as the electrostatic attraction between the nuclei of two atoms and a shared pair of electrons, and describe it in H_2 , O_2 , N_2 , Cl_2 , HCl , CO_2 , NH_3 , CH_4 , C_2H_6 and C_2H_4 .
8. **Understand that Period 3 atoms can expand their octet**, as in SO_2 , PCl_5 and SF_6 .
9. **Describe coordinate (dative covalent) bonding**, including $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4^+$ and the Al_2Cl_6 molecule.
10. **Describe σ bonds (direct overlap) and π bonds (sideways overlap of p orbitals)** in H_2 , C_2H_6 , C_2H_4 , HCN and N_2 , and **use hybridisation** (sp , sp^2 , sp^3).
11. **Define bond energy** (the energy needed to break one mole of a particular covalent bond in the gaseous state) and **bond length** (the distance between the nuclei of two covalently bonded atoms), and use them to compare how reactive covalent molecules are.
12. **State and explain shapes and bond angles with VSEPR theory** for BF_3 , CO_2 , CH_4 , NH_3 , H_2O , SF_6 and PF_5 , and **predict** the shapes and angles of similar molecules and ions.
13. **Describe hydrogen bonding** (only for molecules with N–H or O–H groups, such as ammonia and water) and use it to explain water's unusual properties: high melting and boiling points, high surface tension, and ice being less dense than liquid water.
14. **Use electronegativity to explain bond polarity and the dipole moments of molecules.**

15. **Describe van der Waals' forces** as the general name for all intermolecular forces, of two kinds: instantaneous dipole–induced dipole (id-id, or London dispersion) forces and permanent dipole–permanent dipole (pd-pd) forces. Hydrogen bonding is a special case of pd-pd.
16. **State that ionic, covalent and metallic bonds are generally stronger than intermolecular forces.**
17. **Draw dot-and-cross diagrams** of ionic, covalent and coordinate bonding for all the species above, including expanded octets and species with an odd number of electrons.

Covalent character in ionic compounds is **not** assessed. The names of lattice structures (giant ionic, giant molecular, simple molecular, metallic) belong to States of matter, but questions on this topic use them, so they appear in "How to do it".

Understand it

Electronegativity

Every bond is a tug of war for electrons. **Electronegativity** measures how hard an atom pulls on the electrons in a bond. It depends on how strongly the nucleus attracts the bonding pair:

- **Nuclear charge:** more protons, stronger pull.
- **Atomic radius:** a smaller atom holds the bonding pair closer to the nucleus, stronger pull.
- **Shielding:** inner shells and sub-shells repel the bonding pair and cancel part of the nuclear charge, weaker pull.

Across a period, electronegativity increases (Na 0.9, Mg 1.3, Al 1.5, Si 1.9, P 2.2, S 2.6, Cl 3.0). The nuclear charge rises, the atoms get smaller, and the shielding stays about the same because the electrons go into the same shell.

Down a group, electronegativity decreases (F 4.0, Cl 3.0, Br 2.6; O 3.5, S 2.6). Each atom has an extra shell, so the bonding pair is further away and more shielded. This outweighs the extra nuclear charge.

Predicting the bond type. Look at the **difference** between the two atoms' values. There is no exact cut-off number; use the trend:

- **The larger the difference, the more ionic the bond.** A metal with a low value and a non-metal with a high value: the non-metal takes the electrons completely and ions form. NaCl ($3.0 - 0.9 = 2.1$), MgO ($3.5 - 1.3 = 2.2$) and CaF_2 ($4.0 - 1.0 = 3.0$) are clearly ionic. Smaller differences can still be ionic: KBr ($2.6 - 0.8 = 1.8$) is an ionic solid.
- **The smaller the difference, the more covalent the bond.** The electrons are shared. The bond is **polar** covalent if the difference is not zero (HCl: 0.9), and almost non-polar if it is tiny (C–H: 0.4). Aluminium chloride (1.5) is covalent.
- **Zero** (two atoms of the same element): a pure, non-polar covalent bond.

So use the differences to **compare**: of several compounds, the one with the largest difference is the most ionic and the one with the smallest is the most covalent. If a question tells you a compound is ionic (LiH, difference 1.1, for example), take that as given.

Ionic bonding

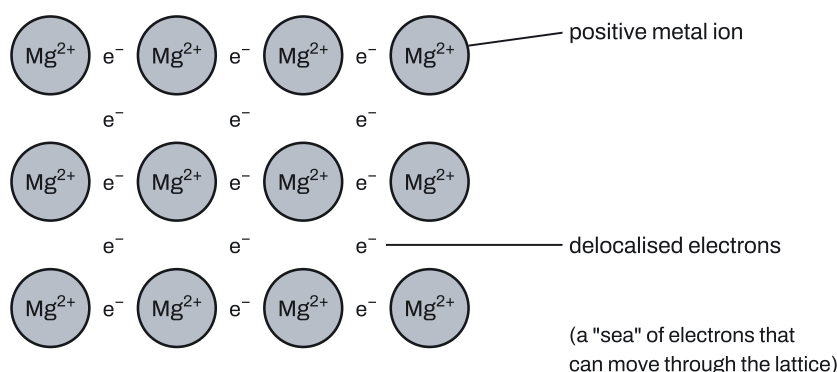
In an **ionic bond**, a metal atom loses its outer electrons to become a positive ion (**cation**) and a non-metal atom gains them to become a negative ion (**anion**). The bond is the **electrostatic attraction between the oppositely charged ions**. It acts in every direction, so each ion attracts many neighbours and the ions pack into a **giant ionic lattice**.

- **Sodium chloride**: Na (2,8,1) gives one electron to Cl (2,8,7): Na^+ and Cl^- , both with full outer shells.
- **Magnesium oxide**: Mg (2,8,2) gives two electrons to O (2,6): Mg^{2+} and O^{2-} .
- **Calcium fluoride**: Ca (2,8,8,2) gives one electron to **each** of two F atoms: Ca^{2+} and 2F^- .

Higher charges and smaller ions attract more strongly, so MgO ($2+$ and $2-$) melts far higher than NaCl ($1+$ and $1-$).

Metallic bonding

In a metal, each atom releases its outer electrons into a shared pool. What is left is a regular lattice of **positive metal ions** in a "**sea**" of **delocalised electrons**, which belong to no single atom and move through the whole structure. **Metallic bonding is the electrostatic attraction between the positive metal ions and the delocalised electrons**.



The bond is stronger (higher melting point) when:

- the ions have a **higher charge**, so there are **more delocalised electrons** per ion (Mg gives two, Na gives one);
- the ions are **smaller**, so the delocalised electrons are closer to the ions' nuclei.

So magnesium melts higher than sodium, and sodium (Na^+ is smaller) melts higher than potassium.

Covalent bonding

A **covalent bond** is a **shared pair of electrons**. Both nuclei attract the same pair, and that electrostatic attraction holds the atoms together. Atoms usually share until each has a full outer shell (8 electrons, or 2 for hydrogen):

Molecule	Bonds	Lone pairs
H ₂ , Cl ₂ , HCl	one single bond	Cl has 3 lone pairs; H has none
O ₂	O=O double bond (2 shared pairs)	2 on each O
N ₂	N≡N triple bond (3 shared pairs)	1 on each N
CO ₂	two C=O double bonds	2 on each O
NH ₃	three N–H	1 on N
CH ₄ , C ₂ H ₆	four single bonds round each C	none
C ₂ H ₄	C=C and four C–H	none

Exceptions to the octet rule.

- **Fewer than 8:** in BF₃ the boron has only **6** outer electrons (3 shared pairs). It has an empty orbital, so it can accept a lone pair.
- **More than 8 (expanded octet):** Period 3 atoms have empty 3d orbitals, so they can use more than four electron pairs. P in PCl₅ has **10** outer electrons (5 bonds); S in SF₆ has **12** (6 bonds); S in SO₂ has **10** (two S=O double bonds and one lone pair). Period 2 atoms (C, N, O, F) can never do this.
- **Odd number of electrons:** NO (11 outer electrons) and NO₂ (17) must have one **unpaired** electron, so they are free radicals.

Coordinate (dative covalent) bonding

In a coordinate bond, **both** electrons of the shared pair come from the **same** atom. One atom needs a **lone pair**; the other needs an **empty orbital** to receive it. It is drawn as an arrow from the donor to the acceptor, for example H₃N → H⁺.

- **Ammonium ion.** In NH₃(g) + HCl(g) → NH₄Cl(s), the N lone pair bonds to H⁺, which has no electrons. Once formed, all four N–H bonds are identical. Solid NH₄Cl has all three bond types: covalent (N–H), coordinate (one of the N–H), and ionic (between NH₄⁺ and Cl[−]).
- **Al₂Cl₆.** In AlCl₃ the Al has only 6 outer electrons. Two AlCl₃ molecules join: a lone pair on a Cl of each molecule forms a coordinate bond to the Al of the other. So Al₂Cl₆ has **two** coordinate bonds, from Cl to Al, and each Al ends up with 8 electrons.

After a coordinate bond forms, the acceptor has one more electron pair around it, so its shape changes: trigonal planar BCl₃ becomes tetrahedral BCl₄[−] (Example 1).

Dot-and-cross diagrams

A dot-and-cross diagram shows where the **outer electrons** are. Electrons from one atom are dots, from the other atoms crosses (or as the question says).

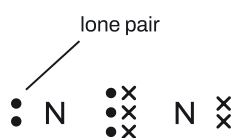


magnesium oxide, MgO

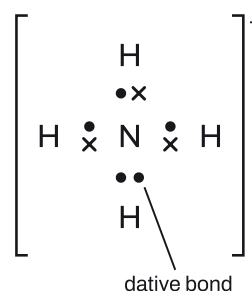
• from O, × from Mg

calcium fluoride, CaF₂

• from F, × from Ca

nitrogen, N₂• from left N,
× from right Ncarbon dioxide, CO₂

• from C, × from O

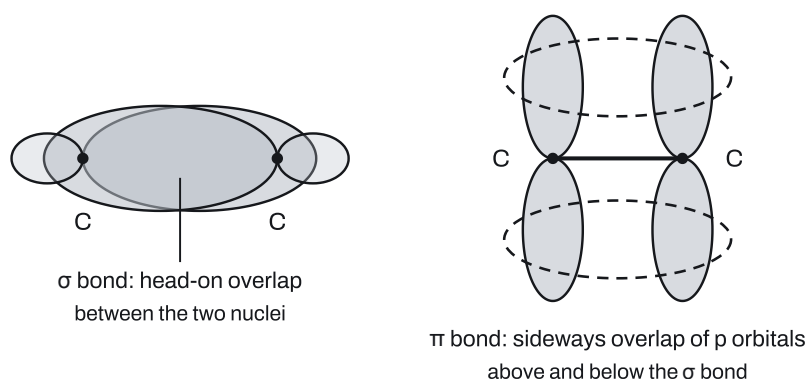
ammonium ion, NH₄⁺

• from N, × from H

Ionic: separate ions in **square brackets** with **charges**, in the right ratio; the transferred electrons show as the other symbol on the anion. Covalent: shared pairs between the atoms and **every lone pair**. Coordinate: a shared pair of **two of the same symbol**. An ion made of several atoms (NH₄⁺, CN⁻) goes in brackets with its charge.

σ and π bonds, and hybridisation

- A **σ (sigma) bond** forms by **direct, head-on overlap** of orbitals **on the line between the two nuclei**. Every single bond is a σ bond.
- A **π (pi) bond** forms by **sideways overlap of two adjacent p orbitals**, with electron density **above and below** the σ bond (not between the nuclei). A double bond is one σ + one π; a triple bond is one σ + **two** π, the two π at right angles to each other.



A π bond is **weaker** than a σ bond (C–C 350 kJ mol⁻¹, but C=C only 610, not 700), because sideways overlap is smaller. Its electrons are also more exposed, above and below the molecule, so reactions of alkenes break the π bond.

Hybridisation. Carbon's 2s and 2p orbitals mix (hybridise) to make new, equal orbitals that point in the directions the bonds need:

Hybrid	Made from	Number	Directions	Left-over p orbitals	Example
sp^3	one s + three p	4	tetrahedral, 109.5°	none	C in CH ₄ , C ₂ H ₆
sp^2	one s + two p	3	trigonal planar, 120°	one (makes one π)	C in C ₂ H ₄
sp	one s + one p	2	linear, 180°	two (make two π)	C and N in HCN; N in N ₂

A quick rule: count the **regions of electron density** round the atom (each bond, single or multiple, counts once; each lone pair counts once). Four regions: sp^3 ; three: sp^2 ; two: sp .

- H₂: two 1s orbitals overlap head-on: one σ bond.
- C₂H₆: each C is sp^3 ; all 7 bonds are σ (sp^3-sp^3 for C–C, sp^3-1s for C–H).
- C₂H₄: each C is sp^2 ; three sp^2 orbitals make σ bonds, and the left-over p orbitals overlap sideways to make the π bond. The molecule is flat.
- HCN: C and N are sp ; σ bonds from head-on overlap, two π bonds from the two pairs of left-over p orbitals. The cyanide ion, CN⁻, has the same C≡N: one σ and two π .
- N₂: each N is sp ; one sp orbital makes the σ bond, the **other holds the lone pair**, and two pairs of p orbitals make two π bonds.

Bond energy and bond length

Bond energy (always positive: breaking bonds takes in energy) and bond length are defined in "What you need to know". A stronger bond is harder to break, so the molecule is less reactive. Shorter bonds are usually stronger, because the shared pair is closer to both nuclei. Values are given in the data section.

- **Down Group 17**, the atoms get larger, so the X–X bond is longer and the orbitals overlap less well: Cl–Cl 242, Br–Br 193, I–I 151 kJ mol⁻¹. The same happens in the hydrogen halides (H–Cl 431, H–Br 366, H–I 299), so HI is the least thermally stable.

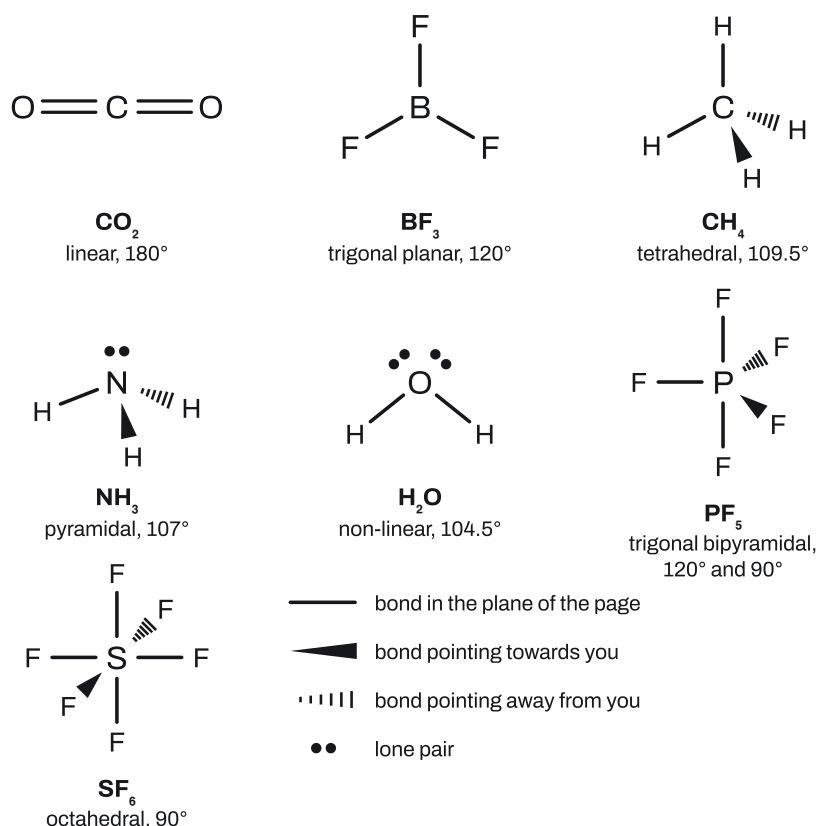
- **Nitrogen is very unreactive** because the $\text{N}\equiv\text{N}$ triple bond is very strong (944 kJ mol^{-1}) and the molecule is non-polar, so nothing is attracted to one end of it.
- **Polarity matters too.** A polar bond has a $\delta+$ end and a $\delta-$ end that attract other reagents, so a polar molecule can be more reactive than its bond energy suggests (carbon monoxide reacts more readily than nitrogen, even though its bond is stronger).

Shapes of molecules: VSEPR

Valence shell electron pair repulsion (VSEPR): the electron pairs around the central atom repel and move as far apart as possible.

The total number of regions (a single, double or triple bond counts as one region, and so does each lone pair) sets the arrangement. **Lone pairs repel more than bonding pairs** (lone pair–lone pair > lone pair–bond pair > bond pair–bond pair), because a lone pair is held by only one nucleus and sits closer to it, so each lone pair squeezes the angles by about 2.5° . The **shape is named from the atoms only**.

Bonding regions	Lone pairs	Shape	Bond angle	Example
2	0	linear	180°	CO_2
3	0	trigonal planar	120°	BF_3
4	0	tetrahedral	109.5°	$\text{CH}_4, \text{NH}_4^+$
3	1	pyramidal	107°	NH_3
2	2	non-linear	104.5°	H_2O
2	1	non-linear	just under 120° (about 118°)	SO_2
5	0	trigonal bipyramidal	120° and 90°	$\text{PF}_5, \text{PCl}_5$
6	0	octahedral	90°	SF_6



Finding the lone pairs. Take the central atom's outer electrons (its group number, or group – 10 for Groups 13–18), subtract one for each electron it uses in bonds (1 per single bond, 2 per double bond), subtract the charge of a positive ion or add the charge of a negative ion, then halve:

$$\text{lone pairs} = \frac{\text{outer electrons} - \text{electrons used in bonds} - \text{charge}}{2}$$

For NH_4^+ : $(5 - 4 - 1) \div 2 = 0$, so tetrahedral. For H_2O : $(6 - 2) \div 2 = 2$, so non-linear.

Analogous molecules. Same group, same shape: SiH_4 is like CH_4 ; PH_3 is like NH_3 ; H_2S is like H_2O ; PCl_6^- and SF_6 are both octahedral.

Polar bonds and polar molecules

When two atoms with different electronegativities share a pair, the pair sits closer to the more electronegative atom. That atom becomes slightly negative (δ^-) and the other slightly positive (δ^+): the bond is **polar** and has a **dipole**. The bigger the difference, the more polar the bond ($\text{O}-\text{H}$, $3.5 - 2.1 = 1.4$, is more polar than $\text{S}-\text{O}$, $3.5 - 2.6 = 0.9$).

A **molecule** has an overall **dipole moment** only if its bond dipoles **don't cancel**:

- **Symmetrical** shapes with identical outer atoms cancel: CO_2 (linear), BF_3 , CH_4 , CF_4 , SF_6 , PF_5 have **no** overall dipole, even though their bonds are polar.
- **Lone pairs** on the central atom or **different outer atoms** break the symmetry: H_2O , NH_3 , SO_2 , CH_3Cl are polar.

Intermolecular forces

Van der Waals' forces is the general name for all forces **between** molecules (as opposed to the bonds inside them). There are two kinds.

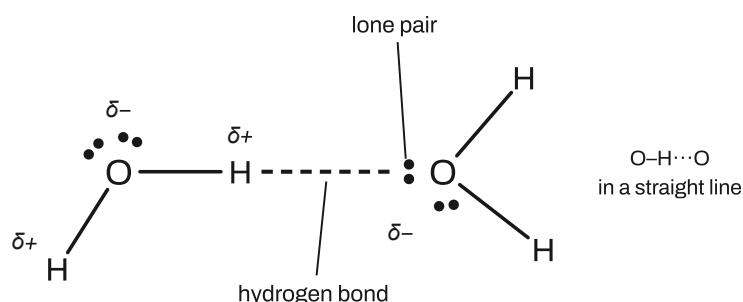
Instantaneous dipole–induced dipole (id-id) forces act between **all** molecules. Moving electrons make a molecule's electron cloud briefly uneven; this temporary dipole induces one in the next molecule, and the two attract. They are stronger when a molecule has **more electrons** (a bigger, more easily distorted cloud) and when molecules are **long and straight** rather than branched, because they touch over a larger area.

Permanent dipole–permanent dipole (pd-pd) forces act between **polar** molecules: the δ^+ end of one attracts the δ^- end of the next. They act **as well as** id-id forces.

Hydrogen bonding is the strongest kind of pd-pd force. It needs:

1. an H atom bonded to N or O (an **N–H or O–H group**), so the H is strongly δ^+ ;
2. a **lone pair** on an N or O atom of a **neighbouring** molecule.

The H is attracted to the lone pair. The three atoms, $X-H\cdots Y$, lie in a straight line (180°).



Water's unusual properties. Each water molecule has **two** H atoms and **two** lone pairs, so it can form up to **four** hydrogen bonds (two on average per molecule).

- **High melting and boiling points.** H_2O boils at $100^\circ C$, but H_2S , with more electrons, is a gas. Many hydrogen bonds must be broken to separate water molecules; H_2S has only weaker pd-pd and id-id forces, because S (2.6) is not electronegative enough.
- **High surface tension.** Hydrogen bonds pull surface molecules strongly inwards and sideways.
- **Ice is less dense than water.** In ice every molecule makes four hydrogen bonds in an open, tetrahedral lattice that holds the molecules further apart than in the liquid. On melting some break and the molecules move closer.

Strength order. Ionic, covalent and metallic bonds are generally much stronger than any intermolecular force. For molecules with similar numbers of electrons, hydrogen bonds > other pd-pd > id-id.

Melting and boiling a simple molecular substance breaks only the intermolecular forces. The covalent bonds inside the molecules stay intact. Giant structures (ionic, metallic, giant covalent such as diamond and SiO_2) melt only when strong bonds break, so their melting points are high.

Key facts and formulas

"Given" means it is printed in the question or the data section when you need it. "Learn it" means you must remember it.

Formula	Status
Pauling electronegativity values (for example H 2.1, C 2.5, N 3.0, O 3.5, F 4.0, Cl 3.0)	Given
Bond energies (for example $\text{N}\equiv\text{N}$ 944, $\text{Cl}-\text{Cl}$ 242, $\text{C}=\text{C}$ 610, $\text{C}-\text{C}$ 350 kJ mol^{-1})	Given
The definitions in "What you need to know" (items 1, 5, 6, 7, 11)	Learn it
Electronegativity rises across a period and falls down a group	Learn it
The larger the electronegativity difference, the more ionic the bond; a smaller difference: polar covalent; zero: non-polar (no exact cut-off)	Learn it
Single bond = 1σ ; double = $1\sigma + 1\pi$; triple = $1\sigma + 2\pi$	Learn it
4 regions: sp^3 ; 3 regions: sp^2 ; 2 regions: sp	Learn it
VSEPR shapes and angles: 180° , 120° , 109.5° , 107° , 104.5° , 120° and 90° , 90° (the table above)	Learn it
Repulsion: lone pair–lone pair > lone pair–bond pair > bond pair–bond pair	Learn it
lone pairs = $\frac{\text{outer electrons} - \text{electrons used in bonds} - \text{charge}}{2}$	Learn it
Hydrogen bond: H of an N–H or O–H group, to a lone pair on N or O; $\text{X}-\text{H}\cdots\text{Y}$ at 180°	Learn it
Strength: ionic, covalent, metallic > hydrogen bond > $\text{pd}-\text{pd}$ > $\text{id}-\text{id}$ (molecules of similar size)	Learn it

How to do it

In the questions we have tagged for this topic (128 questions, 2021–2025, Papers 1 and 2), the types came up this often. 35 of the 128 questions mix more than one type, so the counts add up to 180.

Question type	Questions
Predict and explain a shape and bond angle (VSEPR)	40
Identify intermolecular forces and explain boiling or melting points	36
Identify bonding and structure and explain properties	26
Draw a dot-and-cross diagram	16
Describe σ and π bonds, orbital overlap and hybridisation	13
Identify coordinate bonds and the bond types present	13
Draw a labelled diagram of metallic bonding or a hydrogen bond	10
Define bonding terms; explain and use electronegativity	8
Count electrons, bond pairs, lone pairs or unpaired electrons	8
Use bond energy to explain reactivity	5
Explain bond polarity and dipole moments	5

1. Predict and explain a shape and bond angle (40 questions)

1. Find the **central atom**, count its **bonding regions** (a multiple bond counts once) and work out its **lone pairs** with the formula above.
2. Read the **shape and angle** from the VSEPR table.
3. **Explain:** " n bonding pairs and m lone pairs repel and get as far apart as possible; lone pairs repel more than bonding pairs."

Variations you will meet:

- **Angles inside a bigger molecule** (labelled x, y, z). Treat each labelled atom as a central atom: C with four single bonds 109.5° ; C in C=C or C=O 120° ; C in C $\equiv$ C or C $\equiv$ N 180° ; N with three bonds and a lone pair 107° ; O (or S) with two bonds and two lone pairs 104.5° . In HO-CH₂-C $\equiv$ C-H: C-O-H 104.5° , at the CH₂ carbon 109.5° , at the C $\equiv$ C carbons 180° .
- **A coordinate bond changes the shape.** Recount **both** atoms afterwards: the acceptor gains a bonding pair (120° becomes 109.5° , Example 1); the donor loses a lone pair, so its angles open up (107° becomes 109.5°).
- **Ions:** include the charge. PH₂⁻: $(5 - 2 + 1) \div 2 = 2$ lone pairs, non-linear, about 104.5° . Working backwards: about 107° at an atom with three bonds means one lone pair.
- **Which molecules are planar?** Linear, trigonal planar and non-linear ones, and anything built only of sp² and sp atoms (C₂H₄). Anything tetrahedral, pyramidal or with a CH₃ group is not.
- **Expanded octets:** five bonds, no lone pair: trigonal bipyramidal; six bonds, no lone pair: octahedral.

2. Identify intermolecular forces and explain boiling or melting points (36 questions)

1. For **each** substance list **every** force: id-id always; pd-pd if the molecule is polar; hydrogen bonds only if it has an **N–H** or **O–H** group.
2. Say which substance has the **stronger** forces and why (more electrons; a permanent dipole; hydrogen bonds).
3. Link to the property: stronger forces need **more energy to overcome**, so a higher boiling or melting point.

Variations you will meet:

- **Hydride trends down a group** (H_2O to H_2Te): from S to Te boiling points rise with more electrons (stronger id-id). The first member, H_2O or NH_3 , is anomalously high because of hydrogen bonding.
- **Same number of electrons, different polarity** (Example 2): id-id forces are similar, so the polar one boils higher.
- **Two hydrogen-bonded substances**: compare the strength of each hydrogen bond (O is more electronegative than N) **and** how many form per molecule on average (the smaller of the number of H on N or O and the number of lone pairs on N or O).
- **Isomers**: a straight chain has more surface contact than a branched one, so stronger id-id forces.
- **"What is broken on melting?"** Simple molecular: intermolecular forces (the strongest present); giant covalent: covalent bonds; ionic: ionic bonds.

3. Identify bonding and structure and explain properties (26 questions)

1. **Use the data**: conducts solid and liquid, **metallic**; conducts only molten or in solution, **giant ionic**; never conducts and very high melting point, **giant covalent** (SiO_2 , diamond; graphite is the exception and conducts); never conducts and low melting point, **simple molecular**.
2. **Explain conductivity by mobile charged particles**: delocalised electrons **move through** the structure; ions are **free to move** when molten but fixed in the solid. A covalent liquid such as SiCl_4 has **no ions and no delocalised electrons**.
3. **Explain melting point** by what is broken and how strong it is (the last paragraph of "Understand it").

Variations you will meet:

- **Comparing two metals**: higher ion charge (more delocalised electrons) and a smaller ion give stronger metallic bonding ($\text{Mg} > \text{Na} > \text{K}$).
- **Graphite and diamond**: both giant covalent. Graphite has one delocalised electron per carbon, which moves through the structure; diamond uses all four outer electrons in σ bonds.
- **Electronegativity and structure**: a large difference means ionic bonding, a giant lattice and a high melting point.

4. Draw a dot-and-cross diagram (16 questions)

1. Write each atom's outer electrons; add or remove electrons for the charge.

2. Covalent: shared pairs between atoms (a multiple bond is 2 or 3 pairs), then the remaining electrons as lone pairs. Check every outer electron appears once.
3. Ionic: separate ions in square brackets with charges, in the right ratio. Follow the key the question gives for dots and crosses.

Variations you will meet: a coordinate pair inside a molecule (CO , Al_2Cl_6 , NH_4^+), drawn as two of the donor's symbols; an ion's extra electron (CN^- , SCN^-) in a different symbol; all three bond types together (NH_4Cl : $[\text{NH}_4]^+$ with its dative pair and a separate $[\text{Cl}]^-$); expanded octets (PCl_5).

5. Describe σ and π bonds, orbital overlap and hybridisation (13 questions)

1. **Name the hybridisation** from the regions round the atom (4: sp^3 , 3: sp^2 , 2: sp).
2. Describe each bond with the wording from "Understand it": σ is head-on overlap between the nuclei, naming the orbitals (sp^2 with sp^2 , sp^3 with $1s$); π is sideways overlap of unhybridised p orbitals above and below the σ bond. Label any diagram.

Variations you will meet:

- **Counting σ and π bonds:** one σ per bond; add one π per double bond and two per triple. Propyne, $\text{CH}_3\text{C}\equiv\text{CH}$: 6 σ , 2 π ; ethanenitrile, $\text{CH}_3\text{C}\equiv\text{N}$: 5 σ , 2 π .
- **Why the π bond breaks in reactions:** less overlap, so it is weaker, and its electrons are more exposed.
- **Sketching orbitals:** s is a sphere; p is a dumbbell along one axis.

6. Identify coordinate bonds and the bond types present (13 questions)

1. Look for an atom with an **extra bond**: N with four (NH_4^+), P with four H (PH_4^+), O with three (H_3O^+), Al with four (Al_2Cl_6).
2. Identify the **donor** (a lone pair: N, O, P, Cl, OH^-) and the **acceptor** (an empty orbital: H^+ , B in BF_3 , Al in AlCl_3 , CH_3^+).
3. Count per formula unit: NH_4^+ has one; Al_2Cl_6 has two.

"Which bond types are present?" Check each part separately. Calcium carbonate: ionic between Ca^{2+} and CO_3^{2-} , covalent inside the carbonate, no coordinate bond. Any ammonium salt: all three. Calcium fluoride: ionic only.

7. Draw a labelled diagram of metallic bonding or a hydrogen bond (10 questions)

Metallic bonding: a regular lattice of circles labelled as **positive ions with the right charge** (Mg^{2+}), with e^- between them labelled **"delocalised electrons"** (the third diagram).

Hydrogen bond (the fourth diagram):

1. Draw **two molecules**, including the **O–H or N–H** bond whose H is involved (never an H from a CH_3 group).
2. Mark **δ^+ on the H** and **δ^- on the O or N**.

3. Draw the **lone pair** on the O or N of the second molecule.
4. Draw a **dashed line from the H to that lone pair**, in line with the O–H or N–H bond, and **label it**.

8. Define bonding terms; explain and use electronegativity (8 questions)

1. Use the definitions in "What you need to know", with their key words (**electrostatic attraction**, **delocalised electrons**, **shared pair**, **power of an atom to attract electrons**).
2. **Explain a trend** with all three factors, as in "Understand it": nuclear charge, atomic radius and shielding, and say which one wins.
3. **Predict bond type** from the difference: the larger it is, the more ionic the bond (a metal with a much more electronegative non-metal, as in NaCl or CaF_2); the smallest difference gives the most covalent compound.

9. Count electrons, bond pairs, lone pairs or unpaired electrons (8 questions)

1. Draw (or picture) the dot-and-cross diagram.
2. Count **bond pairs** (a triple bond is 3) and **lone pairs** on **every** atom.

Variations you will meet:

- H_2O : 2 bond pairs, 2 lone pairs; CO_2 : 4 bond pairs, 4 lone pairs; C_2H_4 : 6 bond pairs, none.
- **Incomplete outer shell**: a Group 13 atom with three bonds, or a carbocation carbon (6 electrons each).
- **Total valence electrons**: add each atom's and adjust for charge (CH_3OH : $4 + 4 \times 1 + 6 = 14$). An odd total means an unpaired electron (ClO_2 : $7 + 2 \times 6 = 19$).

10. Use bond energy to explain reactivity (5 questions)

Compare the bond energies (weaker bond, more reactive) and explain the size: **larger atoms** → **longer bond** → **less effective orbital overlap** → **weaker bond** (Si–Si 225 is weaker than C–C 350). Add **polarity** where relevant, as for N_2 and CO in "Understand it".

11. Explain bond polarity and dipole moments (5 questions)

Mark the more electronegative atom δ^- (a bigger difference, a more polar bond), then decide from the shape whether the dipoles cancel, as in "Polar bonds and polar molecules": CF_4 has no dipole; CH_2F_2 is polar.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

Boron trichloride, BCl_3 , is a covalent molecule.

(a) Draw a dot-and-cross diagram of a BCl_3 molecule. Show outer electrons only. **[2]**

(b) State the shape of a BCl_3 molecule and its Cl-B-Cl bond angle. Explain your answer. **[3]**

(c) BCl_3 reacts with a chloride ion to form the BCl_4^- ion. Name the type of bond that forms, explain how it forms, and give the shape and bond angle of BCl_4^- . **[3]**

(d) State the total number of lone pairs in a BCl_4^- ion. **[1]**

(a) B in the middle, bonded to three Cl. Each B–Cl bond is one pair: one dot (from B) and one cross (from Cl).

B1 Each Cl has three more pairs of crosses (lone pairs); B has no lone pair, only 6 outer electrons. **B1**

(b) Trigonal planar **B1**, 120° **B1**. Boron has three bonding pairs and no lone pairs; they repel equally and get as far apart as possible. **B1**

(c) A coordinate (dative covalent) bond. **B1** The Cl^- ion donates a lone pair into the empty orbital on boron (B has only 6 outer electrons). **B1** BCl_4^- has four bonding pairs and no lone pairs on B: tetrahedral, 109.5° . **B1**

(d) Each Cl has 3 lone pairs and B has none: $4 \times 3 = 12$. **B1** (Check: $3 + 4 \times 7 + 1 = 32$ electrons, 16 pairs.)

Example 2

The table shows three compounds. Each molecule has 18 electrons.

Compound	Formula	Boiling point / $^\circ\text{C}$
ethane	CH_3CH_3	–89
fluoromethane	CH_3F	–78
methylamine	CH_3NH_2	–6

(a) Name all the types of intermolecular force in each compound. **[3]**

(b) Explain the order of the boiling points. **[3]**

(c) State the H–N–H bond angle in methylamine. Explain your answer. **[2]**

(a) Ethane: id-id only. **B1** Fluoromethane: id-id and pd-pd. **B1** Methylamine: id-id, pd-pd and hydrogen bonds. **B1**

(b) All three have 18 electrons, so their id-id forces are similar. **M1** The C–F bond is polar (C 2.5, F 4.0), so fluoromethane also has pd-pd forces and needs more energy to separate than ethane. **M1** Methylamine has an N–H group, so it also forms hydrogen bonds, the strongest forces here, and boils highest. **M1**

(c) About 107°. **B1** The N atom has three bonding pairs and one lone pair; the lone pair repels more than the bonding pairs, so the angle is squeezed below 109.5°. **B1**

Example 3

The table shows the properties of three solids, W, X and Y.

Solid	Melting point / °C	Conducts when solid?	Conducts when molten?
W	1263	no	yes
X	660	yes	yes
Y	114	no	no

(a) Name the type of structure and bonding in each solid. **[3]**

(b) Explain why X conducts electricity when solid but W conducts only when molten. **[2]**

(c) Explain why Y has a low melting point. **[2]**

(d) X is aluminium. Sodium melts at 98 °C. Explain why aluminium has the higher melting point. **[2]**

(a) W: giant ionic (ionic bonding). **B1** X: giant metallic (metallic bonding). **B1** Y: simple molecular (covalent bonds within molecules). **B1**

(b) X has delocalised electrons that can move through the structure and carry charge. **B1** In W the ions are held in fixed positions in the solid lattice; when molten they are free to move. **B1**

(c) Melting Y breaks only the weak intermolecular (van der Waals') forces between the molecules **M1**; the covalent bonds inside the molecules are not broken, so little energy is needed. **M1**

(d) Aluminium forms Al^{3+} ions and gives three delocalised electrons per atom, while sodium forms Na^+ and gives one. **M1** Al^{3+} is also smaller, so the electrostatic attraction between the cations and the delocalised electrons is stronger in aluminium and needs more energy to overcome. **M1**

Example 4

Propenenitrile has the structure $\text{CH}_2=\text{CH}-\text{C}\equiv\text{N}$.

(a) State the number of σ bonds and the number of π bonds in one molecule. **[2]**

(b) State the hybridisation of each carbon atom, numbering them C1 (CH_2), C2 (CH) and C3 (the C of $\text{C}\equiv\text{N}$). Give the $\text{H}-\text{C1}-\text{H}$ and $\text{C2}-\text{C3}-\text{N}$ bond angles. [3]

(c) Describe, in terms of orbital overlap, how the bonds in $\text{C}\equiv\text{N}$ are formed. State where the lone pair on the nitrogen atom is. [3]

(d) Bond energies: $\text{C}-\text{C}$ 350 kJ mol^{-1} and $\text{C}=\text{C}$ 610 kJ mol^{-1} . Use them to estimate the energy of the π bond in $\text{C}=\text{C}$, and explain why the $\text{C}=\text{C}$ bond energy is less than twice the $\text{C}-\text{C}$ bond energy. [2]

(a) Bonds: three $\text{C}-\text{H}$, one $\text{C}=\text{C}$, one $\text{C}-\text{C}$ and one $\text{C}\equiv\text{N}$. Each bond has one σ : **6 σ** . **B1** $\text{C}=\text{C}$ has one π and $\text{C}\equiv\text{N}$ two: **3 π** . **B1**

(b) C1 and C2 each have three regions: **sp^2** . C3 has two regions: **sp** . **B1** $\text{H}-\text{C1}-\text{H}$: 120° **B1**. $\text{C2}-\text{C3}-\text{N}$: 180° **B1**.

(c) σ : head-on overlap of an sp orbital on C3 with an sp orbital on N, between the nuclei. **M1** Two π : sideways overlap of the two pairs of unhybridised p orbitals, at right angles to each other. **M1** The N lone pair is in N's other sp orbital, pointing away from C3. **B1**

(d) π bond $\approx 610 - 350 = 260 \text{ kJ mol}^{-1}$, less than the σ bond (350). **B1** A double bond is one σ plus one π , not two σ bonds, and sideways overlap is smaller than head-on overlap, so $610 < 2 \times 350 = 700$. **B1**

Common mistakes

- **Using a bond angle that ignores lone pairs.** A central atom with two bonds is not always linear: SCl_2 -type molecules also have two lone pairs, so they are non-linear, 104.5° . Count the lone pairs first.
- **Naming the shape from the electron pairs.** NH_3 has four pairs, but its shape is **pyramidal**, not tetrahedral.
- **Drawing a hydrogen bond to the wrong H.** Only an H on N or O can hydrogen bond, never one from a CH_3 group.
- **Listing only the strongest intermolecular force.** Every molecule has id-id forces as well.
- **Saying boiling breaks covalent bonds.** Boiling water breaks hydrogen bonds, not O-H bonds.
- **"Free electrons" or "sea of electrons" on its own.** Say **delocalised** electrons, and that they **move through the structure**. Draw lattice ions with the right charge (Mg^{2+} , not Mg).
- **Coordinate bonds the wrong way round** in Al_2Cl_6 (from Cl to Al), and **missing lone pairs** on Cl atoms.
- **Forgetting an ion's charge or brackets** in a dot-and-cross diagram, such as the extra electron in SCN^- .
- **Miscounting σ bonds.** Every single bond, including each C-H, is a σ bond, and each multiple bond has exactly one.
- **Saying a molecule is polar without saying why.** Give the electronegativity difference and whether the dipoles cancel.

How to get the marks

- **Definitions** are marked on key phrases, so learn the wording in "What you need to know" (electrostatic attraction, delocalised electrons, shared pair, power to attract electrons).
- **"State the shape"** needs the name; **"predict the bond angle"** needs a number (107° , not "less than 109.5° "). **"Explain"** a shape needs the numbers of bonding and lone pairs, and that lone pairs repel more.
- **A difference in boiling points** needs the forces in **each** substance, which is **stronger** and why, and **"more energy"** needed to overcome.
- **Properties from structure**: name the particles that move for conductivity and what is broken for melting point, for **both** substances.
- **Orbital overlap**: "head-on, between the nuclei" for σ and "sideways overlap of p orbitals, above and below" for π ; "the orbitals overlap" alone earns little.
- **Dot-and-cross diagrams**: every outer electron, lone pairs on every atom, brackets and charges on ions, and the key followed.
- **Labelled diagrams** (hydrogen bond, metallic lattice): the marks are for the labels, dipoles and lone pairs as much as the drawing.
- **"Suggest"** means apply the ideas to an unfamiliar species, by analogy or from its data.
- **Hypothesis questions** ("true, false, or not enough information?"): give a verdict **and** justify it from the data.

Quick check

1. Bromine (2.6) is less electronegative than chlorine (3.0). Explain why. Then use the values K 0.8, Al 1.5 and Br 2.6 to find the electronegativity difference in potassium bromide and in aluminium bromide, and state which of the two has the more ionic bonding.
2. Predict the shape and bond angle of: (a) a gaseous BeCl_2 molecule; (b) NF_3 ; (c) the AsF_6^- ion.
3. Which of these molecules have an overall dipole moment: SiF_4 , PF_3 , BeCl_2 , CHF_3 ? Explain.
4. The bond energies of C–Cl and C–I are 340 and 240 kJ mol^{-1} . Explain the difference in terms of bond length, and state which bond is broken more easily when a halogenoalkane reacts.

Answers

1. Br has an extra shell, so its atomic radius is larger and its bonding pair is more shielded from the nucleus. This outweighs its larger nuclear charge, so it attracts the bonding pair less strongly. KBr: $2.6 - 0.8 = 1.8$; AlBr_3 : $2.6 - 1.5 = 1.1$. KBr has the larger difference, so its bonding is more ionic (it is an ionic solid); aluminium bromide, with the smaller difference, is (polar) covalent.
2. (a) Be has 2 outer electrons, used in two bonds: 2 bonding pairs, 0 lone pairs, linear, 180° . (b) N: 3 bonding pairs and $(5 - 3) \div 2 = 1$ lone pair, pyramidal, about 107° . (c) As: $5 + 1 = 6$ electrons, all used in six bonds: $(5 - 6 + 1) \div 2 = 0$ lone pairs, octahedral, 90° .
3. PF_3 and CHF_3 . SiF_4 (tetrahedral) and BeCl_2 (linear) are symmetrical with identical outer atoms, so their bond dipoles cancel. PF_3 is pyramidal (a lone pair on P), so the three P–F dipoles don't cancel; in CHF_3 the C–H bond is much less polar than the C–F bonds, so they don't cancel.
4. The I atom is much larger than the Cl atom, so the C–I bond is longer; the bonding pair is further from the nuclei and the orbitals overlap less, so the bond is weaker (by $340 - 240 = 100 \text{ kJ mol}^{-1}$). The C–I bond needs less energy, so it breaks more easily.

Past questions to try

- **9701/24/O/N/25 Q1**: parts (b) and (d), halogen bond strengths, the shape of NI_3 and why liquid MgCl_2 conducts but SiCl_4 does not.
- **9701/21/M/J/24 Q5**: defining a covalent bond, then σ and π bonds in $\text{C}=\text{C}$ with labelled overlap diagrams.
- **9701/23/M/J/24 Q4**: part (a), ionic, covalent and coordinate bonds in NH_4Cl .
- **9701/22/M/J/25 Q1**: graphite and diamond, and the strongest force broken when five oxides melt (parts (a) and (d)).