

Unit 2

AP Chemistry

Compound Structure and Properties

Bonds · Energy curves · Ionic solids · Alloys · Lewis diagrams · Formal charge · Shapes

Unit 1

7–9%
Atomic structure

Unit 2

7–9%
Compound structure

Unit 3

18–22%
Substances and mixtures

Unit 4

7–9%
Chemical reactions

Unit 5

7–9%
Kinetics

Unit 6

7–9%
Thermochemistry

Unit 7

7–9%
Equilibrium

Unit 8

11–15%
Acids and bases

Unit 9

7–9%
Thermodynamics, electrochemistry

Unit 2 is 7–9% of the multiple-choice section. On the free-response section, most years ask you to draw or finish a Lewis diagram and then give its shape, a bond angle, the hybridization or the formal charges.

2.1 Bond types



2.2 Potential energy curves



2.3 Ionic solids



2.4 Metals and alloys



2.5 Lewis diagrams



2.6 Resonance and formal charge



2.7 VSEPR and hybridization



Electron domains

0 lone pairs

1 lone pair

2 lone pairs

3 lone pairs

2 Linear 180° sp



An **electron domain** is a bond or a lone pair on the central atom. A double or triple bond points at one neighbor, so it counts as one domain: O=C=O and H-C≡N each have two domains, and both are linear. Each lone pair (yellow lobe) is one domain too.

3 Trigonal planar 120° sp²



Bent just under 120° SO₂ 119.5°



A lone pair, held by one nucleus, spreads wider than a bond and squeezes the bonds together. A single unpaired electron is also a domain: NO₂ (three domains) is bent, NO₂⁺ (two) is linear.

4 Tetrahedral 109.5° sp³



Trigonal pyramidal below 109.5° NH₃ 106.7°



Bent below the NH₃ angle H₂O 104.5°



The name of the shape describes only where the atoms are. NH₃ has four domains, one of them a lone pair, so it is **trigonal pyramidal**.

5 Trigonal bipyramidal 90°, 120°, 180°



Seesaw just under 90°, 120°, 180° SF₄



T-shaped just under 90° ClF₃



Linear 180° XeF₂, I₃⁻



6 Octahedral 90°, 180°



Square pyramidal just under 90° BrF₅



Square planar 90°, 180° XeF₄



For five or six domains, the exam asks only for the **shape**, not the hybridization.

Each row's name is its electron geometry, the way all the domains point. An angle after a formula is a measured value.

What the exam asks most

- A complete Lewis diagram.** Count the valence electrons, adjusting for any charge. Show every bond and lone pair, with 8 electrons around C, N, O and F. (2.5)
- The shape and the hybridization.** Count the domains on the central atom and find the shape in the grid above. Two, three and four domains mean sp, sp² and sp³. (2.7)
- Why two bond angles differ.** More domains mean a smaller angle; with equal domains, more lone pairs squeeze it further. (2.7)
- The better of two Lewis diagrams.** Find every atom's formal charge in both diagrams; the one closer to all zeros is better. (2.6)
- A potential energy curve.** The distance at the lowest point is the bond length; its depth below zero is the bond energy. (2.2)
- Explaining a property with particles.** Solid salts don't conduct because their ions are locked in place; metals do because their electrons move. Compare radii to classify an alloy. (2.3) (2.4)

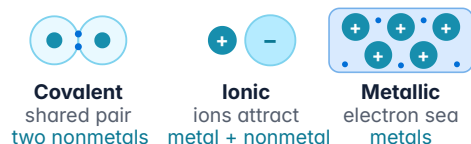
Not on the equation sheet: know these by heart

For this unit, the equation sheet gives only Coulomb's law, $F \propto q_1q_2/r^2$ (q = charges, r = distance between centers).

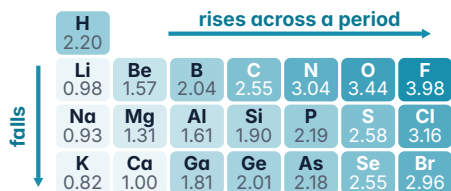
- Formal charge** = valence electrons – nonbonding electrons – $\frac{1}{2}$ bonding electrons.
- Bonds:** a double bond is one σ and one π bond; a triple bond is one σ and two π . Each extra bond between two atoms makes the bond shorter and stronger.
- Electronegativity** increases across a period and decreases down a group. Fluorine has the highest.

2.1 Types of chemical bonds

The type of bond depends on where the valence electrons go.



Electronegativity is how strongly an atom pulls on shared electrons



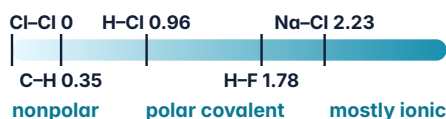
Darker is higher. Ga and Ge break the downward trend slightly. The exam's periodic table lists none of these values.

Explaining the trend

Across a period: F and O both share electrons in their second shell, but F has one more proton, so it pulls a shared pair harder.

Down a group: Cl shares electrons in its third shell, farther from its nucleus, so it pulls less than F. Saying only "top right corner" gets no credit.

Polar and nonpolar bonds



The bigger the electronegativity difference, the more polar the bond. The more electronegative atom gets a partial negative charge (δ^-): in H-Cl, Cl is δ^- and H is δ^+ .

Worked example

Which bond in CBrClF_2 is the most polar: C-Br, C-Cl or C-F?

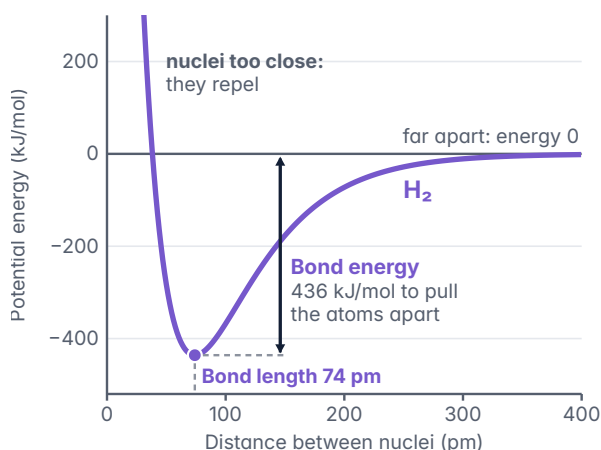
- Subtract C's 2.55 from each: C-F 1.43, C-Cl 0.61, C-Br 0.41.

C-F. Without values, say F is more electronegative than Cl and Br.

Exam trap

Treating "over 1.7 means ionic" as a rule. No number separates ionic from covalent, so go by the elements: H-F differs by 1.78, but both are nonmetals, so HF is covalent.

2.2 Intramolecular force and potential energy



A bond is an intramolecular force: it holds atoms together inside a molecule. Each nucleus attracts the shared electrons, and the two nuclei repel each other. The atoms settle where these balance, at the lowest point on the curve.

Reading the curve

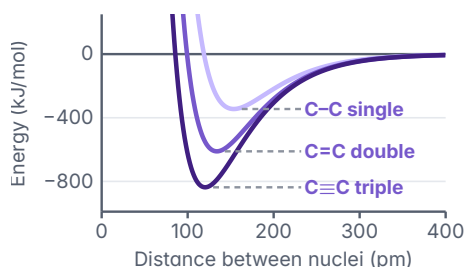
- The distance at the lowest point is the **bond length**.
- The depth of the lowest point below zero is the **bond energy**, written as a positive number: $0 - (-436) = 436 \text{ kJ/mol}$.
- Breaking a bond **takes in** that much energy, and forming it **releases** the same amount.

Exam trap

Writing the bond energy as -436 kJ/mol . Bond energy is the energy needed to pull the atoms apart, so it is positive: 436 kJ/mol .

Reading the bond length where the steep left side crosses zero (about 40 pm here), where the nuclei already repel. It is at the lowest point: 74 pm.

Double and triple bonds are shorter and stronger



Bond	Order	Length	Energy
C-C	1	154 pm	345 kJ/mol
C=C	2	134 pm	611 kJ/mol
C≡C	3	120 pm	837 kJ/mol

Each extra shared pair pulls the two nuclei closer and holds them more tightly. Bond energies are averages over many compounds.

Bigger atoms make longer bonds

H-F	92 pm
H-Cl	127 pm
H-Br	141 pm
H-I	161 pm

Why H-Br is longer than H-F

Br has two more occupied electron shells than F, so the shared pair sits farther from the Br nucleus. The H-Br bond is longer.

Charge and distance in ionic compounds

Bigger charges, or closer ion centers, mean a stronger attraction. In Coulomb's law, r is the distance between the ion centers: about the cation's radius plus the anion's. Radii in pm, to scale:



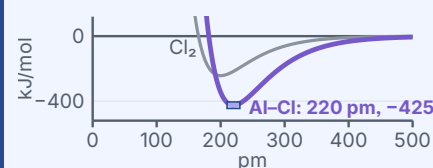
Worked example

Lattice enthalpy (energy to separate a solid into ions): Na_2O 2481, Rb_2O 2163 kJ/mol. Explain using Coulomb's law.

- Both have $+1$ and 2^- ions, so only the sizes differ: Rb^+ 152 pm, Na^+ 102 pm.

Rb_2O 's ion centers are farther apart, so less energy separates its ions.

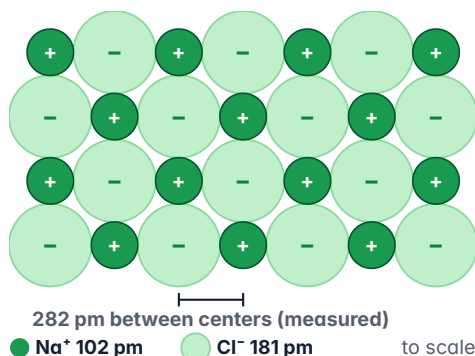
Drawing the curve for a new bond



Given the Cl_2 curve, sketch Al-Cl: bond length 220 pm, bond energy 425 kJ/mol (lowest point at -425). One point was for the distance (within 10 pm), one for the depth (within 20 kJ/mol).

2.3 Structure of ionic solids

An ionic solid is a lattice: a huge 3-D grid of cations and anions with no separate molecules. Each ion is surrounded by ions of the opposite charge, so attractions are strong and like charges stay apart.



✎ Drawing an ionic solid for full credit

- 1 Surround each ion with opposite charges, so no two like charges touch.
- 2 Keep the formula's ratio: CaCl₂ needs two Cl⁻ for every Ca²⁺.
- 3 Draw the sizes correctly (table below) and label the charges.
- 4 Draw no lines between ions: there is no "NaCl molecule" to draw.

Which ion is bigger? Why

A cation is smaller than its atom	It has lost electrons, often its whole outer shell.
An anion is larger than its atom	Its extra electrons repel each other, with no extra protons to hold them.
Ions get bigger down a group	Each has one more occupied shell.
Among ions with the same number of electrons, fewer protons means bigger	Fewer protons hold the electrons less tightly: O ²⁻ 140 > F ⁻ 133 > Na ⁺ 102 > Mg ²⁺ 72 pm.

🔍 Quick check

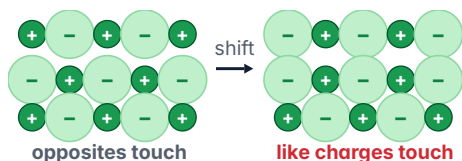
Sketch solid Li₂O. How many of each ion, and which do you draw larger?

Answer: two Li⁺ for every O²⁻, each touching only opposite charges. Draw O²⁻ (140 pm) much larger than Li⁺ (76 pm).

⊘ Not on the exam

The names of crystal structures, such as rock salt or cesium chloride.

Why salt shatters



Shift one layer by one ion and like charges meet. They repel and the crystal splits, so salts are hard but brittle.

State	Ions	Conducts?
Solid	locked in place	no
Melted	free to move and carry current	yes
Dissolved	free to move and carry current	yes

📊 Worked example

In a particle diagram of MgO, which ion is drawn larger, and why?

- 1 Both have 10 electrons; Mg²⁺ has 12 protons, O²⁻ only 8.

O²⁻, at 140 pm to the 72 pm of Mg²⁺: fewer protons pull on the same 10 electrons.

⚠ Exam trap

Drawing Na-Cl pairs joined by lines. An ionic solid has no molecules: each ion is attracted to all the oppositely charged ions around it. Draw no lines between ions.

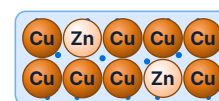
2.4 Structure of metals and alloys

A metal is a lattice of **positive ions** (atoms without their valence electrons) in a **sea of valence electrons** that move through the whole piece.

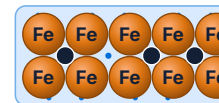


Property	Because
Conducts as a solid	Its valence electrons move freely.
Malleable and ductile (hammers into sheets, draws into wire)	Layers slide and the electron sea follows, so it bends instead of breaking.

Two kinds of alloy



Substitutional
 brass: Zn (142 pm) replaces Cu (145 pm)



Interstitial
 steel: C (67 pm) fits between Fe (156 pm)

When two metals have similar radii, one takes the other's place in the lattice, making a **substitutional** alloy. A much smaller atom fits into the gaps instead, making an **interstitial** alloy. The tiny blue dots are the electron sea; the black dots in steel are C atoms.

✎ What alloying changes

- The small atoms in an **interstitial alloy** get in the way when layers try to slide, so the metal is harder and less malleable and ductile. That is why steel is harder than pure iron.
- **Both kinds still conduct**, because the electron sea is still there.

🔗 Classify an alloy from radii

Sterling silver: Ag 165 pm, Cu 145 pm.

The radii of Ag and Cu are **similar**, so Cu atoms **replace** Ag atoms in the lattice, making a substitutional alloy. Cu is too large to fit in the gaps between Ag atoms.

📊 Worked example

An Fe-C steel wire is 0.81% carbon by mass. About how many Fe atoms are there per C atom?

- 1 Take 100 g of steel: it holds $0.81 \div 12.01 = 0.0674$ mol C and $99.19 \div 55.85 = 1.776$ mol Fe.
- 2 Divide: $1.776 \div 0.0674 = 26$

About 26 Fe atoms for every C atom. Carbon is under 1% of the mass but nearly 4% of the atoms, because C atoms are so light.

⚠ Exam trap

Explaining an alloy type with mass.

"Brass is substitutional because Cu and Zn have similar masses" gets no credit: whether an atom fits depends on its size. Write "Cu (145 pm) and Zn (142 pm) have similar radii, so Zn atoms replace Cu atoms."

Drawing a metal as alternating + and - ions. A metal has no anions. Draw Mg as rows of 2+ ions with electrons between them, as in the picture of Mg.

2.5 Lewis diagrams

The five steps, shown for the NO_2^+ ion

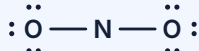
1 Count electrons

16 e^- 5 + 2(6) - 1 = 16
minus 1 for the + charge

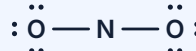
2 Connect the atoms

N in the middle
lower electronegativity

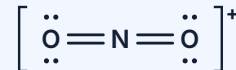
3 Fill the outer atoms

each O gets 6 more
all 16 are used

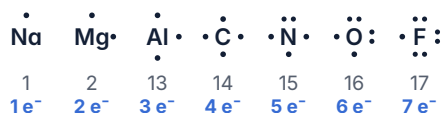
4 Check the center

no electrons left
and N has only 4

5 Make double bonds

each O turns a lone pair
into a second bond to N

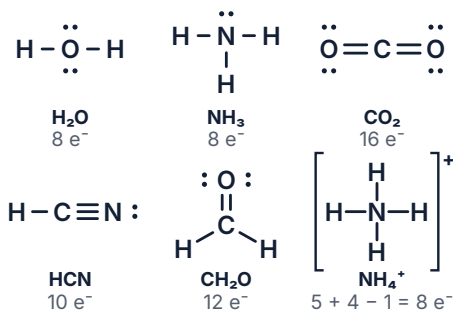
A Lewis diagram shows all the valence electrons: each line is a shared pair (2 electrons) and each dot an unshared electron.



The last digit of the group number gives the valence electrons. Put one dot per side before pairing.

Six Lewis diagrams to know

A Lewis diagram shows which atoms are joined, not the shape: water is drawn straight here, but it is bent.



⚠ Exam trap

Adding an electron for a positive charge. A + charge means the ion has lost an electron, so subtract: NO_2^+ has 5 + 12 - 1 = 16 valence electrons, not 18.

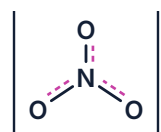
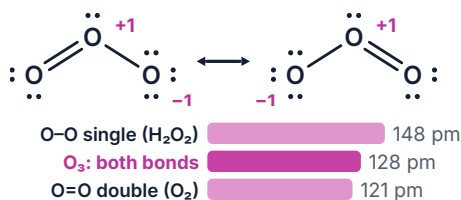
Leaving the lone pair off N in NH_3 . N would have only 6 electrons. Count: 8 valence electrons, the 3 bonds use 6, so the last 2 sit on N as a lone pair.

⚡ Atoms that don't end up with 8

- Fewer than 8:** B and Al can stop at 6, as in BF_3 and AlCl_3 .
- An odd number of electrons:** NO (11) and NO_2 (17) must leave one unpaired.
- More than 8:** only atoms in period 3 or below, such as P, S, Cl and Xe (SF_6 has 12 around S). Never C, N, O or F.

2.6 Resonance and formal charge

When equally good diagrams differ only in **where the electrons sit**, draw them all, with $\leftrightarrow$ between. The real molecule is always a blend: ozone's two O-O bonds are identical, each $1\frac{1}{2}$ bonds, with a length between single and double.



In nitrate, three equal diagrams blend into three identical N-O bonds; the dashes show one extra pair shared by all three. On the exam, draw the three Lewis diagrams with $\leftrightarrow$ between them, not this sketch.

⚠ Exam trap

Thinking the molecule flips between its resonance forms. It has one structure, a blend of the forms, all the time. So write that ozone's two bonds are the same length (128 pm), not that one is single and one double.

📖 Bond order when there is resonance

In one drawing, count the bonds from the central atom (a double bond counts as 2) and divide by the number of outer atoms. O_3 : $3 \div 2 = 1\frac{1}{2}$. NO_3^- and CO_3^{2-} : $4 \div 3 = 1\frac{1}{3}$.

Atom	Shared pairs	Lone pairs	Formal charge
C	4	0	0
N	3	1	0
N	4	0	+1
O	2	2	0
O	1	3	-1
O	3	1	+1
Halogen	1	3	0

📖 Formal charge

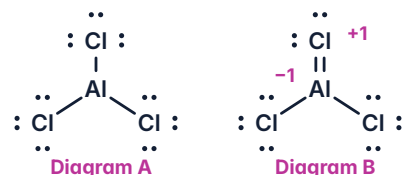
The charge an atom would have if shared pairs were split evenly:

$$\text{FC} = \text{valence } e^- - \text{nonbonding } e^- - \frac{1}{2} \text{ bonding } e^-$$

Or subtract every dot and line on the atom from its valence electrons.

📖 Worked example

Which AlCl_3 diagram is better? Show the formal charges.



A Al: $3 - 0 - \frac{6}{2} = 0$; Cl: $7 - 6 - \frac{2}{2} = 0$

B Al: $3 - 0 - \frac{8}{2} = -1$; double-bonded Cl: $7 - 4 - \frac{4}{2} = +1$

Diagram A: all zeros, even though Al has only 6 electrons.

📖 Picking the best diagram

- It has the right electron count, and no more than 8 around C, N, O or F.
- Its formal charges are as close to zero as possible.
- The better diagram puts any negative formal charge on the more electronegative atom: for SCN^- , -1 on N beats -1 on S.

⚠ Exam trap

Giving one formal charge for the whole molecule. Formal charge belongs to each atom, so " AlCl_3 has 0" loses the point; write "Al 0, each Cl 0."

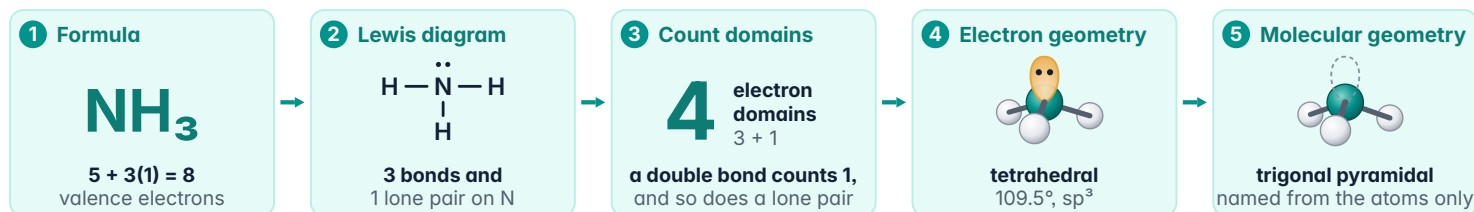
🔍 Quick check

Draw a resonance form of CrO_4^{2-} with formal charge 0 on two O atoms and -1 on the other two.

Answer: two Cr=O and two Cr-O. Cr has 6 valence electrons. O in Cr=O: $6 - 4 - 2 = 0$; O in Cr-O: $6 - 6 - 1 = -1$; Cr: $6 - 0 - 6 = 0$. The total, -2, is the ion's charge.

2.7 VSEPR and hybridization

VSEPR (valence shell electron pair repulsion): domains spread as far apart as they can



Electron geometry is the way all the domains on the central atom point. **Molecular geometry**, the shape, describes only where the atoms are.

Worked example

Predict the shape and the O–S–O angle in SO₂.

- 1 $6 + 2(6) = 18$ electrons: O=S–O or O=S=O, with one lone pair on S
- 2 3 domains on S (2 O + 1 lone pair): electron geometry trigonal planar, sp²

Bent, just under 120° (measured: 119.5°)

Worked example

Predict the shape of XeF₄. Is it polar?

- 1 $8 + 4(7) = 36$ electrons. The four Xe–F bonds and the F lone pairs use 32, so the last 4 form 2 lone pairs on Xe
- 2 6 domains, so the electron geometry is octahedral, with the two lone pairs opposite each other

Square planar. The four Xe–F bond dipoles cancel, so XeF₄ is nonpolar.

Explaining a difference in bond angle

Different numbers of domains: NO₂⁺ has 2 on N, so it is linear (180°). NO₂ has 3, because its unpaired electron counts, so it is bent. Its angle, 134°, is wider than 120° because one electron pushes less than a pair.

Same number, more lone pairs: H₂O and NH₃ both have 4 domains; water's two lone pairs squeeze its bonds more (104.5°; ammonia 106.7°).

Bigger outer atoms: in CBrClF₂, the angle between the large Br and Cl (112.3°) is wider than between the small F atoms (106.8°).

Sigma and pi bonds

Bond	σ	π	Example
Single	1	0	C–H, O–H
Double	1	1	C=O, C=C
Triple	1	2	C≡N, N≡N

The first bond between two atoms is a σ bond, on the line between the nuclei; each extra bond is a π bond, above and below it. A π bond is weaker, so C=C (611 kJ/mol) is less than twice C–C (345).

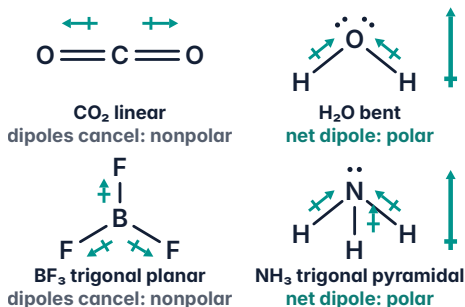
Worked example

In acetonitrile, CH₃–C≡N, count σ and π bonds and give each central atom's hybridization.

- 1 Every single, double or triple bond has one σ: 3 C–H + 1 C–C + 1 C≡N = 5 σ. The triple bond adds 2 π.
- 2 CH₃ carbon: 4 domains, sp³, tetrahedral
- 3 C of C≡N: 2 domains, sp, linear

5 σ and 2 π bonds; the CH₃ carbon is sp³ and the other carbon is sp.

Polar and nonpolar molecules



Each small arrow is a bond dipole, pointing toward the more electronegative atom; its crossed tail is the partly positive end. The bold arrow is the whole molecule's dipole.

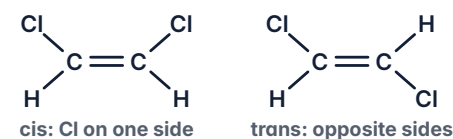
Shape	Polarity
Linear, trigonal planar, tetrahedral, trigonal bipyramidal, square planar, octahedral, with identical outer atoms (as in CCl ₄)	nonpolar: the dipoles cancel
Bent, trigonal pyramidal, seesaw, T-shaped, square pyramidal	polar
Any shape with different outer atoms, such as CHCl ₃	usually polar

Exam trap

Explaining a shape with polarity.

Polarity results from the shape. "NO₂ is bent because it is polar" gets no credit; "because N has three electron domains" does.

A double bond can't rotate



Turning one end of the C=C would break the π bond, so these are two different compounds with different properties.

Exam trap

Counting shared pairs instead of domains. C in CH₂O shares four pairs, but the C=O double bond is one domain, so C has three domains and is sp², not sp³.

Giving a shape when asked for hybridization. For C in CH₄, the answer is "sp³", not "tetrahedral".

Mixing up electron geometry and shape. "Tetrahedral pyramidal" mixes the electron geometry with the shape. Give only the shape: trigonal pyramidal.

Not on the exam

Hybridization for five or six domains, how hybrid orbitals form, and molecular orbital diagrams.

What the free-response questions ask

✓ earns the point ✗ lost the point on past exams

2.1 "Which bond in CBrClF₂ is the most polar?"

- ✓ Name the bond whose atoms differ most in electronegativity, and say why: C–F, because F is more electronegative than Cl and Br.
- ✗ Naming the right bond with no electronegativity comparison. The point needs both.

2.2 "Draw the curve for the new bond"

- ✓ Put the lowest point at the given distance and depth, and bring the curve back up toward zero at large distances.
- ✗ The lowest point at the wrong distance or depth. One question allowed 10 pm and 20 kJ/mol either way; others may differ.

2.2 "Explain why one bond is longer"

- ✓ A higher bond order makes a shorter bond: the P≡P triple bond in P₂ (189 pm) is shorter than the P–P single bonds in P₄ (221 pm). A bigger atom makes a longer bond: H–Br (141 pm) is longer than H–F (92 pm) because Br has two more occupied shells.
- ✗ "Br is bigger", with no reason from its electron shells.

2.2 "Use Coulomb's law to explain why Na₂O has a larger lattice enthalpy than Rb₂O"

- ✓ Both have +1 and 2– ions, but Rb⁺ is larger, so the ion centers in Rb₂O are farther apart, the attraction is weaker and less energy separates the ions.
- ✗ Using one ion's radius as *r*. In Coulomb's law, *r* is the distance between the ion centers.

2.5 "Complete the Lewis diagram"

- ✓ Keep the atoms and bonds you're given and add the missing bonds and lone pairs until the total matches the valence electrons.
- ✗ Redrawing the skeleton, or adding an electron for a positive ion.

2.6 "Use formal charge to justify the better AlCl₃ diagram"

- ✓ Give the formal charge of every atom in each diagram, then compare: all zeros (diagram A on page 4) is better than Al –1 and Cl +1.
- ✗ "The molecule's formal charge is zero", or giving Al +3 and Cl –1 (ion charges) instead.

2.6 "Draw a resonance structure with these formal charges"

- ✓ Keep the same atoms in the same places and the same number of electrons. Move only bonds and lone pairs until each atom's formal charge matches.
- ✗ Changing the number of electrons or moving an atom.

2.7 "Do you agree that the bond angles in NO₂ and NO₂⁺ differ?"

- ✓ Say whether you agree, then explain with both: N has three domains in NO₂ and two in NO₂⁺, so NO₂ is bent and NO₂⁺ is linear.
- ✗ Blaming polarity or "single versus double bonds", or saying a linear molecule has a 120° angle.

2.7 "Which AlCl₃ diagram can't be trigonal planar?"

- ✓ If a diagram puts a lone pair on Al, Al has four domains, so the shape would be trigonal pyramidal, not trigonal planar.
- ✗ Rejecting it only because of the octet or the electron count. The question is about shape, so the answer has to count domains.

Words that cost points

Write	Instead of
unpaired electron	"lone pair"
electron domains (to explain a shape)	"bonds"
more occupied electron shells	"more orbitals"
distance between the ion centers	"the radius"
formal charge on each atom	"the molecule's formal charge"

Where Unit 2 comes back

Unit 3 A molecule's shape and polarity determine its intermolecular forces. Boiling separates whole molecules; covalent bonds stay intact.

Unit 6 Bond energies give ΔH : the energy to break bonds minus the energy released forming new ones.

Unit 8 An acid is stronger when its conjugate base's charge sits on a more electronegative atom or spreads out by resonance.

One question, worked through: NF₃ and BF₃

Both molecules have three F atoms around a central atom.

Part	An answer that earns the point	What loses it
Draw the Lewis diagram of NF ₃ .	5 + 3(7) = 26 electrons: three N–F single bonds, one lone pair on N and three on each F.	Leaving off the lone pair on N, or using the wrong total.
Give the shape of each molecule.	NF ₃ is trigonal pyramidal. BF ₃ is trigonal planar.	"Tetrahedral" for NF ₃ : that names where its electrons point, not the shape.
Explain why the F–N–F angle is smaller than the F–B–F angle.	N has four electron domains (three bonds and a lone pair), so its angles start near 109.5°, and the lone pair squeezes them further. B has three domains and no lone pair, so its angles are 120°.	Describing only one molecule, or saying the angle is smaller because NF ₃ is polar.
Which molecule is polar? Justify your answer.	NF ₃ . Its three bond dipoles point down the sides of a pyramid and don't cancel. In flat BF ₃ the three dipoles cancel.	Calling BF ₃ polar because its bonds are polar.
Find the formal charge of each atom in your NF ₃ diagram.	N: 5 – 2 – 3 = 0. Each F: 7 – 6 – 1 = 0.	Giving one formal charge for the whole molecule.
Give the hybridization of N and of B.	N is sp ³ (four domains). B is sp ² (three domains).	Giving a shape or a count of σ bonds instead.