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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Structure 2.2 The Covalent Model

Booklet 2 · Understandings 2.2.4 – 2.2.7

Molecular shape, bond and molecular polarity, and covalent network structures

Complete revision notes & memory points · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Core material. Every student must know this.

HL Extension material. Understandings 2.2.4–2.2.7 are common to SL and HL. HL panels add the AHL ideas examiners combine with them: shapes with five and six electron domains (2.2.13), hybridization (2.2.16), and resonance and delocalization (2.2.11). See section 1.

Inside this booklet

1. What the syllabus actually asks of you
2. 2.2.4 — Shapes of molecules: the VSEPR model
3. 2.2.5 — Bond polarity and electronegativity
4. 2.2.6 — Molecular polarity
5. 2.2.7 — Covalent network structures
6. Bringing the four together
7. Key points to memorise
8. Check yourself
9. Answers

1 What the syllabus actually asks of you

Booklet 1 taught you to draw the Lewis formula of a molecule. This booklet turns that flat diagram into a three-dimensional shape, uses the shape and the electronegativities of the atoms to decide whether the molecule is polar, and then looks at substances in which covalent bonds never stop: the giant networks of carbon and silicon. Every question in 2.2.4–2.2.6 starts from a correct Lewis formula, so if you are unsure of one, go back to Booklet 1 first.

Understanding	What it says	What you must be able to do	Level
2.2.4 VSEPR model	The valence shell electron pair repulsion (VSEPR) model predicts the shape of a molecule from the number of electron domains around a central atom.	Predict the **electron-domain geometry** and the **molecular geometry** for species with up to four electron domains. Predict how lone pairs and multiple bonds change bond angles.	SL+HL
2.2.5 Bond polarity	Bond polarity results from the difference in electronegativity of the bonded atoms.	Deduce whether a covalent bond is polar from electronegativity values (data booklet). Show bond dipoles with partial charges (δ+, δ−) or with vectors.	SL+HL
2.2.6 Molecular polarity	Molecular polarity depends on both bond polarity and molecular geometry.	Deduce the net dipole moment of a molecule or ion from its bond polarities and its shape, including examples where bond dipoles cancel and where they do not.	SL+HL
2.2.7 Covalent networks	Carbon and silicon form covalent network structures.	Describe the structures and explain the properties of silicon, silicon dioxide and the carbon allotropes diamond, graphite, fullerenes and graphene.	SL+HL

The SL / HL question, answered properly

All four understandings are examined at both levels. HL papers extend them in three predictable places, each flagged with an HL panel:

- **Five and six electron domains** — trigonal bipyramidal and octahedral geometries and their lone-pair variants (seesaw, T-shaped, square pyramidal, square planar) → *AHL 2.2.13*. At SL, no central atom will have more than four domains.
- **Hybridization** — sp^3, sp^2 and sp, which match the four-, three- and two-domain geometries and explain the bonding in diamond, graphite and graphene → *AHL 2.2.16*.
- **Resonance and delocalization** — why all the bonds in CO_3^{2-} or NO_3^- are equal, and why graphite conducts → *AHL 2.2.11*.

Linking questions and Nature of Science attached to 2.2.4 – 2.2.7

- **Nature of Science:** how useful is the VSEPR model at predicting molecular geometry? → section 2.6
- **Structure 2.1:** what properties of ionic compounds might be expected in compounds with polar covalent bonding? → section 3.4
- HL **Structure 3.2:** what features of a molecule make it "infrared (IR) active"? → section 4.4
- **Structure 3.1:** why are silicon–silicon bonds generally weaker than carbon–carbon bonds? → section 5.6

What the data booklet gives you — and what it does not

- **Given:** electronegativity values (**Section 9**); covalent bond lengths (Section 11) and bond enthalpies (Section 12), which you need for the silicon–carbon comparisons in section 5.
- **Not given:** any shape, any bond angle, or which molecules are polar. Bond angles such as 109.5°, 107° and 104.5° must be memorised.

A warning about worksheets and older textbooks. Many UK resources quote older Pauling values, for example H 2.1, C 2.5, O 3.5 and Cl 3.0. The IB data booklet uses **H 2.2, C 2.6, O 3.4 and Cl 3.2**. The conclusions rarely change, but in an IB exam always use the booklet's values. This booklet does throughout.

2 Structure 2.2.4 — Shapes of molecules: the VSEPR model SL + HL

2.1 Electron domains

Around the central atom of a molecule there are regions where electron density is concentrated. Each one is called an **electron domain**. The crucial counting rule is that a domain is a *region*, not a pair:

- a lone pair counts as **one** domain;
- a single bond counts as **one** domain;
- a double bond counts as **one** domain;
- a triple bond counts as **one** domain.

So CO_2, with two double bonds and no lone pairs on carbon, has **two** electron domains, not four. Getting this count right is half of every shape question.

This is why the name "electron *pair* repulsion" is a slight over-simplification. The four or six electrons of a multiple bond are held in the same region between two atoms and act as a single unit, so the model is more accurately a theory of electron *domain* repulsion.

The VSEPR model in two sentences — learn them

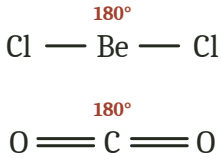
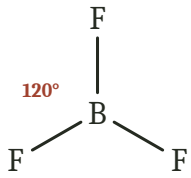
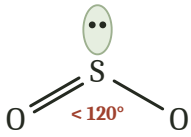
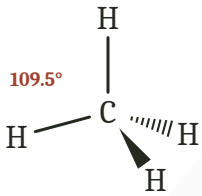
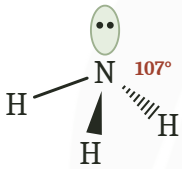
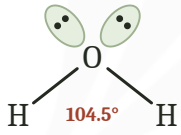
Electron domains around a central atom are regions of negative charge, so they **repel one another** and arrange themselves as **far apart as possible**, which minimises repulsion. The arrangement of the domains gives the **electron-domain geometry**; the arrangement of the **atoms** alone gives the **molecular geometry**, which is the shape you name.

The two geometries are the same when every domain is a bond. They differ as soon as there is a lone pair. The lone pair still occupies a position and pushes the bonds away, but it is invisible when you name the shape, because the shape describes where the atoms are.

2.2 The method

1. **Draw the Lewis formula** (Booklet 1), including every lone pair on the central atom.
2. **Count the electron domains** on the central atom: bonds (single, double or triple, one each) plus lone pairs.
3. **Write the electron-domain geometry**: 2 → linear, 3 → triangular planar, 4 → tetrahedral.
4. **Name the molecular geometry** from the positions of the atoms only (table in section 2.3).
5. **Adjust the bond angle**: each lone pair, and to a smaller extent each multiple bond, squeezes the remaining bond angles below the ideal value.

2.3 The six shapes you must know at SL

Domains	Bonding / lone	Electron-domain geometry	Shape	Molecular geometry	Angle	Examples
2	2 / 0	linear		**linear**	180°	$BeCl_2$, CO_2, HCN, C_2H_2 (each C)
3	3 / 0	triangular planar		**triangular planar**	120°	BF_3, BCl_3, CO_3^{2-}, NO_3^-, C_2H_4 (each C)
	2 / 1	triangular planar		**bent** (V-shaped)	< 120° ≈ 117° in O_3, ≈ 119° in SO_2	SO_2, O_3, NO_2^-
4	4 / 0	tetrahedral		**tetrahedral**	109.5°	CH_4, CCl_4, $SiCl_4$, NH_4^+, BF_4^-
	3 / 1	tetrahedral		**trigonal pyramidal**	107°	NH_3, PCl_3, H_3O^+
	2 / 2	tetrahedral		**bent** (V-shaped)	104.5°	H_2O, H_2S, OF_2, NH_2^-

A note on names. The IB guide and the Pearson text use **triangular planar**; many other books say *trigonal planar*. Both mean the same shape and both earn the mark. This booklet uses the IB term. Note that *trigonal pyramidal* keeps the word "trigonal". For five domains (HL) Pearson uses *triangular bipyramidal*, also widely called *trigonal bipyramidal*; both are accepted.

How to read the 3D drawings: a plain line lies in the plane of the paper, a solid wedge points towards you, and a hashed wedge points away from you. Shaded lobes are lone pairs. You are expected to draw shapes this way in Paper 2, and to show lone pairs on the central atom.

2.4 Why lone pairs and multiple bonds squeeze bond angles

A lone pair is held by only one nucleus, so it sits closer to the central atom and spreads out more widely than a bonding pair, which is pulled out between two nuclei. A lone pair therefore repels more strongly. The order of repulsion is:

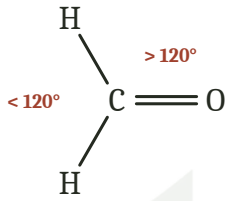
lone pair – lone pair > lone pair – bonding pair > bonding pair – bonding pair

Put more simply, the strength of repulsion from a single domain falls in the order **non-bonding pair > multiple bond > single bond**. A lone pair has a higher concentration of charge because it is not shared between two atoms; a multiple bond has a higher concentration of charge because it contains two or three pairs.

This explains the classic sequence for four domains. Each lone pair that replaces a bond reduces the angle between the remaining bonds by roughly 2–2.5°:

Molecule	CH_4	NH_3	H_2O
Bonding pairs / lone pairs	4 / 0	3 / 1	2 / 2
Electron-domain geometry	tetrahedral	tetrahedral	tetrahedral
Molecular geometry	tetrahedral	trigonal pyramidal	bent
Bond angle	**109.5°**	**107°**	**104.5°**

A **multiple bond** has a higher electron density than a single bond (four or six electrons rather than two), so it too repels more strongly. In methanal, HCHO, carbon has three domains and the geometry is triangular planar, but the C=O domain pushes the two C–H bonds together. The H–C–H angle closes to slightly **less** than 120° (about 116°), and the two H–C=O angles open to slightly **more** than 120°. Ethene shows the same effect: its H–C–H angle is about 117°.



Methanal: the double bond takes more room.

What you actually have to know about distorted angles

You do **not** need to memorise exact distorted angles such as 116° or 119°. You **do** need to state whether an angle will be slightly **more** or slightly **less** than the ideal VSEPR value (180°, 120° or 109.5°) and explain why. The exceptions are the three classic values — **109.5°, 107° and 104.5°** — which are quoted so often that you should know them.

Model answer: "Explain why the bond angle in water is smaller than in methane." (3 marks)

Both molecules have four electron domains around the central atom, so the electron-domain geometry is tetrahedral in each **[1]**. Oxygen in water has two lone pairs and two bonding pairs, whereas carbon in methane has four bonding pairs **[1]**. Lone pairs repel more strongly than bonding pairs, so the O–H bonds are pushed closer together, reducing the angle from 109.5° to 104.5° **[1]**.

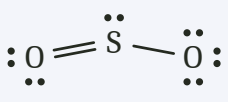
2.5 Worked examples

Worked example — sulfur dioxide, SO_2

Lewis formula: 6 + (2 × 6) = 18 valence electrons. Following the octet rule, sulfur forms one double bond and one single bond to oxygen and keeps **one lone pair**.

Domains on S: double bond (1) + single bond (1) + lone pair (1) = **3** → electron-domain geometry **triangular planar**.

Shape: two atoms and one lone pair → **bent**, with an angle **slightly less than 120°** because the lone pair repels more strongly.

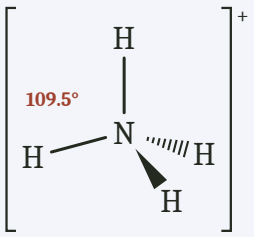


SO_2 (octet structure)

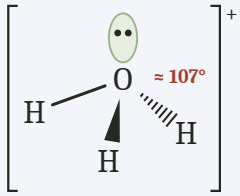
HL An expanded-octet structure, O=S=O with one lone pair on S, is also drawn at HL and is favoured by formal charge. It too has three domains, so the predicted shape is the same either way.

Worked example — ions: the ammonium ion, NH_4^+, and the hydronium ion, H_3O^+

NH_4^+: 5 + 4 − 1 = 8 electrons, four N–H bonds, no lone pair → 4 domains, all bonding → **tetrahedral, 109.5°**. Once formed, the coordination bond is identical to the others, so the ion is a perfect tetrahedron.



H_3O^+: 6 + 3 − 1 = 8 electrons, three O–H bonds and one lone pair → 4 domains → **trigonal pyramidal, ≈ 107°**. It has the same shape as NH_3, with which it is isoelectronic.



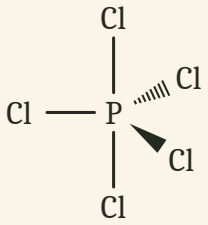
Worked example — molecules with more than one central atom: ethanoic acid, CH_3COOH

Apply VSEPR to **each** central atom separately. The methyl carbon has four single bonds, so it is tetrahedral (109.5°). The carboxyl carbon has three domains (C–C, C=O and C–O), so it is triangular planar (≈ 120°). The O of the O–H group has two bonds and two lone pairs, so it is bent (≈ 104.5°). The same approach describes ethene (each C triangular planar), ethyne (each C linear) and every organic molecule you meet in Structure 3.2.

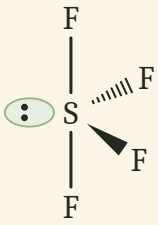
HL

Five and six electron domains (links to AHL 2.2.13 — expanded octets)

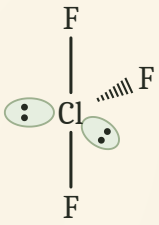
Central atoms from period 3 onwards can hold five or six electron domains. The domains adopt a **trigonal bipyramidal** or an **octahedral** arrangement. In the trigonal bipyramid, lone pairs always occupy the **equatorial** positions, where they have more room (only two neighbours at 90°, compared with three for an axial position). In the octahedron, a second lone pair goes **opposite** the first.



PCl_5 · trigonal bipyramidal
5 / 0 · 90°, 120°, 180°



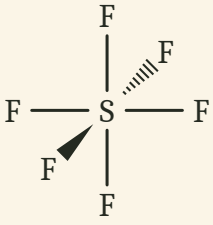
SF_4 · seesaw
4 / 1 · < 90°, < 120°



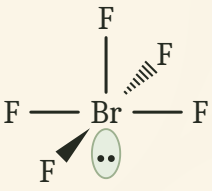
ClF_3 · T-shaped
3 / 2 · < 90°



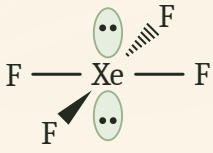
XeF_2 · linear
2 / 3 · 180°



SF_6 · octahedral
6 / 0 · 90°



BrF_5 · square pyramidal
5 / 1 · < 90°



XeF_4 · square planar
4 / 2 · 90°

Numbers under each name are **bonding / lone pairs** and the bond angles. Count domains exactly as at SL: each lone pair and each bond, of any order, is one domain.

Hybridization (AHL 2.2.16) maps directly onto the SL geometries: 4 domains ↔ sp^3 (tetrahedral), 3 domains ↔ sp^2 (triangular planar), 2 domains ↔ sp (linear). Work out the number of domains first and the hybridization follows.

2.6 How useful is the VSEPR model?

Nature of Science — how useful is the VSEPR model at predicting molecular geometry?

VSEPR is a simple, powerful model. With nothing more than a count of electron domains it predicts shapes that are confirmed by X-ray crystallography and spectroscopy. It is still a model, not an explanation. It does not say *why* lone pairs repel more, it only gives approximate angles, and it fails for most transition element complexes. HL hybridization theory offers a deeper account of the same shapes, and HL resonance (2.2.11) explains some discrepancies — for example, why both bonds in O_3 and SO_2 are the same length. A fair verdict is that VSEPR is **very reliable for the overall shape** but needs extra ideas to explain **small differences in angle** between similar molecules.

Examiner traps in 2.2.4

- Counting a double or triple bond as two or three domains. It is **one**.
- Naming the electron-domain geometry when the question asks for the shape. NH_3 is *trigonal pyramidal*, not tetrahedral; H_2O is *bent*, not tetrahedral.
- Forgetting the lone pairs on the central atom in the Lewis formula, then predicting the wrong shape: "NH_3 is triangular planar".
- Drawing a 3D shape without wedges, or without the lone pairs on the central atom.
- Quoting 109.5° for NH_3 or H_2O. Lone pairs reduce the angle: 107° and 104.5°.
- Explaining a smaller angle with "lone pairs take up more space" alone. The marking point is that lone pairs **repel more strongly** than bonding pairs.

3 Structure 2.2.5 — Bond polarity SL + HL

3.1 Electronegativity

Definition to learn word-for-word

Electronegativity — a measure of the ability of an atom to attract a shared pair of electrons in a covalent bond. The Pauling scale is relative and has no units; fluorine, the most electronegative element, is 4.0.

Group	1	2	13	14	15	16	17
Period 1	H 2.2						
Period 2	Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Period 3	Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
Period 4	K 0.8	Ca 1.0					Br 3.0
Period 5							I 2.7
Period 6	Cs 0.8						

Electronegativity values from the IB data booklet (Section 9). Darker cells are more electronegative.

Across a period, electronegativity increases. Nuclear charge rises while the shielding by inner electrons stays the same, so effective nuclear charge increases and the atomic radius falls. The bonding pair is attracted more strongly.

Down a group, electronegativity decreases. The bonding pair is in a shell further from the nucleus and is more shielded, so it is attracted less strongly, even though the nuclear charge increases.

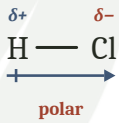
The most electronegative elements (F, O, N, Cl) sit at the top right of the periodic table, and the least (Cs, K, Na) at the bottom left. Noble gases are not usually assigned values.

3.2 Polar and non-polar bonds

When the two atoms in a bond have the **same** electronegativity, as in H–H, Cl–Cl and C–C, the shared pair is shared **equally**. The bond is **non-polar**. When the electronegativities **differ**, the more electronegative atom pulls the shared pair towards itself. It gains a **partial negative charge, δ−**, and the other atom is left with a **partial positive charge, δ+**. The bond is **polar** and has a **bond dipole**.

The guide accepts two notations for the dipole, and you should be able to use both:

- **Partial charges**: δ+ on the less electronegative atom, δ− on the more electronegative atom.
- **A dipole vector**: an arrow pointing **towards the more electronegative atom** (the δ− end). It is often drawn with a small cross at the tail, so the tail looks like a + sign at the δ+ end.



H 2.2, Cl 3.2: $\Delta\chi = 1.0$

3.3 How polar? Using Δχ

The size of the bond dipole increases with the **difference in electronegativity**, $\Delta\chi$, between the two atoms. Bonds form a continuum, from non-polar covalent ($\Delta\chi = 0$), through increasingly polar covalent bonds, to predominantly ionic bonding when $\Delta\chi$ is large. The only truly non-polar bonds are between atoms of the same element, as in H_2, F_2 and O_2; these are sometimes called **pure covalent** bonds. Every other bond has some degree of polarity. The bonding triangle of Structure 2.4 refines this picture.

Bond	C–C	C–H	C–Cl	N–H	C–O	H–Cl	O–H	C–F	H–F
Δχ	0	0.4	0.6	0.8	0.8	1.0	1.2	1.4	1.8
δ− atom	—	C	Cl	N	O	Cl	O	F	F

The C–H bond ($\Delta\chi = 0.4$) is only slightly polar, and in organic chemistry it is usually treated as essentially non-polar. That is why hydrocarbons behave as non-polar substances. H–F has $\Delta\chi = 1.8$ yet is a covalent molecule, so no single cut-off value of $\Delta\chi$ separates "covalent" from "ionic"; the bonding triangle also takes account of the *average* electronegativity of the two atoms.

Worked example — which bond is the most polar: N–H, O–H or F–H? Mark the partial charges.

$\Delta\chi$: N–H = 3.0 − 2.2 = 0.8; O–H = 3.4 − 2.2 = 1.2; F–H = 4.0 − 2.2 = 1.8. The **F–H** bond is the most polar. In every case hydrogen is **δ+** and the other atom is **δ−**. These are exactly the three bonds that make hydrogen bonding possible (2.2.8).

Hydrogen halide	HF	HCl	HBr	HI
Electronegativity of halogen (H = 2.2)	4.0	3.2	3.0	2.7
$\Delta\chi$	**1.8**	1.0	0.8	0.5
Bond polarity	decreases from HF to HI →			

Can oxygen ever be δ+? Only one element is more electronegative than oxygen: fluorine. So in OF_2 (O 3.4, F 4.0) oxygen carries the partial **positive** charge. In every other compound of oxygen it is δ−.

3.4 Polar bonds give covalent substances some ionic character

This answers the Structure 2.1 linking question: *what properties of ionic compounds might be expected in compounds with polar covalent bonding?* The partial separation of charge in a polar bond gives a covalent substance some **ionic character**, and the more polar the bond, the more ionic-like the behaviour. Hydrogen chloride is the standard example. It is a covalent molecular gas, but its H–Cl bond is polar enough that in water it forms $H^+(aq)$ and $Cl^-(aq)$ ions, so the solution **conducts electricity**, which is a property we normally associate with ionic compounds. Polar molecules also tend to dissolve in polar solvents such as water. This is why chemists describe substances as "predominantly" covalent or ionic, or as "strongly" polar, rather than forcing every substance into one box.

	Pure covalent	Polar covalent	Ionic
$\Delta\chi$	zero	small to moderate	large
Electrons	shared equally	shared unequally (partial transfer)	transferred completely
Examples	Cl_2, H_2, O_2	HF, HCl, HBr, HI, H_2O	NaCl, LiF, K_2O
Particles	discrete molecules		lattice of oppositely charged ions

Nature of Science — continuous rather than discrete

Many properties in nature vary continuously rather than in steps, so a spectrum of intermediates exists and classification is difficult. Polar covalent bonds are exactly such intermediates. Recognising them does not weaken the ionic and covalent models; it deepens our understanding of how electrons behave when bonds form.

Examiner traps in 2.2.5

- Confusing electronegativity (an atom in a *bond*) with electron affinity (a gaseous atom gaining an electron).
- Putting the dipole arrow the wrong way round. It points to the **more** electronegative atom, the δ− end.
- Calling a bond polar because "one atom has more electrons". The cause is a **difference in electronegativity**.
- Using older electronegativity values from other books instead of the data booklet.

4 Structure 2.2.6 — Molecular polarity SL + HL

4.1 Bond dipoles add like vectors

A molecule is polar if it has a **net dipole moment**: an overall separation of charge, with one end δ+ and the other δ−. Each bond dipole has a size and a direction, so the net dipole is the **vector sum** of all the bond dipoles. Two things therefore decide molecular polarity, and you must mention **both** in an answer:

1. **Bond polarity** — are there polar bonds at all?
2. **Molecular geometry** — are the bond dipoles arranged **symmetrically**, so that they cancel?

Q1. Does the molecule contain any polar bonds?
No → **non-polar** (H_2, Cl_2, CS_2).

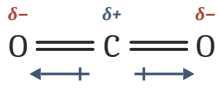
→

Q2. *Yes.* Is the shape symmetrical, with **identical** outer atoms and **no lone pairs** on the central atom?

→

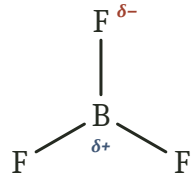
Yes → dipoles cancel → **non-polar** (CO_2, BF_3, CH_4, CCl_4).
No → net dipole → **polar** (H_2O, NH_3, $CHCl_3$, SO_2).

4.2 Gallery: where dipoles cancel and where they do not



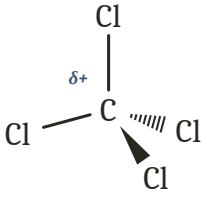
dipoles cancel · non-polar

CO_2 · linear
two equal dipoles at 180°



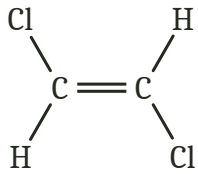
dipoles cancel · non-polar

BF_3 · triangular planar
three equal dipoles at 120°



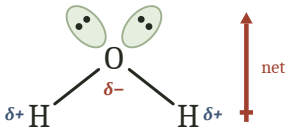
dipoles cancel · non-polar

CCl_4 · tetrahedral
four equal dipoles cancel in 3D



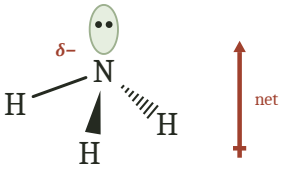
trans · non-polar

trans-1,2-dichloroethene
C–Cl dipoles opposite



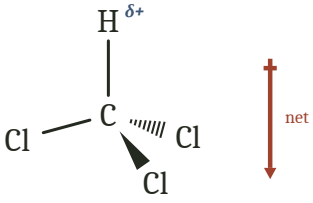
polar

H_2O · bent
dipoles add towards O



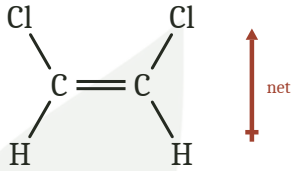
polar

NH_3 · trigonal pyramidal
dipoles add towards N



polar

$CHCl_3$ · tetrahedral
outer atoms not identical

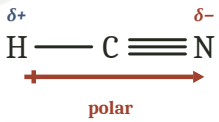


cis · polar

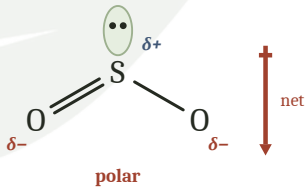
cis-1,2-dichloroethene
C–Cl dipoles on one side

Two of these pairs are worth studying closely. **CCl_4 and $CHCl_3$** have the same tetrahedral shape, but replacing one Cl with H breaks the symmetry, so $CHCl_3$ is polar. **cis- and trans-1,2-dichloroethene** have identical bonds, but in the trans isomer the two C–Cl dipoles point in opposite directions and cancel. This shows clearly that geometry alone can decide polarity.

Linear is not automatically non-polar. HCN is linear, but its two ends are different: H is δ+ and N is δ−, so the dipoles reinforce one another rather than cancel. The same applies to any linear molecule with different end atoms, such as OCS. Likewise **bent** SO_2 is polar while **linear** CO_2 is not: the lone pair on sulfur bends the molecule, and the two S–O dipoles no longer point in opposite directions.



polar



polar

4.3 Evidence and consequences

A thin stream of a polar liquid, such as water, is **deflected** towards a charged rod, whichever the sign of the charge, because the molecules turn so that their oppositely charged end faces the rod. A stream of a non-polar liquid, such as hexane, is not deflected. Molecular polarity is the key to Booklet 3. It decides which intermolecular forces act between molecules (2.2.8), and therefore boiling points, solubility ("like dissolves like") and behaviour in chromatography.

Polarity of ions. The guide asks you to consider ions as well as molecules. The same vector argument applies: NH_4^+, CO_3^{2-} and SO_4^{2-} are symmetrical, so their bond dipoles cancel, while H_3O^+ (trigonal pyramidal) has a net dipole.

HL **Polarity of five- and six-domain molecules**

Apply the same rule. With identical outer atoms and no lone pairs, PCl_5 and SF_6 are **non-polar**. Lone pairs placed symmetrically also cancel: XeF_2 (linear) and XeF_4 (square planar) are **non-polar**. Asymmetric shapes — SF_4 (seesaw), ClF_3 (T-shaped) and BrF_5 (square pyramidal) — are **polar**.

HL **4.4 Polarity and infrared (IR) activity (links to Structure 3.2)**

A molecule absorbs infrared radiation when a vibration **changes its dipole moment**; such a molecule is described as **IR active**. A diatomic molecule has only one vibration, stretching, so it is IR active only if its bond is **polar**: HCl absorbs IR, but N_2, O_2 and H_2 do not. Polyatomic molecules have several modes of vibration, including bending, and a vibration can create a changing dipole even in a molecule with **no overall dipole**. CO_2 and CH_4 are non-polar, yet both are IR active, which is why they act as greenhouse gases. You will use this idea in IR spectroscopy in Structure 3.2.

Model answer: "Explain why carbon dioxide is non-polar although its bonds are polar." (2 marks)

Each C=O bond is polar, because oxygen (3.4) is more electronegative than carbon (2.6) **[1]**. The molecule is linear, so the two equal bond dipoles point in opposite directions and cancel, leaving no net dipole moment **[1]**.

Examiner traps in 2.2.6

- Concluding "polar bonds, so polar molecule". Shape decides whether the dipoles cancel.
- Mentioning only symmetry, or only electronegativity. Full marks need **both**.
- Treating a molecule with different outer atoms ($CHCl_3$, CH_2Cl_2, HCN) as symmetrical.
- Judging symmetry from a flat Lewis formula. CH_2Cl_2 drawn flat with the Cl atoms opposite looks non-polar, but the real molecule is tetrahedral and polar.

5 Structure 2.2.7 — Covalent network structures SL + HL

5.1 Molecules or networks? Know which structure type you are dealing with

Covalent bonding produces two very different kinds of substance. In a **simple molecular** substance, the covalent bonds stop at the edge of each small molecule, and the molecules are held to one another only by weak intermolecular forces. In a **covalent network** (giant covalent) structure, the covalent bonds continue in every direction through the whole crystal. There are no molecules at all, only one enormous lattice of atoms. Being able to sort substances into the five structure types is one of the most useful skills in the whole Structure strand:

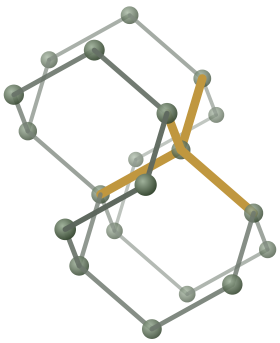
Structure type	Particles	Held together by	Typical substances	Formula represents
Monatomic	atoms	very weak London forces	group 18 elements: He, Ne, Ar …	the atom
Simple molecular	molecules	covalent bonds *within* molecules; weak intermolecular forces *between* them	most non-metal elements (H_2, O_2, P_4, S_8, C_{60}) and compounds of non-metals (H_2O, CO_2, glucose)	one molecule (molecular formula)
Covalent network	atoms	strong covalent bonds throughout	diamond, graphite, graphene, silicon, SiO_2, SiC	the ratio of atoms (empirical formula)
Ionic	ions	electrostatic attraction between ions	most metal + non-metal compounds (NaCl, MgO, $CaCO_3$)	the ratio of ions
Metallic	cations + delocalized electrons	metallic bonding	metals and alloys	the element

Ammonium salts such as NH_4Cl are ionic, even though they contain no metal: the ammonium ion behaves as a cation. Fullerene C_{60} is an allotrope of carbon, but it is **molecular**, not a network. Both appear in exams for exactly this reason.

Definition

Allotropes — different structural forms of the same element in the same physical state. Because their atoms are bonded in different patterns, allotropes have different physical properties and can have different chemical reactivity. Oxygen, O_2, and ozone, O_3, both gases, are allotropes of oxygen. Carbon has four allotropes on the IB course: diamond, graphite, graphene and fullerenes.

5.2 Diamond



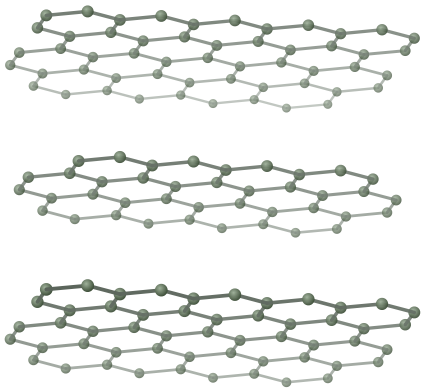
Diamond: each C bonded to four others. One tetrahedral unit is highlighted.

Structure. Every carbon atom is covalently bonded to **four** other carbon atoms, arranged tetrahedrally at **109.5°**. The pattern repeats in three dimensions throughout the crystal. All four valence electrons of each atom are used in localized C–C bonds (154 pm).

- **Very high melting point; extremely hard** — melting or scratching diamond means breaking very many strong covalent bonds in a rigid 3D network.
- **Does not conduct electricity** — every valence electron is held in a bond, so there are no mobile charged particles.
- **Insoluble** in every solvent — no solvent interaction can release energy comparable to breaking the covalent bonds.
- **Excellent thermal conductor** — better than metals — because the rigid lattice passes vibrations efficiently. It is unusual in being a good thermal conductor but an electrical insulator.
- **Transparent and lustrous**, but **brittle**: a sharp blow can split the crystal.

Uses: jewellery; tools and machinery for cutting, drilling and grinding, including cutting glass.

5.3 Graphite



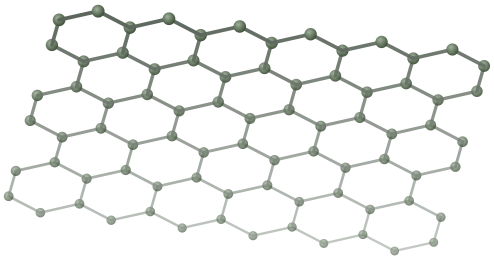
Graphite: hexagonal layers. The layer spacing is drawn wider than scale for clarity.

Structure. Each carbon atom is covalently bonded to **three** others in flat **hexagonal layers**, at **120°**. The fourth valence electron of each atom is **delocalized** across the whole layer. The layers are held together only by weak **London (dispersion) forces**, and they are far apart: 335 pm between layers, compared with 142 pm between bonded atoms within a layer.

- **Conducts electricity** (along the layers) — the delocalized electrons are free to move through each layer.
- **Soft and slippery** — the layers slide over one another because only weak forces hold them together. Graphite is used as a lubricant and in pencils.
- **Very high melting point** — melting still requires breaking the strong covalent bonds *within* the layers.
- **Less dense than diamond** (about 2.2 compared with 3.5 g cm^{-3}) because of the large gaps between layers.
- **Grey, non-lustrous and brittle**; it conducts heat well only along the layers.

Uses: dry lubricant, pencil "lead", electrodes for electrolysis. Graphite is the **most stable** allotrope of carbon: diamond → graphite has $\Delta H = -2$ kJ mol^{-1}, although the conversion is so slow at room temperature that diamonds last indefinitely.

5.4 Graphene



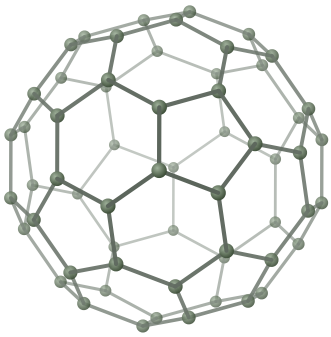
Graphene: a single layer, one atom thick.

Structure. A **single layer** of graphite: a two-dimensional hexagonal sheet just **one atom thick**. Each carbon is bonded to three others at 120°, with one delocalized electron per atom.

- **Exceptionally strong** — about 100 times stronger than steel of the same thickness — because it is a continuous network of strong covalent bonds with no weak links between layers.
- **Excellent conductor** of electricity — delocalized electrons move freely across the sheet — and the best thermal conductor known, better even than diamond.
- **Almost transparent and very flexible**, because it is only one atom thick.
- **Very high melting point** — strong covalent bonds must be broken.

Uses: TEM (transmission electron microscopy) grids, photovoltaic cells, touchscreens and high-performance electronics; many applications are still being developed, because producing large sheets of pure graphene economically is difficult. Graphene rolled into a cylinder gives a **carbon nanotube**; closed into a cage it gives a **fullerene**.

5.5 Fullerenes: buckminsterfullerene, C_{60}



C_{60}: 12 pentagons and 20 hexagons.

Structure. Fullerenes are hollow, closed cages of carbon atoms. In C_{60}, 60 carbon atoms form a sphere of **12 pentagons and 20 hexagons**, like a football. Each carbon is bonded to **three** others, and some electrons are delocalized over the cage. Crucially, C_{60} is a **molecule**: the bonding stops at the surface of each sphere, and neighbouring molecules are held together only by **London forces**.

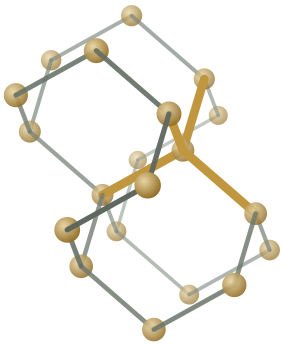
- **Much lower melting and sublimation point** than diamond or graphite — only weak intermolecular forces are overcome, and no covalent bonds break.
- **Soft**, and **soluble in some non-polar solvents** such as methylbenzene — unlike any network structure.
- **Poor electrical conductor** — electrons are delocalized within each molecule but cannot pass easily from one molecule to the next.
- **Very low thermal conductivity**; a black powder in bulk.
- More reactive than diamond or graphite; it readily accepts electrons and reacts with potassium to form a superconducting crystalline solid.

Uses: lubricants, medical and industrial devices that bind specific target molecules; related forms make nanotubes used in electronics (capacitors) and as catalysts.

Nature of Science — discovery through collaboration

Both newer carbon allotropes came from scientists combining different expertise. In 1985 Harold Kroto (University of Sussex), who studied how carbon might form in the atmospheres of giant stars, worked with Robert Curl and Richard Smalley (Rice University, Texas), who used lasers to vaporise materials; vaporising graphite produced stable C_{60} and C_{70} clusters, and the three shared the 1996 Nobel Prize in Chemistry. In 2004 Andre Geim and Konstantin Novoselov (University of Manchester), originally supervisor and research student, isolated graphene by repeatedly peeling graphite flakes apart with adhesive tape. They shared the 2010 Nobel Prize in Physics.

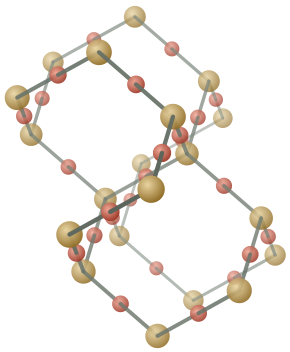
5.6 Silicon and silicon dioxide



Silicon: the diamond structure.

Silicon is in group 14, directly below carbon, and adopts the **same structure as diamond**: each Si atom is covalently bonded to four others tetrahedrally. Silicon atoms are larger, so the Si–Si bond is longer and weaker than C–C. The data booklet gives Si–Si as 232 pm and 226 kJ mol^{-1}, against 154 pm and 346 kJ mol^{-1} for C–C. As a result, silicon has a high melting point (about 1410 °C), but well below that of diamond. Silicon is also a **semiconductor**. Its electrons are held less tightly than in diamond, so a small rise in temperature lets a few of them break free and conduct. Each freed electron leaves behind a positive "hole", which other electrons can move into, so conductivity rises sharply with temperature. Silicon can also be **doped** with tiny amounts of group 13 or group 15 elements (such as boron or arsenic) to increase its conductivity. This is the basis of microchips and solar cells.

Linking question (Structure 3.1): why are Si–Si bonds weaker than C–C bonds? A silicon atom is larger than a carbon atom, because its valence electrons are in the third shell. The shared pair in Si–Si is therefore further from both nuclei and more shielded, so the electrostatic attraction between the nuclei and the shared pair is weaker. This is also why carbon forms long chains and rings (catenation) and multiple bonds, while silicon is more often found bonded to oxygen: Si–O (466 kJ mol^{-1}) is much stronger than Si–Si (226 kJ mol^{-1}).



SiO_2: Si (gold) and O (red).

Silicon dioxide (silica; quartz and sand) is a covalent network in which each **Si atom is bonded to four O atoms** tetrahedrally and each **O atom is bonded to two Si atoms**. This gives the ratio Si : O = 1 : 2, so SiO_2 is an **empirical** formula, not a molecule. The Si–O bonds are strong (466 kJ mol^{-1}, 163 pm in the data booklet) and run throughout the lattice, so silica has a **high melting point** (about 1700 °C), is **hard**, is **insoluble** in water, and **does not conduct** electricity, since it has no mobile charged particles.

The classic comparison: CO_2 is a gas made of small molecules, whereas SiO_2 is a hard solid network. Carbon forms strong C=O double bonds, so one C atom is satisfied by two O atoms in a discrete molecule. Silicon, being larger, does not form strong double bonds with oxygen, so it forms four single Si–O bonds that link into a giant lattice.

5.7 Comparing the structures

Property	Diamond	Graphite	Graphene	C_{60}	Silicon	SiO_2
Type	3D network	layered network	2D network	**molecular**	3D network	3D network
Atoms bonded to each atom	4 C	3 C	3 C	3 C	4 Si	Si: 4 O; O: 2 Si
Geometry / angle	tetrahedral, 109.5°	triangular planar, 120°	triangular planar, 120°	cage (approx. 120° / 108°)	tetrahedral, 109.5°	tetrahedral at Si
Electrical conductivity	none	good (along layers)	excellent	poor	semiconductor	none
Thermal conductivity	very high	good along layers only	highest known	very low	moderate	low
Appearance	transparent, lustrous	grey, non-lustrous	almost transparent	black powder	grey, shiny	colourless (quartz)
Melting point	very high	very high	very high	low (sublimes)	high	high
Hardness	hardest natural substance	soft, slippery	very strong, flexible	soft	hard, brittle	hard
HL Hybridization	sp^3	sp^2	sp^2	sp^2 (approx.)	sp^3	Si sp^3

Model answer: "Explain why graphite conducts electricity but diamond does not." (3 marks)

In graphite each carbon atom forms only three covalent bonds, so one valence electron per atom is delocalized across the layer **[1]**. These delocalized electrons are free to move and carry charge through the structure **[1]**. In diamond each carbon atom forms four covalent bonds, so all its valence electrons are held in localized bonds and none are free to move **[1]**.

HL **Delocalization and hybridization in carbon allotropes (links to AHL 2.2.11 and 2.2.16)**

In diamond each carbon is **sp^3** hybridized and forms four σ bonds. In graphite and graphene each carbon is **sp^2** hybridized: three σ bonds lie in the plane, and the remaining unhybridized p orbital on each atom overlaps sideways with its neighbours, forming a delocalized π system that extends over the whole sheet. This is why graphite's in-layer C–C bond (142 pm) is shorter than a single bond (154 pm) but longer than a double bond (134 pm) — it is the same averaging of bond character that you meet in benzene and in resonance structures.

Examiner traps in 2.2.7

- Saying diamond or graphite has a high melting point because of "strong intermolecular forces". There are **no molecules** in a network; covalent bonds must be broken.
- Saying graphite's layers are held by covalent bonds. They are held by weak London forces, which is why graphite is soft.
- Saying graphite conducts because of "free ions" or "free atoms". It is the **delocalized electrons**.
- Calling C_{60} a giant covalent structure. It is molecular.
- Describing SiO_2 as a molecule, or saying each O bonds to four Si. Each Si bonds to **four** O, and each O to **two** Si.

6 Bringing the four understandings together SL + HL

One chain of reasoning for any molecule

Lewis formula (Booklet 1) → **count domains** on each central atom → **electron-domain geometry** → **molecular geometry** and bond angle → **bond polarities** from $\Delta\chi$ → do the dipoles **cancel**? → **polar or non-polar molecule** → intermolecular forces and physical properties (next booklet).

For a substance rather than a molecule, ask first whether it is a **network** or **molecular**. If it is a network, its properties come from breaking covalent bonds, and polarity and intermolecular forces are irrelevant.

Species	Domains on central atom	Shape	Angle	Polar bonds?	Polar molecule?
CO_2	2	linear	180°	yes	no — symmetrical
HCN	2	linear	180°	yes	**yes** — different ends
BF_3	3	triangular planar	120°	yes	no — symmetrical
SO_2	3	bent	< 120°	yes	**yes**
CH_4	4	tetrahedral	109.5°	barely (C–H)	no
CH_2Cl_2	4	tetrahedral	≈ 109.5°	yes	**yes**
NH_3	4	trigonal pyramidal	107°	yes	**yes**
H_2O	4	bent	104.5°	yes	**yes**

7 Key points to memorise

Everything below is either wording that examiners look for, or knowledge that is *not* in the data booklet. If you learn nothing else from this booklet, learn this section.

A Definitions — near-verbatim

- **Electron domain:** a region of electron density around a central atom — a lone pair, or a single, double or triple bond (each counts as one).
- **VSEPR:** electron domains repel one another and adopt the arrangement that keeps them as far apart as possible.
- **Electron-domain geometry** describes all the domains; **molecular geometry** describes the positions of the atoms only.
- **Electronegativity:** a measure of the ability of an atom to attract a shared pair of electrons in a covalent bond.
- **Polar bond:** a covalent bond between atoms of different electronegativity, so that the electrons are shared unequally (δ+ / δ−).
- **Polar molecule:** a molecule with a net dipole moment, because its bond dipoles do not cancel.
- **Allotropes:** different structural forms of the same element in the same physical state.

B Shapes and angles (not in the data booklet)

Domains	Bond / lone	Shape	Angle	Example
2	2 / 0	linear	180°	CO_2
3	3 / 0	triangular planar	120°	BF_3
3	2 / 1	bent	< 120°	SO_2
4	4 / 0	tetrahedral	109.5°	CH_4
4	3 / 1	trigonal pyramidal	107°	NH_3
4	2 / 2	bent	104.5°	H_2O

Repulsion: LP–LP > LP–BP > BP–BP; a multiple bond repels more than a single bond. HL 5 domains: trigonal bipyramidal, seesaw, T-shaped, linear. 6 domains: octahedral, square pyramidal, square planar.

C Polarity in three lines

- Electronegativity rises across a period and falls down a group; F = 4.0 is the highest.
- Bond polarity comes from Δχ; the dipole arrow points to the δ− (more electronegative) atom.
- Molecular polarity needs polar bonds **and** an unsymmetrical shape (or different outer atoms, or lone pairs on the central atom).

D Networks in one glance

- **Diamond**: 4 bonds, tetrahedral — hard, very high melting point, non-conductor.
- **Graphite**: 3 bonds, layers, delocalized electrons, London forces between layers — conducts, soft, very high melting point.
- **Graphene**: one layer of graphite — very strong, excellent conductor, nearly transparent.
- **C_{60}**: a molecule of 60 atoms — low melting point, soluble in non-polar solvents, poor conductor.
- **Si**: diamond structure, weaker bonds — high melting point, semiconductor. **SiO_2**: Si bonded to 4 O, O to 2 Si — hard, high melting point, insulator.

E Command terms — what each one is actually asking for

- **Predict / Deduce** (a shape) — give the shape *name* and the bond angle; show the domain count if working is asked for.
- **Sketch** (a shape) — a 3D drawing with wedges and lone pairs on the central atom.
- **Explain** (an angle) — domains → repulsion → lone pairs repel more → smaller angle.
- **Explain** (polarity) — bond polarity ($\Delta\chi$) **and** whether the geometry makes the dipoles cancel.
- **Describe** (a structure) — what each atom is bonded to, how many neighbours, the geometry, and how far the bonding extends.

8 Check yourself

Cover the answers on the next pages. Marks are shown in brackets where a question is written in exam style. Questions marked HL are for Higher Level students.

1. State what is meant by an *electron domain*, and state how many domains surround the central atom in (a) CO_2 (b) HCN (c) SO_2 (d) NH_3. **[3]**

2. Predict the electron-domain geometry, the molecular geometry and the bond angle of: (a) PCl_3 (b) H_2S (c) BCl_3 (d) SiH_4 (e) NO_3^- (f) NH_2^- (g) HCN (h) NO_2^+ (i) $SnCl_3^-$

3. Which molecule has the smallest bond angle?
   A CH_4 B NH_3 C H_2O D BF_3

4. Explain why the H–N–H bond angle in NH_3 is smaller than the H–N–H bond angle in NH_4^+. **[2]**

5. In methanal, HCHO, state whether the H–C–H bond angle is more or less than 120°, and explain why. **[2]**

6. State the shape around each carbon atom in propene, $CH_3CH{=}CH_2$.

7. Using the data booklet, place these bonds in order of increasing polarity and identify the δ− atom in each: C–Cl, O–H, C–H, N–H, C–F.

8. Use the data booklet to decide which bond in each pair is more polar: (a) C–H or C–Cl (b) Si–Li or Si–Cl (c) N–Cl or N–Mg. Name a compound in which oxygen carries a partial positive charge.

9. Which molecule is non-polar but contains polar bonds?
   A H_2O B Cl_2 C CCl_4 D $CHCl_3$

10. For each molecule, state whether it is polar or non-polar, with a reason: (a) BCl_3 (b) PCl_3 (c) CS_2 (d) CH_2F_2 (e) OCS (f) Br_2

11. Explain why sulfur dioxide is polar whereas carbon dioxide is not. **[3]**

12. Identify the structure type (monatomic, simple molecular, covalent network, ionic or metallic) of: (a) quartz, SiO_2 (b) iodine, I_2 (c) argon (d) silicon carbide, SiC (e) ammonium nitrate (f) C_{60} (g) copper (h) sucrose (i) graphite (j) phosphorus, P_4

13. Describe the structure of diamond and explain why it has a very high melting point. **[3]**

14. Explain why graphite is used as a lubricant, and why it has a high melting point despite being soft. **[3]**

15. Explain why C_{60} has a much lower melting point than graphite, although both contain only carbon atoms each bonded to three others. **[2]**

16. Describe the bonding in silicon dioxide and explain why it is a solid with a high melting point while carbon dioxide is a gas. **[3]**

17. Silicon has a lower melting point than diamond. Use data from the data booklet to explain why. **[2]**

18. HL Predict the shape and state whether the molecule is polar: (a) SF_4 (b) XeF_4 (c) ClF_3 (d) PF_5

19. HL Explain why carbon dioxide absorbs infrared radiation although it is a non-polar molecule, whereas nitrogen does not. **[2]**

20. HL State the hybridization of carbon in diamond and in graphite, and explain how the hybridization in graphite leads to electrical conductivity. **[3]**

9 Answers

1. A region of electron density around the central atom — a lone pair, or a single, double or triple bond, each counting as one **[1]**. (a) 2 (b) 2 (c) 3 (two bonds + one lone pair) (d) 4 (three bonds + one lone pair) **[2 for all four; 1 for two or three]**.
2. (a) tetrahedral; trigonal pyramidal; ≈ 107° (less than 109.5°) (b) tetrahedral; bent; ≈ 104.5° (less than 109.5°) (c) triangular planar; triangular planar; 120° (d) tetrahedral; tetrahedral; 109.5° (e) triangular planar; triangular planar; 120° (f) tetrahedral; bent; ≈ 104.5° (g) linear; linear; 180° (h) linear; linear; 180° (O=N=O, two domains) (i) tetrahedral; trigonal pyramidal; less than 109.5° (three Sn–Cl bonds + one lone pair).
3. **C.** Water has two lone pairs on the central atom (104.5°).
4. NH_3 has one lone pair and three bonding pairs, whereas NH_4^+ has four bonding pairs and no lone pair **[1]**. The lone pair repels the bonding pairs more strongly, pushing the N–H bonds closer together (107° compared with 109.5°) **[1]**.
5. Less than 120°. The C=O double bond has a greater electron density than the C–H single bonds **[1]**, so it repels them more strongly, pushing the two C–H bonds closer together **[1]**.
6. CH_3 carbon: tetrahedral (four single bonds). Each of the two C=C carbons: triangular planar (three domains).
7. C–H (0.4, δ− C) < C–Cl (0.6, δ− Cl) < N–H (0.8, δ− N) < O–H (1.2, δ− O) < C–F (1.4, δ− F).
8. (a) C–Cl (0.6 vs 0.4) (b) Si–Cl (1.3 vs 0.9) (c) N–Mg (1.7 vs 0.2). Oxygen is δ+ in OF_2, because fluorine is the only element more electronegative than oxygen.
9. **C.** The C–Cl bonds are polar, but the tetrahedral arrangement of four identical bonds is symmetrical, so the dipoles cancel. Cl_2 has no polar bonds at all.
10. (a) non-polar: triangular planar, three identical B–Cl dipoles cancel (b) polar: trigonal pyramidal (lone pair on P), so the dipoles do not cancel (c) non-polar: linear and symmetrical, and C–S has $\Delta\chi = 0$ with booklet values (d) polar: tetrahedral, but with two different kinds of outer atom (e) polar: linear, but the ends are different (O and S) (f) non-polar: identical atoms, $\Delta\chi = 0$.
11. Both contain polar bonds, because O (3.4) is more electronegative than S (2.6) and C (2.6) **[1]**. CO_2 is linear, so its two bond dipoles point in opposite directions and cancel **[1]**. SO_2 has a lone pair on S, which makes it bent, so its bond dipoles do not cancel and there is a net dipole **[1]**.
12. (a) covalent network (b) simple molecular (c) monatomic (d) covalent network (e) ionic (f) simple molecular (g) metallic (h) simple molecular (i) covalent network (j) simple molecular.
13. Each carbon atom is covalently bonded to four other carbon atoms, tetrahedrally **[1]**, in a three-dimensional network extending throughout the crystal **[1]**. Melting requires breaking many strong covalent bonds, which needs a large amount of energy **[1]**.
14. The layers are held together only by weak London forces, so they slide over one another easily **[1]**. Within each layer, every carbon is covalently bonded to three others **[1]**, and melting requires breaking these strong covalent bonds **[1]**.
15. C_{60} consists of separate molecules held together by weak London forces, which need little energy to overcome **[1]**. Melting graphite requires breaking strong covalent bonds throughout its network **[1]**.
16. Each Si atom is covalently bonded to four O atoms, and each O atom to two Si atoms, in a giant three-dimensional network **[1]**. Melting it requires breaking many strong Si–O covalent bonds **[1]**. CO_2 consists of small discrete molecules held by weak intermolecular forces, which are easily overcome **[1]**.
17. Si atoms are larger than C atoms, so the Si–Si bond is longer (232 pm against 154 pm) **[1]** and weaker (226 kJ mol^{-1} against 346 kJ mol^{-1}), so less energy is needed to break the bonds in the network **[1]**.
18. (a) seesaw, polar (b) square planar, non-polar (c) T-shaped, polar (d) trigonal bipyramidal, non-polar.
19. Some vibrations of CO_2 (asymmetric stretching and bending) change the dipole moment of the molecule, so it is IR active **[1]**. N_2 has only one vibration, stretching its non-polar bond, which produces no change in dipole, so it is IR inactive **[1]**.

20. Diamond sp^3; graphite sp^2 **[1]**. In graphite, each carbon has one unhybridized p orbital, which overlaps with those of its neighbours to form a delocalized π system across the layer **[1]**. The delocalized electrons are free to move through the layer and carry charge **[1]**.

Where this booklet leads

- **Structure 2.2 — Booklet 3**: intermolecular forces and physical properties (2.2.8–2.2.9), chromatography (2.2.10), and HL resonance and benzene (2.2.11–2.2.12). Molecular polarity from section 4 is the starting point.
- HL **Structure 2.2 — Booklet 4**: expanded octets, formal charge, σ and π bonds and hybridization (2.2.13–2.2.16).
- **Structure 2.4** — the bonding triangle, which extends the Δχ continuum of section 3.3.
- **Structure 3.2** — shapes around each atom of organic molecules and functional groups.

These notes are a free resource, shared for any IB student who finds them useful.

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