Structure 2: Models of Bonding and Structure

IB SL Study Guide

Contents

IB TIP

Data booklet: You can use the IB Chemistry Data Booklet in the exam — all constants, the periodic table, and key equations are provided.

IB Chemistry SL — Bonding & Structure

Complete Study Guide

Topics Covered

1. Ionic Bonding & Ionic Compounds (S2.1)
2. Covalent Bonding (S2.2)
3. Metallic Bonding (S2.3)
4. VSEPR Theory & Molecular Geometry (S2.2)
5. Electronegativity & Polarity (S2.2)
6. Intermolecular Forces (S2.2)
7. The Bonding Triangle (S2.1–S2.3)

Videos on this page: Ionic Bonding — Ion Formation & Structure · Covalent Bonding, Lewis Structures, VSEPR & Polar Bonds

Prerequisites: This guide assumes familiarity with electron configurations, sublevels (s, p, d), and periodic trends (electronegativity, atomic radius). If you need a refresher, start with **Atomic Structure & Periodicity (Structure 1)** before continuing here.

1. Ionic Bonding & Ionic Compounds (S2.1)

Ionic bonding occurs when one or more electrons are **transferred** from a metal atom to a non-metal atom. The metal loses electrons to form a positive ion (cation) and the non-metal gains electrons to form a negative ion (anion). The oppositely charged ions attract each other through strong electrostatic forces, forming an **ionic bond**. Ionic bonding typically occurs between elements with a large difference in electronegativity — usually a metal from Groups 1 or 2 and a non-metal from Groups 16 or 17.

Electron Transfer and Ion Formation

The driving force for ionic bonding is that both atoms achieve a more stable electron configuration (usually a full outer shell, like a noble gas). For example, in sodium chloride:

- Sodium (Na, Group 1) has 1 valence electron. It loses this electron to form Na^+ (with the electron configuration of neon).

- Chlorine (Cl, Group 17) has 7 valence electrons. It gains 1 electron to form Cl^- (with the electron configuration of argon).

$$\begin{gathered}\mathrm{Na} \rightarrow \mathrm{Na}^+ + e^- \\ \mathrm{Cl} + e^- \rightarrow \mathrm{Cl}^- \\ \mathrm{Na} + \mathrm{Cl} \rightarrow \mathrm{Na}^+\mathrm{Cl}^-\end{gathered}$$

For magnesium oxide, magnesium (Group 2) loses 2 electrons and oxygen (Group 16) gains 2:

$$\begin{gathered}\mathrm{Mg} \rightarrow \mathrm{Mg}^{2+} + 2e^- \\ \mathrm{O} + 2e^- \rightarrow \mathrm{O}^{2-}\end{gathered}$$

Key Ion Charges to Know

Element	Group	Ion Formed	Example
Sodium	1	Na^+	NaCl
Potassium	1	K^+	KBr
Magnesium	2	Mg^{2+}	MgO
Calcium	2	Ca^{2+}	$\mathrm{CaCl_2}$
Fluoride	17	F^-	NaF
Chloride	17	Cl^-	NaCl
Oxide	16	O^{2-}	MgO
Sulfide	16	S^{2-}	$\mathrm{Na_2S}$

MEMORISE THIS

Charge rule from group number:

- Groups 1, 2, 3 → lose electrons → charge = +group number (e.g. Group 2 → 2+)
- Groups 15, 16, 17 → gain electrons → charge = group number − 18 (e.g. Group 17 → 17 − 18 = −1; Group 16 → −2)

The Ionic Lattice

Ions do not exist as simple pairs in a solid ionic compound. Instead, they form a giant **ionic lattice** — a regular three-dimensional arrangement in which each positive ion is surrounded by negative ions and vice versa. In sodium chloride, each Na^+ is surrounded by six Cl^- ions, and each Cl^- is surrounded by six Na^+ ions.

The strong electrostatic attractions between all the ions in the lattice hold it together. This is why we write NaCl as an **empirical formula** (the simplest ratio of ions) rather than a molecular formula — there are no discrete molecules in an ionic solid.

Properties of Ionic Compounds

Property	Explanation
High melting and boiling points	Strong electrostatic forces between many ions require a large amount of energy to overcome
Brittle	Layers of ions can shift so that like charges align; repulsion then shatters the crystal
Do not conduct electricity as solids	Ions are fixed in the lattice and cannot move
Conduct electricity when molten or dissolved in water	Ions are free to move and carry charge
Often soluble in water	Polar water molecules attract and surround ions, pulling them away from the lattice

⚠ **EXAM ALERT**

A very common exam question asks: “Why does solid NaCl not conduct electricity, but molten NaCl does?” The key word is **mobile ions**. In the solid, ions are locked in the lattice. When melted or dissolved, ions become free to move and can carry a current. NEVER say “there are no ions in solid NaCl” — there are plenty of ions, they just cannot move.

WORKED EXAMPLE

Worked Example: Drawing a dot-and-cross diagram for an ionic compound

Draw the formation of $MgCl_2$ showing electron transfer.

Step 1: Identify the atoms and their valence electrons.

- Mg (Group 2): 2 valence electrons
- Cl (Group 17): 7 valence electrons. Two Cl atoms are needed.

Step 2: Mg transfers 1 electron to each Cl atom.

- Mg loses 2 electrons → forms Mg^{2+} (8 electrons remaining, neon configuration)
- Each Cl gains 1 electron → forms Cl^{-} (8 electrons in outer shell, argon configuration)

Step 3: Write the result.

$$\mathrm{Mg} + 2\mathrm{Cl} \rightarrow \mathrm{Mg}^{2+} + 2\mathrm{Cl}^{-}$$

Charge check: Overall charge = (2+) + 2(1−) = 0. The compound is electrically neutral. ✓

▸**Watch: Ionic Bonding — Ion Formation & Structure** VIDEO

2. Covalent Bonding (S2.2)

Covalent bonding occurs when two non-metal atoms **share pairs of electrons**. Neither atom has enough tendency to completely take electrons from the other (both have high electronegativity), so instead they share, and both atoms gain a complete outer shell through the shared electrons. The shared pair of electrons is attracted simultaneously to both nuclei, holding the atoms together.

Single, Double, and Triple Bonds

Atoms can share more than one pair of electrons:

Bond Type	Shared Pairs	Bond Strength	Bond Length	Example
Single bond	1 pair (2 electrons)	Weakest	Longest	H_2, F_2, CH_4
Double bond	2 pairs (4 electrons)	Intermediate	Intermediate	O_2, CO_2, C_2H_4
Triple bond	3 pairs (6 electrons)	Strongest	Shortest	N_2, C_2H_2

Key relationship: As the number of shared pairs increases, bond energy increases and bond length decreases. More electron density between the nuclei = stronger attraction = shorter, stronger bond.

MEMORISE THIS

Bond energy and length trend:

Triple bond > Double bond > Single bond (for bond energy/strength)

Triple bond < Double bond < Single bond (for bond length)

Think: more bonds = shorter and stronger.

Lewis (Dot-and-Cross) Structures

A **Lewis structure** shows all valence electrons in a molecule — both the bonding pairs (shared) and the lone pairs (non-bonding). Drawing Lewis structures correctly is one of the most tested skills in IB Chemistry SL.

Rules for drawing Lewis structures:

1. Count the total number of valence electrons from all atoms.
2. Arrange atoms — the least electronegative atom is usually the central atom (hydrogen is always terminal).
3. Place single bonds between connected atoms first (each bond uses 2 electrons from your total).
4. Complete octets on outer atoms first, then the central atom.
5. If the central atom still lacks an octet, convert lone pairs on outer atoms into double or triple bonds.
6. Check: total electrons used = total valence electrons counted in Step 1.

Valence electron counts by group:

Group	Valence electrons	Common atoms
1	1	H, Na (but Na rarely in Lewis structures)
14	4	C
15	5	N, P
16	6	O, S
17	7	F, Cl, Br, I

⚠ **EXAM ALERT**

The most common Lewis structure errors:

1. **Forgetting lone pairs** on atoms like O, N, and halogens. Oxygen in a single bond has 2 lone pairs; in a double bond it has 1 lone pair.
2. **Not satisfying the octet** on the central atom — always count electrons around the central atom after you finish.
3. **Giving hydrogen more than 2 electrons** — H is always an exception to the octet rule; it needs only 2 (a duet).

WORKED EXAMPLE

Worked Example: Drawing the Lewis structure of H_2O

Step 1: Count valence electrons.

- O (Group 16): 6 electrons
- H × 2 (Group 1): $1 \times 2 = 2$ electrons
- **Total: 8 electrons**

Step 2: Arrange atoms. O is the central atom; both H atoms are terminal.

Step 3: Place single bonds: O–H and O–H. This uses $2 \times 2 = 4$ electrons. Remaining: $8 - 4 = 4$ electrons.

Step 4: Complete octets on outer atoms. H already has 2 electrons (full duet). Place remaining 4 electrons on O as 2 lone pairs.

Step 5: Check. O has: 2 bonding pairs (4 electrons) + 2 lone pairs (4 electrons) = 8 electrons = full octet. ✓

Result: Water has 2 O–H single bonds and 2 lone pairs on oxygen.

WORKED EXAMPLE

Worked Example: Drawing the Lewis structure of CO_2

Step 1: Count valence electrons.

- C (Group 14): 4 electrons
- O × 2 (Group 16): 6 × 2 = 12 electrons
- **Total: 16 electrons**

Step 2: C is the central atom; both O atoms are terminal (O=C=O).

Step 3: Place single bonds: C–O and C–O. Uses 4 electrons. Remaining: 12 electrons.

Step 4: Complete octets on O atoms first. Each O needs 6 more electrons = 3 lone pairs each. That would use 12 electrons total. But C only has 4 electrons (2 from each single bond) = not a full octet.

Step 5: C needs 4 more electrons to complete its octet. Convert 1 lone pair from each O into a bonding pair with C → forms two C=O double bonds.

Final check: C has 4 + 4 = 8 electrons (two double bonds). Each O has 1 double bond (4 electrons) + 2 lone pairs (4 electrons) = 8 electrons. Total electrons used: 4 (bonds) + 4 (bonds) + 4 (lone pairs on each O × 2) = 16. ✓

Result: CO_2 has two C=O double bonds and 2 lone pairs on each oxygen.

WORKED EXAMPLE

Worked Example: Drawing the Lewis structure of NH_3

Step 1: Count valence electrons.

- N (Group 15): 5 electrons
- H × 3 (Group 1): 1 × 3 = 3 electrons
- **Total: 8 electrons**

Step 2: N is the central atom; all H atoms are terminal.

Step 3: Place single bonds: N–H, N–H, N–H. Uses 6 electrons. Remaining: 2 electrons.

Step 4: Complete octets on H (already done — each has 2). Place remaining 2 electrons on N as 1 lone pair.

Step 5: N has: 3 bonding pairs (6e) + 1 lone pair (2e) = 8 electrons = full octet. ✓

Result: Ammonia has 3 N–H single bonds and 1 lone pair on nitrogen.

WORKED EXAMPLE

Worked Example: Drawing the Lewis structure of PCl_3

Step 1: Count valence electrons.

- P (Group 15): 5 electrons
- Cl × 3 (Group 17): 7 × 3 = 21 electrons
- **Total: 26 electrons**

Step 2: P is the central atom; all Cl atoms are terminal.

Step 3: Place 3 P–Cl single bonds. Uses 6 electrons. Remaining: 20 electrons.

Step 4: Complete octets on each Cl: each needs 6 more electrons = 3 lone pairs. 3 × 6 = 18 electrons used. Remaining: 2 electrons.

Step 5: Place remaining 2 electrons on P as 1 lone pair.

Check: P has 3 bonding pairs + 1 lone pair = 8 electrons = full octet. Each Cl has 3 lone pairs + 1 bonding pair = 8 electrons. Total: 6 + 18 + 2 = 26. ✓

Result: PCl_3 has 3 P–Cl single bonds, 1 lone pair on P, and 3 lone pairs on each Cl.

⚠ EXAM ALERT

Octet rule exceptions — IB exam traps

Not all atoms obey the octet rule:

- **Electron-deficient molecules:** Boron in BF_3 has only 6 electrons around it (a stable sextet). Do NOT force a B=F double bond trying to complete the octet — BF_3 is stable with 6 electrons on B.
- **Expanded octets:** Period 3 and beyond can hold more than 8 electrons. PCl_5 has 10 electrons around phosphorus; SF_6 has 12 electrons around sulfur. This is allowed because d orbitals are available.

If you see a central atom from Period 3+ bonded to 5 or 6 atoms, it likely has an expanded octet — don't try to reduce it to 4 bonds.

Dative (Coordinate) Covalent Bonds

In a **dative bond** (also called a coordinate bond), both electrons in the shared pair come from the **same atom**. Once formed, a dative bond is identical to a regular covalent bond — the difference is only in how it formed, not in its nature. In Lewis structures, dative bonds are sometimes shown with an arrow pointing from the donor atom to the acceptor atom.

Example: The ammonium ion NH_4^+ forms when NH_3 donates its lone pair to a proton H^+. The N atom donates both electrons to form a fourth N–H bond. All four N–H bonds in NH_4^+ are equivalent.

IB TIP

In IB exams, you may be asked to identify or draw a dative bond. The key phrase is: “both electrons come from one atom.” The NH_4^+ ion (ammonium) is the classic example. Once formed, all four N–H bonds in ammonium are identical — dative bonds are NOT weaker than regular covalent bonds.

3. Metallic Bonding (S2.3)

In metals, the valence (outer) electrons are not fixed to individual atoms. Instead, they become **delocalised** — they leave their atoms and move freely throughout the entire metallic structure. The result is a **lattice of positive metal ions** (also called cations or metal cores) surrounded by a “sea” of freely moving electrons. The electrostatic attraction between the positive metal ions and the delocalised electrons holds the metal together. This is the **metallic bond**.

The Sea of Electrons Model

- **Positive ion lattice:** The metal atoms lose their valence electrons and become positive ions arranged in a regular, closely packed pattern.
- **Delocalised electron cloud:** The released valence electrons move freely through the entire lattice, not belonging to any single atom.
- **Metallic bond:** The overall electrostatic attraction between the positive ions and the electron cloud.

The strength of the metallic bond depends on:

1. **Charge on the metal ion** — higher charge = stronger attraction (e.g. Mg^{2+} forms stronger bonds than Na^+)
2. **Number of delocalised electrons per atom** — more electrons = stronger bond
3. **Size of the metal ion** — smaller ion = closer approach of electrons = stronger attraction

Properties of Metals Explained by the Model

Property	Explanation using sea of electrons model
High melting and boiling points	Strong electrostatic attraction between positive ion lattice and delocalised electrons requires much energy to break
Good electrical conductivity	Delocalised electrons can move freely and carry charge throughout the lattice; this happens in both solid and liquid state
Good thermal conductivity	Delocalised electrons can rapidly transfer kinetic energy through the lattice
Malleability and ductility	Layers of positive ions can slide past each other without breaking bonds; the electron sea re-adjusts around the new positions
Lustre (shiny appearance)	Delocalised electrons absorb and re-emit light across all visible wavelengths

MEMORISE THIS

Key contrast — why ionic solids are brittle but metals are malleable:

In an ionic lattice, shifting a layer brings like charges together → repulsion → the crystal shatters.

In a metal, shifting a layer just moves positive ions through the electron sea → no repulsion → the metal bends.

IB TIP

When explaining electrical conductivity of metals, always say **delocalised electrons** carry the charge. Never say “free electrons” without explaining why they are free. Also: metals conduct in both solid AND liquid state (unlike ionic compounds which only conduct when molten or dissolved).

4. VSEPR Theory & Molecular Geometry (S2.2)

Valence Shell Electron Pair Repulsion (VSEPR) theory predicts the shape of a molecule based on the idea that electron pairs (both bonding pairs and lone pairs) around a central atom repel each other and arrange themselves as far apart as possible to minimise repulsion.

Electron Domains

An **electron domain** (or electron group) is any region of electron density around the central atom. Each of the following counts as ONE electron domain:

- A single bond
- A double bond (counts as 1 domain, even though it contains 2 pairs)
- A triple bond (counts as 1 domain)
- A lone pair

The number of electron domains determines the **electron geometry** (the arrangement of all electron pairs). The **molecular shape** (or molecular geometry) only counts the positions of atoms, not lone pairs — so lone pairs affect bond angles but do not appear in the shape name.

Lone Pair vs Bonding Pair Repulsion

Not all electron pairs repel equally:

Lone pair–lone pair repulsion > Lone pair–bonding pair repulsion > Bonding pair–bonding pair repulsion

This is because lone pairs are closer to the nucleus and occupy more space. The presence of lone pairs compresses the bond angles below the ideal geometry.

- In CH_4 (no lone pairs): bond angle = 109.5° (perfect tetrahedral)
- In NH_3 (1 lone pair): bond angle ≈ 107° (compressed from 109.5°)
- In H_2O (2 lone pairs): bond angle ≈ 104.5° (compressed further)

VSEPR Shapes — Complete Reference Table

MEMORISE THIS

Electron domains	Bonding pairs	Lone pairs	Shape	Bond angle	Example
2	2	0	Linear	180°	CO_2, $BeCl_2$
3	3	0	Trigonal planar	120°	BF_3, SO_3
3	2	1	Bent (V-shaped)	~117°	SO_2
4	4	0	Tetrahedral	109.5°	CH_4, CCl_4
4	3	1	Trigonal pyramidal	~107°	NH_3, PCl_3
4	2	2	Bent (V-shaped)	~104.5°	H_2O
5	5	0	Trigonal bipyramidal	90°, 120°, 180°	PCl_5
6	6	0	Octahedral	90°	SF_6

EXAM ALERT

Two common shape mistakes:

1. **Giving the electron geometry instead of the molecular shape.** NH_3 has a tetrahedral electron geometry but a **trigonal pyramidal** molecular shape. You must name the shape based on atom positions only.
2. **Forgetting that double bonds count as ONE electron domain.** CO_2 has 2 double bonds = 2 domains = linear, NOT bent.

WORKED EXAMPLE

Worked Example: Determining the shape of H_2O

Step 1: Draw the Lewis structure. O is central, 2 O–H bonds, 2 lone pairs on O.

Step 2: Count electron domains around O.

- 2 bonding pairs (the two O–H bonds)
- 2 lone pairs
- **Total: 4 electron domains**

Step 3: Electron geometry = tetrahedral (4 domains).

Step 4: Determine molecular shape. 2 lone pairs are present — they occupy two corners of the tetrahedron but are not "seen" in the shape name. The shape described by the 3 atoms (O and 2 H atoms) is **bent** (V-shaped).

Step 5: Bond angle. Two lone pairs compress the H–O–H angle below 109.5°. Bond angle ≈ 104.5°.

Answer: H_2O is bent/V-shaped with a bond angle of approximately 104.5°.

WORKED EXAMPLE

Worked Example: Determining the shape of NH_3

Step 1: Lewis structure: N is central, 3 N–H single bonds, 1 lone pair on N.

Step 2: Electron domains around N: 3 bonding pairs + 1 lone pair = **4 domains**.

Step 3: Electron geometry = tetrahedral.

Step 4: Molecular shape: 1 lone pair is invisible in shape name. The 4 atoms (N + 3 H) form a **trigonal pyramidal** shape — like a triangular base with N at the apex.

Step 5: Bond angle: 1 lone pair pushes bonding pairs slightly → H–N–H angle ≈ 107°.

Answer: NH_3 is trigonal pyramidal with bond angles of approximately 107°.

IB TIP

A reliable method for VSEPR in the exam:

1. Draw the Lewis structure first.
2. Count ALL electron domains (lone pairs + bonds) around the central atom.
3. State the electron geometry.
4. Remove lone pairs from the name → molecular shape.
5. Adjust bond angle down from ideal if lone pairs are present.

Watch: Covalent Bonding, Lewis Structures, VSEPR & Polar Bonds VIDEO

5. Electronegativity & Polarity (S2.2)

Electronegativity is the ability of an atom to attract the shared pair of electrons in a covalent bond towards itself. It is a property of an atom within a molecule (not a property of an isolated atom). The most widely used electronegativity scale is the **Pauling scale**, where fluorine (F) is the most electronegative element at 4.0.

Electronegativity Trends

Trend	Direction	Reason
Across a period (left → right)	Increases	More protons → greater nuclear charge → stronger pull on bonding electrons
Down a group (top → bottom)	Decreases	More electron shells → greater shielding → bonding electrons further from nucleus

Most electronegative elements: F (4.0) > O (3.5) > N (3.0) ≈ Cl (3.0)

Polar Covalent Bonds

When two atoms with **different electronegativities** share electrons, the electrons are not shared equally. The more electronegative atom pulls the shared pair towards itself, developing a partial negative charge ($\delta-$), while the less electronegative atom develops a partial positive charge ($\delta+$). This is a **polar covalent bond**.

- If the electronegativity difference is **small** (roughly < 0.4): the bond is considered non-polar (e.g. C–H).
- If the electronegativity difference is **intermediate** (roughly 0.4–1.7): the bond is polar covalent (e.g. O–H, N–H, C–Cl).
- If the electronegativity difference is **large** (roughly > 1.7): the bond is considered ionic.

Example: In HCl, chlorine is more electronegative than hydrogen:

$$\delta + \mathrm{H} - \mathrm{Cl}\delta -$$

Polar vs Non-Polar Molecules

A molecule is **polar** if it has a net **dipole moment** — an overall separation of charge. A molecule is **non-polar** if the individual bond dipoles cancel out due to symmetrical molecular geometry.

A molecule is non-polar if:

- All bonds are non-polar (e.g. Cl_2), OR
- All bonds are polar BUT the molecule is symmetrical so dipoles cancel (e.g. CO_2, BF_3, CCl_4)

A molecule is polar if:

- It has polar bonds AND an asymmetrical shape so dipoles do NOT cancel (e.g. H_2O, NH_3, HCl, $CHCl_3$)

MEMORISE THIS

Non-polar despite polar bonds (shape makes them cancel):

- CO_2 — linear, two C=O dipoles point in opposite directions → cancel
- BF_3 — trigonal planar, three B–F dipoles at 120° → vector sum = 0 → cancel
- CCl_4 — tetrahedral, four C–Cl dipoles arranged symmetrically → cancel
- PCl_5 — trigonal bipyramidal, symmetrical → cancel

Polar molecules (asymmetrical, dipoles do NOT cancel):

- H_2O — bent shape, two O–H dipoles do not cancel → net dipole
- NH_3 — trigonal pyramidal, three N–H dipoles + lone pair → net dipole
- HCl — diatomic with different atoms → always polar
- $CHCl_3$ — tetrahedral but asymmetric (3 Cl, 1 H) → net dipole

EXAM ALERT

The most common polarity mistake: Stating that a molecule with polar bonds must be a polar molecule. This is WRONG. You must consider the shape. CO_2 has two polar C=O bonds but is a non-polar molecule because the linear shape causes the dipoles to point in exactly opposite directions and cancel. Similarly, CCl_4 has four polar C–Cl bonds but is non-polar due to its tetrahedral symmetry. Always draw or state the shape before determining polarity.

WORKED EXAMPLE

Worked Example: Predicting the polarity of CCl_4

Step 1: Identify the bonds. C–Cl: carbon (EN ≈ 2.5) and chlorine (EN ≈ 3.0). The difference is ~0.5. Each C–Cl bond is polar ($C^{\delta+}$–$Cl^{\delta-}$).

Step 2: Determine the shape. C is the central atom, 4 Cl atoms, 0 lone pairs → **tetrahedral**, bond angles 109.5°.

Step 3: Assess symmetry. The four C–Cl dipoles point symmetrically from the centre to the four corners of the tetrahedron. By vector addition, the four equal and symmetrically arranged dipoles cancel exactly.

Answer: CCl_4 has polar C–Cl bonds but is a **non-polar molecule** because its tetrahedral symmetry causes the bond dipoles to cancel. It has no net dipole moment.

6. Intermolecular Forces (S2.2)

Intermolecular forces (IMFs) are the attractive forces **between** molecules. They are much weaker than the covalent bonds within a molecule (intramolecular bonds). IMFs are responsible for the physical properties of molecular substances — particularly boiling points, melting points, and solubility.

⚠ **EXAM ALERT**

The single most common error in IB Chemistry: Confusing intermolecular forces with intramolecular bonds.

- **Intramolecular bonds** (covalent bonds, ionic bonds) hold atoms together within a molecule or formula unit. These are strong (hundreds of kJ/mol).
- **Intermolecular forces** hold molecules together in a bulk sample. These are much weaker (a few to tens of kJ/mol).

When a molecular substance melts or boils, the **intermolecular forces** are overcome — NOT the covalent bonds. The molecules themselves remain intact. For example, when water boils, the O–H covalent bonds are NOT broken; only the hydrogen bonds between molecules are disrupted.

Three Types of Intermolecular Forces

London (Dispersion) Forces

London dispersion forces (a type of van der Waals force, also called induced-dipole–induced-dipole forces) are present in **all molecules** — polar and non-polar alike. They arise from **temporary dipoles**:

- At any instant, the electrons in a molecule may be unevenly distributed, creating a temporary (instantaneous) dipole.
- This temporary dipole induces a dipole in a neighbouring molecule.
- The two temporary dipoles then attract each other.

Factors affecting strength of London forces:

- **More electrons (higher molar mass / larger molecule) → stronger London forces** — more electrons means larger, more easily polarised electron clouds and bigger temporary dipoles.
- **Shape** — more elongated/flat molecules have greater surface area contact → stronger London forces than compact/spherical molecules of the same molar mass.

Effect on boiling point: As London forces increase (larger molecules, more electrons), more energy is needed to separate molecules → higher boiling point. This explains why boiling points increase down a group for non-polar molecules (e.g. noble gases: He < Ne < Ar < Kr < Xe).

Dipole–Dipole Forces

Dipole–dipole forces act between **polar molecules**. Because polar molecules have a permanent $\delta+$ end and a $\delta-$ end, neighbouring molecules align so that the positive end of one molecule is attracted to the negative end of another.

Dipole–dipole forces are stronger than London forces for molecules of similar size.

Example: HCl molecules attract each other via dipole–dipole forces (in addition to London forces which are present too).

Hydrogen Bonding

Hydrogen bonding is a special, particularly strong type of dipole–dipole interaction. It forms when:

1. A hydrogen atom is **directly bonded** to a highly electronegative atom: **N, O, or F**.
2. The $\delta+$ H is then attracted to a lone pair on an N, O, or F atom of a neighbouring molecule.

Hydrogen bonds are approximately **5–10 times stronger** than typical dipole–dipole forces and significantly stronger than London forces (for small molecules).

Examples of molecules that form hydrogen bonds: H_2O, NH_3, HF, alcohols (-OH), amines (-NH_2)

⚠ **EXAM ALERT**

The H-bonding conditions exam trap: The hydrogen atom must be **directly bonded** to N, O, or F — not simply near one. For example, in methanol (CH_3OH), the H attached to O forms hydrogen bonds, but the H atoms attached to C do NOT form hydrogen bonds (C is not electronegative enough). Always trace the bond: H–N, H–O, or H–F.

Effect of IMFs on Boiling Points

The stronger the intermolecular forces, the higher the energy needed to separate molecules into a gas, and therefore the **higher the boiling point**.

Ranking IMF strength (weakest → strongest):

London forces < Dipole–dipole forces < Hydrogen bonds

However, London forces can overcome H-bonding if the molecule is very large (many electrons). For instance, large alkanes with hundreds of electrons have higher boiling points than small hydrogen-bonded molecules like water.

IMF Comparison Table

MEMORISE THIS

Force	Present in	Relative strength	Origin
London (dispersion)	ALL molecules	Weakest (increases with molar mass)	Temporary dipoles from uneven electron distribution
Dipole–dipole	Polar molecules only	Intermediate	Permanent dipoles align between molecules
Hydrogen bonding	Molecules with H directly bonded to N, O, or F	Strongest (for small molecules)	Partial positive H attracted to lone pair on N, O, or F

WORKED EXAMPLE

Worked Example: Why does HF have a higher boiling point than HCl?

Step 1: Identify the IMFs in each molecule.

- HF: H is directly bonded to F → **hydrogen bonds** form between HF molecules (in addition to London forces and dipole–dipole forces).
- HCl: H is bonded to Cl. Cl is not N, O, or F → no hydrogen bonds. HCl is polar → **dipole–dipole forces** + London forces.

Step 2: Compare IMF strength.

- Hydrogen bonds (in HF) are significantly stronger than dipole–dipole forces (in HCl).
- HF has a smaller molar mass (20 g/mol) than HCl (36.5 g/mol), so HCl has stronger London forces — but this is outweighed by the hydrogen bonding in HF.

Step 3: Conclusion.

- More energy is needed to overcome the hydrogen bonds in HF than the dipole–dipole forces in HCl.
- Therefore HF has a higher boiling point than HCl despite having a lower molar mass.

Answer: HF has a higher boiling point than HCl because HF molecules form hydrogen bonds (H directly bonded to F), which are much stronger than the dipole–dipole forces between HCl molecules.

Water's Anomalous Properties

Water's unusually high boiling point (100°C) for such a small molecule ($M_r = 18$) is entirely due to its extensive hydrogen bonding. Each water molecule can form up to 4 hydrogen bonds (2 from the two O–H hydrogen atoms, 2 using the 2 lone pairs on oxygen as acceptors). This network of hydrogen bonds gives water its high boiling point, high surface tension, and ability to dissolve many ionic and polar substances.

Properties Comparison: Ionic, Covalent, Metallic

MEMORISE THIS

Property	Ionic	Simple molecular covalent	Giant covalent	Metallic
Examples	NaCl, MgO	H_2O, CO_2, CH_4	Diamond, SiO_2	Fe, Cu, Na
Melting point	High	Low	Very high	Variable; generally high
Electrical conductivity (solid)	No	No	No (except graphite)	Yes
Electrical conductivity (liquid/solution)	Yes (mobile ions)	No (no ions or free electrons)	No	Yes
Solubility in water	Often soluble	Variable	Insoluble	Insoluble
Structural unit	Ions in lattice	Discrete molecules	Atoms in network lattice	Positive ions + delocalised electrons
Force to overcome	Electrostatic (ionic)	Intermolecular forces	Covalent bonds	Metallic bonding

7. The Bonding Triangle (S2.1–S2.3)

The three types of bonding — **ionic, covalent, and metallic** — are not completely distinct categories. They form a continuum, and real substances often sit somewhere between pure ionic, pure covalent, or pure metallic character. The **bonding triangle** (also called Van Arkel–Ketelaar triangle) represents this continuum graphically, with each corner representing one idealised bond type.

Position on the Triangle

The position of a compound or element on the bonding triangle depends on two factors:

1. **Electronegativity difference between the atoms** — determines how ionic or covalent a bond is.
2. **Average electronegativity of the atoms** — determines how metallic or covalent the substance is.

High electronegativity difference + high average electronegativity	→ Ionic (e.g. NaF, MgO, CsF)
Low electronegativity difference + high average electronegativity	→ Covalent (e.g. F_2, Cl_2, CO_2, H_2O)
Low average electronegativity (metals)	→ Metallic (e.g. Na, Fe, Cu, Mg)

Key Points About the Bonding Triangle

- **Pure ionic bonds** do not truly exist — all "ionic" bonds have some covalent character because electrons are never 100% transferred; there is always some sharing.
- **Pure covalent bonds** between identical atoms (e.g. Cl_2, H_2) come closest to pure covalent character, since there is no electronegativity difference.
- **Ionic vs covalent character** is a spectrum based on electronegativity difference. Greater electronegativity difference → more ionic character.
- **Polar covalent bonds** sit between pure covalent and ionic on the triangle.

Examples of Where Substances Sit

Substance	Position on triangle	Reason
CsF	Most ionic corner	Cs (low EN) + F (highest EN) = largest possible EN difference
NaCl	Ionic (EN diff ≈ 2.1)	Large EN difference between Na and Cl
HCl	Polar covalent	Intermediate EN difference; H and Cl are both non-metals
Cl_2	Non-polar covalent	Identical atoms; EN difference = 0
Na	Metallic corner	Metal with low EN; electrons delocalised
Fe	Metallic corner	Transition metal; delocalised electrons
MgO	Ionic (closer to ionic than NaCl)	Mg^{2+} and O^{2-} — large charge; high EN difference

IB TIP

The IB may ask you to place a substance on the bonding triangle or to compare the ionic/covalent/metallic character of different compounds. The key relationship to remember: **larger electronegativity difference = more ionic character**. A compound like CsF (electronegativity difference ≈ 3.3) is much more ionic than HF (difference ≈ 1.9).

EXAM ALERT

Do not state that ionic compounds have 100% ionic bonding and covalent compounds have 100% covalent bonding. IB examiners expect you to recognise the **continuum** of bonding. Even a compound like NaCl has some covalent character, and the IB first assessment of 2025 explicitly tests this idea through the bonding triangle.

Written Question Answers

W1 — Ionic lattice conductivity: Solid NaCl does not conduct electricity because the ions are fixed in the lattice and cannot move. When melted or dissolved in water, the ions become free to move and can carry charge, so the melt/solution conducts electricity.

W2 — Bond energy and length: Bond energy increases and bond length decreases as bond order increases (single < double < triple). More shared electron pairs create greater electron density between the nuclei, resulting in stronger, shorter bonds.

W3 — Polarity of $CHCl_3$ vs CCl_4: CCl_4 is non-polar because its tetrahedral symmetry causes all four equal C–Cl dipoles to cancel. $CHCl_3$ is polar because substituting one Cl with H (which has a different electronegativity) breaks the symmetry, so the bond dipoles no longer cancel and a net dipole moment exists.

W4 — Hydrogen bonding in water: Water's anomalously high boiling point is due to extensive hydrogen bonding. Each water molecule has two O–H bonds (H-bond donors) and two lone pairs on oxygen (H-bond acceptors), allowing each molecule to form up to four hydrogen bonds. Overcoming this hydrogen bond network requires significantly more energy than overcoming the weaker London/dipole–dipole forces in molecules of similar molar mass.

Mixed Practice — Exam Style

> **IB TIP**
>
> **How to use this section:** Unlike topic-specific practice, these questions are interleaved — they mix all topics from this guide in random order. Before answering, identify *which concept or topic area* the question is testing. This is exactly the skill you need on Paper 1 and Paper 2, where you don't know in advance which topic each question covers.

1. **[Bonding Types]** Which of the following compounds contains both ionic and covalent bonds?

   A. $NaCl$

   B. NH_3

   C. Na_2SO_4 (sodium sulfate)

   D. HCl

2. **[VSEPR — Molecular Shape]** The molecule SF_4 has 4 bonding pairs and 1 lone pair around the central sulfur atom. Its molecular geometry is:

   A. Tetrahedral

   B. Trigonal bipyramidal

   C. See-saw (sawhorse)

   D. Square planar

3. **[Electronegativity and Polarity]** CCl_4 has four polar C–Cl bonds but is a non-polar molecule overall. The reason is:

   A. The electronegativity difference between C and Cl is too small to create dipoles

   B. The four C–Cl bond dipoles are arranged symmetrically (tetrahedral), and they cancel exactly, giving a zero net dipole moment

   C. Carbon's electronegativity equals chlorine's, so no dipoles exist

   D. The molecule is non-polar because it contains no lone pairs on the central atom

4. **[Intermolecular Forces]** Arrange the following in order of increasing boiling point: CH_4, HF, CH_3CH_2OH.

   A. $HF < CH_4 < CH_3CH_2OH$

   B. $CH_4 < HF < CH_3CH_2OH$

   C. $CH_4 < CH_3CH_2OH < HF$

   D. $CH_3CH_2OH < HF < CH_4$

5. **[VSEPR — Molecular Geometry]** Using VSEPR theory, predict the molecular geometry and bond angle of CO_2.

   A. Bent, ~104.5°

   B. Linear, 180°

   C. Trigonal planar, 120°

   D. Trigonal pyramidal, ~107°

6. **[Bonding Triangle — Distractor]** A student places NaCl at the "100% ionic" corner of the Van Arkel–Ketelaar bonding triangle, arguing that the electronegativity difference (2.1) is very large and therefore the bonding is purely ionic. Evaluate this claim:

   A. Correct — any electronegativity difference above 1.7 defines purely ionic bonding

   B. Incorrect — the bonding triangle represents a continuum. NaCl has predominantly ionic character but still has some covalent character. No real compound occupies the pure ionic corner.

   C. Correct — NaCl is always placed at the ionic apex because its ions are fully charge-separated

D. Incorrect — NaCl should be placed at the metallic corner because sodium is a metal

7. **[VSEPR — Bond Angles]** The H–N–H bond angle in NH_3 is approximately 107°, less than the ideal tetrahedral angle of 109.5°. The reason is:

   A. Nitrogen is more electronegative than hydrogen, which compresses the bonding pairs

   B. The lone pair on nitrogen occupies more space than a bonding pair (greater electron-electron repulsion), pushing the bonding pairs closer together

   C. Three hydrogen atoms are larger than nitrogen, forcing the bond angles to decrease

   D. The molecule is trigonal planar, so the angle must be less than 120°

8. **[Intermolecular Forces — Hydrogen Bonding]** Which of the following can act as a hydrogen bond donor?

   A. CH_4

   B. F_2

   C. NH_3

   D. CCl_4

9. **[Structure and Properties]** Diamond and graphite are both allotropes of carbon. Diamond is hard and does not conduct electricity, while graphite is soft and conducts electricity. The difference in properties arises because:

   A. Diamond is ionic, graphite is metallic

   B. In diamond, each carbon forms 4 covalent bonds in a rigid 3D network; in graphite, each carbon forms 3 covalent bonds in flat layers, leaving one delocalised electron per carbon that can conduct electricity; the layers are held by weak London dispersion forces, allowing them to slide

   C. Graphite has a higher molar mass than diamond, making it denser and more conductive

   D. Diamond has a molecular (simple covalent) structure; graphite has a giant ionic lattice

10. **[Electronegativity — Bond Character]** Place the following bonds in order from most ionic to most covalent character: C–H, Na–Cl, O–H, F–F.

    A. Na–Cl > O–H > C–H > F–F

    B. F–F > C–H > O–H > Na–Cl

C. O–H > Na–Cl > F–F > C–H

D. Na–Cl > C–H > O–H > F–F

▶ Show Answers

Structure 2: Models of Bonding and Structure — IB Chemistry SL

May 2026 Prediction Questions

⚠ **EXAM ALERT**

These are NOT official IB questions. These are trend-based practice questions written to reflect the topic areas and question styles most likely to appear on the May 2026 IB Chemistry SL Paper 2. Based on recent exam patterns (2022-2025), expect heavy weighting on: VSEPR shapes and bond angles, molecular polarity, intermolecular forces explaining physical properties, and comparison of bonding types.

WORKED EXAMPLE

Question 1 [VSEPR and Polarity] [~7 marks]

Consider the molecules BF_3, NF_3, and CF_4.

(a) Draw the Lewis (electron dot) structure for each molecule, showing all bonding pairs and lone pairs.

(b) Use VSEPR theory to predict the shape and bond angle of each molecule.

(c) Explain why NF_3 is polar but BF_3 and CF_4 are non-polar, despite all three containing polar X-F bonds.

▶ Show Solution

WORKED EXAMPLE

Question 2 [Intermolecular Forces and Properties] [~7 marks]

The boiling points of the hydrogen halides are:

Compound	Boiling point (K)
HF	293
HCl	188
HBr	206
HI	238

(a) Explain the general trend in boiling points from HCl to HI.

(b) Explain why HF has an anomalously high boiling point compared to the other hydrogen halides.

(c) Explain why $NaCl$ (melting point 1074 K) has a much higher melting point than any of the hydrogen halides.

▶ Show Solution

WORKED EXAMPLE

Question 3 [Bonding Types Comparison] [~6 marks]

Compare and contrast the bonding and structure in diamond, graphite, and ice (H_2O), and explain how the bonding in each substance accounts for its electrical conductivity.

▶ Show Solution