Structure and Bonding

Matthew Williams • Chemistry • May 15, 2026

Structure and Bonding

The melting point, solubility, electrical conductivity, and hardness of a substance all follow directly from how its particles are bonded and arranged. This page covers how bonds form, what structures result, and the properties those structures produce.

Why Atoms Bond

Most atoms are chemically reactive because their outer electron shells are incomplete. Atoms achieve stability — an arrangement matching the electron configuration of a noble gas — by:

- **losing** electrons (forming positive ions)
- **gaining** electrons (forming negative ions)
- **sharing** electrons (forming covalent bonds)

Noble gases (Group 0) are stable and unreactive precisely because their outer shells are already full.

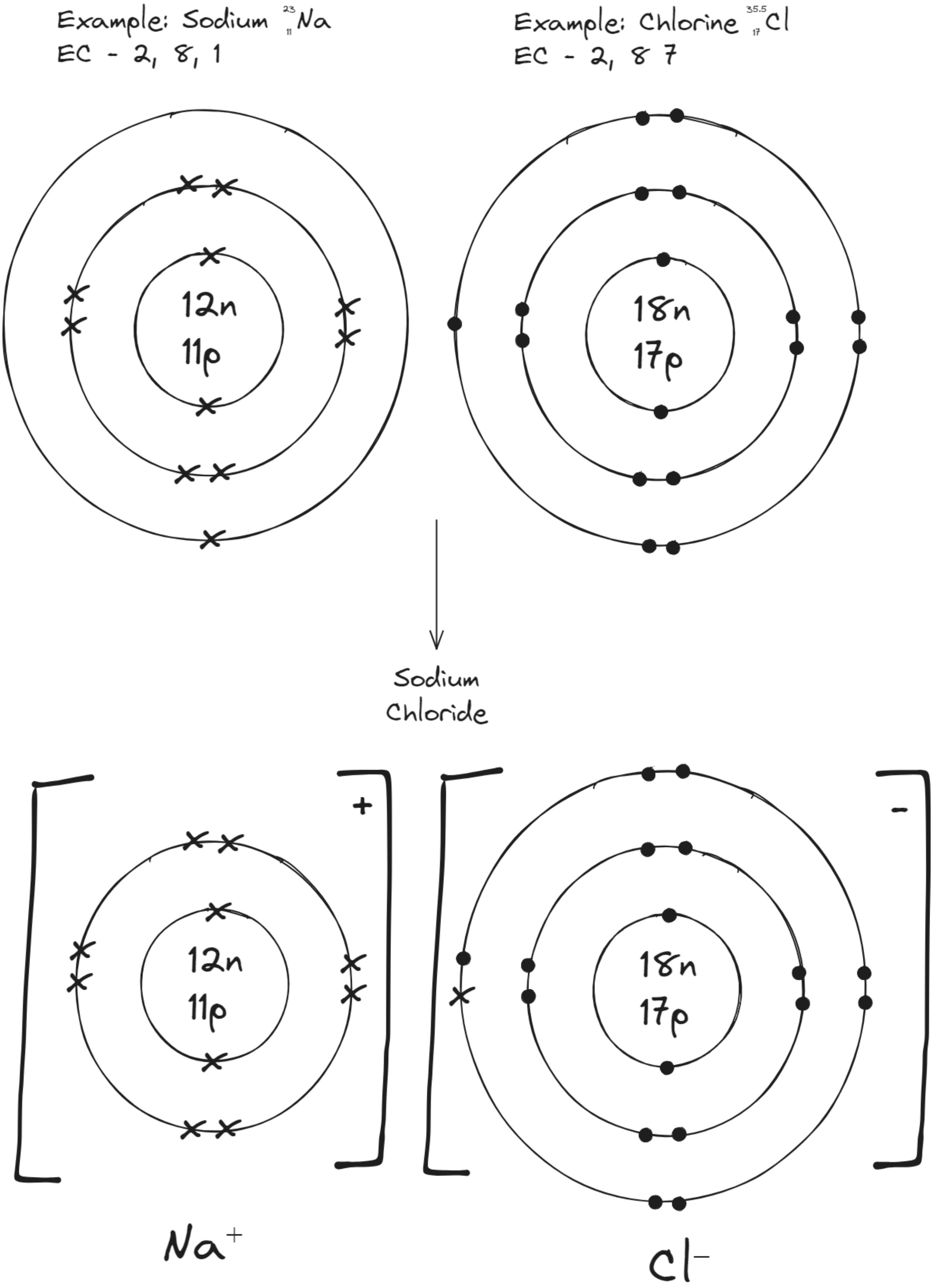
Ionic Bonding

Ionic bonding is the strong electrostatic attraction between oppositely charged ions. It forms when one or more electrons are transferred from a metal atom to a non-metal atom. The metal becomes a positive cation; the non-metal becomes a negative anion. Both achieve a full outer shell.

Formation of Sodium Chloride

Sodium (2,8,1) has one electron in its outer shell. Chlorine (2,8,7) needs one more to complete its outer shell.

- 1. Na ’Naz+ e{
- 2. Cl + e{’Cl{
- 3. The oppositely charged ions attract each other — this electrostatic attraction is the ionic bond.



Dot-and-cross diagram of the formation of NaCl, showing Na losing its outer electron to Cl, resulting in Nazand Cl{ ions with full outer shells

Other common ionic compounds:

Compound	Formula	Ions
Magnesium oxide	MgO	Mg^{2+} and O^{2-}
Calcium chloride	$CaCl_2$	Ca^{2+} and $2Cl^-$
Aluminium oxide	Al_2O_3	$2Al^{3+}$ and $3O^{2-}$
Potassium bromide	KBr	K^+ and Br^-

Predicting Bond Type

The type of bond that forms depends on the types of atoms involved:

Combination	Bond type	Reason	Example
Metal + non-metal	Ionic	Metal loses electrons; non-metal gains them	$NaCl$, MgO, $CaBr_2$
Non-metal + non-metal	Covalent	Neither atom easily loses electrons; they share instead	H_2O, CO_2, NH_3

Formulae of Ions and Molecules

An ionic compound formula is written so that the total positive charge equals the total negative charge — the compound is electrically neutral overall. To balance charges, use the charge of each ion as the subscript of the other (cross-multiply, then simplify).

Common ions to know:

Cation	Formula	Anion	Formula
Sodium	Na^+	Chloride	Cl^-
Potassium	K^+	Bromide	Br^-
Magnesium	Mg^{2+}	Oxide	O^{2-}
Calcium	Ca^{2+}	Sulfide	S^{2-}
Aluminium	Al^{3+}	Nitride	N^{3-}
Ammonium	NH_4^+	Hydroxide	OH^-
		Nitrate	NO_3^-

		Carbonate	CO_3^{2-}
		Sulfate	SO_4^{2-}

📋 **Example**

Aluminium oxide: Al is 3+ and O is 2−. To balance: 2 × Al^{3+} gives 6+; 3 × O^{2-} gives 6−. Formula: Al_2O_3

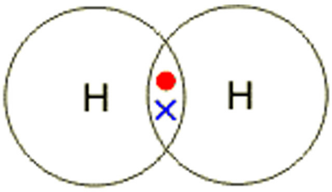
Covalent Bonding

Covalent bonding forms when two non-metal atoms share one or more pairs of electrons. Each shared pair is one covalent bond. Both atoms achieve a full outer shell through sharing. Electrons not involved in bonding sit as **lone pairs** on each atom.

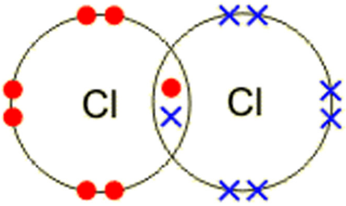
Bond type	Shared pairs	Examples
Single bond	1	H_2, HCl, H_2O, NH_3, CH_4
Double bond	2	O_2, CO_2
Triple bond	3	N_2

Dot-and-Cross Diagrams

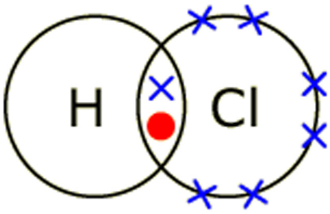
In a dot-and-cross diagram, one atom's electrons are shown as dots and the other's as crosses. Shared pairs sit between the two atoms; lone pairs sit on the outer edge of each atom.



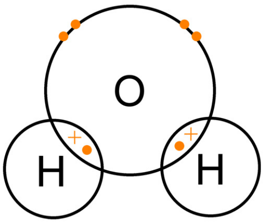
i. Hydrogen



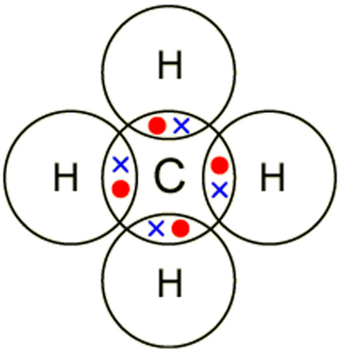
ii. Chlorine



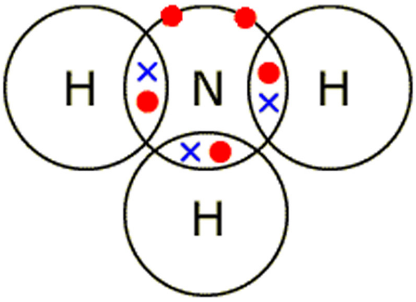
iii. Hydrogen Chloride



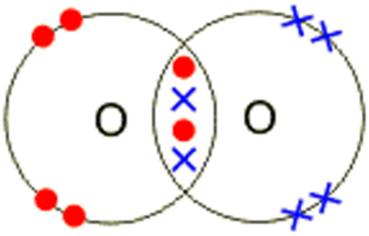
iv. Water



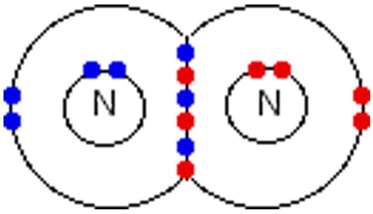
v. Methane



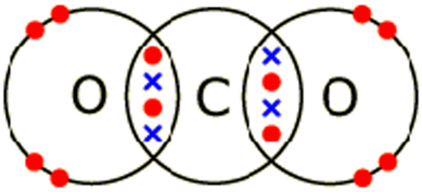
vi. Ammonia



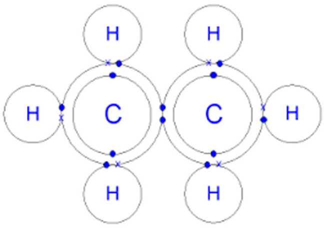
vii. Oxygen



viii. Nitrogen



ix. Carbon Dioxide



x. ethane

Dot-and-cross diagrams for H_2, Cl_2, HCl, H_2O, NH_3, CH_4, O_2, CO_2, and N_2

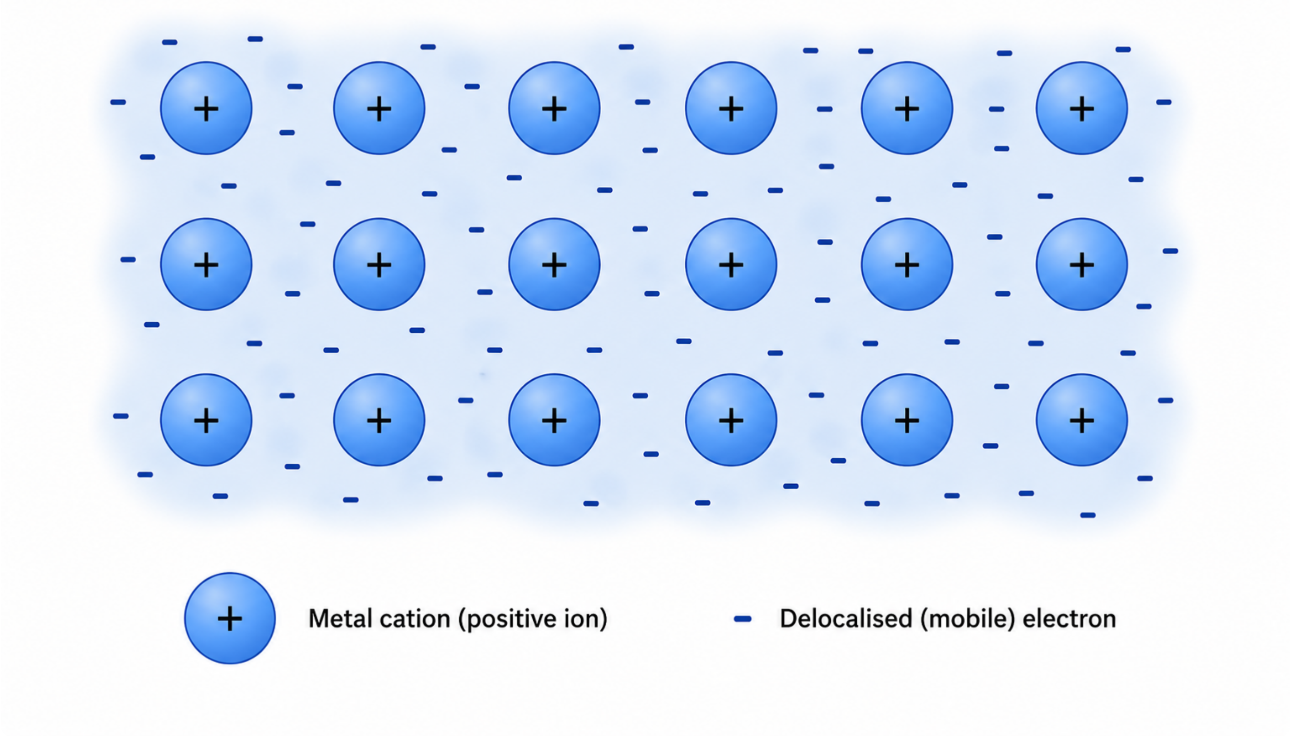
Key molecules to know for CSEC:

Molecule	Bond type	Lone pairs
H,	1 single bond	None
Cl,	1 single bond	3 lone pairs per Cl
H,O	2 single bonds	2 lone pairs on O
NHƒ	3 single bonds	1 lone pair on N
CH„	4 single bonds	None
O,	1 double bond	2 lone pairs per O
CO,	2 double bonds	2 lone pairs per O
N,	1 triple bond	1 lone pair per N

Metallic Bonding

In a metal, atoms release their valence electrons into a shared pool. The result is a regular arrangement of **positive metal cations** surrounded by a **sea of delocalised electrons** free to move throughout the entire structure.

Metallic bonding is the electrostatic attraction between those positive cations and the mobile electron sea.



Mettalic bonding diagram showing positive metal cations in a regular lattice surrounded by a sea of delocalised electrons

The properties of metals follow directly from this structure:

Property	Explanation
High melting point	Strong attraction between cations and delocalised electrons requires substantial energy to overcome
Good electrical conductor	Delocalised electrons carry charge freely through the structure
Good thermal conductor	Delocalised electrons transfer kinetic energy rapidly
Malleable	Layers of cations slide over each other while the electron sea maintains the bond
Ductile	Atoms rearrange without the metallic bonds breaking
Shiny lustre	Delocalised electrons reflect light

Types of Crystal Structures

Solids form three main types of crystal structure, with very different properties.

Ionic Crystals

An ionic crystal is a giant three-dimensional lattice of alternating positive and negative ions held together by strong electrostatic attractions in all directions. There are no separate molecules — the entire crystal is one continuous ionic structure.

In sodium chloride, each Na^+ is surrounded by 6 Cl^- ions, and each Cl^- is surrounded by 6 Na^+ ions.

3D ball-and-stick model of the sodium chloride ionic lattice, showing the cubic arrangement of alternating Na^+ and Cl^- ions

Key: purple spheres = Na^+ ions, green spheres = Cl^- ions

Simple Molecular Crystals

These are composed of small covalent molecules held in the solid state by **weak intermolecular forces** (van der Waals forces). The covalent bonds within each molecule are strong, but the attractions between molecules are weak. Examples: iodine (I_2), ice (H_2O), dry ice (CO_2), wax.

Giant Molecular (Giant Covalent) Structures

All atoms are joined by covalent bonds throughout the entire solid — there are no separate small molecules. Diamond, graphite, and silicon dioxide all have this structure.

Ionic versus Molecular Solids

Property	Ionic solids	Simple molecular solids
Melting point	High — strong ionic bonds throughout the lattice	Low — only weak intermolecular forces to overcome
Electrical conductivity	Conduct when molten or dissolved (ions become mobile); solid does not conduct	Do not conduct (no mobile ions or free electrons)
Solubility in water	Usually soluble — water molecules surround and separate the ions	Usually insoluble (especially non-polar molecules)

Solubility in organic solvents	Usually insoluble	Often soluble
Physical character	Hard, brittle crystals	Soft or waxy solids, liquids, or gases

📝 **Exam Tip**

A very common exam error is saying ionic compounds conduct electricity in the solid state. They do not — ions are fixed in the lattice. Conductivity only occurs when molten or dissolved, because that is when ions become free to move and carry charge.

Structure, Properties, and Uses

Sodium Chloride (NaCl)

Sodium chloride has a giant ionic lattice. Strong electrostatic attractions between Nazand Cl{ ions in all directions give it a high melting point (801 °C). It is hard but brittle — when layers are forced to shift, like charges align, causing sudden repulsion that shatters the crystal. NaCl dissolves readily in water and conducts electricity when molten or dissolved.

Uses: food seasoning and preservation, production of chlorine and sodium hydroxide by electrolysis, de-icing roads.

Diamond

Diamond is a carbon allotrope in which every carbon atom forms **four covalent bonds** to four other carbon atoms in a giant three-dimensional tetrahedral network. There are no double bonds and no free electrons.

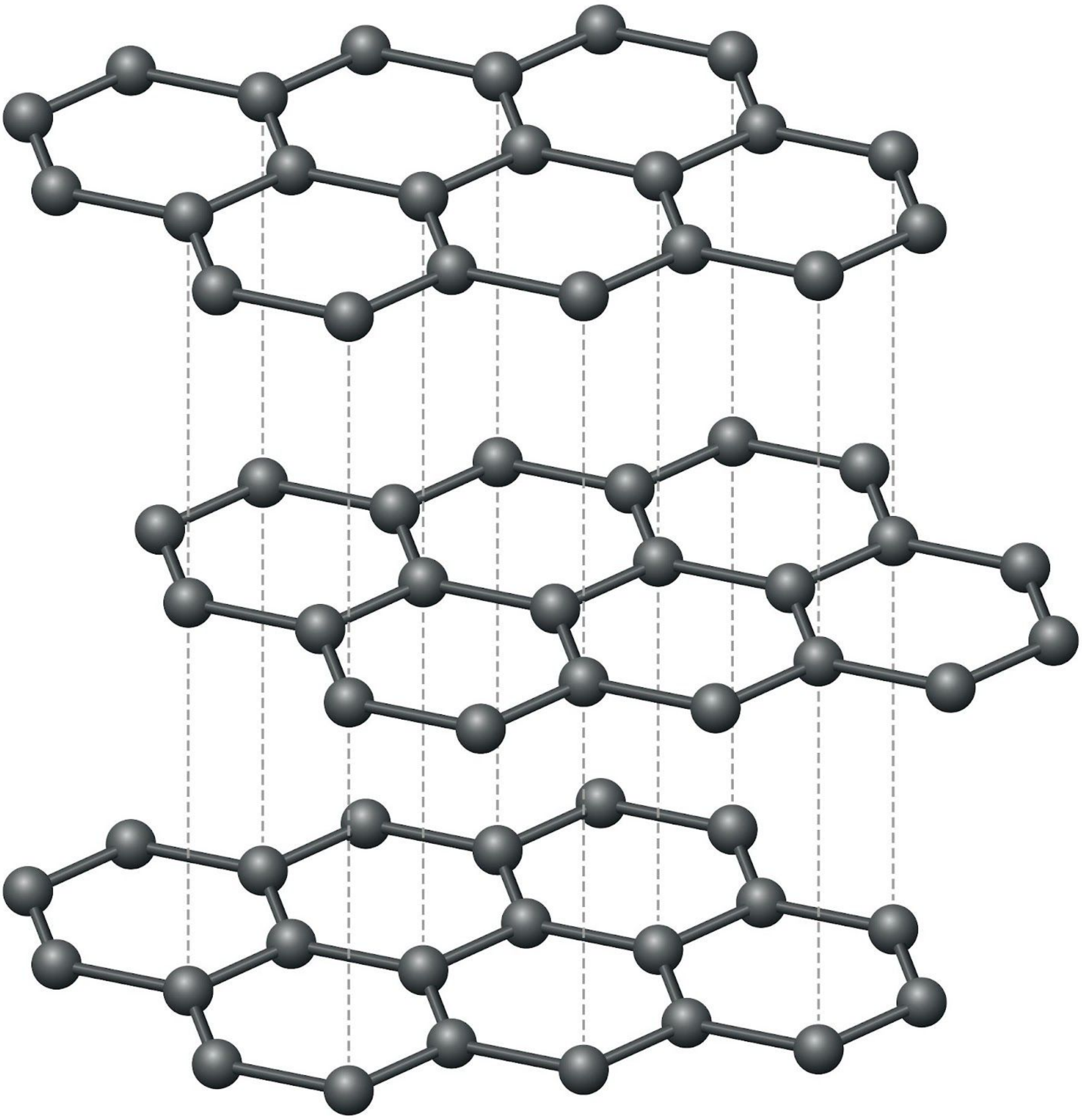
Ball-and-stick model of the diamond crystal lattice showing the tetrahedral bonding network, where each carbon atom bonds to four others

Property	Reason
Extremely hard	Every atom is locked in place by four strong covalent bonds throughout the giant structure
Very high melting point (>3500 °C)	An enormous number of strong covalent bonds must be broken
Does not conduct electricity	All outer electrons are used in bonding; there are no free electrons
Insoluble in water	Giant non-polar covalent structure

Uses: cutting tools and drill bits (hardest natural substance), abrasives, gemstones.

Graphite

Graphite is also a carbon allotrope, but with a very different structure. Each carbon atom forms **three covalent bonds** to three neighbours in flat hexagonal layers. The fourth outer electron per carbon is **delocalised** and free to move between layers.



Ball-and-stick model of the graphite crystal structure showing flat hexagonal layers of bonded carbon atoms in ABAB stacking

Property	Reason

Soft and slippery	Layers are held together only by weak van der Waals forces and slide over each other easily
High melting point	Strong covalent bonds within each layer require considerable energy to break
Conducts electricity	Delocalised electrons flow through the structure and carry charge
Insoluble in water	Giant covalent structure

Uses: lubricant (especially at high temperatures where oils break down), electrodes in electrolysis cells and batteries, pencil leads (mixed with clay).

> 🧠 **Remember**
> The key contrast: diamond has every carbon bonded to **four** others — no free electrons, maximum hardness, no conductivity. Graphite has every carbon bonded to **three** others — one electron delocalised per carbon, soft slippery layers, conducts electricity.

Allotropy

Allotropy is the existence of an element in two or more structurally different forms in the same physical state.

Diamond and graphite are both pure carbon in the solid state — both consist entirely of carbon atoms with only covalent bonds. The difference is entirely in how those atoms are arranged and connected. Because the arrangement differs, the properties are dramatically different despite being the same element.