

Chemical bonding

A-Level Chemistry 9701 · Topic 3 revision notes

Everything from the 9 lessons of this topic on a few pages: the key idea of each lesson, the diagram to remember, the points to learn and the traps examiners set. Read it, cover it, test yourself.

1. Electronegativity and bonding.

2. Ionic bonding.

3. Metallic bonding.

4. Covalent and dative covalent bonding.

5. Sigma and pi bonds, hybridisation

6. Shapes of molecules

7. Intermolecular forces

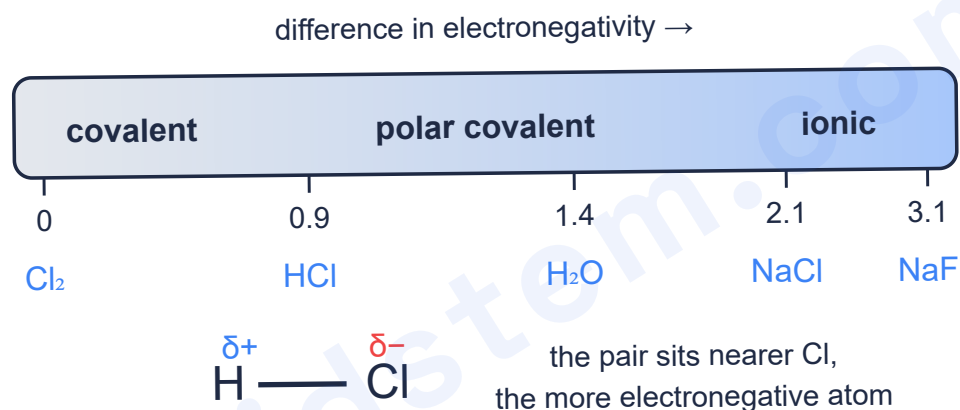
8. Hydrogen bonding and bond properties

9. Dot-and-cross diagrams

1 Electronegativity and bonding

Syllabus 3.1

Electronegativity is the power of an atom to attract the pair of electrons in a covalent bond towards itself. It increases with nuclear charge and decreases as the atomic radius and shielding increase, so it rises across a period and falls down a group; fluorine (4.0 on the Pauling scale) is the most electronegative element. When two bonded atoms have **equal** electronegativities the bond is pure covalent. A small difference gives a **polar** covalent bond with $\delta+$ and $\delta-$ ends; a large difference (roughly 2 or more) means the electron transfer is essentially complete and the bonding is ionic.



bonding is a spectrum, not three boxes

Fig. 1: electronegativity difference turns covalent bonding gradually into ionic bonding

Nuclear charge against distance and shielding.

- ✓ **Definition:** attraction for the bonding pair in a covalent bond.
- ✓ **Across a period:** increases; **down a group:** decreases.
- ✓ **Difference 0:** non-polar; **small:** polar; **large:** ionic.
- ✓ **$\delta-$** goes on the more electronegative atom.

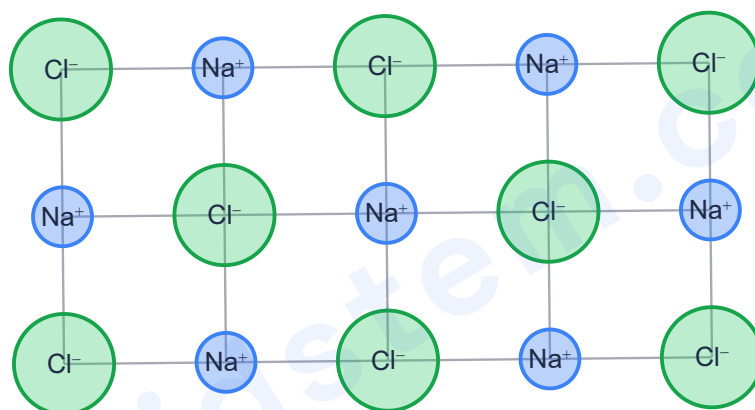
Watch out!

Electronegativity is about electrons **in a bond**. Do not confuse it with electron affinity or ionisation energy, which describe isolated gaseous atoms. And a molecule can contain polar bonds yet be non-polar overall if the dipoles cancel by symmetry, as in CO_2 ; that idea returns in the intermolecular forces lesson.

2 Ionic bonding

Syllabus 3.2

Ionic bonding is the **electrostatic attraction between oppositely charged ions**. The ions form when electrons transfer, usually from a metal to a non-metal, leaving each with a stable (often noble gas) configuration. The ions arrange themselves in a giant lattice in which every ion is surrounded by ions of opposite charge, so the attraction acts in all directions. The bonding is stronger when the ions carry larger charges and when they are smaller, because the charges get closer together; that is why MgO melts at about 2850°C and NaCl at 801°C .



each ion is surrounded by ions of opposite charge
attraction acts in every direction: a giant lattice, not molecules

Fig. 1: one layer of the sodium chloride lattice

Bigger charges, smaller ions, stronger bonding.

- ✓ **Definition:** electrostatic attraction between oppositely charged ions.
- ✓ **Giant lattice:** attraction in all directions, no molecules.
- ✓ **Strength:** rises with ionic charge, falls with ionic radius.
- ✓ **Conducts** only when the ions can move: molten or in solution.

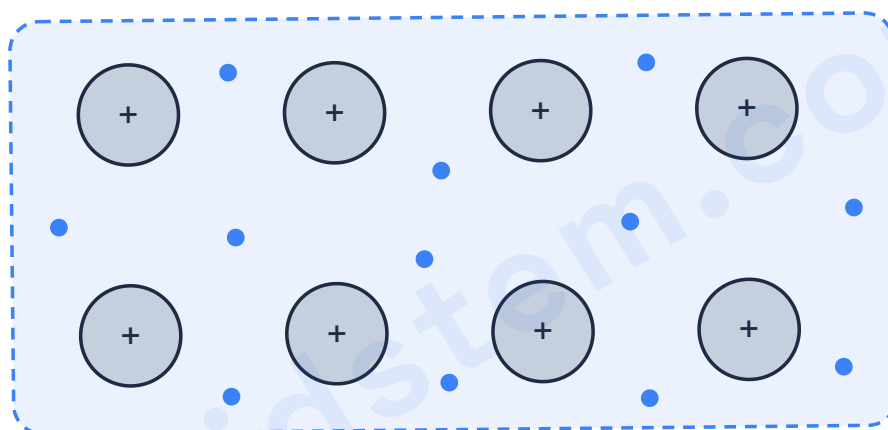
Watch out!

Ionic solids conduct nothing because the ions cannot move, and it is **ions**, not electrons, that carry the current in a melt or solution. Writing "free electrons" for molten NaCl loses the mark. Also avoid the word "molecule" for an ionic compound: NaCl is a formula unit in a giant lattice.

3 Metallic bonding

Syllabus 3.3

Metallic bonding is the **electrostatic attraction between positive metal ions and delocalised electrons**. The outer electrons of each atom are released into a “sea” that belongs to the whole giant lattice. The bonding gets stronger as more electrons per atom are delocalised, as the ionic charge rises and as the ions get smaller, so melting points rise from sodium to magnesium to aluminium. The mobile electrons carry current and heat, and because the bonding is non-directional the layers of ions can slide without the lattice breaking.



positive ions in a lattice, attracted to delocalised electrons
the electrons belong to the whole lattice and are free to move

Fig. 1: positive ions held together by a sea of delocalised electrons

Ions in a sea of delocalised electrons.

- ✓ **Definition:** attraction between positive ions and delocalised electrons.
- ✓ **Stronger** with more electrons per atom, higher charge, smaller ions.
- ✓ **Conducts:** delocalised electrons move.
- ✓ **Malleable:** layers slide, bonding is non-directional.

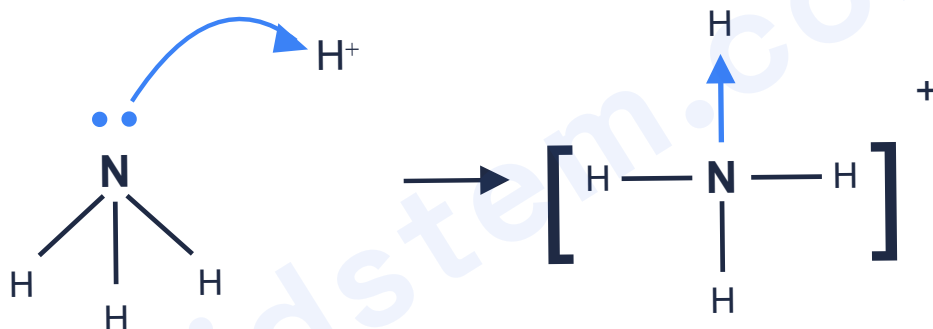
Watch out!

Say **positive ions** in a metal lattice, not atoms, and describe the electrons as **delocalised**.
Answers that describe “metal atoms sharing electrons” or “ions attracting each other” do not score: the bond is ion to electron sea.

4 Covalent and dative covalent bonding

Syllabus 3.4

A covalent bond is the **electrostatic attraction between two nuclei and a shared pair of electrons**. Atoms can share one, two or three pairs to form single, double and triple bonds. In a **dative covalent (coordinate) bond** both electrons of the shared pair come from the same atom, which must have a lone pair, and the other atom must have an empty orbital to receive it; it is drawn as an arrow from donor to acceptor. Once formed, a dative bond is identical to any other covalent bond. Period 3 elements can use d orbitals to hold more than eight outer electrons (PCl_5 , SF_6), while boron in BF_3 has only six.



the arrow shows both electrons came from the nitrogen lone pair
once formed, all four N–H bonds in NH_4^+ are identical

Fig. 1: ammonia donates its lone pair to H^+ , forming the ammonium ion

A shared pair holds two nuclei together.

- ✓ **Covalent:** attraction of two nuclei to a shared pair.
- ✓ **Dative:** both electrons from one atom; arrow from donor to acceptor.
- ✓ **Donor** needs a lone pair; **acceptor** needs an empty orbital.
- ✓ **Period 3** atoms can exceed eight; B in BF_3 has six.

Watch out!

When you draw a dative bond, the arrow points **from** the atom that donates the lone pair **to** the atom that accepts it. In a dot-and-cross diagram, show both electrons of the dative pair with the donor's symbol. And remember the finished bond is not weaker or different: all four bonds in NH_4^+ have the same length and strength.

5 Sigma and pi bonds, hybridisation

Syllabus 3.4

A **σ (sigma) bond** forms by direct, end-on overlap of orbitals, putting the electron density on the line between the nuclei. A **π (pi) bond** forms by sideways overlap of two p orbitals, putting the density above and below that line. A single bond is one σ bond; a double bond is one σ + one π ; a triple bond is one σ + two π . Carbon mixes its 2s and 2p orbitals into **hybrid orbitals**: sp^3 (four, tetrahedral, 109.5°), sp^2 (three, trigonal planar, 120° , one p left for a π bond) and sp (two, linear, 180° , two p left). The π bond is weaker than the σ bond because sideways overlap is smaller.

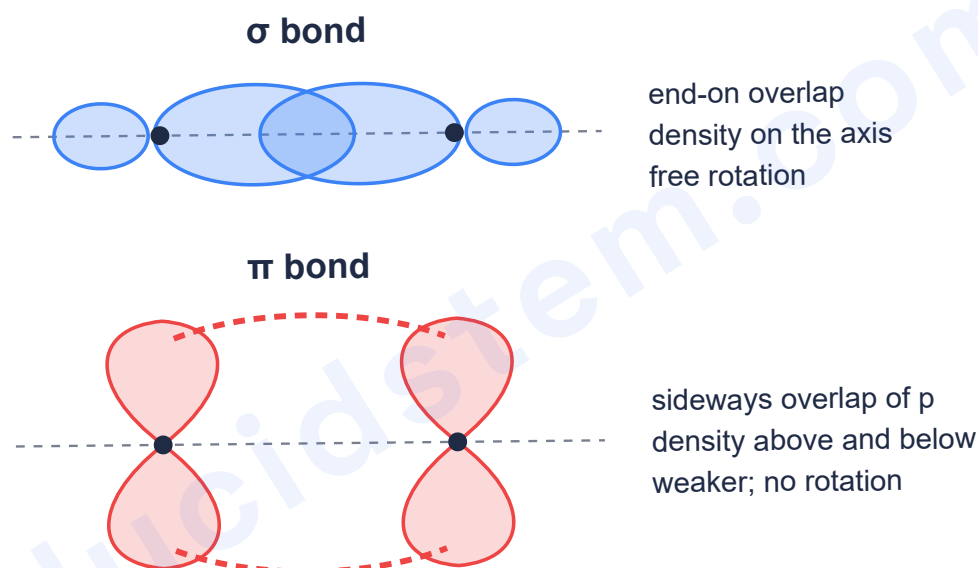


Fig. 1: a σ bond from end-on overlap and a π bond from sideways overlap of p orbitals

Every bond has one σ ; extras are π .

- ✓ **σ :** end-on overlap, density on the axis.
- ✓ **π :** sideways p overlap, density above and below.
- ✓ **sp^3** 109.5° , **sp^2** 120° , **sp** 180° .
- ✓ **π is weaker** and blocks rotation.

Watch out!

Count σ and π bonds in a whole molecule by counting **every** bond as one σ , then adding one π for each double bond and two for each triple bond. Ethene, $H_2C=CH_2$, has 5 σ and 1 π , not 4 and 2. Forgetting the C-H bonds is the most common slip.

6 Shapes of molecules

Syllabus 3.5

In **VSEPR theory** the electron pairs around a central atom repel one another and arrange themselves as far apart as possible, which fixes the shape. Count the bonding pairs (a double or triple bond counts as one region) and the lone pairs. Lone pairs are held closer to the nucleus and repel more than bonding pairs: lone pair–lone pair > lone pair–bond pair > bond pair–bond pair. Each lone pair closes the remaining bond angles by roughly 2.5° , so CH_4 is 109.5° , NH_3 107° and H_2O 104.5° . The shape is named from the positions of the **atoms**, not the lone pairs.

bond pairs	lone pairs	shape	angle	e.g.
2	0	linear	180°	CO_2
3	0	trigonal planar	120°	BF_3
4	0	tetrahedral	109.5°	CH_4
3	1	pyramidal	107°	NH_3
2	2	non-linear (bent)	104.5°	H_2O
5	0	trigonal bipyramidal	$90^\circ, 120^\circ$	PCl_5
6	0	octahedral	90°	SF_6

each lone pair squeezes the bond angle by about 2.5°

Fig. 1: the shapes and bond angles you need, by number of bonding and lone pairs

Count pairs, then name the shape from the atoms.

- ✓ **Pairs:** outer electrons + one per bond, adjusted for charge, halved.
- ✓ **Repulsion:** lone-lone > lone-bond > bond-bond.
- ✓ **Double bond:** one region of electron density.
- ✓ **Angles:** $109.5^\circ \rightarrow 107^\circ \rightarrow 104.5^\circ$ as lone pairs are added.

Watch out!

When explaining a bond angle, always name **both** the number of bonding pairs and the number of lone pairs, and state that lone pairs repel more than bonding pairs. "Water is bent because of its lone pairs" is not enough for the marks. Also, name the shape from the atoms: NH_3 is pyramidal, not tetrahedral, even though its four pairs point towards the corners of a tetrahedron.

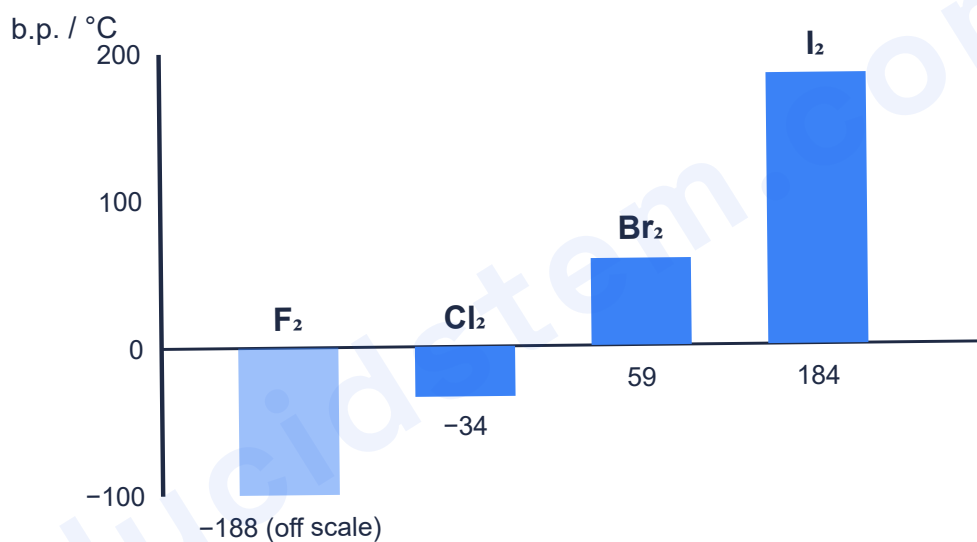
7 Intermolecular forces

Syllabus 3.6

Intermolecular forces act **between** molecules and are much weaker than covalent bonds.

Instantaneous dipole-induced dipole (id-id) forces act between all molecules: a momentary uneven electron distribution in one molecule induces a dipole in its neighbour. They grow with the number of electrons and with the contact area between molecules.

Permanent dipole-permanent dipole (pd-pd) forces act as well between polar molecules, whose $\delta+$ end of one attracts the $\delta-$ end of the next. A molecule with polar bonds is only polar if the bond dipoles do not cancel by symmetry. The strongest intermolecular force, hydrogen bonding, has its own lesson.



electrons per molecule: 18, 34, 70, 106, so stronger id-id forces

Fig. 1: boiling points of the halogens rise with the number of electrons per molecule

Name every force, then compare.

- ✓ **id-id:** all molecules; stronger with more electrons and more contact.
- ✓ **pd-pd:** polar molecules only, in addition to id-id.
- ✓ **Polar molecule:** polar bonds that do not cancel by symmetry.
- ✓ **Boiling** overcomes intermolecular forces, not covalent bonds.

Watch out!

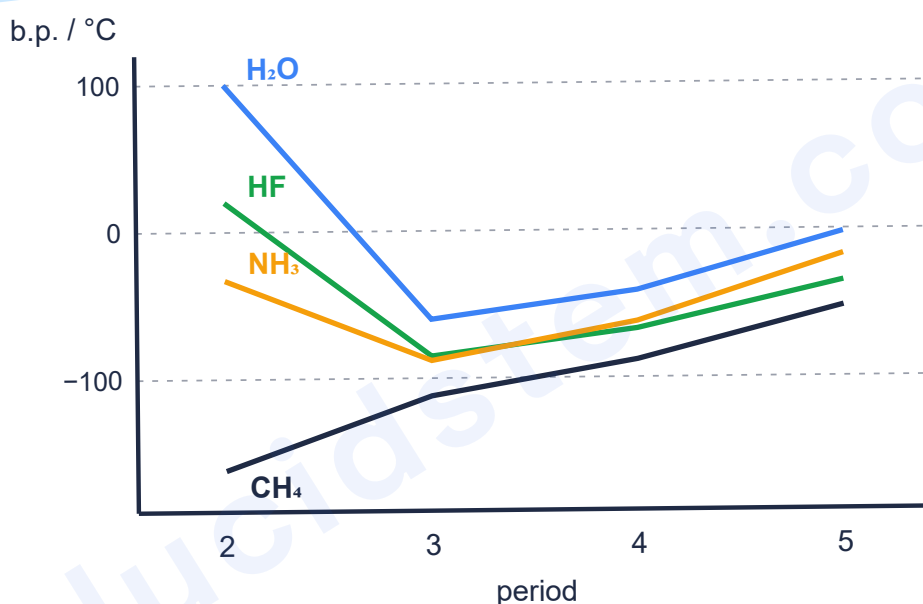
Never say that boiling a simple molecular substance "breaks covalent bonds". It overcomes **intermolecular forces**. And do not assume the more polar molecule always boils higher: id-id forces grow so quickly with electron number that they often dominate, as HBr against HCl shows.

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8 Hydrogen bonding and bond properties

Syllabus 3.6

A **hydrogen bond** forms between a hydrogen atom bonded to a very electronegative atom (N, O or F) in one molecule and a **lone pair** on N, O or F in a neighbouring molecule. The H carries a large $\delta+$ charge and is tiny, so the lone pair can approach it closely. Hydrogen bonds are the strongest intermolecular forces, which explains the unusually high boiling points of H_2O , HF and NH_3 and the open, less dense structure of ice. Inside molecules, **bond energy** is the energy needed to break one mole of a covalent bond in the gas phase; shorter bonds are generally stronger, so $\text{C}\equiv\text{C} > \text{C}=\text{C} > \text{C}-\text{C}$.



period 2 hydrides of N, O and F break the trend: hydrogen bonding

Fig. 1: boiling points of the hydrides of Groups 14 to 17

H on N, O or F, pointing at a lone pair.

- ✓ **Donor:** H covalently bonded to N, O or F (large δ^+).
- ✓ **Acceptor:** a lone pair on N, O or F in another molecule.
- ✓ **Draw:** partial charges, lone pair, dashed line in a straight line.
- ✓ **Shorter bond, stronger bond:** $C\equiv C > C=C > C-C$.

Watch out!

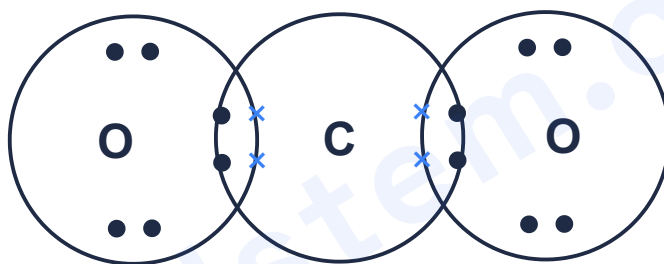
A hydrogen bond is an **intermolecular** force, not a covalent bond; the covalent O-H bond inside water is about 20 times stronger. When drawing hydrogen bonding you must show the lone pair and the partial charges, and the dashed line must go from the lone pair to the hydrogen, not to the other atom.

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9 Dot-and-cross diagrams

Syllabus 3.7

Dot-and-cross diagrams show outer-shell electrons only, using dots for one atom's electrons and crosses for the other's. For an **ionic** compound, draw each ion separately in square brackets with its charge, showing the electrons transferred. For a **covalent** molecule, place each shared pair in the overlap between atoms (one dot and one cross), two pairs for a double bond and three for a triple; draw lone pairs on the atom they belong to. In a **dative** bond both electrons of the shared pair carry the donor's symbol. Check the finish: every atom (except H, B and expanded-octet atoms) has eight outer electrons, and the total matches the number you started with, adjusted for any charge.



x carbon electrons, • oxygen electrons (outer shells only)
two double bonds, two lone pairs on each O: every atom has eight

Fig. 1: the dot-and-cross diagram of carbon dioxide

Outer electrons only, then check the count.

- ✓ **Ionic:** separate ions, square brackets, charges.
- ✓ **Covalent:** shared pairs in the overlap, lone pairs on their atom.
- ✓ **Dative:** both electrons of the pair use the donor's symbol.
- ✓ **Total:** sum of outer electrons, minus the charge.

Watch out!

Count the total before and after. NH_4^+ has $5 + 4 - 1 = 8$ outer electrons, which is four pairs and no lone pairs. The most common mistakes are leaving the lone pairs off the outer atoms, forgetting the square brackets and charge on an ion, and drawing a dot and a cross in a dative bond where both electrons should come from the donor.