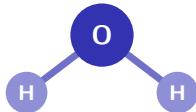


Chemical Bonding

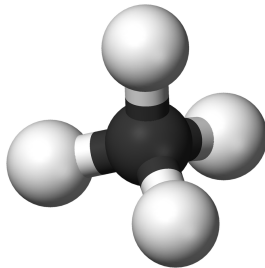
A-Level Chemistry ■ Topic 3

A-Level Chemistry



What we will cover

- 1 Electronegativity
- 2 Ionic & metallic bonding
- 3 Covalent & coordinate bonds
- 4 Sigma & pi bonds
- 5 Shapes of molecules
- 6 Intermolecular forces



molecules in 3-D



Electronegativity

Electronegativity is the power of an atom to attract bonding electrons towards itself. The difference in **Pauling electronegativity** values predicts the bond type.

nuclear charge

more protons pull the
pair more strongly

atomic radius

closer to the nucleus
⇒ stronger pull

shielding

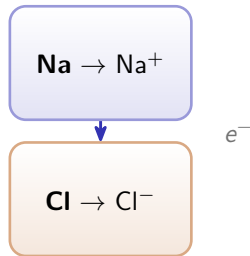
inner shells
weaken the pull

It **rises** across a period and **falls** down a group. Fluorine is the most electronegative element. A large difference ⇒ ionic; a small difference ⇒ covalent.

Ionic bonding

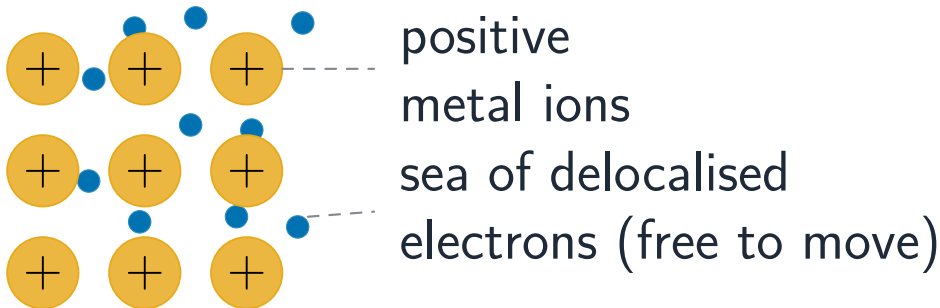
Ionic bonding – positive **cations** and negative **anions** – is the electrostatic attraction between oppositely charged **ions**. A metal gives electrons to a non-metal. The ions pack into a giant **lattice** held in every direction.

NaCl: sodium transfers its outer electron to chlorine $\rightarrow \text{Na}^+$ and a full-octet Cl^- .



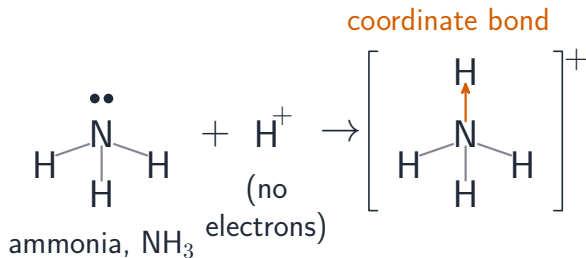
Metallic bonding

Metallic bonding is the attraction between positive metal ions and a sea of **delocalised electrons**. The free electrons explain why metals conduct and are strong.





Covalent and coordinate bonds



A **covalent bond** is a shared pair of electrons.

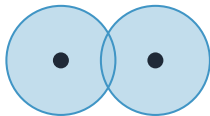
In a **coordinate bond** (dative) **both** electrons come from one atom – e.g. N's lone pair makes NH_4^+ .

Sigma and pi bonds

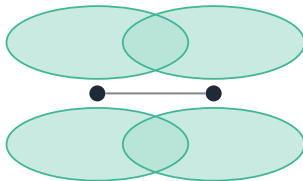
Covalent bonds form when **orbitals** overlap:

- a **sigma bond** σ – head-on **overlap**
- a **pi bond** π – sideways p-orbital overlap

Double = $1\sigma + 1\pi$; triple = $1\sigma + 2\pi$. **Hybridisation** : sp , sp^2 , sp^3 .



σ bond:
head-on overlap



π bond: sideways overlap
of p orbitals

Bond energy and bond length

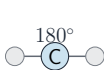
- **bond energy** is the energy needed to break one mole of a particular covalent bond in the gas state.
- **bond length** is the distance between the centres of the two bonded atoms.

Shapes of molecules

VSEPR theory : electron pairs push apart as far as possible.

A **lone pair** repels harder than a bonding pair, squeezing the **bond angle** ($\sim 2.5^\circ$ each) – so NH_3 is 107° , H_2O 104.5° .

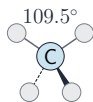
Shapes of molecules, in detail



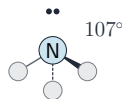
linear
CO2



trigonal planar
BF3



tetrahedral
CH4



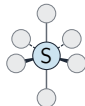
pyramidal
NH3



bent
H2O



trig. bipyramidal
PF5, 120° & 90°



octahedral
SF6, 90°

The seven shapes from VSEPR theory; the lone pairs on NH3 and H2O push harder, squeezing the bond angle below 109.5°

The shapes, and the angles that come with them

Linear

CO_2 – two bonding regions, 180° .

Trigonal planar

BF_3 – three regions, 120° , all in one plane.

Tetrahedral

CH_4 – four regions, 109.5° . Lone pairs push harder: NH_3 is pyramidal at 107° , H_2O bent at 104.5° .

Trigonal bipyramidal

PF_5 – five regions, 120° in the plane and 90° to the axis.

Octahedral

SF_6 – six regions, all 90° .

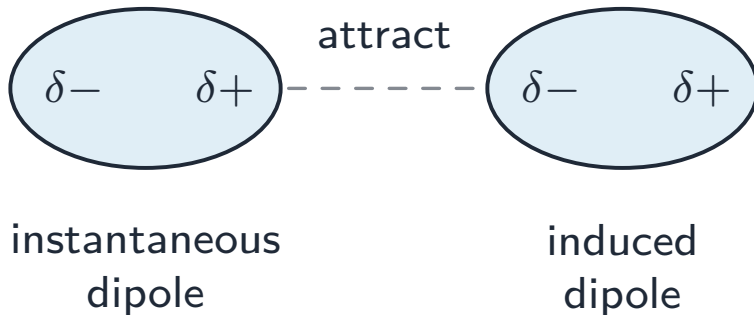
Count the regions of electron density around the central atom, then subtract for lone pairs.
The last two shapes need an atom that can **expand the octet**.



Intermolecular forces

A bond is **polar** if the atoms differ in electronegativity (a **dipole**).

- **London forces** – between all molecules; stronger with more electrons
- permanent dipole forces – between polar molecules



Van der Waals' forces, and when a molecule is polar

Van der Waals' forces

Van der Waals' forces is the general name for all intermolecular forces. There are two main types.

London dispersion

The instantaneous dipole-induced dipole force, or **London dispersion force**. Moving electrons make a brief **instantaneous dipole**, which induces a matching **induced dipole** next door. It acts between **all** molecules.

Permanent dipole

Between molecules that are always polar, because each carries a **permanent dipole**.

Dipole moment

If the bond dipoles in a molecule do **not** cancel, the whole molecule has a **dipole moment** and is polar – which is why CO_2 is not, and H_2O is.

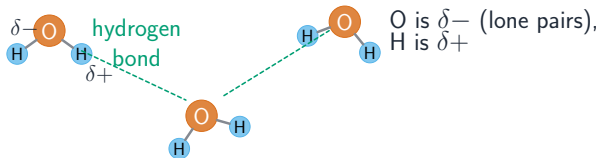
Shape decides polarity: identical bonds arranged symmetrically cancel, however polar each bond is on its own.

Bond polarity and dipoles

When two atoms with different electronegativity share a bond, the electrons sit closer to the more electronegative atom.

The bond then has a **polarity** : one end is slightly negative ($\delta-$) and the other slightly positive ($\delta+$).

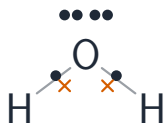
Hydrogen bonding



Hydrogen bonding forms when H is bonded to N, O or F and meets a lone pair on N, O or F.

It gives water a high **boiling point**, high **surface tension**, and makes ice **less dense** (so it floats).

Dot-and-cross diagrams



H₂O: 2 bonding pairs
+ 2 lone pairs

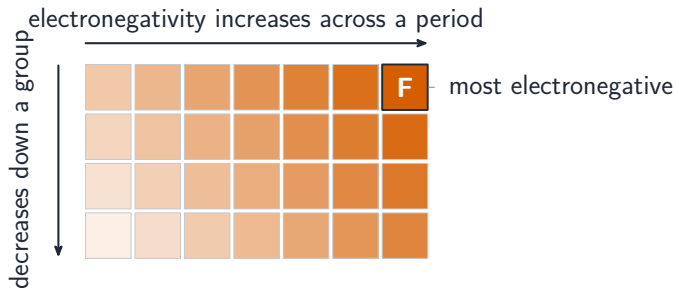


N₂: triple bond
(3 shared pairs)

- electron from one atom
- x electron from the other

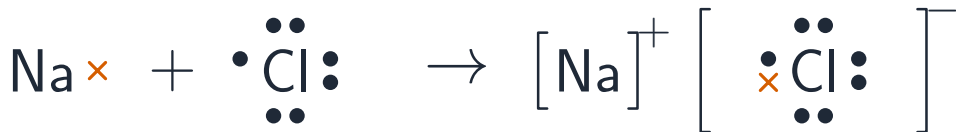
Covalent dot-and-cross: each shared pair is one electron from each atom. Water has two bonding pairs and two lone pairs; nitrogen shares three pairs (a triple bond)

Electronegativity, continued



Electronegativity rises across a period and falls down a group, so fluorine is the most electronegative element

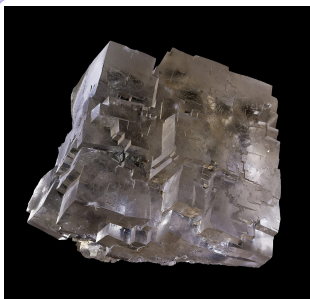
Ionic bonding, continued



sodium gives its outer electron (×) to chlorine

ionic bonding in NaCl: sodium transfers its single outer electron to chlorine, giving Na^+ and a full-octet Cl^-

Ionic bonding, continued (2)



A real crystal of rock salt (halite, NaCl); the cubic shapes mirror the giant ionic lattice inside

Exam tips

- For **shapes**, count bonding pairs and lone pairs, name the shape, then give the exact bond angle (e.g. NH_3 : pyramidal, 107°) – each lone pair lowers the angle by about 2.5° .
- A **dative (coordinate) bond** has both electrons from one atom (e.g. NH_4^+ , H_3O^+); draw the arrow from the lone pair.
- Name the intermolecular force precisely: **hydrogen bonding** needs H bonded to N, O or F; otherwise it is permanent-dipole or induced-dipole (van der Waals). Never call van der Waals forces “bonds”.
- Explain a physical property by stating **which forces are broken**, not just “strong bonds”.



Worked example – shape of a molecule

Predict the shape and bond angle of NH_3 .

- N has 5 outer electrons: 3 bond pairs + 1 **lone pair** = 4 electron pairs.
- 4 pairs repel to a **tetrahedral** arrangement (109.5°).
- A lone pair repels **more** than a bond pair, so the angle closes.
- Name the shape from the **atoms only**: **pyramidal**, $\approx 107^\circ$.

Count electron **pairs** for the arrangement, then name the shape from the **atoms**. Each lone pair closes the angle by about 2.5° .

Topic 3 – key takeaways

1

Bond types

ionic, metallic, covalent (& coordinate)

2

Sigma & pi

single 1σ ; double $1\sigma 1\pi$; triple $1\sigma 2\pi$

3

Shapes

VSEPR; lone pairs squeeze the angle

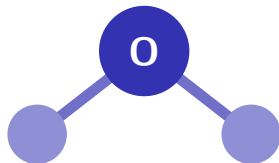
4

Forces

London, dipole, hydrogen bonding

Check yourself

- 1 Which is the most electronegative element?
- 2 In a coordinate bond, where do both electrons come from?
- 3 How many sigma and pi bonds in a triple bond?
- 4 Why does ice float on water?



Answers 1. fluorine 2. one atom (a lone pair) 3. one σ and two π 4. hydrogen bonds hold an open structure (less dense)