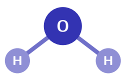
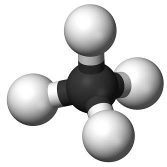
Chemical Bonding

A-Level Chemistry ▪ Topic 3

A-Level Chemistry



What we will cover



molecules in 3-D

1 Bonding types

Electronegativity

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

nuclear charge
more protons pull the pair more strongly

atomic radius
closer to the nucleus ⇒ stronger pull

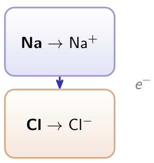
shielding
inner shells weaken the pull

> 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

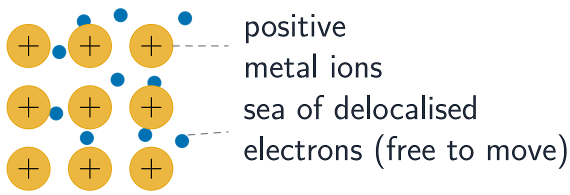
Ionic bonding 离子键 – positive **cations** 阳离子 and negative **anions** 阴离子 – is the electrostatic attraction between oppositely charged **ions** 离子. A metal gives electrons to a non-metal. The ions pack into a giant **lattice** 晶格 held in every direction.

> NaCl: sodium transfers its outer electron to chlorine → Na^+ and a full-octet Cl^-.



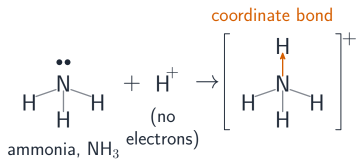
Metallic bonding

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



2 Covalent bonding

Covalent and coordinate bonds



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

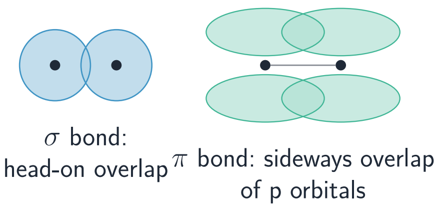
> In a **coordinate bond** 配位键 (dative) both electrons come from one atom – e.g. N's lone pair makes NH_4^+.

Sigma and pi bonds

Covalent bonds form when **orbitals** 轨道 overlap:

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

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



Bond energy and bond length

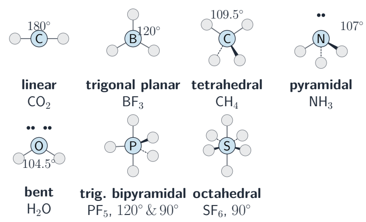
- **bond energy 键能** is the energy needed to break one mole of a particular covalent bond in the gas state.
- **bond length 键长** is the distance between the centres of the two bonded atoms.

Shapes of molecules

VSEPR theory 价层电子对互斥理论: electron pairs push apart as far as possible.

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

Shapes of molecules, in detail



> The seven shapes from VSEPR theory; the lone pairs on NH_3 and H_2O 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 $90°$.

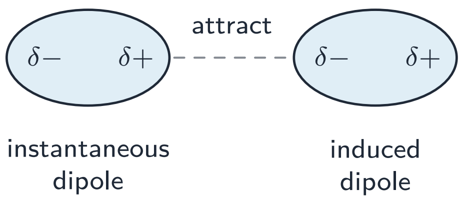
> Count the regions of electron density around the central atom, then subtract for lone pairs. The last two shapes need an atom that can expand the octet.

3 Intermolecular forces

Intermolecular forces

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

- **London forces** 伦敦色散力 – 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** 范德华力 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** 伦敦色散力. Moving electrons make a brief **instantaneous dipole** 瞬时偶极, which induces a matching **induced dipole** 诱导偶极 next door. It acts between all molecules.
Permanent dipole	Between molecules that are always polar, because each carries a **permanent dipole** 永久偶极.
Dipole moment	If the bond dipoles in a molecule do not cancel, the whole molecule has a **dipole moment** 偶极矩 and is polar – which is why CO_2 is not, and H_2O is.

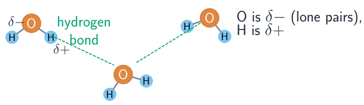
> Shape decides polarity: identical bonds arranged symmetrically cancel, however polar each bond is on its own.

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** 极性: one end is slightly negative ($\delta-$) and the other slightly positive ($\delta+$).

Hydrogen bonding



Hydrogen bonding 氢键 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** 沸点, high **surface tension** 表面张力, and makes ice less dense (so it floats).

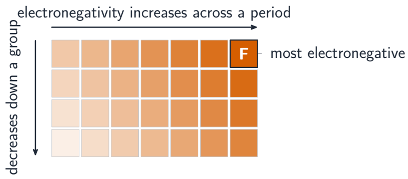
Dot-and-cross diagrams

H_2O: 2 bonding pairs + 2 lone pairs

N_2: triple bond (3 shared pairs)

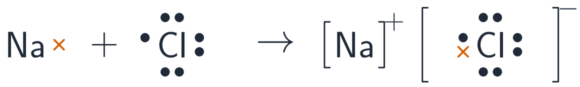
> Covalent dot-and-cross: each shared pair is one electron from each atom. Water has two bonding pairs and two lone pairs; nitrogen shares three pairs (a triple bond)

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".

4 Recap

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** 孤对电子 = 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, $\approx 107°$.

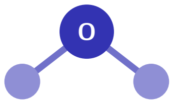
Count electron pairs for the arrangement, then name the shape from the atoms. Each lone pair closes the angle by about $2.5°$.

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)