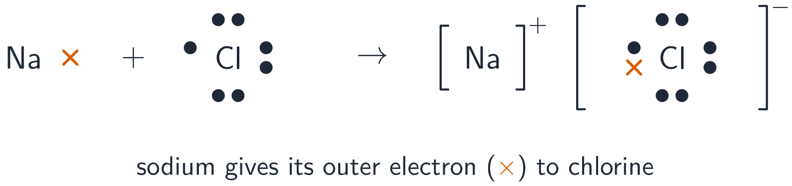
Compound Structure and Properties

AP Chemistry

Types of Chemical Bonds

A **chemical bond** 化学键 is an attraction that holds atoms together. Which type forms depends on the atoms' electronegativities:



Ionic bonding: a metal transfers its outer electrons to a non-metal

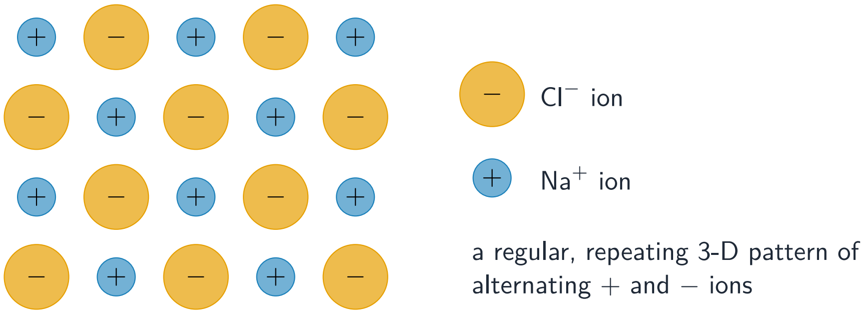
- **Ionic bond** 离子键: electrons transfer from a metal to a nonmetal (large electronegativity difference).
- **Covalent bond** 共价键: nonmetals **share** electrons (small difference). A big-but-not-huge difference gives a **polar covalent** 极性共价 bond.
- **Metallic bond** 金属键: metal atoms share a ”sea” of mobile electrons.

Intramolecular Force and Potential Energy

As two atoms approach, a **potential energy** 势能 curve captures the balance of attraction and repulsion. It dips to a minimum at the **bond length** 键长 (the stable separation) whose depth is the **bond energy** 键能. Shorter, stronger bonds sit in deeper, tighter wells; more shared pairs (double, triple bonds) give shorter, stronger bonds.

Structure of Ionic Solids

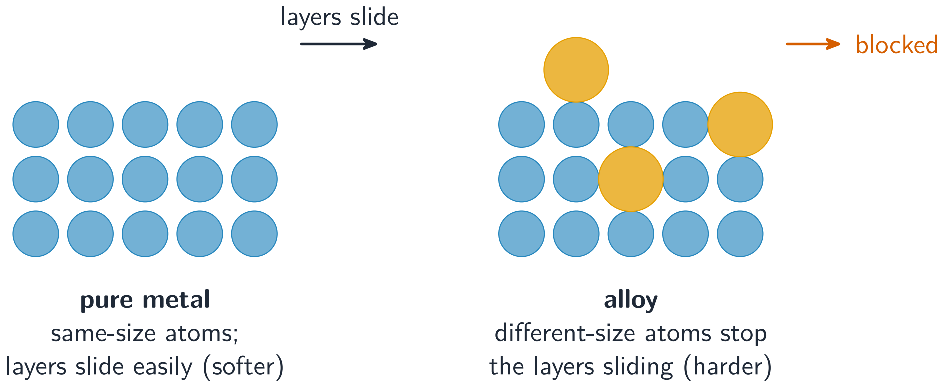
An **ionic solid** 离子固体 is a repeating 3-D **lattice** 晶格 of alternating cations and anions, held by strong electrostatic attraction. This explains their high melting points, brittleness, and why they conduct only when molten or dissolved (ions freed to move). The **lattice energy** rises with larger ion charges and smaller ions, so MgO (both $2+/2-$) melts far higher than NaCl (both $1+/1-$).



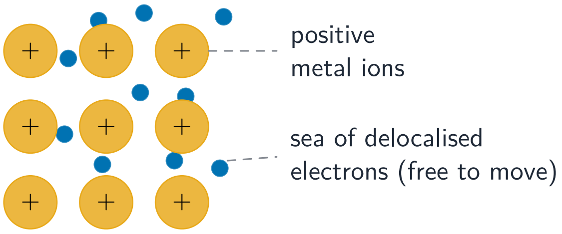
Ions pack into a giant lattice of alternating positive and negative ions

Structure of Metals and Alloys

In a metal, cations sit in a lattice bathed in **delocalized** 离域 electrons, which explains conductivity, malleability, and luster. An **alloy** 合金 mixes metals: a **substitutional** alloy swaps in similar-sized atoms; an **interstitial** alloy (like steel) fits small atoms into the gaps, making it harder.



Different-sized atoms in an alloy stop the layers sliding, so it is harder

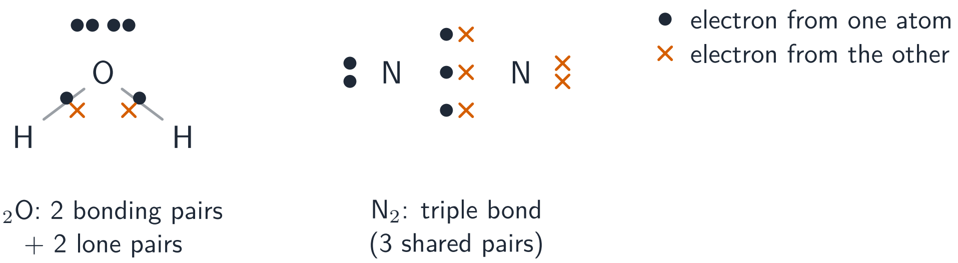


Metallic bonding: positive ions in a sea of delocalised electrons

Lewis Diagrams

A **Lewis diagram** 路易斯结构 shows valence electrons as bonding pairs and **lone pairs** 孤对电子, giving most atoms an **octet** 八隅体 (8 valence electrons; H wants 2). Steps:

count total valence electrons, connect atoms with single bonds, complete octets on outer atoms, then form multiple bonds if the central atom is short.



H_2O: 2 bonding pairs + 2 lone pairs

N_2: triple bond (3 shared pairs)

Dot-and-cross diagrams show the bonding pairs and lone pairs in a molecule

Worked example. Draw carbon dioxide, CO_2. Total valence electrons $= 4 + 2(6) = 16$. Put C in the centre; single bonds to each O use 4 electrons and leave the outer O atoms short. Completing octets forces two **double** bonds, $O = C = O$: each O then has two lone pairs, C has none, and all 16 electrons are placed with every atom at an octet.

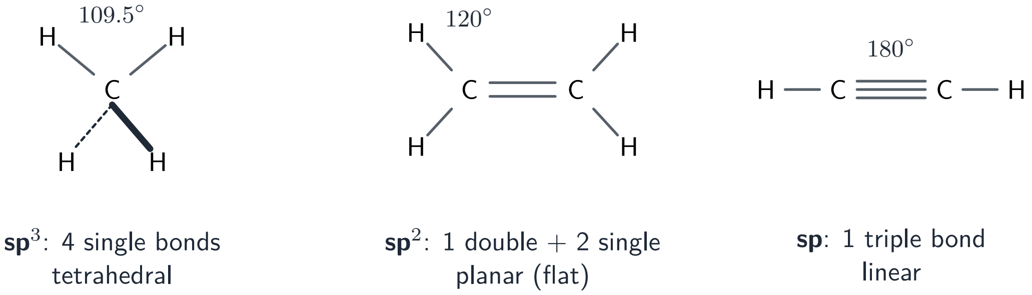
Resonance and Formal Charge

When two or more valid Lewis diagrams differ only in electron placement, the true structure is an average –**resonance** 共振. **Formal charge** 形式电荷 (valence electrons minus lone-pair electrons minus half the bonding electrons) picks the best structure: the one with formal charges closest to zero, and any negative charge on the most electronegative atom.

Worked example. Assign formal charges in the nitrate ion, NO_3^- (one double bond, two single bonds). For N (4 bonds, no lone pairs): $5 - 0 - 4 = +1$. For the double-bonded O (2 lone pairs): $6 - 4 - 2 = 0$. For each single-bonded O (3 lone pairs): $6 - 6 - 1 = -1$. The total is $+1 + 0 + (-1) + (-1) = -1$, matching the ion's overall charge –a good check that the structure is drawn correctly. Because the three O atoms are equivalent by resonance, the real ion has three identical bonds.

VSEPR and Bond Hybridization

VSEPR 价层电子对互斥 theory predicts shape: electron pairs (bonds and lone pairs) around a central atom spread out as far apart as possible. Counting electron domains gives the geometry (linear, trigonal planar, tetrahedral, $\cdots$); lone pairs push bonds closer, bending the shape. **Hybridization** 杂化 (sp, sp^2, sp^3) describes the mixed orbitals matching that geometry. Molecular shape and bond polarity together decide whether the whole molecule is polar.

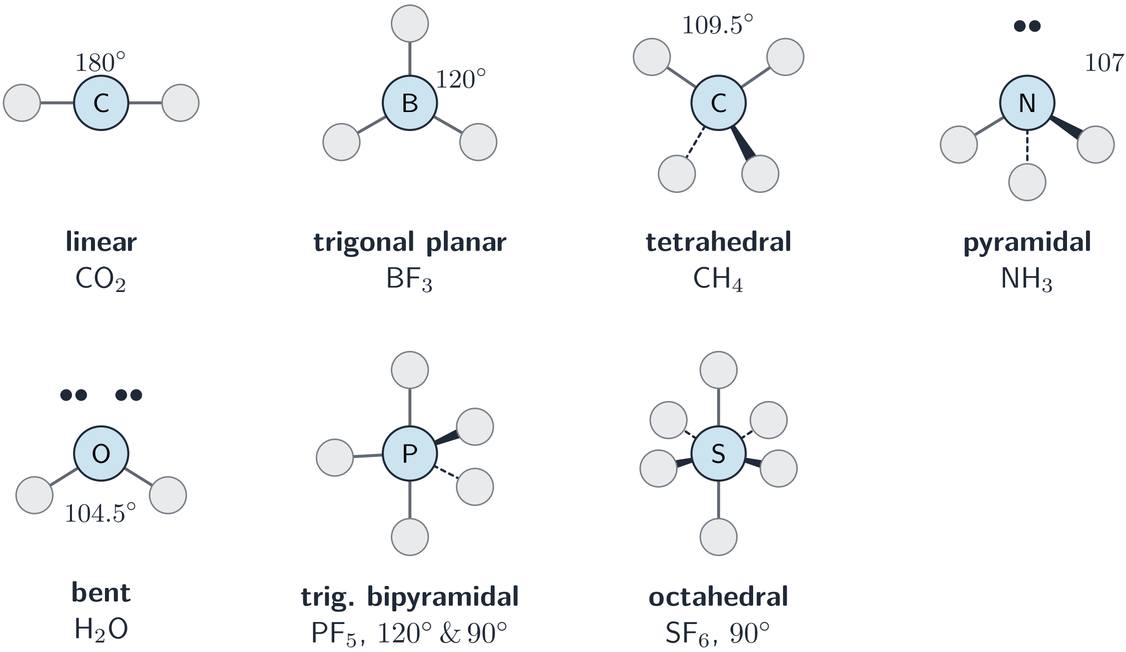


sp^3: 4 single bonds
tetrahedral

sp^2: 1 double + 2 single
planar (flat)

sp: 1 triple bond
linear

Hybridisation and shape: sp3 tetrahedral, sp2 planar, sp linear



linear
CO_2

trigonal planar
BF_3

tetrahedral
CH_4

pyramidal
NH_3

bent
H_2O

trig. bipyramidal
PF_5, 120° & 90°

octahedral
SF_6, 90°

The common VSEPR shapes and their bond angles

Worked example. Predict the shape of ammonia, NH_3. Nitrogen has 3 bonding pairs and 1 lone pair –four electron domains, so the electron geometry is tetrahedral and the hybridization is sp^3. The lone pair is invisible in the shape but still pushes the bonds together, so the **molecular** shape is trigonal pyramidal with a bond angle of about 107° (a little less than the ideal 109.5°). The three N–H dipoles do not cancel, so the molecule is polar.

Exam tips

- Decide the bond type from the atoms: **ionic** (metal + non-metal, electron transfer), **covalent** (non-metals, sharing), **metallic** (sea of delocalised electrons).
- An ionic solid conducts only when **molten or dissolved** (ions free to move), never as a solid.
- Draw **Lewis structures** to satisfy the octet (H wants 2), then use **VSEPR** —electron pairs spread as far apart as possible —to predict the shape.

- **Lone pairs** take up space and push bonds closer, bending the shape (water is bent, ammonia pyramidal).
- A molecule can have polar bonds yet be **non-polar overall** if its symmetry makes the dipoles cancel (CO_2).