

Ionic and Covalent Bonds

What is the difference between ionic and covalent bonds?

●○○ Beginner

Bonding & Molecular Structure

Last reviewed 3 October 2026

Before you start: it helps to understand [Inside the Atom: Protons, Neutrons and Electrons](#) and [Electron Configuration](#).

What is it?

A **chemical bond** is the attraction that holds atoms together. Atoms bond because the bonded arrangement has **lower energy** than the separate atoms. There are two main types:

- **Ionic bond:** one atom **transfers** electrons to another, forming oppositely charged **ions** that attract each other. Typical between a **metal** and a **non-metal**, e.g. sodium chloride, NaCl.
- **Covalent bond:** two atoms **share** a pair of electrons. Typical between **non-metals**, e.g. the H-Cl bond in hydrogen chloride.

🔑 KEY IDEA

Ionic bonding = electron **transfer** and attraction between ions. Covalent bonding = electron **sharing**. The difference in electronegativity between the two atoms tells you which is more likely.

Why does it matter?

- **It explains properties.** Salt melts at 801 °C, but sugar melts (and starts to caramelise) below 200 °C. The difference comes from the type of bonding.
- **It predicts formulas.** Knowing ion charges lets you write formulas such as MgCl₂ and Al₂O₃.
- **It is the starting point for molecular structure.** Lewis structures and molecular shapes build on covalent bonding.

How does it work?

1. Ionic bonding

Metals have low ionization energies and lose electrons easily; non-metals gain electrons. Each atom tends to reach a **noble-gas configuration**:



The ions Na^+ and Cl^- attract strongly. In a solid, each ion is surrounded by ions of the opposite charge in a repeating 3D **lattice**. There are no separate “molecules” of NaCl; the formula gives the **ratio** of ions.

Writing formulas: the total positive charge must equal the total negative charge.

Ions	Formula
Mg^{2+} and Cl^-	MgCl_2
Al^{3+} and O^{2-}	Al_2O_3
Ca^{2+} and N^{3-}	Ca_3N_2

2. Covalent bonding

Two non-metal atoms both attract electrons strongly, so neither gives one up. Instead they **share** a pair of electrons, and both nuclei attract the shared pair. Atoms can share one, two or three pairs: **single**, **double** and **triple** bonds. More shared pairs make a bond **shorter and stronger**.

3. Polar covalent bonds and electronegativity

Electronegativity (EN) is how strongly an atom attracts the shared electrons in a bond. When the two atoms have different electronegativities, the electrons are shared **unequally**: the more electronegative atom gets a partial negative charge (δ^-), the other a partial positive charge (δ^+). This is a **polar covalent** bond.

The **electronegativity difference** (ΔEN) gives a rough guide:

ΔEN	Bond type	Example
less than about 0.4	nonpolar covalent	C-H, $2.55 - 2.20 = 0.35$
about 0.4 to 1.8	polar covalent	O-H, $3.44 - 2.20 = 1.24$
more than about 1.8	ionic	Na-Cl, $3.16 - 0.93 = 2.23$

∞ THINK OF IT LIKE THIS

Think of a tug of war over a rope (the electron pair). Two equally strong players: the rope stays in the middle (nonpolar). One player a bit stronger: the rope moves towards them (polar). One player far stronger: they take the rope completely (ionic).

Q MORE PRECISELY

Bonding is a continuum, not three separate boxes. Even NaCl has a little covalent character, and the cut-off values are only guides; textbooks use slightly different numbers. A third type, **metallic bonding** (metal cations in a “sea” of shared, delocalised electrons), holds metals together and is covered separately.

4. Properties compared

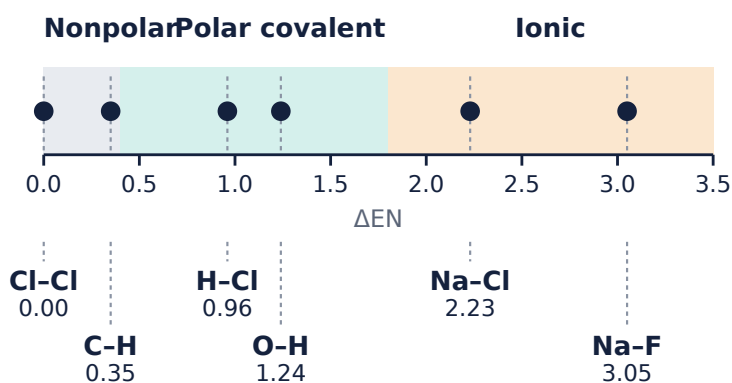
Property	Ionic compounds	Covalent (molecular) substances
Melting point	high	usually low
State at room temperature	solid	often gas or liquid
Conducts electricity	when molten or dissolved	usually not
Solubility in water	often soluble	varies

Ionic solids have high melting points because many strong attractions between ions must be overcome. Molecular substances melt easily because only the weak forces **between** molecules break; the covalent bonds inside the molecules stay intact.

Visualise it

Bond type and electronegativity difference

ΔEN = difference in Pauling electronegativity



The cut-offs (about 0.4 and 1.8) are approximate and vary between textbooks: bond character changes gradually.

Bond type changes gradually as the electronegativity difference grows.

Worked example

WORKED EXAMPLE: Classifying bonds

Question: Classify the bonds in H-F, C-O and K-Br. (EN: H 2.20, F 3.98, C 2.55, O 3.44, K 0.82, Br 2.96)

Step 1 H-F: $3.98 - 2.20 = 1.78$, **polar covalent** (close to the boundary)

Step 2 C-O: $3.44 - 2.55 = 0.89$, **polar covalent**, with O δ^- and C δ^+

Step 3 K-Br: $2.96 - 0.82 = 2.14$, **ionic**

WORKED EXAMPLE: Writing an ionic formula

Question: Write the formula of the compound formed by aluminium and sulfur.

Step 1 Al is in group 13 and forms Al^{3+} ; S is in group 16 and forms S^{2-} .

Step 2 Lowest common multiple of 3 and 2 is 6: $2 \times (+3) = +6$ and $3 \times (-2) = -6$.

Step 3 Formula: Al_2S_3

Common mistake

COMMON MISTAKE: Calling NaCl a molecule

Ionic compounds form lattices, not molecules. NaCl is a **formula unit** showing a 1 : 1 ratio of ions.

COMMON MISTAKE: Thinking melting breaks covalent bonds

When ice or wax melts, the covalent bonds inside the molecules don't break; only the attractions **between** molecules are overcome.

COMMON MISTAKE: Using metal/non-metal as an absolute rule

The metal + non-metal rule is a good first guess, but it has exceptions. For example, AlCl_3 has considerable covalent character. Use ΔEN when the values are given.

NOTATION NOTE

- δ^+ and δ^- ("delta plus", "delta minus") mean **partial** charges, smaller than a full charge.
- ΔEN is always taken as a positive number (larger EN minus smaller EN).

Remember this

REMEMBER THIS

- **Ionic:** electrons transferred, ions in a lattice, usually metal + non-metal.
- **Covalent:** electron pairs shared, usually non-metal + non-metal.
- ΔEN below about 0.4: nonpolar; 0.4 to 1.8: polar covalent; above about 1.8: ionic.
- Ionic compounds have high melting points and conduct when molten or dissolved.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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