

Topic 2.1: Types of Chemical Bonds

Revision Notes · AP Chemistry

UNIT 2 · COMPOUND STRUCTURE AND PROPERTIES

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Why Bonds Form At All

Atoms bond because the bonded arrangement has lower potential energy than the separated atoms. Electrons and nuclei arrange themselves into the configuration where attractions are maximized and repulsions are minimized, and that arrangement holds together.

Every bond type in this topic is the same story told with different electron behaviour. What changes is what the valence electrons do: they are transferred, shared, or pooled. The type of bond follows from the atoms involved, and the properties of the substance follow from the type of bond.

Ionic Bonding: Electrons Transferred

An ionic bond forms between a metal and a nonmetal when valence electrons are transferred from the metal to the nonmetal, as covered in Topic 1.8. The resulting oppositely charged ions attract each other electrostatically.

The ions do not pair off into molecules. They pack into a three dimensional lattice in which every ion is surrounded by ions of the opposite charge, and the attraction operates in all directions at once. That is why ionic compounds are described by formula units, not molecules.

Strength of the attraction: Coulomb again. The force between ions follows Coulomb's law: it grows with the product of the charges and shrinks with the square of the distance between ion centres. Small, highly charged ions produce the strongest attractions, the largest lattice energies, and the highest melting points. MgO ($2+$ and $2-$ ions, both small) melts around $2,850\text{ }^{\circ}\text{C}$; NaCl ($1+$ and $1-$, larger) melts at $801\text{ }^{\circ}\text{C}$.

Covalent Bonding: Electrons Shared

A covalent bond forms between nonmetals when atoms share pairs of valence electrons. Each shared pair sits between the two nuclei and attracts both, holding the atoms together.

Sharing lets each atom count the shared electrons toward its own outer shell, so both atoms can reach a full shell without the energy cost of fully transferring electrons. Discrete groups of covalently bonded atoms are molecules, and the attractions inside a molecule (the bonds) are far stronger than the attractions between molecules.

Bond character is a continuum. If the two atoms pull the shared pair equally, the bond is nonpolar covalent, as in H_2 or Cl_2 . If one atom pulls harder, the pair shifts toward it and the bond is polar covalent, with a partial negative charge (δ^-) on the stronger pulling atom and δ^+ on the other, as in HCl . Push the difference far enough and the pair is effectively transferred, which is the ionic end of the same scale.

Electronegativity difference is the guide. The larger the difference in electronegativity between the two atoms, the more polar the bond. Identical atoms give a difference of zero and a purely nonpolar bond. Differences typical of metal plus nonmetal pairs are large, matching electron transfer; differences between two nonmetals are smaller, matching sharing.

Metallic Bonding: Electrons Pooled

In a metal, the valence electrons are not attached to any one atom. The structure is a lattice of positive metal ions sitting in a shared pool, often called a sea, of delocalized electrons that move through the whole piece of metal.

This one model explains the classic metal properties. The delocalized electrons carry charge, so metals conduct electricity. They also carry thermal energy, so metals conduct heat. The ions can slide past each other while the electron sea keeps holding them, so metals are malleable and ductile rather than brittle. And the sea absorbs and re-emits light across wavelengths, giving metals their lustre.

Telling the Bond Types Apart From Properties

Melting point. Ionic lattices and metallic structures generally melt high (every particle is held by strong attractions in all directions). Molecular covalent substances melt low, because melting only has to overcome the weak attractions between molecules, not the covalent bonds inside them. Network covalent substances such as diamond and SiO_2 are the exception that melts extremely high, because their covalent bonds run through the whole structure.

Conductivity. Ionic compounds conduct when molten or dissolved (the ions are free to move) but not as solids (the ions are locked in place). Metals conduct as solids, thanks to the mobile electron sea. Molecular substances do not conduct in any state, having neither ions nor delocalized electrons.

Solubility and brittleness. Many ionic compounds dissolve in water, separating into ions. Ionic solids are brittle: sliding a layer shifts ions so that like charges align, and repulsion shatters the lattice. Metals bend instead, because the electron sea adjusts as the ions move.

Traps That Cost Marks in Topic 2.1

Treating bond type as three sealed boxes. Bond character runs continuously from nonpolar covalent through polar covalent to ionic. Exam questions reward the continuum picture: state where on the scale the pair sits and why, using electronegativity difference.

Saying ionic compounds contain molecules. They contain ions in a lattice. NaCl is a formula unit, the lowest ratio, not a molecule, and any explanation that gives NaCl discrete molecules loses the thread of its properties (high melting point, brittleness, conductivity when molten).

Claiming covalent bonds break when a molecular substance melts or boils. Melting ice or boiling water overcomes the attractions between molecules. The covalent bonds inside the molecules survive intact, which is exactly why molecular melting points are low while the molecules themselves are stable.

Judging bond strength from charge alone in lattice comparisons. Coulomb's law has two variables. A $2+/2-$ compound with large ions can have a smaller lattice energy than a $2+/2-$ compound with small ions. Compare charges first, then sizes, and say both.

Explaining metallic properties with "free electrons" and stopping there. The marks sit in the mechanism: the electrons are delocalized across the whole structure, they move under a voltage, and the bonding persists while ion layers slide. Name the sea, then the property it produces.

Key Relationships to Remember

- Bonds form because the bonded state has lower potential energy than the separated atoms
- Ionic: electrons transferred, metal + nonmetal, ions in a lattice; attraction strength follows Coulomb's law (charge product up, ion size down, stronger)
- Covalent: electrons shared between nonmetals, forming molecules; bond polarity rises with electronegativity difference (0 = nonpolar, large = ionic end of the continuum)
- Metallic: valence electrons delocalized in a sea around metal cations; explains conductivity, malleability and lustre
- Properties diagnose bond type: conductivity as solid and in melt, melting point, brittleness, whether the unit is a molecule or a formula unit

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