

Topic 2.1: Types of Chemical Bonds

Worksheet · AP Chemistry

Unit 2 · Compound Structure and Properties

On ExamFitter.com: [Topic 2.1 Revision Notes](#) · [Topic 2.1 Quiz](#) · [Previous Worksheet: Topic 1.8](#) ·
[Next Worksheet: Topic 2.2](#) · [Unit 2: All Topics](#) · [AP Chemistry Home](#)

How to Use This Worksheet

Work through the sections in order. Section A checks the ideas, Section B builds classification and comparison fluency, and Section C mixes everything the way an exam would.

Keep the continuum in mind throughout: bonds are not three sealed boxes but a scale from nonpolar covalent to ionic, set by electronegativity difference, and every property question is answered by asking what particles are present and how they are held.

Do not look at the answer key until you have attempted every question. The key gives full working, so compare your method, not just your final answer.

Section A: Recall and Concepts

A1. Explain why atoms bond at all, in terms of potential energy and the forces between electrons and nuclei.

A2. Describe what happens to the valence electrons in each of the three bond types: ionic, covalent and metallic.

A3. Explain why an ionic compound is described by a formula unit rather than a molecular formula.

A4. State what the symbols δ^+ and δ^- mean in a polar covalent bond, and state which atom in HCl carries each.

A5. A solid does not conduct electricity, but the same substance conducts well when molten. State its likely bond type and explain both observations.

Section B: Core Practice

B1. Classify the bond in each pair as ionic, polar covalent or nonpolar covalent, using the elements involved: (a) K and Br, (b) H and H, (c) C and O, (d) Mg and F, (e) N and H.

B2. Rank the bonds in H–F, H–Cl and H–H in order of increasing polarity, and justify the order.

B3. Using Coulomb's law, explain why CaO (Ca^{2+} , O^{2-}) has a much higher melting point than KCl (K^+ , Cl^-).

B4. Both NaF and MgO contain ions of similar size. State which compound has the larger lattice energy and explain why, referring to the charges.

.....

.....

B5. Explain, using the electron sea model, why a metal wire can be drawn out thin without breaking, while an ionic crystal shatters when hammered.

.....

.....

B6. A substance melts at 1,500 °C, conducts electricity as a solid, and is malleable. State its bond type and the evidence for your choice.

.....

.....

Section C: Exam Style

C1. A student claims: "When water boils, the covalent bonds in the water molecules break, which is why steam forms." (a) Identify the error. (b) State what actually breaks when water boils, and why boiling water needs so much less energy than breaking its molecules apart would require.

.....

.....

C2. Consider the two ionic compounds LiF and LiI. Both contain Li^+ . (a) State which has the larger lattice energy. (b) Justify your answer using Coulomb's law, identifying which variable differs and which are the same. (c) Predict which compound has the higher melting point.

.....

.....

C3. Substance X is a brittle solid with a melting point of 801 °C. It does not conduct electricity as a solid, but conducts when molten or in aqueous solution. Substance Y is a solid that conducts electricity at room temperature and can be hammered into sheets. (a) Identify the bond type of X and of Y. (b) Explain X's lack of conductivity as a solid and its conductivity when molten, in terms of its particles. (c) Explain why X is brittle, referring to what happens when its layers are shifted.

.....

.....

Answer Key

A1. When atoms approach, their electrons are attracted to the other atom's nucleus while electrons repel electrons and nuclei repel nuclei. At the bonded distance, the attractions outweigh the repulsions and the total potential energy of the system is lower than when the atoms are separate. Systems settle into the lowest energy arrangement available, so the bonded state persists: the bond is that energy minimum.

A2. Ionic: valence electrons are transferred from the metal atom to the nonmetal atom, producing oppositely charged ions that attract. Covalent: valence electrons are shared in pairs between two nonmetal atoms, each shared pair attracted to both nuclei. Metallic: valence electrons are given up to a shared, delocalized pool (a sea) that surrounds a lattice of metal cations and moves through the whole structure.

A3. An ionic compound contains no discrete molecules: each ion is attracted to all the oppositely charged ions around it in a continuous lattice, and there is no natural grouping of one cation with one anion. The formula records only the lowest whole number ratio of the ions, which is a formula unit, not the composition of a separate particle.

A4. They mark partial charges in a polar bond: the shared electron pair is pulled closer to the more electronegative atom, giving it a partial negative charge, δ^- , and leaving the other atom with a partial positive charge, δ^+ . In HCl, chlorine is far more electronegative than hydrogen, so Cl carries δ^- and H carries δ^+ .

A5. Ionic. As a solid its ions are locked in the lattice and cannot move to carry charge, so it does not conduct. On melting, the lattice breaks down and the ions become free to move toward the electrodes, so the liquid conducts. (A metallic solid would conduct in both states, and a molecular substance in neither.)

B1. (a) K + Br: ionic (metal + nonmetal, large electronegativity difference). (b) H + H: nonpolar covalent (identical atoms, difference zero). (c) C + O: polar covalent (two nonmetals, unequal pull). (d) Mg + F: ionic. (e) N + H: polar covalent.

B2. $\text{H}-\text{H} < \text{H}-\text{Cl} < \text{H}-\text{F}$. Polarity increases with electronegativity difference. H–H has a difference of zero. Fluorine is the most electronegative element, so H–F has the largest difference and the most polar bond; chlorine is less electronegative than fluorine but more than hydrogen, placing H–Cl in the middle.

B3. Two Coulomb factors both favour CaO. The charge product is $(2 \times 2) = 4$ for CaO against $(1 \times 1) = 1$ for KCl, so at the same separation the attraction would be four times stronger. In addition Ca^{2+} and O^{2-} are smaller ions than K^+ and Cl^- , so their centres sit closer together, strengthening the attraction further. The much stronger lattice needs far more thermal energy to break apart, hence the much higher melting point (about 2,570 °C against 770 °C).

B4. MgO. Its ions carry charges of +2 and -2 , giving a charge product of 4, while NaF's +1 and -1 give a product of 1. With distances similar, Coulomb's law says the attraction, and so the lattice energy, is much larger for MgO (roughly fourfold from the charge factor alone).

B5. In the metal, the positive ions sit in a sea of delocalized electrons. When the wire is drawn, layers of ions slide past each other, but the electron sea flows around them and the attraction between the ions and the sea is maintained in the new positions, so the structure holds together. In the ionic crystal, shifting a layer brings ions of the same charge face to face; the resulting repulsion pushes the layers apart and the crystal cleaves and shatters.

B6. Metallic bonding. Conducting as a solid rules out ionic (locked ions) and molecular covalent (no charge carriers); the high melting point also argues against a molecular substance. Malleability is the signature of the electron sea model: layers can slide while bonding persists.

C1. (a) The error: boiling does not break the covalent O–H bonds inside water molecules; steam consists of intact H_2O molecules, as the chemistry of condensation straight back to liquid water shows. (b) What breaks is the set of attractions between the molecules (intermolecular forces, including hydrogen bonds). These are far weaker than the covalent bonds within a molecule, so separating the molecules from each other takes much less energy than tearing the atoms inside them apart would.

C2. (a) LiF. (b) The charges are identical in both compounds (+1 and -1 , charge product 1), and the cation is the same. Only the distance differs: F^- is a much smaller ion than I^- (it has one fewer shell), so the ion centres in LiF sit closer together. Coulomb's law says attraction rises steeply as distance falls (inverse square), so LiF has the stronger attraction and the larger lattice energy. (c) LiF, because more thermal energy is needed to break down its more strongly bound lattice (in fact LiF melts at 845°C , LiI at 469°C).

C3. (a) X is ionic (the description matches NaCl exactly). Y is metallic. (b) X as a solid holds its ions in fixed lattice positions; with no mobile charge carriers it cannot conduct. When molten, the lattice collapses and the ions move freely, carrying charge to the electrodes. Dissolving does the same by separating the ions in water. (c) In the lattice, each ion is surrounded by opposite charges. Shifting a layer by one ion spacing brings like charges face to face across the shift; powerful repulsion then forces the layers apart along that plane, so the crystal splits cleanly instead of deforming: it is brittle.

Continue Revising

[Topic 2.1 Revision Notes](#)

[Topic 2.1 Quiz: Types of Chemical Bonds](#)

[Previous Worksheet: Topic 1.8 Valence Electrons and Ionic Compounds](#)

[Next Worksheet: Topic 2.2 Intramolecular Force and Potential Energy](#)