The ionic model

Structure 2.1 · SL and HL

Why does salt conduct only when it melts?

> Salt is made of charged particles, locked in place in the solid and free to move in the liquid.

- Solid salt does not conduct; molten or dissolved salt does.
- Melting mobilizes ions; dissolving adds water that solvates them.
- This unit: ions $\rightarrow$ bonds and formulas $\rightarrow$ lattices and lattice enthalpy $\rightarrow$ properties.

How ions form

Sodium $\rightarrow$ Na^+

[Ne] $3s^1$ loses one electron.

Na^+ is [Ne].

Chlorine $\rightarrow$ Cl^-

[Ne] $3s^2$ $3p^5$ gains one electron.

Cl^- is [Ar].

Predict the charge from the group

Group	Ion charge	Example
1, 2, 13	$+1$, $+2$, $+3$	Li^{+}, Mg^{2+}, Al^{3+}
15, 16, 17	-3, -2, -1	N^{3-}, O^{2-}, F^{-}
transition	variable	iron(II) Fe^{2+}, iron(III) Fe^{3+}

Polyatomic ions to learn

Ion	Formula	Ion	Formula
ammonium	NH_4^+	carbonate	CO_3^{2-}
hydroxide	OH^-	hydrogencarbonate	HCO_3^-
nitrate	NO_3^-	sulfate	SO_4^{2-}
		phosphate	PO_4^{3-}

The ionic bond

- **Electrostatic attraction** between oppositely charged ions.
- It acts in every direction, not between fixed pairs.
- The bigger the **electronegativity** difference, the more ionic the bond.

Ionic bond the electrostatic attraction between oppositely charged ions

Writing formulas that balance charge

Ions	Balance	Formula
Mg^{2+}, Cl^{-}	$+2 + 2(-1) = 0$	$MgCl_2$
Al^{3+}, O^{2-}	$2(+3) + 3(-2) = 0$	Al_2O_3
Ca^{2+}, NO_3^{-}	$+2 + 2(-1) = 0$	$Ca(NO_3)_2$
NH_4^{+}, SO_4^{2-}	$2(+1) + (-2) = 0$	$(NH_4)_2SO_4$

Use the evidence

Name the **cation** first; a simple **anion** ends in -ide: $FeCl_3$ is iron(III) chloride.

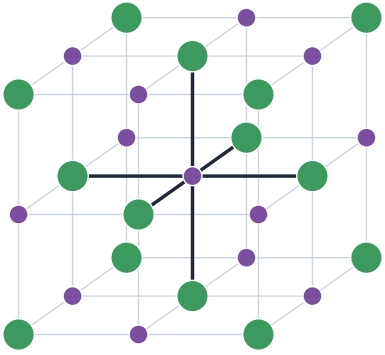
Check: write the formula

Your turn

What is the formula of aluminium sulfate?

A. $AlSO_4$
B. Al_2SO_4
C. $Al_2(SO_4)_3$
D. $Al_3(SO_4)_2$

A giant three-dimensional lattice



Na^+
Cl^-

bold lines: the central Na^+ and its six Cl^- neighbours

The formula of an ionic compound gives the ratio of ions in the lattice, not a molecule.

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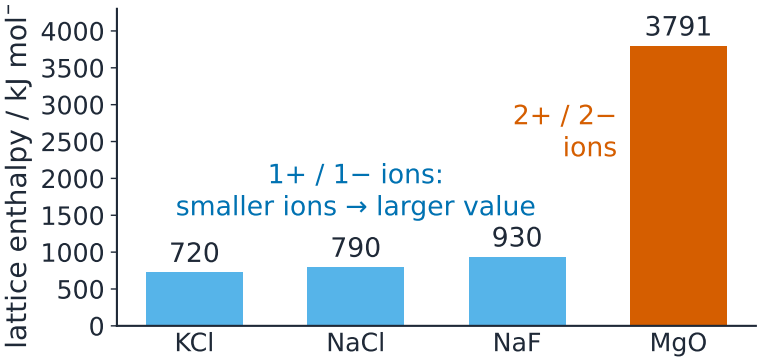
Lattice enthalpy measures bond strength

- Lattice enthalpy is the energy to separate one mole of solid into gaseous ions.
- Dissociation is positive; the reverse formation process is negative.

$$\mathrm{NaCl(s)} \rightarrow \mathrm{Na^+(g)} + \mathrm{Cl^-(g)} \qquad \Delta H = +790\ \mathrm{kJ\,mol^{-1}}$$

Compare charge first, then radius; both change the **electrostatic attraction**.

Compare charge and radius with conditions



For these supplied examples, MgO has the largest lattice dissociation enthalpy.

Worked example: MgO against NaCl

1. **Given.** Supplied melting points: MgO 2852 °C; NaCl 801 °C. Explain the difference using their ionic lattices.
2. **Structure.** Both are giant ionic lattices; melting overcomes attractions between ions.
3. **Ions.** Mg^{2+} and O^{2-} have double the charges and are smaller.
4. **Attraction.** Much stronger **electrostatic attraction**; lattice enthalpy 3791 against 790 kJ mol^{-1}.

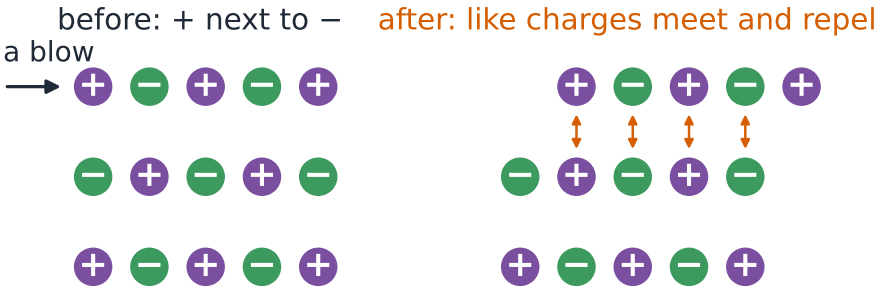
Decision

More energy is generally needed to disrupt the lattice; molten ions still interact.

Properties come from the structure

Property	**Behaviour**	**Reason**
melting point	high (low **volatility**)	many strong attractions in a giant lattice
conductivity	only molten or dissolved	ions fixed in the solid; free to move in liquid
solubility	varies: not all salts dissolve	solvation, lattice interactions and entropy all matter
brittleness	shatters	a shifted layer lines up like charges

Why ionic crystals shatter



Use a pattern of evidence: high melting point AND no conduction as a solid AND conduction when molten.

Exam tips

> "Transfer of electrons" explains how ions form, not what the **ionic bond** is.

- Ionic bond: **electrostatic attraction** between oppositely charged ions.
- Never call NaCl a molecule; conduction is by mobile ions.
- Lattice enthalpy: solid $\rightarrow$ gaseous ions, positive; compare charge, then radius.
- Bracket polyatomic ions: $Ca(OH)_2$, not $CaOH_2$.

Melting is not lattice dissociation

Melting

Ordered solid $\rightarrow$ ionic liquid.

Ions remain close and interact, but can move.

Lattice dissociation

One mole solid $\rightarrow$ separate gaseous ions.

Positive enthalpy in the convention used here.

Worked: infer a metal charge

1. **Anions.** $Fe_2(SO_4)_3$: three SO_4^{2-} ions give total charge -6.
2. **Cations.** Two Fe ions must supply $+6$; each is Fe^{3+}.
3. **Name and count.** Iron(III) sulfate; three sulfate ions contain 12 O atoms.

Decision

Formula-unit counts, atom counts and total ion counts are different.

Balancing charges in a formula



total charge 0
$\Rightarrow$ **Al_2O_3**

$2 \times (+3) = +6$ $\qquad$ $3 \times (-2) = -6$

smallest whole-number ratio that makes the charges sum to zero

So the formula is Al_2O_3

Solubility is not an ionic test by itself

> The ionic model predicts charge carriers; it does not guarantee high water solubility.

- AgCl and $CaCO_3$ are sparingly soluble despite their ionic structures.
- Sugar can dissolve as neutral molecules; dissolution alone does not show mobile ions.
- Assess the full property pattern and the experimental detection limit.

Worked: a distance comparison

1. **Model.** $|F| \propto 1/r^2$ for the same charge product.
2. **Ratio.** $F_A/F_B = (300/240)^2 = 1.5625$.
3. **Limit.** Pair potential energy scales as $1/\mathrm{r}$; neither ratio fixes a melting-point ratio.

Decision

State which physical quantity a proportional model describes.

Test an aqueous model fairly

> Compare observations under controlled conditions before assigning a structure.

- Use approved safe samples, a low-voltage instrument and a water blank.
- Match concentration, temperature, electrodes and instrument settings; rinse between samples.
- A low reading can reflect little dissolved material or the detection limit.

Check yourself

Try these

Write the formulas of calcium fluoride, iron(III) hydroxide and ammonium carbonate.

Name K_2S, Mg_3N_2 and $NaHCO_3$.

Put NaCl, CaO and KBr in order of increasing lattice enthalpy and explain.

Explain why sodium chloride conducts when dissolved but not when solid.