The ionic model

International Baccalaureate · IB Diploma · Chemistry · HL

Scope and prerequisites

This handout teaches Structure 2.1, The ionic model, for the course first assessed in 2025. All of it is SL and HL content.

Prerequisites: electron configurations and the periodic table groups from Structure 1.3. You will use lattice enthalpy values from the data booklet.

Ions and their charges

Why does salt conduct only when it melts?

Push two electrodes into solid table salt and no current flows. Melt the salt, or dissolve it in water, and it conducts at once. Nothing has been added: the same particles are there. The **ionic model** explains this with one idea: salt is made of charged particles that are locked in place in the solid and free to move in the liquid.

How ions form

An **ion** is an atom or group of atoms with an electric charge. Atoms form ions by losing or gaining electrons until they reach a stable arrangement, usually the configuration of a noble gas.

- Metal atoms **lose** electrons and form positive ions, called **cations**. Sodium, $[\mathrm{Ne}]\,3s^1$, loses one electron to form Na^+, [Ne].
- Non-metal atoms **gain** electrons and form negative ions, called **anions**. Chlorine, $[\mathrm{Ne}]\,3s^2\,3p^5$, gains one electron to form Cl^-, [Ar].

In an ionic compound the electrons lost by the metal are the electrons gained by the non-metal, so the total charge is zero.

Predicting the charge from the group

For main-group elements, the group number tells you the charge:

Group	1	2	13	15	16	17
Electrons lost or gained	lose 1	lose 2	lose 3	gain 3	gain 2	gain 1
Ion charge	+1	+2	+3	−3	−2	−1
Example	Li^+	Mg^{2+}	Al^{3+}	N^{3-}	O^{2-}	F^-

Transition elements can form ions with different charges, so the name gives the charge as a Roman numeral: iron(II) is Fe^{2+} and iron(III) is Fe^{3+}; copper(I)

is Cu^{+} and copper(II) is Cu^{2+}.

Polyatomic ions

A **polyatomic ion** is a group of covalently bonded atoms that carries an overall charge. Learn these:

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-}

Ionic bonding, formulas and names

The ionic bond

An **ionic bond** is the **electrostatic attraction** between oppositely charged ions. The attraction acts in every direction, so each ion attracts all the oppositely charged ions around it, not just one partner.

Ionic bonding happens between elements with a large difference in **electronegativity** : typically a metal on the left of the periodic table and a non-metal on the right. The larger the difference, the more ionic the bond. (Structure 2.4 shows bonding as a scale, not three separate boxes.)

Writing the formula

The formula shows the simplest ratio of ions that makes the total charge zero.

1. Write the two ions with their charges.
2. Find the smallest numbers of each that make the charges add up to zero.
3. Write the ratio as subscripts. Put a polyatomic ion in brackets if you need more than one.

Ions	Balance	Formula	Name
Na^+, Cl^-	$1 \times (+1) + 1 \times (-1) = 0$	$NaCl$	sodium chloride
Mg^{2+}, Cl^-	$1 \times (+2) + 2 \times (-1) = 0$	$MgCl_2$	magnesium chloride
Al^{3+}, O^{2-}	$2 \times (+3) + 3 \times (-2) = 0$	Al_2O_3	aluminium oxide
Ca^{2+}, NO_3^-	$1 \times (+2) + 2 \times (-1) = 0$	$Ca(NO_3)_2$	calcium nitrate
NH_4^+, SO_4^{2-}	$2 \times (+1) + 1 \times (-2) = 0$	$(NH_4)_2SO_4$	ammonium sulfate

Naming ionic compounds

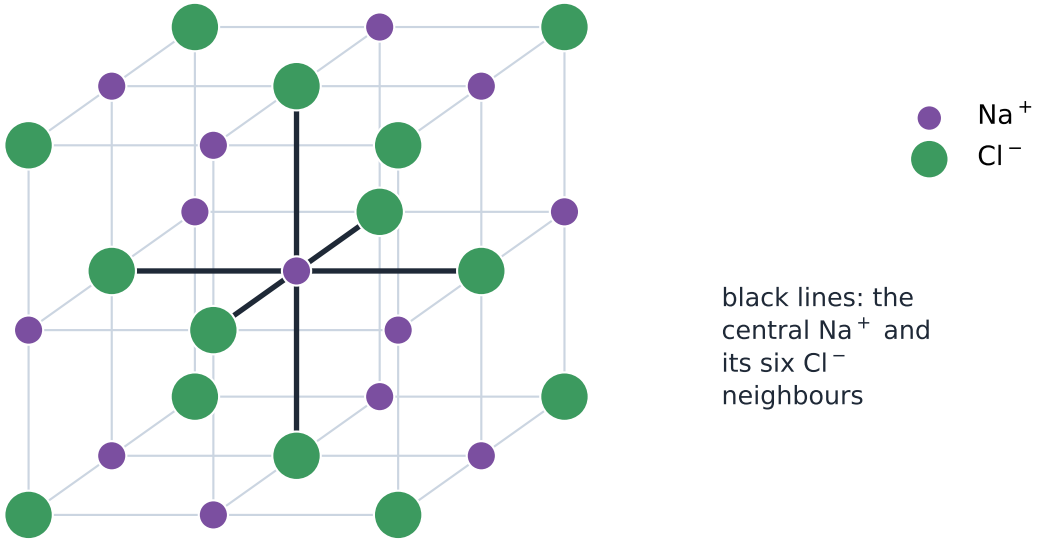
Name the cation first, then the anion. A simple anion ends in **-ide**: chloride, oxide, nitride, sulfide. A polyatomic ion keeps its own name: sulfate, nitrate. Give a transition-metal charge as a Roman numeral: $FeCl_3$ is iron(III) chloride.

Ionic lattices and lattice enthalpy

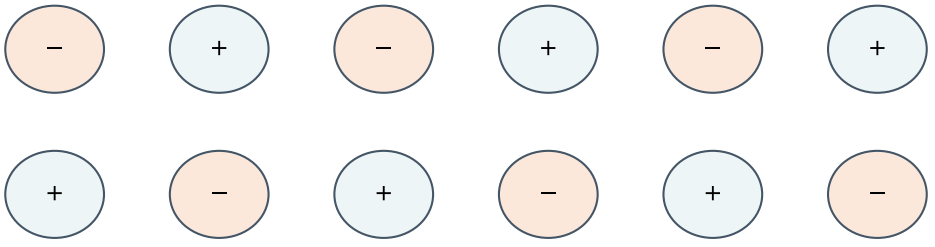
A giant three-dimensional lattice

Ionic compounds do not form molecules. The ions pack into a regular, repeating three-dimensional arrangement called a **lattice** . The formula NaCl gives only the **ratio** of ions, 1 : 1, not a molecule.

In sodium chloride each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion by six Na^+ ions.



In NaCl every ion has six nearest neighbours of the opposite charge



Alternating charges in an ionic lattice

Each ion is surrounded by ions of opposite charge: ionic bonding acts in every direction

Lattice enthalpy measures the strength of the bonding

The **lattice enthalpy** is the enthalpy change when one mole of a solid ionic compound is separated into its gaseous ions:

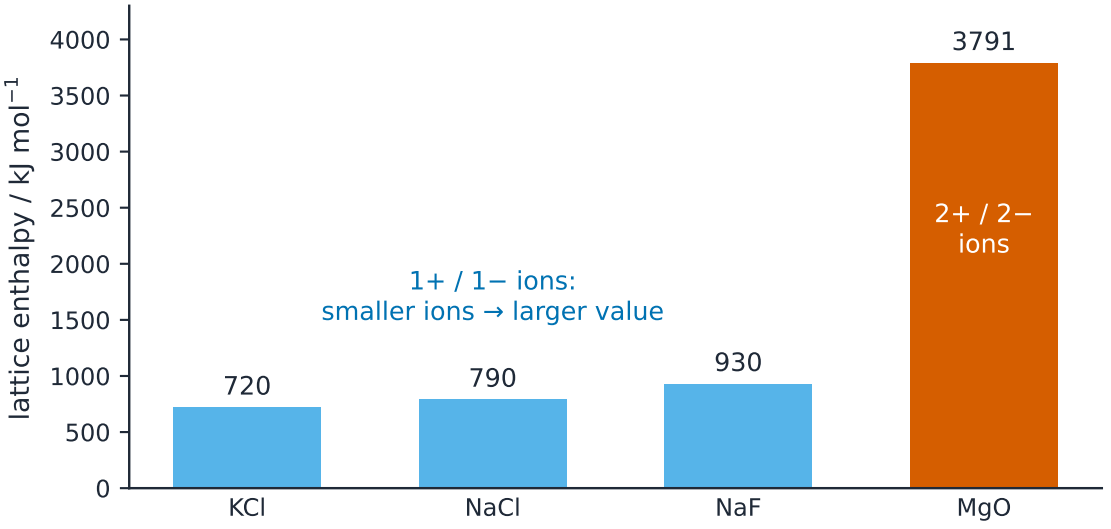
$$\mathrm{NaCl(s)} \rightarrow \mathrm{Na^+(g)} + \mathrm{Cl^-(g)} \qquad \Delta H_{\mathrm{lattice}} = +790 \text{ kJ mol}^{-1}$$

Breaking the attractions always needs energy, so lattice enthalpy is always positive (endothermic). A larger value means stronger ionic bonding.

Two factors decide its size, because the attraction between two ions is stronger when the charges are larger and the ions are closer together:

- **Ionic charge:** larger charges attract more strongly, so the lattice enthalpy is larger.
- **Ionic radius:** smaller ions sit closer together, so the lattice enthalpy is larger.

Compound	Ions	Lattice enthalpy / kJ mol^{-1}	Reason
KCl	K^+, Cl^-	720	K^+ is larger than Na^+
NaCl	Na^+, Cl^-	790	reference
NaF	Na^+, F^-	930	F^- is smaller than Cl^-
MgO	Mg^{2+}, O^{2-}	3791	both charges are doubled



Doubling both charges has a far bigger effect than changing the radius

Worked example. Explain why magnesium oxide melts at 2852 °C but sodium chloride melts at 801 °C.

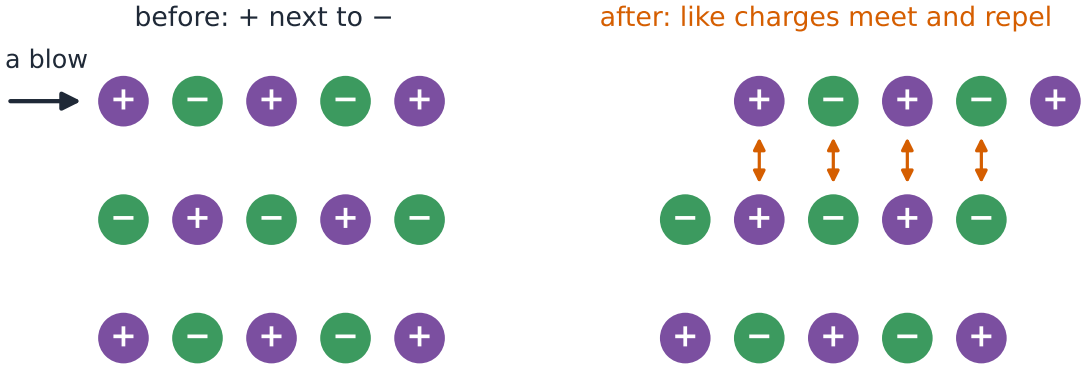
Both are giant ionic lattices, so melting means overcoming electrostatic attractions between ions. Mg^{2+} and O^{2-} carry twice the charges of Na^{+} and Cl^{-}, and they are also smaller. So the attraction between the ions is much stronger, the lattice enthalpy is much larger (3791 against 790 kJ mol^{-1}), and more energy is needed to separate the ions.

Properties of ionic compounds

■ Explaining properties from the structure

To explain any property, name the **structure**, the **particles**, the **forces** between them, and whether there are **mobile charged particles**.

Property	Typical behaviour	Explanation
melting and boiling point	high; low **volatility**	many strong electrostatic attractions in a giant lattice must be overcome
electrical conductivity	none as a solid; conducts when molten or in solution	ions are fixed in the solid; they can move and carry charge when molten or dissolved
solubility	often soluble in water, insoluble in non-polar solvents such as hexane	polar water molecules attract and surround the ions; non-polar molecules cannot
brittleness	shatters when hit	a blow shifts a layer so that like charges line up and repel



Shift one layer by one ion and like charges meet: the lattice splits apart

Note that an ionic compound conducts by **moving ions**, not by moving electrons. Electrolysis of a molten salt shows this: the ions move to the electrodes and are discharged there.

Use a pattern of evidence

One property alone rarely proves a structure. A giant covalent solid such as silicon dioxide also has a very high melting point, and some simple molecular substances dissolve in water. Look for the full pattern: high melting point, **and** no conduction as a solid, **and** conduction when molten.

> **Worked example.** Substance X melts at 801 °C, does not conduct as a solid, conducts when molten, and dissolves in water. Substance Y melts at 60 °C and does not conduct as a solid or as a liquid. Which is ionic?

X fits every part of the ionic pattern: a high melting point from strong attractions in a lattice, and conduction only when the ions are free to move. Y does not: its low melting point and lack of conduction suggest a simple molecular substance with weak forces between neutral molecules.

Practice

1. Give the charge on the ion formed by: (a) potassium, (b) sulfur, (c) aluminium, (d) nitrogen.
2. Write the formulas of: (a) calcium fluoride, (b) sodium oxide, (c) aluminium sulfate, (d) iron(III) hydroxide, (e) ammonium carbonate.
3. Name: (a) K_2S, (b) Mg_3N_2, (c) $CuSO_4$ (copper is +2), (d) $NaHCO_3$.
4. Write the equation for the lattice enthalpy of calcium oxide, including state symbols.
5. Put these in order of increasing lattice enthalpy and explain: NaCl, CaO, KBr.
6. Explain why sodium chloride conducts electricity when dissolved in water but not when solid.
7. Explain why an ionic crystal shatters when hit with a hammer.

Answers.

1. (a) +1, (b) −2, (c) +3, (d) −3.
2. (a) CaF_2, (b) Na_2O, (c) $Al_2(SO_4)_3$, (d) $Fe(OH)_3$, (e) $(NH_4)_2CO_3$.
3. (a) potassium sulfide, (b) magnesium nitride, (c) copper(II) sulfate, (d) sodium hydrogencarbonate.
4. $CaO(s) \rightarrow Ca^{2+}(g) + O^{2-}(g)$
5. KBr < NaCl < CaO. KBr has the largest ions with single charges, so the weakest attraction. CaO has 2+ and 2− ions, so by far the strongest attraction.
6. In the solid the ions are held in fixed positions in the lattice. In solution the ions are free to move, so they can carry charge through the liquid.
7. The blow shifts one layer of ions by one position. Ions of the same charge then line up next to each other, they repel, and the crystal splits along that plane.

Exam tips

- Define an ionic bond as "electrostatic attraction between **oppositely charged ions**". "Transfer of electrons" describes how ions form, not the bond.
- An ionic formula is a **ratio**: never call NaCl a molecule.
- Ionic compounds conduct by **mobile ions**, never by "free electrons".
- For lattice enthalpy, compare **charge** first, then **radius**, and say that both change the electrostatic attraction.
- Write lattice enthalpy as the solid going to **gaseous** ions, with a positive value.

- Put brackets around a polyatomic ion when you need two or more of it: $Ca(OH)_2$, not $CaOH_2$.