

# 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.
- The particles are the same; only their freedom to move changes.
- This unit: ions  $\rightarrow$  bonds and formulas  $\rightarrow$  lattices and lattice enthalpy  $\rightarrow$  properties.

## How ions form

Sodium  $\rightarrow \text{Na}^+$

$[\text{Ne}] 3s^1$  loses one electron.

$\text{Na}^+$  is  $[\text{Ne}]$ .

Chlorine  $\rightarrow \text{Cl}^-$

$[\text{Ne}] 3s^2 3p^5$  gains one electron.

$\text{Cl}^-$  is  $[\text{Ar}]$ .

## Predict the charge from the group

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

## Polyatomic ions to learn

| Ion       | Formula         | Ion               | Formula            |
|-----------|-----------------|-------------------|--------------------|
| ammonium  | $\text{NH}_4^+$ | carbonate         | $\text{CO}_3^{2-}$ |
| hydroxide | $\text{OH}^-$   | hydrogencarbonate | $\text{HCO}_3^-$   |
| nitrate   | $\text{NO}_3^-$ | sulfate           | $\text{SO}_4^{2-}$ |
|           |                 | phosphate         | $\text{PO}_4^{3-}$ |

# 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                      |
|-------------------------------------|---------------------|------------------------------|
| $\text{Mg}^{2+}, \text{Cl}^{-}$     | $+2 + 2(-1) = 0$    | $\text{MgCl}_2$              |
| $\text{Al}^{3+}, \text{O}^{2-}$     | $2(+3) + 3(-2) = 0$ | $\text{Al}_2\text{O}_3$      |
| $\text{Ca}^{2+}, \text{NO}_3^{-}$   | $+2 + 2(-1) = 0$    | $\text{Ca}(\text{NO}_3)_2$   |
| $\text{NH}_4^{+}, \text{SO}_4^{2-}$ | $2(+1) + (-2) = 0$  | $(\text{NH}_4)_2\text{SO}_4$ |

### Use the evidence

Name the **cation** first; a simple **anion** ends in -ide:  $\text{FeCl}_3$  is iron(III) chloride.

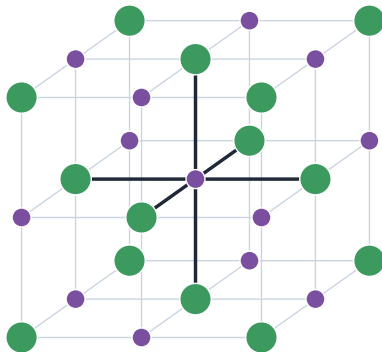
## Check: write the formula

### Your turn

What is the formula of aluminium sulfate?

- A.  $\text{AlSO}_4$
- B.  $\text{Al}_2\text{SO}_4$
- C.  $\text{Al}_2(\text{SO}_4)_3$
- D.  $\text{Al}_3(\text{SO}_4)_2$

## A giant three-dimensional lattice

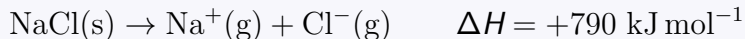


black lines: the  
central  $\text{Na}^+$  and  
its six  $\text{Cl}^-$   
neighbours

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

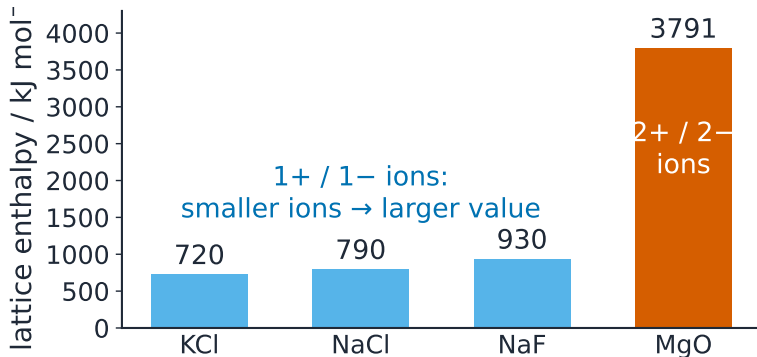
## Lattice enthalpy measures bond strength

- Lattice enthalpy is the energy to separate one mole of solid into gaseous ions.
- It is always positive; larger charges and smaller ions make it larger.



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

## Charge beats radius



Doubling both charges has a far bigger effect than any change of radius.

## Worked example: MgO against NaCl

- 1 Structure.** Both are giant ionic lattices; melting overcomes attractions between ions.
- 2 Ions.**  $\text{Mg}^{2+}$  and  $\text{O}^{2-}$  have double the charges and are smaller.
- 3 Attraction.** Much stronger **electrostatic attraction**; lattice enthalpy 3791 against  $790 \text{ kJ mol}^{-1}$ .

### Decision

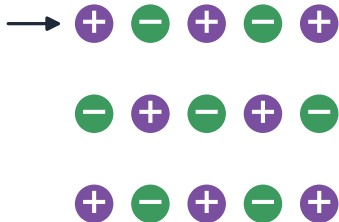
More energy is needed to separate the ions, so MgO melts far higher.

## Properties come from the structure

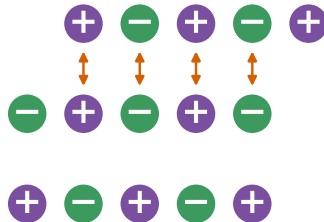
| Property      | Behaviour                     | Reason                                          |
|---------------|-------------------------------|-------------------------------------------------|
| melting point | high (low <b>volatility</b> ) | many strong attractions in a giant lattice      |
| conductivity  | only molten or dissolved      | ions fixed in the solid; free to move in liquid |
| solubility    | often soluble in water        | polar water molecules surround the ions         |
| brittleness   | shatters                      | a shifted layer lines up like charges           |

# Why ionic crystals shatter

before: + next to -  
a blow



after: like charges meet and repel



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:  $\text{Ca}(\text{OH})_2$ , not  $\text{CaOH}_2$ .

## Check yourself

### Try these

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

Name  $\text{K}_2\text{S}$ ,  $\text{Mg}_3\text{N}_2$  and  $\text{NaHCO}_3$ .

Put  $\text{NaCl}$ ,  $\text{CaO}$  and  $\text{KBr}$  in order of increasing lattice enthalpy and explain.

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