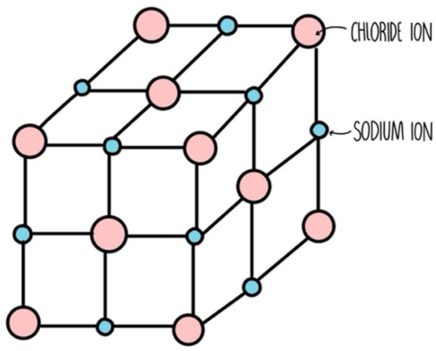
Lattice Enthalpy

Lattice enthalpy



- Ionic compounds form huge, regular structures called **giant lattice.** The ions in the lattice are held together by **ionic bonds**.
- **Lattice enthalpy** is the energy change associated with the formation of **one mole** of an **ionic lattice** from its **gaseous ions** under standard conditions (298 K and 100 kPa).
- The **stronger the ionic bonds** in the ionic compound, the **more negative** the lattice enthalpy.

Factors that affect ionic bond strength

The **strength** of ionic bonds (and therefore **lattice enthalpy**) depends on two factors:

1. Ionic charge
2. Ionic radius

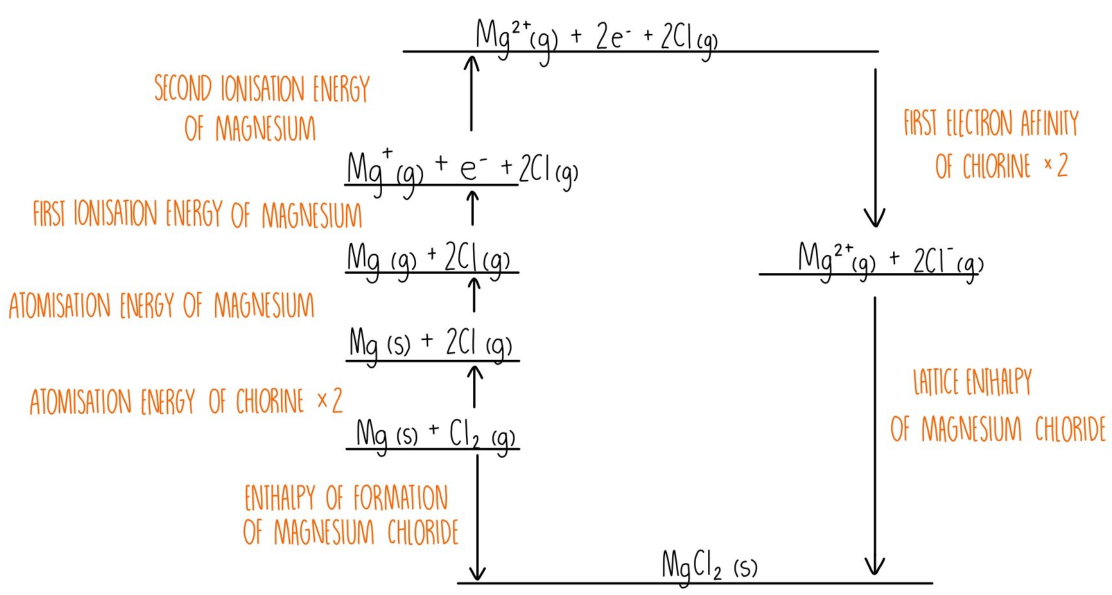
The **higher the charge** on the ions, the **stronger the electrostatic attraction**. This means **more energy is released** when the lattice forms — lattice enthalpy becomes more negative.

The **smaller the ionic radius**, the **closer together the ions** are in the lattice and the **stronger the ionic bonds** that form between them. That means that for smaller ions, there is a **more negative** lattice enthalpy.

Born-Haber cycles

Lattice enthalpy **cannot be directly calculated** so we use **Born-Haber cycles** instead. Bohn-Haber cycles are where the various enthalpy changes are linked together. It relies on **Hess's law** – the idea that the total enthalpy change of a reaction is the same regardless of the route taken.

If you look at the Born-Haber cycle for magnesium chloride below, you can see that all of the arrows along the left hand side (enthalpy of formation + atomisation energies + ionisation energies) is equal to the arrows along the right hand side (electron affinity + lattice enthalpy).

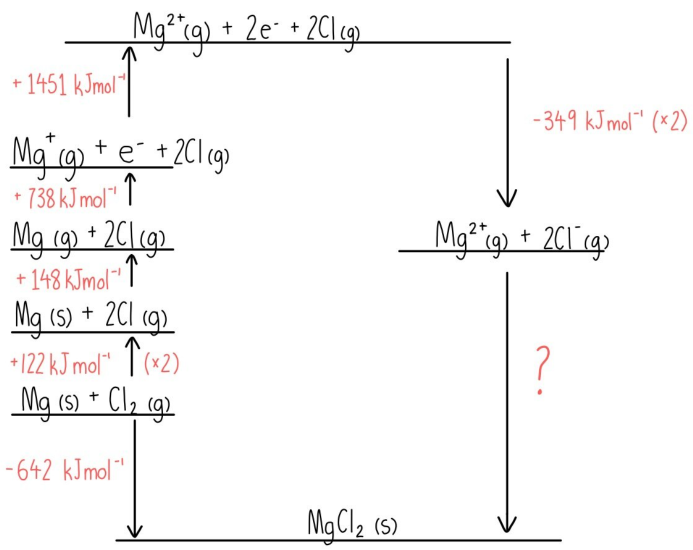


There are three things to remember when calculating lattice enthalpy from a Born-Haber cycle:

1. **Ignore the minus sign for enthalpy of formation**.
2. **Multiply the enthalpy value by the number of particles**.
3. **Add a negative sign to your final answer**.

Worked example: calculating lattice enthalpy from Born-Haber cycles

Calculate the lattice enthalpy for magnesium chloride using the Born-Haber cycle.



- Formation of $MgCl_2$ + (atomisation of Cl x 2) + atomisation of Mg + first ionisation energy of Mg + second ionisation energy of Mg = (first electron affinity of Cl x 2) + lattice enthalpy of $MgCl_2$
- 642 + (122 x 2) + 148 + 738 + 1451 = (349 x 2) + LE
- 3223 = 698 + LE
- LE = 3223 – 698 = 2525
- LE = -2525 kJ mol^{-1}

Enthalpy of solution

The enthalpy change of **solution** is the energy change that takes place when **one mole of solute dissolves in water**.

Dissolving an ionic lattice in water involves **two stages**:

1. The **ionic bonds** in the lattice **break** to form **gaseous ions**. This is an **endothermic** process.

2. The gaseous ions form bonds with water molecules and become **hydrated**. This is an **exothermic** process and is known as the **enthalpy change of hydration**.

Substances that **release more energy during hydration** than the energy required to break the ionic bonds in the first stage are likely to be **soluble**. For soluble compounds, the enthalpy change of solution will be **negative** (exothermic).

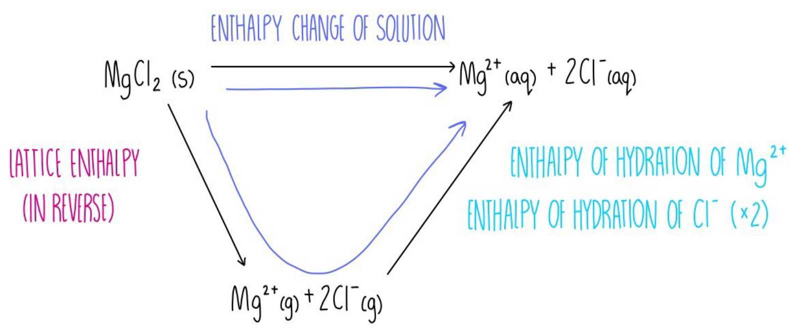
Calculating enthalpy change of solution

You can calculate the enthalpy change of solution using a **Hess cycle**, as shown in the worked example below.

Worked example – calculating enthalpy change of solution using a Hess cycle

Calculate the enthalpy change of solution of magnesium chloride, given the following data.

- Lattice enthalpy of $MgCl_2$ = -2526 kJ mol^{-1}
- Enthalpy of hydration of Mg^{2+} = -1920 kJ mol^{-1}
- Enthalpy of hydration of Cl^- = -364 kJ mol^{-1}



ENTHALPY CHANGE OF SOLUTION = –(LATTICE ENTHALPY) + ENTHALPY OF HYDRATION OF Mg^{2+} + (ENTHALPY OF HYDRATION OF Cl^- × 2)

ENTHALPY CHANGE OF SOLUTION = –(-2526) + -1920 + (-364 × 2)

ENTHALPY CHANGE OF SOLUTION = -122 kJ mol^{-1}

Factors which affect the enthalpy change of hydration

The second stage of dissolving a compound involves the **gaseous ions becoming hydrated** (they form bonds with water molecules.) The energy released when this happens is known as the **enthalpy change of hydration** as is defined as:

Enthalpy change of hydration - the energy change that takes place when **one mole** of **gaseous ions dissolves in water**.

The enthalpy change of hydration depends on **two factors**:

1. **The charge on the ion** – the larger the charge, the better the ion is at attracting water molecules, forming a stronger electrostatic attraction between them. More energy is released when these stronger bonds are formed, making the enthalpy change of hydration more exothermic (more negative).
2. **Ionic radius** – smaller ions have their charge concentrated in a smaller area (they have a higher charge density). Smaller ions can therefore attract water molecules more easily than larger ones, resulting in a more exothermic enthalpy change of hydration.