

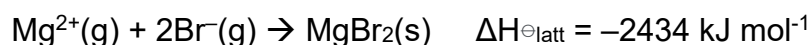
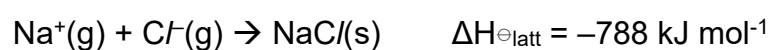
Group II Compound Solubilities

The solubility of an ionic compound depends on the **lattice enthalpy** of the compound and the **hydration enthalpies** of the anion and the cation.

Lattice Enthalpy, $\Delta H_{\ominus \text{latt}}$

The lattice enthalpy is defined as **the energy given out when one mole of a solid is formed by the coming together of separate ions (gaseous)**.

- When ions are separated from each other, they can be thought of as being in the gaseous state; when separate ions come together to form a lattice, they are in the solid state.
- So, the lattice enthalpy can be defined as the enthalpy change involved in processes such as:



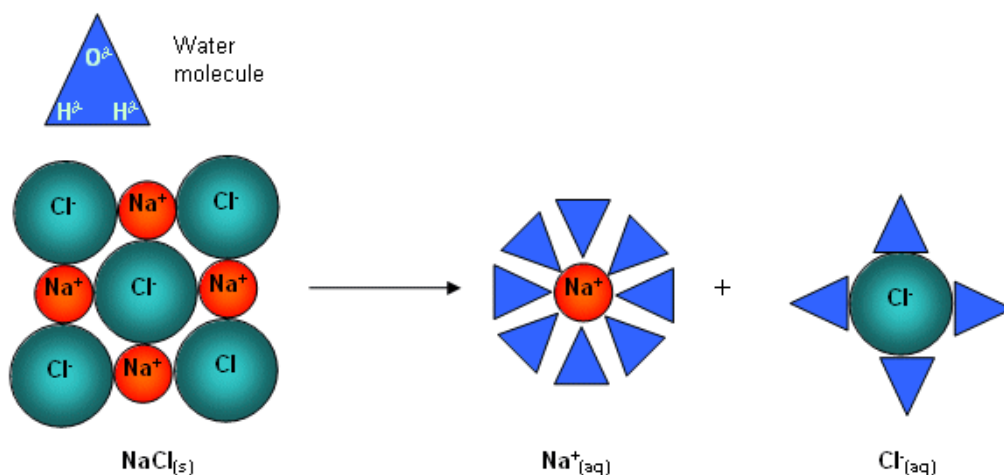
- Forming a lattice from gaseous ions is always an exothermic process, and so ΔH_{latt} is always negative.
- In order for an ionic substance to dissolve in water, the ions in the lattice must be separated; this requires energy to overcome the electrical attraction between the ions.
- The strength of the ionic attractions in the lattice is measured using the lattice energy.

$$\text{Lattice energy} \propto (q_1 q_2) / r^2$$

- The lattice energy must be supplied in order to break down a lattice and enable the ionic substance to dissolve in water.
- This energy tends to stop substances dissolving, unless the energy is paid back in later.

Hydration Enthalpy, $\Delta H_{\text{hyd}}^{\ominus}$

- Even though energy is needed to break up the lattice, many ionic substances do dissolve.
- Therefore, something else must happen to supply that energy.
- Water molecules possess a dipole and are attracted to the ions on the surface of the lattice when an ionic compound is added to water.
- The water molecules pull the ions from the surface of the compound; the ions in the solution become hydrated; they have water molecules bound to them:
- The smaller the ion and the higher its charge, the more water molecules it attracts.
- The strength of the attractions between the ions and the water molecules is measured by the enthalpy of hydration ($\Delta H_{\text{hyd}}^{\ominus}$).



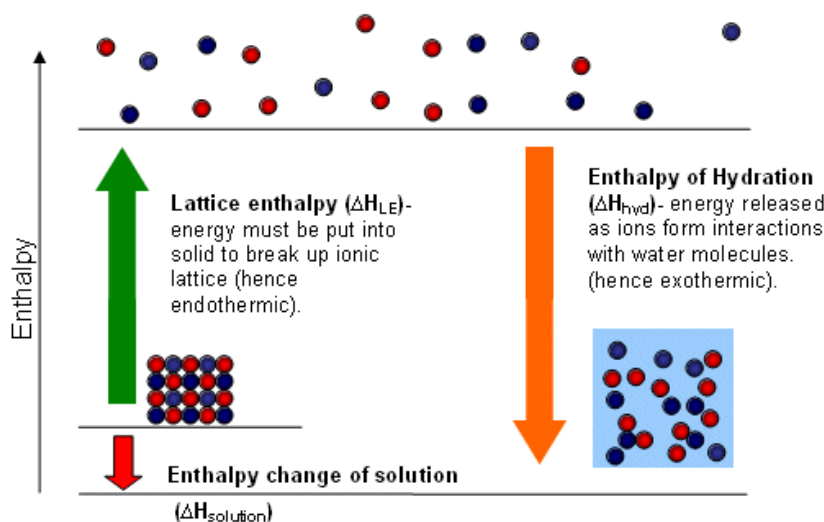
- The smaller the ion and the higher its charge, the more water molecules it attracts.
- The strength of the attractions between the ions and the water molecules is measured by the enthalpy of hydration ($\Delta H_{\text{hyd}}^{\ominus}$).
- The enthalpy change of hydration is defined as **the enthalpy change for the production of a solution of ions from one mole of gaseous ions.**
- For example:



- Enthalpies of hydration depend upon the concentration of the solution produced.
- Values quoted refer to an infinitely dilute solution.
- **Enthalpies of hydration are always negative (exothermic).**
- Once again, the most exothermic values occur when the **ionic radii decreases** and the **ionic charge increases**.
- The hydration of ions favours dissolving and helps to supply the energy needed to separate the ions from a lattice.

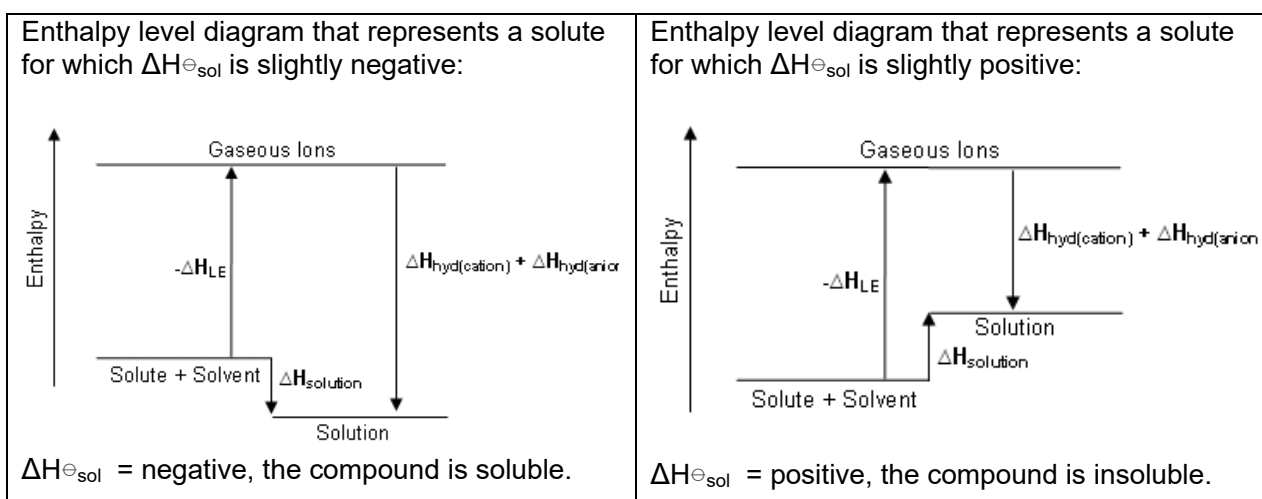
Enthalpy change of solution, $\Delta H_{\ominus \text{sol}}$

- The difference between the enthalpies of hydration of ions and the lattice enthalpy gives the **enthalpy change of solution** ($\Delta H_{\ominus \text{sol}}$).
- This can be measured experimentally; **it is the enthalpy change when one mole of a solute dissolves to form an infinitely dilute solution.**
- The process of dissolving can be broken down into two steps:
 1. The solid ionic compound is split up into a gas of ions (requires the energy equal to the lattice energy to be supplied).
 2. The ions dissolve in solution and become hydrated; the energy equal to the hydration energies of the two ions is given out.



Therefore, the enthalpy of solution is equal to:

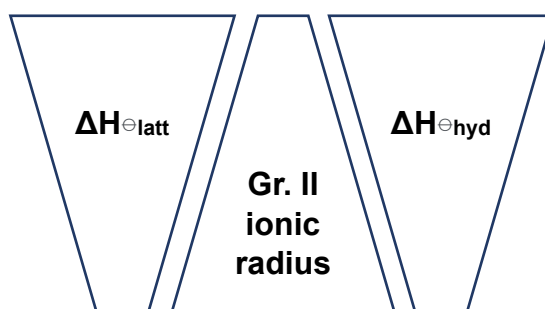
$$\Delta H_{\ominus \text{sol}} = \Delta H_{\ominus \text{hyd}(\text{cation})} + \Delta H_{\ominus \text{hyd}(\text{anion})} + -\Delta H_{\ominus \text{latt}}$$



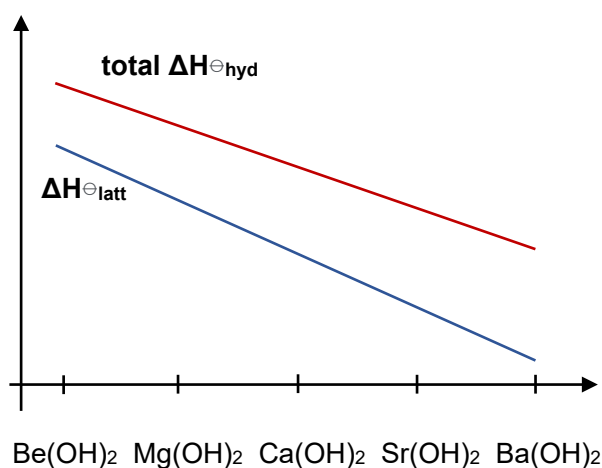
- The enthalpy change of solution can also be used to explain the trends in the solubility of group two compounds.
- The solubility of the group 2 hydroxides increases down the group, whereas the solubility of group 2 sulfates and carbonates decreases down the group.

Solubility of group II hydroxide

As the ionic radius increase down the group, the $\Delta H_{\ominus\text{hyd}}$ of group II metal cations become less exothermic. The $\Delta H_{\ominus\text{latt}}$ of the ionic compound also becomes less exothermic



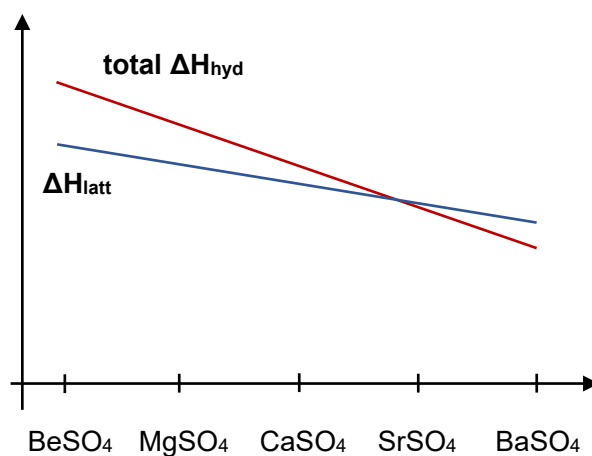
Because OH^- is a small ion, the change of the distance between the nuclei of the cation and anion is quite significant down the group, thus the $\Delta H_{\ominus\text{latt}}$ becomes significantly less exothermic.



As a result, the $\Delta H_{\ominus\text{sol}}$ becomes more exothermic, and the hydroxide becomes more soluble down the group.

Solubility of group II sulfates and carbonates

Because SO_4^{2-} and CO_3^{2-} are large ions, the change of the distance between the nuclei of the cation and anion becomes less significant down the group, thus the $\Delta H_{\ominus\text{latt}}$ only becomes slightly less exothermic, and the $\Delta H_{\ominus\text{hyd}}$ acts as the dominant factor that determines the solubility.



As a result, the $\Delta H_{\ominus\text{sol}}$ becomes less exothermic, and the sulfates/carbonates become less soluble down the group.