

Enthalpy of Solution | Lattice vs Hydration

WHY CAN DISSOLVING A SALT ABSORB OR RELEASE HEAT?



Some salts cool the water
(Endothermic).



Some salts heat the water
(Exothermic).

- Dissolving a salt isn't just one simple step.
- It is a microscopic tug-of-war between two immense thermodynamic forces.

Step 1: Breaking the Crystal

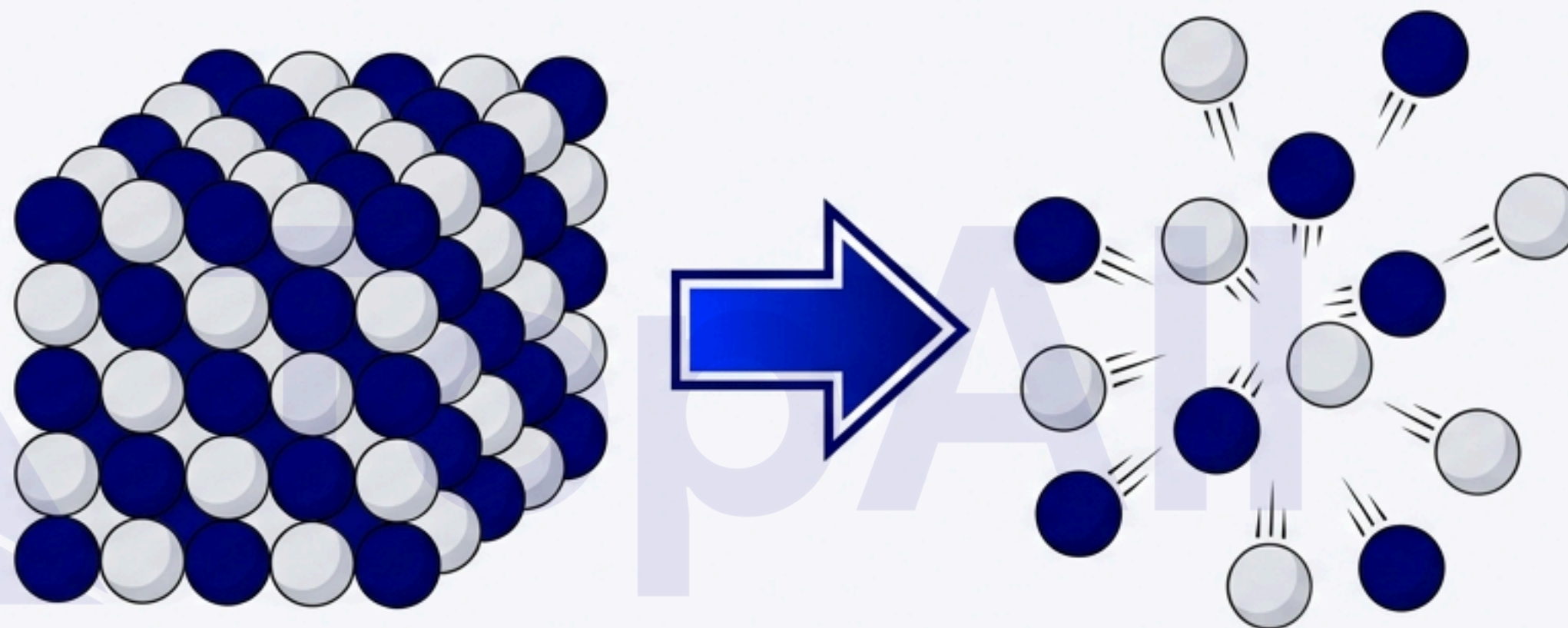
LATTICE ENTHALPY: PULLING IONS APART

The Process:

Solid Crystal $\rightarrow$ Gaseous /
Separated Ions

The Logic:

Ions are held together by extremely strong electrostatic forces. Shattering this crystal lattice requires a massive energy input.



Energy Required: $\Delta_{lattice}H^\circ > 0$ (Endothermic)

Step 2: Solvation by Water

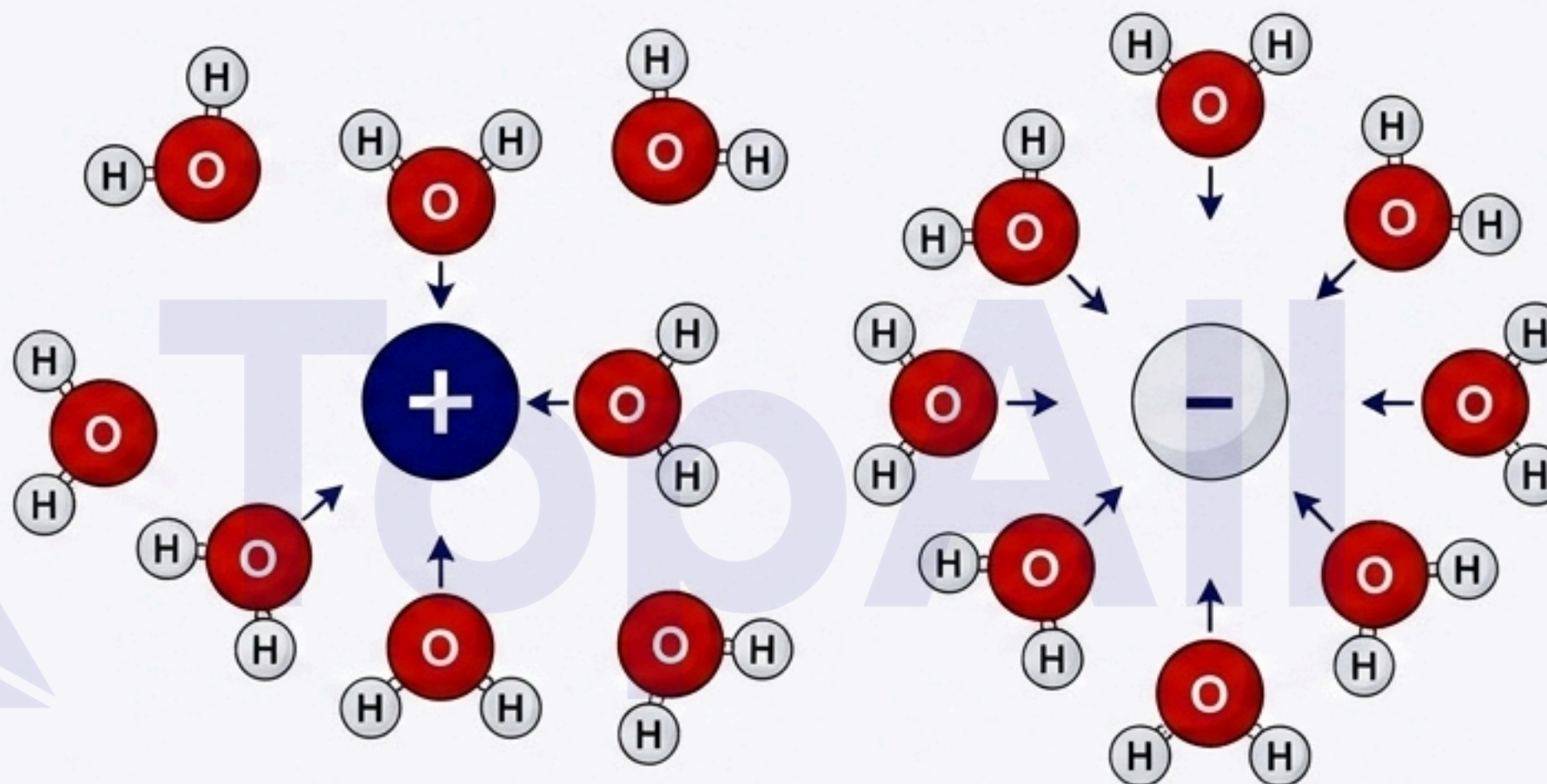
HYDRATION ENTHALPY: SURROUNDING THE IONS

The Process:

Gaseous Ions + Water $\rightarrow$
Aqueous / Hydrated Ions

The Logic:

Water dipoles form new electrostatic bonds with the free ions. Forming these new bonds releases massive amounts of energy to the surroundings.



Energy Released: $\Delta_{hyd}H^\circ < 0$ (Exothermic)



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$$E=mc^2$$



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The Net Energy Change

ENTHALPY OF SOLUTION ($\Delta_{\text{sol}}H^\circ$)

$$\Delta_{\text{sol}}H^\circ = \Delta_{\text{lattice}}H^\circ + \Delta_{\text{hyd}}H^\circ$$

Always Positive (+)

(Breaking bonds consumes energy)

Always Negative (-)

(Forming bonds releases energy)

The Net Balance

Can be positive or negative
depending on which thermodynamic
force wins the tug-of-war.

Real-World Example: Dissolving NaCl

A THERMODYNAMIC TUG-OF-WAR

Lattice Enthalpy:
 $+788 \text{ kJ mol}^{-1}$

Hydration Enthalpy:
 $+(-784 \text{ kJ mol}^{-1})$

Net Enthalpy ($\Delta_{\text{sol}}H^\circ$):
 $= +4 \text{ kJ mol}^{-1}$

- **Why is the net change so small?**
- The massive energy required to shatter the salt crystal (+788) is almost perfectly 'paid back' by the energy released when water stabilizes the ions (-784).
- **Result:** Dissolving table salt is only very slightly endothermic.

MASTERING ENTHALPY OF SOLUTION



The Trap

Assuming $\Delta_{\text{sol}}H^\circ$ must be negative (exothermic) for a salt to successfully dissolve.

The Reality

Many salts (like NaCl, $+4 \text{ kJ mol}^{-1}$) have a slightly positive $\Delta_{\text{sol}}H^\circ$ but still dissolve easily.

Why? Because Entropy (randomness) increases immensely when a rigid crystal dissolves, driving the process forward regardless of the slight heat penalty!



1. $\Delta_{\text{lattice}}H^\circ > 0$: Energy is consumed to separate crystal ions.
2. $\Delta_{\text{hyd}}H^\circ < 0$: Energy is released when water stabilizes ions.
3. $\Delta_{\text{sol}}H^\circ = \Delta_{\text{lattice}}H^\circ + \Delta_{\text{hyd}}H^\circ$
4. If **Lattice > Hydration**, dissolution is **Endothermic (+)**.
5. If **Hydration > Lattice**, dissolution is **Exothermic (-)**.



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