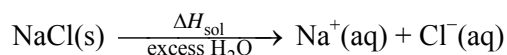


## Resource 1 – Chapter 5

### Enthalpy change of solution

The enthalpy change of solution ( $\Delta H_{\text{sol}}$ ) is the enthalpy change when one mole of solute is dissolved in excess solvent to form a solution of ‘infinite dilution’ under standard conditions. ‘Infinite dilution’ means that any further dilution of the solution produces no further enthalpy change. The particles are assumed not to interact with each other in the solution.

Let us now consider dissolving common salt (sodium chloride) in water. The equation for this process is:

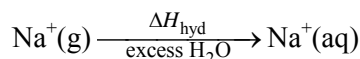


This can be broken down into two separate processes:

**stage 1:** lattice enthalpy ( $\Delta H_{\text{latt}}$ ); breaking apart the lattice into gaseous ion. This is an **endothermic** process

**stage 2:** hydration of the ions ( $\Delta H_{\text{hyd}}$ ); the surrounding of the gaseous ions by water molecules.

Enthalpy change of hydration ( $\Delta H_{\text{hyd}}$ ) is the enthalpy change when one mole of gaseous ions are surrounded by water molecules to form an ‘infinitely dilute solution’ under standard conditions. For example:



$$\Delta H_{\text{hyd}} = -390 \text{ kJ mol}^{-1}$$

This process involves formation of electrostatic interactions between the ions and water molecules. This is an **exothermic** process. The dissolving process thus involved an endothermic process ( $\Delta H_{\text{latt}}$ ) and an exothermic process ( $\Delta H_{\text{hyd}}$ ).

Overall, the enthalpy change of solution will be exothermic if the total enthalpy change of hydration is more negative than the lattice enthalpy is positive. If we work out the value of the enthalpy change of solution for sodium chloride:

$$\begin{aligned}\Delta H_{\text{sol}}(\text{NaCl}) &= \Delta H_{\text{latt}} + \Delta H_{\text{hyd}}(\text{Na}^+) + \Delta H_{\text{hyd}}(\text{Cl}^-) \\ &= 771 + (-406) + (-364) \\ &= +1 \text{ kJ mol}^{-1}\end{aligned}$$

The dissolving of sodium chloride in water is thus an endothermic process. We know, however, that the dissolution of sodium chloride is a spontaneous process at 298 K. Therefore,  $\Delta G^\circ$  must be negative.

$\Delta G^\circ$  can be negative (for an endothermic reaction) only if the reaction involves an increase in entropy. The dissolving process is therefore driven by the increase in entropy of the system, resulting from aqueous sodium and chloride ions being a more disordered situation than crystalline sodium chloride and pure water.

The fact that the entropy change for this process is positive means that  $-T\Delta S$  is negative. As the temperature increases,  $-T\Delta S$  becomes more negative, and therefore the dissolving of sodium chloride in water becomes more favourable. This can be contrasted with a gas such as carbon dioxide, for which the dissolving process involves a decrease in entropy ( $\Delta S$  negative) and therefore becomes less favourable as the temperature increases.