A solution is a homogeneous mixture of two or more substances and can either be in the gas phase, the liquid phase, or the solid phase. The enthalpy change of solution refers to the amount of heat that is released or absorbed during the dissolving process (at constant pressure). This enthalpy of solution ($\Delta H_{\text{solution}}$) can either be positive (endothermic) or negative (exothermic). When understanding the enthalpy of solution, it is easiest to think of a hypothetical three-step process happening between two substances. One substance is the solute, let’s call that $A$. The other substance is the solvent, let’s call that $B$.

### Step 1: Breaking up the Solute

The first process that happens deals only with the solute, $A$, which requires breaking all intramolecular forces holding it together. This means the solute molecules separate from each other. The enthalpy of this process is called $\Delta H_1$. This since this is always an endothermic process (requiring energy to break interactions), then $\Delta H_1 > 0$.

\[
\ce{A (s) ->[\text{energy in}] A (g)} \nonumber
\]

### Step 2: Breaking up the Solvent

The second process is very similar to the first step. Much like how the solute, $A$, needed to break apart from itself, the solvent, $B$, also needs to overcome the intermolecular forces holding it together. This causes the solvent molecules separate from each other. The enthalpy of this process is called $\Delta H_2$. Like the first step, this reaction is always endothermic ($\Delta H_2 > 0$) because energy is required to break the interaction between the $B$ molecules.

\[
\ce{B (l) ->[\text{energy in}] B (g)} \nonumber
\]

At this point, let us visualize what has happened so far. The solute, $A$, has broken from the intermolecular forces holding it together and the solvent, $B$, has broken from the intermolecular forces holding it together as well. It is at this time that the third process happens. We also have two values $\Delta H_1$ and $\Delta H_2$, that are both greater than zero (endothermic).

### Step 3: Combining the Two Together

The third process is when substance $A$ and substance $B$ mix to for a solution. The separated solute molecules and the separated solvent molecules join together to form a solution. This solution will contain one mole of the solute $A$ in an infinite amount of the solvent $B$. The enthalpy of combining these two substances to form the solution is $\Delta H_3$ and is an exothermic reaction (releasing heat since interactions are formed) with $\Delta H_3 < 0$.

\[
\ce{A (g) + B (g) ->[\text{energy out}] A(sol)} \nonumber
\]

The enthalpy of solution can expressed as the sum of enthalpy changes for each step:

\[
\Delta H_{\text{solution}} = \Delta H_1 + \Delta H_2 + \Delta H_3. \label{eq1}
\]

So the enthalpy of solution can either be endothermic, exothermic or neither ($\Delta H_{\text{solution}} = 0$), depending on how
much heat is required or release in each step. If \(\Delta H_{solution} = 0\), then the solution is called an **ideal solution** and if \(\Delta H_{solution} > 0\) or \(\Delta H_{solution} < 0\), then these solutions are called **non-ideal solutions**.

Figure \(\PageIndex{1}\): Energy Diagram for Endothermic Dissolving Process (where \(\Delta H_{solution} > 0\)). The combined magnitude of Steps 1 and 2 is greater than the magnitude of Step 3. This is also a non-ideal solution.

The diagrams below can be used as visuals to help facilitate the understanding of this concept. Figure \(\PageIndex{1}\) is for an endothermic reaction, where \(\Delta H_{solution} > 0\). Figure \(\PageIndex{2}\) is for an exothermic reaction, where \(\Delta H_{solution} < 0\). Figure \(\PageIndex{3}\) is for an ideal solution, where \(\Delta H_{solution} = 0\).

Figure \(\PageIndex{2}\): Energy Diagram for Exothermic Dissolving Process (where \(\Delta H_{solution} < 0\)). The combined magnitude of Steps 1 and 2 is less than the magnitude of Step 3. This is also a non-ideal solution.

### Ideal Solutions

The enthalpy of solution depends on the strengths of intermolecular forces of the solute and solvent and solvent (Equation \ref{eq1}). If the solution is ideal, and \(\Delta H_{solution} = 0\), then

\[
\begin{align*}
\Delta H_{solution} &= \Delta H_1 + \Delta H_2 + \Delta H_3 \quad &\text{=} 0. \label{eq2}
\end{align*}
\]

This means the forces of attraction between like (the solute-solute and the solvent-solvent) and unlike (solute-solvent) molecules are the same (Figure \(\PageIndex{3}\)). If the solution is non-ideal, then either \(\Delta H_{1}\) added to \(\Delta H_{2}\) is greater than \(\Delta H_{3}\) or \(\Delta H_{3}\) is greater than the sum of \(\Delta H_{1}\) and \(\Delta H_{2}\). The first case means the forces of attraction of unlike molecules is greater than the forces of attraction between like molecules. The second case means the forces of attraction between like molecules is greater than the forces of attraction between unlike molecules (Figure \(\PageIndex{2}\)).

Example \(\PageIndex{1}\): Table salt
Table salt (\(\ce{NaCl}\)) dissolves readily in water. In solid (\(\ce{NaCl}\)), the positive sodium ions are attracted to the negative chloride ions. The same is true of the solvent, water; the partially positive hydrogen atoms are attracted to the partially negative oxygen atoms. While (\(\ce{NaCl}\)) dissolves in water, the positive sodium cations and chloride anions are stabilized by the water molecule electric dipoles. Thus, the intermolecular interactions (i.e., ionic bonds) between (\(\ce{NaCl}\)) are broken and the salt is dissolved. The overall chemical equation for this reaction is as follows:

\[
\ce{NaCl (s) ->[H_2O] Na^+ (aq) + Cl^- (aq)}
\]

Enthalpy of solution is only one part of the driving force in the formation of solutions; the other part is the entropy of solution.

References


Contributors and Attributions

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