The standard enthalpy of formation is defined as the change in enthalpy when one mole of a substance in the standard state (1 atm of pressure and 298.15 K) is formed from its pure elements under the same conditions.

### Introduction

The standard enthalpy of formation is a measure of the energy released or consumed when one mole of a substance is created under standard conditions from its pure elements. The symbol of the standard enthalpy of formation is $\Delta H_f$.

- $\Delta$ = A change in enthalpy
- $^o$ = A degree signifies that it's a standard enthalpy change.
- $f$ = The f indicates that the substance is formed from its elements

The equation for the standard enthalpy change of formation (originating from Enthalpy's being a State Function), shown below, is commonly used:

$$\Delta H_{\text{reaction}}^o = \sum \Delta H_{f}^o(\text{products}) - \sum \Delta H_{f}^o(\text{Reactants})$$

This equation essentially states that the standard enthalpy change of formation is equal to the sum of the standard enthalpies of formation of the products minus the sum of the standard enthalpies of formation of the reactants.

Example \(\PageIndex{1}\)

Given a simple chemical equation with the variables A, B and C representing different compounds:

$$(A + B \leftrightharpoons C)$$

and the standard enthalpy of formation values:

- $\Delta H_f[\text{A}] = 433 \text{ KJ/mol}$
- $\Delta H_f[\text{B}] = -256 \text{ KJ/mol}$
- $\Delta H_f[\text{C}] = 523 \text{ KJ/mol}$

the equation for the standard enthalpy change of formation is as follows:

$$\Delta H_{\text{reaction}}^o = \Delta H_f^o[\text{C}] - (\Delta H_f^o[\text{A}] + \Delta H_f^o[\text{B}])$$

$$\Delta H_{\text{reaction}}^o = (1 \text{ mol})(523 \text{ kJ/mol}) - ((1 \text{ mol})(433 \text{ kJ/mol}) + (1 \text{ mol})(-256 \text{ kJ/mol}))$$

Because there is one mole each of A, B and C, the standard enthalpy of formation of each reactant and product is multiplied by 1 mole, which eliminates the mol denominator:

$$\Delta H_{\text{reaction}}^o = 346 \text{ kJ}$$

The result is 346 kJ, which is the standard enthalpy change of formation for the creation of variable "C".
The standard enthalpy of formation of a pure element is in its reference form its standard enthalpy formation is zero. Carbon naturally exists as graphite and diamond. The enthalpy difference between graphite and diamond is too large for both to have a standard enthalpy of formation of zero. To determine which form is zero, the more stable form of carbon is chosen. This is also the form with the lowest enthalpy, so graphite has a standard enthalpy of formation equal to zero. Table 1 provides sample values of standard enthalpies of formation of various compounds.

**Table 1: Sample Table of Standard Enthalpy of Formation Values.**

<table>
<thead>
<tr>
<th>Compound</th>
<th>ΔH_f^o</th>
</tr>
</thead>
<tbody>
<tr>
<td>O_2(g)</td>
<td>0 kJ/mol</td>
</tr>
<tr>
<td>C(graphite)</td>
<td>0 kJ/mol</td>
</tr>
<tr>
<td>CO(g)</td>
<td>-110.5 kJ/mol</td>
</tr>
<tr>
<td>CO_2(g)</td>
<td>-393.5 kJ/mol</td>
</tr>
<tr>
<td>H_2(g)</td>
<td>0 kJ/mol</td>
</tr>
<tr>
<td>H_2O(g)</td>
<td>-241.8 kJ/mol</td>
</tr>
<tr>
<td>HF(g)</td>
<td>-271.1 kJ/mol</td>
</tr>
<tr>
<td>NO(g)</td>
<td>90.25 kJ/mol</td>
</tr>
<tr>
<td>NO_2(g)</td>
<td>33.18 kJ/mol</td>
</tr>
<tr>
<td>N_2O_4(g)</td>
<td>9.16 kJ/mol</td>
</tr>
<tr>
<td>SO_2(g)</td>
<td>-296.8 kJ/mol</td>
</tr>
<tr>
<td>SO_3(g)</td>
<td>-395.7 kJ/mol</td>
</tr>
</tbody>
</table>

All values have units of kJ/mol and physical conditions of 298.15 K and 1 atm, referred to as the "standard state." These are the conditions under which values of standard enthalpies of formation are typically given. Note that while the majority of the values of standard enthalpies of formation are exothermic, or negative, there are a few compounds such as NO(g) and N_2O_4(g) that actually require energy from its surroundings during its formation; these endothermic compounds are generally unstable.

Example \(\PageIndex{2}\)

Between Br_2(l) and Br_2(g) at 298.15 K, which substance has a nonzero standard enthalpy of formation?

**SOLUTION**

Br_2(l) is the more stable form, which means it has the lower enthalpy; thus, Br_2(l) has ΔH_f = 0. Consequently, Br_2(g) has a nonzero standard enthalpy of formation.
Note: that the element phosphorus is a unique case. The reference form in phosphorus is not the most stable form, red phosphorus, but the less stable form, white phosphorus.

Recall that standard enthalpies of formation can be either positive or negative.

Example \(\PageIndex{3}\)

The enthalpy of formation of carbon dioxide at 298.15K is \(\Delta H = -393.5\) kJ/mol \(\text{CO}_2(g)\). Write the chemical equation for the formation of \(\text{CO}_2\).

**SOLUTION**

This equation must be written for one mole of \(\text{CO}_2(g)\). In this case, the reference forms of the constituent elements are \(\text{O}_2(g)\) and graphite for carbon.

\[
\text{O}_2(g) + \text{C(graphite)} \rightleftharpoons \text{CO}_2(g)
\]

The general equation for the standard enthalpy change of formation is given below:

\[
\Delta H_{\text{reaction}}^\circ = \sum \Delta H_{f}^\circ(\text{products}) - \sum \Delta H_{f}^\circ(\text{Reactants})
\]

Plugging in the equation for the formation of \(\text{CO}_2\) gives the following:

\[
\Delta H_{\text{reaction}}^\circ = \Delta H_{f}^\circ[\text{CO}_2(g)] - (\Delta H_{f}^\circ[\text{O}_2(g)] + \Delta H_{f}^\circ[\text{C(graphite)}])
\]

Because \(\text{O}_2(g)\) and \(\text{C(graphite)}\) are in their most elementally stable forms, they each have a standard enthalpy of formation equal to 0:

\[
\Delta H_{\text{reaction}}^\circ = -393.5\) kJ = \Delta H_{f}^\circ[\text{CO}_2(g)] - ((1 \text{ mol})(0 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}))
\]

\[
\Delta H_{f}^\circ[\text{CO}_2(g)] = -393.5\text{ kJ}
\]

Example \(\PageIndex{4}\)

Using the values in the above table of standard enthalpies of formation, calculate the \(\Delta H_{\text{reaction}}^\circ\) for the formation of \(\text{NO}_2\).

**SOLUTION**

\(\text{NO}_2(g)\) is formed from the combination of \(\text{NO}_2\) and \(\text{O}_2\) in the following reaction:

\[
2\text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g)
\]

To find \(\Delta H_{\text{reaction}}^\circ\), use the formula for the standard enthalpy change of formation:

\[
\Delta H_{\text{reaction}}^\circ = \sum \Delta H_{f}^\circ(\text{products}) - \sum \Delta H_{f}^\circ(\text{Reactants})
\]

The relevant standard enthalpy of formation values from Table 1 are:
Kirchhoff's Law describes the enthalpy of a reaction's variation with temperature changes. In general, enthalpy of any substance increases with temperature, which means both the products and the reactants' enthalpies increase. The overall enthalpy of the reaction will change if the increase in the enthalpy of products and reactants is different.

Kirchoff's Law - Enthalpy is Temperature Dependent

At constant pressure, the heat capacity is equal to change in enthalpy divided by the change in temperature.

\[ c_p = \frac{\Delta H}{\Delta T} \]  \(\text{Equation 1}\)

Therefore, if the heat capacities do not vary with temperature then the change in enthalpy is a function of the difference in temperature and heat capacities. The amount that the enthalpy changes by is proportional to the product of temperature change and change in heat capacities of products and reactants. A weighted sum is used to calculate the change in heat capacity to incorporate the ratio of the molecules involved since all molecules have different heat capacities at different states.

\[ H_{T_f} = H_{T_i} + \int_{T_i}^{T_f} c_{p} \, dT \]  \(\text{Equation 2}\)

If the heat capacity is temperature independent over the temperature range, then Equation \(\text{ref}{1}\) can be approximated as

\[ H_{T_f} = H_{T_i} + c_{p} \, (T_f - T_i) \]  \(\text{Equation 3}\)

with

- \( c_{p} \) is the (assumed constant) heat capacity and
- \( H_{T_i} \) and \( H_{T_f} \) are the enthalpy at the respective temperatures.

Equation \(\text{ref}{3}\) can only be applied to small temperature changes, (<100 K) because over a larger temperature change, the heat capacity is not constant. There are many biochemical applications because it allows us to predict enthalpy changes at other temperatures by using standard enthalpy data.
Contributors

- Janki Patel (UCD), Kostia Malley (UCD), Jonathan Nguyen (UCD), Garrett Larimer (UCD)