Hess’s Law

We’ve seen that energy changes can be measured in a calorimeter. But, how can anyone measure all of the energy changes in the world? What if the reaction is unpleasant, or is an explosion, or is very complex – how can we measure the energy change then? .....We use Hess’s Law to calculate the energy change from other systems.

Hess’s Law:
*If a reaction is carried out in a series of steps, $\Delta H$ for the reaction will be equal to the sum of enthalpy ($H$) changes for the individual steps.*

**Example 1**

Calculate the enthalpy change, $\Delta H$ (in KJ), for the reaction

$$2\text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + \text{Al}_2\text{O}_3(s) \quad \Delta H = ?$$

Use the enthalpy changes for the combustion of aluminum and iron:

a. $2\text{Al} (s) + 3/2 \text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \quad \Delta H = -1669.8\text{kJ}$

b. $2\text{Fe} (s) + 3/2 \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \quad \Delta H = -824.2\text{kJ}$

**Solution**

Notice that to achieve the desired reaction, equation (b) must be written in reverse. When this is done, be sure to change the sign of $\Delta H$ since the energy is now flowing in the opposite direction. Call this equation (c)

(c) $\text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + 3/2 \text{O}_2(g) \quad \Delta H = +824.2\text{kJ}$

Now, add equations (a) and (c) to obtain the final answer. As in algebra, like terms cancel out.

$$2\text{Al} (s) + 3/2 \text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \quad \Delta H = -1669.8\text{kJ}$$

$$\text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + 3/2 \text{O}_2(g) \quad \Delta H = +824.2\text{kJ}$$

$$2\text{Al}(s) + \text{Fe}_2\text{O}_3 \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3 \quad \Delta H = -854.6\text{kJ}$$

**Key Points**

- If a reaction is reversed, the sign of $\Delta H$ is also reversed.
- The magnitude of $\Delta H$ is directly proportional to the quantities of reactants and products in a reaction. If the coefficients in a balanced reaction are multiplied by an integer, the value of $\Delta H$ is multiplied by the same integer.
**Example 2**

<table>
<thead>
<tr>
<th>Reaction</th>
<th>ΔH</th>
</tr>
</thead>
<tbody>
<tr>
<td>S(s) + O(_2)(g) → SO(_2)(g)</td>
<td>-395.2 kJ</td>
</tr>
<tr>
<td>2SO(_2)(g) + O(_2)(g) → 2SO(_3)(g)</td>
<td>-198.2 kJ</td>
</tr>
</tbody>
</table>

From the following information:

(a) \(S(s) + 3/2O_2(g) → SO_3(g)\) \(ΔH = -395.2\ kJ\)
(b) \(2SO_2(g) + O_2(g) → 2SO_3(g)\) \(ΔH = -198.2\ kJ\)

**Solution**

To obtain the reactants and products in the desired reaction, we need to reverse equation (b) and multiply it by \(\frac{1}{2}\). This action reverses the sign and cuts the amount of energy by a factor of 2:

\[
\frac{1}{2}[2SO_3(g) → 2SO_2(g) + O_2(g)] \quad ΔH = +198.2\ kJ
\]

OR

\[
SO_3(g) → SO_2(g) + 1/2O_2(g) \quad ΔH = +99.1\ kJ
\]

Now we add this reaction to the first reaction:

\[
\begin{align*}
S(s) + 3/2O_2(g) & → SO_3(g) \quad ΔH = -395.2\ kJ \\
& \quad \text{or} \quad SO_2(g) \rightarrow SO_2(g) + 1/2O_2(g) \quad ΔH = +99.1\ kJ
\end{align*}
\]

\[
S(s) + O_2(g) \rightarrow SO_2(g) \quad ΔH = -296.1\ kJ
\]