Hess's Law

When a final reaction is the sum of one or more individual reactions, ΔH for the final reaction is the sum of the ΔH values for the individual reactions. Changing the direction of the reaction flips the sign of ΔH, so that an **exothermic** reaction would become **endothermic**. Multiplying the reaction by a factor affects ΔH by the same factor. The example below explains the decision making process for each reaction in a final reaction that is the sum of four individual steps.

Example: Calculate ΔH for this final reaction

$2B(s) + (3/2)O_2(g) \rightarrow B_2O_3(s)$ ΔH = ?

using these individual reactions

Reaction 1: $B_2O_3(s) + 3H_2O(g) \rightarrow 3O_2(g) + B_2H_6(g)$ ΔH = +2035 kJ
Reaction 2: $H_2O(l) \rightarrow H_2O(g)$ ΔH = +44 kJ
Reaction 3: $H_2(g) + (1/2)O_2(g) \rightarrow H_2O(l)$ ΔH = −286 kJ
Reaction 4: $2B(s) + 3H_2 \rightarrow B_2H_6(g)$ ΔH = +36 kJ

Solution: Note what side the reactants and products are on in the final reaction as well as the coefficients as you review each individual reaction. Key individual reactants and products are displayed in bold text.

With Reaction 1, $B_2O_3(s)$ appears as a product in the final reaction with a coefficient of one, so reverse the direction of Reaction 1 and change the sign of ΔH. Note that $B_2O_3(s)$ does not appear in any other individual reaction, so this is a good reaction to begin the process.

Reaction 1: $\mathbf{B_2O_3(s)} + 3H_2O(g) \rightarrow 3O_2(g) + B_2H_6(g)$ ΔH = +2035 kJ
Reaction 1a: $3O_2(g) + B_2H_6(g) \rightarrow \mathbf{B_2O_3(s)} + 3H_2O(g)$ ΔH = −2035 kJ

For Reaction 2, H_2O does not appear in the final reaction so skip this reaction for now.
Reaction 2: $H_2O(l) \rightarrow H_2O(g)$ ΔH = +44 kJ

In Reaction 3, $(1/2)O_2(g)$ appears as a reactant, and it is needed as a reactant in the final reaction. Reaction 1a already provides $3O_2(g)$ as a reactant, yet only (3/2) $O_2(g)$ are needed in the final reaction. Because the third reaction is the only other source of $O_2(g)$, reverse the direction of Reaction 3 and multiply it by three to leave the correct moles of $O_2(g)$ as a reactant after adding Reaction 1a to Reaction 3a.

Reaction 3: $H_2(g) + \mathbf{(1/2)O_2}(g) \rightarrow H_2O(l)$ ΔH = −286 kJ
Reaction 3a: $3H_2O(l) \rightarrow 3H_2(g) + \mathbf{(3/2)O_2}(g)$ ΔH = +858 kJ

The first intermediate result (I1) and ΔH value after adding Reaction 1a to Reaction 3a are

$3O_2(g) + B_2H_6(g) \rightarrow B_2O_3(s) + 3H_2O(g)$ ΔH = −2035 kJ
$3H_2O(l) \rightarrow 3H_2(g) + (3/2)O_2(g)$ ΔH = +858 kJ

I1: $3H_2O(l) + (3/2)\ O_2(g) + B_2H_6(g) \rightarrow B_2O_3(s) + 3H_2O(g) + 3H_2(g)$ $\Delta H = -1177$ kJ

Looking at Reaction 4, 2B(s) appears as a reactant, and it appears as a reactant in the final reaction with a coefficient of two. Note also that $3H_2$ and $B_2H_6(g)$ will cancel when Reaction 4 is added to I1.

Reaction 4: $\mathbf{2B(s)} + 3H_2 \rightarrow B_2H_6(g)$ $\Delta H = +36$ kJ

The second intermediate result (I2) and ΔH value after adding Reaction I1 to Reaction 4 are

I1: $3H_2O(l) + (3/2)\ O_2(g) + \mathbf{B_2H_6(g)} \rightarrow B_2O_3(s) + 3H_2O(g) + \mathbf{3H_2(g)}$ $\Delta H = -1177$ kJ

4: $\mathbf{2B(s) + 3H_2} \rightarrow \mathbf{B_2H_6(g)}$ $\Delta H = +36$ kJ

I2: $3H_2O(l) + 2B(s) + (3/2)\ O_2(g) \rightarrow B_2O_3(s) + 3H_2O(g)$ $\Delta H = -1141$ kJ

Comparing I2 to the final reaction, three water molecules appear on each side in I2 but not in the final reaction. Reaction 2 has $H_2O(l)$ as a reactant with a coefficient of one. I2 has $H_2O(l)$ as a reactant with a coefficient of three, so reverse Reaction 2 and multiply the reactiona and ΔH value by three so that $H_2O(l)$ will cancel. Note that $H_2O(g)$ will also cancel leaving the final reaction and the final ΔH value.

Reaction 2: $\mathbf{H_2O(l)} \rightarrow H_2O(g)$ $\Delta H = +44$ kJ

Reaction 2a: $3H_2O(g) \rightarrow \mathbf{3H_2O(l)}$ $\Delta H = -132$ kJ

Add I2 to Reaction 2a to get the final answer and ΔH value.

I2: $3H_2O(l) + 2B(s) + (3/2)\ O_2(g) \rightarrow B_2O_3(s) + 3H_2O(g)$ $\Delta H = -1141$ kJ

2a: $3H_2O(g) \rightarrow \mathbf{3H_2O(l)}$ $\Delta H = -132$ kJ

Final: $2B(s) + (3/2)O_2(g) \rightarrow B_2O_3(s)$ $\Delta H = -1273$ kJ

The final reaction and ΔH value are

$2B(s) + (3/2)O_2(g) \rightarrow B_2O_3(s)$ $\Delta H = -1273$ kJ