In Hess's Law problems, you add, subtract, multiply, and divide chemical equations in order to create the equation you are interested in. Whatever you do to any of the equations, you also do to the ΔH for that equation. If you reverse an equation, you change the sign on the ΔH.

**Post-lab 8 problem 4:** Given the following thermodynamic data,

$2\ H_2O\ (g) \rightarrow O_2\ (g) + 2\ H_2\ (g)$ $\quad \Delta H_{rxn} = 483.6$ kJ

$H_2O\ (g) \rightarrow H_2O\ (l)$ $\quad \Delta H_{rxn} = -44.0$ kJ

What is the molar heat of formation for liquid water, $H_2O\ (l)$?

**Answer:**

The first thing you need to know to solve this is what equation you are trying to build.

The molar heat of formation is the ΔH for the formation of one mole of product from its elements in their standard states. So when calculating the molar heat of formation, your product must be one mole of the compound in question, *and nothing else*, and your reactants can only be pure elements (no compounds), in their standard states.

So the equation you are trying to build in this case is:

$\frac{1}{2}\ O_2\ (g) + H_2\ (g) \rightarrow H_2O\ (l)$ $\quad \Delta H_f = ?$ kJ

These are the equations you are given:

$2\ H_2O\ (g) \rightarrow O_2\ (g) + 2\ H_2\ (g)$ $\quad \Delta H_{rxn} = 483.6$ kJ

$H_2O\ (g) \rightarrow H_2O\ (l)$ $\quad \Delta H_{rxn} = -44.0$ kJ

The reactants you want in your final equation are the products in the first reaction, so you will want to flip that equation around. When you reverse the equation, you change the sign on the ΔH value. Reversing the first equation gives you 2 moles of $H_2O\ (g)$ as a product, so you will want to multiply the second equation by 2 to cancel out those $H_2O\ (g)$ molecules. When you multiply the second equation by 2, you also multiply the ΔH by 2:

$O_2\ (g) + 2\ H_2\ (g) \rightarrow 2\ H_2O\ (g)$ $\quad \Delta H_{rxn} = -483.6$ kJ
$2\ H_2O\ (g) \rightarrow 2\ H_2O\ (l)$ $\quad \Delta H_{rxn} = -88.0$ kJ

These add up to:

$O_2\ (g) + 2\ H_2\ (g) \rightarrow 2\ H_2O\ (l)$ $\quad \Delta H_{rxn} = -571.6$ kJ

But the heat of formation is for one mole of product, not two, so you need to divide everything by 2:

**$\frac{1}{2}\ O_2\ (g) + H_2\ (g) \rightarrow H_2O\ (l)$** $\quad$ **$\Delta H_f = -285.8$ kJ** $\quad$ And that's your answer.