CHEM 101A Chapter 5 Thermochemistry – Hess' Law Example *Explained*

(given equation) $C_6H_{14}(OH)_2(aq) + H_2O_2(aq) \longrightarrow C_6H_4O_2(aq) + 2H_2O(l)$ **ΔH_{rxn} = ?**

Calculate ΔH for this reaction from the following data:

(1)	$C_6H_{14}(OH)_2(aq) \longrightarrow$	$C_6H_4O_2(aq) + H_2(g)$	$\Delta H_1 = + 177.4$ kJ
(2)	$H_2(g) + O_2(g) \longrightarrow$	$H_2O_2(aq)$	$\Delta H_2 = - 191.2$ kJ
(3)	$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow$	$H_2O(g)$	$\Delta H_3 = - 241.8$ kJ
(4)	$H_2O(g) \longrightarrow$	$H_2O(l)$	$\Delta H_4 = - 43.8$ kJ

The goal is to manipulate the equations so that they add up to the given equation. We look at the given reaction and identify species in the other reactions. Here is what we observe:

- In Equation (1), The $C_6H_{14}(OH)_2(aq)$ species is on the reactant side, and also $C_6H_4O_2(aq)$ species is on the product side. Since that is how they appear exactly in the given equation, we keep Equation (1) as is, and its ΔH stays the same.
- In Equation (2), the product is $H_2O_2(aq)$. Equation (2) is the only equation with that species, and the given equation has $H_2O_2(aq)$ on the reactant side, so we choose to reverse this and thus change the sign of enthalpy to + 191.2 kJ
- Equation (4) is the only equation that contains $H_2O(l)$. It is on the product side as in the given equation but the coefficient is 1, thus we multiply Equation (4) by 2 and also its enthalpy (ΔH = -87.6 kJ).
- Equation (3) does not contain any species in the given equation. It must be provided to eliminate species introduced from equations (1), (2), and (4).

The manipulated equations now look like this, let's sum them and their ΔH values:

(1)	$C_6H_{14}(OH)_2(aq) \longrightarrow$	$C_6H_4O_2(aq) + H_2(g)$	$\Delta H_1 = + 177.4$ kJ
(2*)	$H_2O_2(aq) \longrightarrow$	$H_2(g) + O_2(g)$	$\Delta H_{2*} =$ **+** 191.2 kJ
(4*)	$2H_2O(g) \longrightarrow$	$2H_2O(l)$	$\Delta H_{4*} = $ - 87.6.8 kJ

(1,2*,4*) $C_6H_{14}(OH)_2(aq) + H_2O_2(aq) + 2H_2O(g) \longrightarrow$

$C_6H_4O_2(aq) + 2H_2(g) + O_2(g) + 2H_2O(l)$ $\Delta H = 281.0$ kJ

No we need to manipulate Equation (3) so that it cancels species we don't want. The equation (1,2*,4*) is rewritten to indicate the species not in the given equation:

(1,2*,4*) $C_6H_{14}(OH)_2(aq) + H_2O_2(aq) + \boxed{2H_2O(g)} \longrightarrow$

$$C_6H_4O_2(aq) + \boxed{2H_2(g)} + \boxed{O_2(g)} + 2H_2O(l) \qquad \Delta H_{1,2^*,4^*} = 281.0\ kJ$$

Looking at Equation (3), again, we can see that all of the same species as in (1,2,4), but on other side of the equation. We want this, because they will cancel when we sum the equation. However, We need to multiply Equation (3) by 2 in order to match coefficients, and also multiply its ΔH by 2. It will now look like:

(3*) $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \qquad \Delta H_{3^*} = -483.6\ kJ$

Now, when we add (1,2,4) and (3) the species in Equation (3) will cancel and it will match the given equation. Now, all that is left to do is add the last two enthalpy values for the given reaction enthalpy:

ΔH_{rxn} = 281.0 kJ + (-483.6 kJ) = **- 202.6 kJ**