Values of H_f^o from Standard Enthalpies of Combustion

To measure directly the heat of formation of sucrose, $C_{12}H_{22}O_{11}$, you would have to carry out the following reaction:

$$12\ C(s) + 11\ H_2(g) + 5\tfrac{1}{2}\ O_2(g) \text{ ------> } C_{12}H_{22}O_{11}$$

But no one has ever been able to figure out how to make this reaction occur directly under any conditions, so there is no direct way to measure the H_f^o of sucrose. How, then, can we obtain values of H_f^o for compounds such as sucrose?

If the compound in question can be burned - which is usually far easier to do than make it from its elements - then we have a source of energy data that we can use to calculate its H_f^o. This is because the products of combustion are nearly always compounds whose values of H_f^o are known or can be measured by direct means. The combustion of sucrose in an atmosphere of pure oxygen proceeds by the following equation:

$$C_{12}H_{22}O_{11}(s) + 12\ O_2(g) \text{ -----> } 12\ CO_2(g) + 11\ H_2O(l)$$

If the standard enthalpy change for this reaction can be measured, and if we can look up the values of H_f^o for three of the four chemicals in the equation, then we can use the Hess law equation to find the H_f^o of the remaining substance, sucrose.

The above equation is a combustion equation and your data tables have H_c^o values (standard enthalpy of combustion). Putting these into a Hess's Law equation you should get:

$$H_c^o = [(12)CO_2(g) + (11)H_2O(l)] - [C_{12}H_{22}O_{11}(s) + (12)O_2(g)]$$

All values are in kJ/moles.

$$-5639.7 = [(12)\text{-}393.5 + (11)\text{-}285.8] - [C_{12}H_{22}O_{11}(s) + (12)0]$$
$$-5639.7 = -4722 - 3143.8 - [C_{12}H_{22}O_{11}]$$
$$2226.1 = -[C_{12}H_{22}O_{11}]$$

Therefore the H_f^o of sucrose, $C_{12}H_{22}O_{11}$, is -2226.1 kJ/mole

Sample problem

One of the "building blocks" for proteins such as those in muscles and sinews is an amino acid called glycine, $C_2H_5NO_2$. The equation for its combustion is

$$4\ C_2H_5NO_2(s) + 9\ O_2(g) \text{--->} 8\ CO_2(g) + 10\ H_2O(l) + 2\ N_2(g)$$

The value of H_c^o for glycine is -973.49 kJ/mole. Using this information and the values of H_f^o calculate the H_f^o for glycine.

Solution

For this problem, Hess's law equation becomes

$$H^o = [(8)CO_2(g) + (10)H_2O(l) + (2)N_2(g)]-[(4)C_2H_5NO_2(s) + (9)O_2(g)]$$

No, I didn't forget the H_c^o. The first term H^o, is obtained from the standard heat of combustion of glycine. Since the chemical equation for this reaction is for the combustion of *four* moles of glycine, we have to multiple $H^o_{combustion}$ by four.

$$H^o = 4 \text{ mol x } -973.49 \text{ kJ/mol} = -3894.0 \text{ kJ}$$

Now we can substitute into Hess's law equation the correct values.

$$-3894.0 \text{ kJ} = [(8)\text{-}393.5 + (10)\text{-}285.8 + (2)0] - [(4)C_2H_5NO_2 + (9)0]$$
$$-3894.0 = -3148.0 -2858 -[(4)C_2H_5NO_2]$$
$$2112.0 = -[(4)C_2H_5NO_2]$$

Therefore by rearranging we get

$$H_f^o \text{ for } C_2H_5NO_2 = \frac{-2112.0 \text{ kJ/mole}}{4 \text{ moles}} = -528.0 \text{ kJ/mole}$$

Thus, the standard heat of formation of glycine is -528.0 kJ/mol, and we have seen how we can determine this quantity without making glycine directly from its elements.

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If the compound in question can be burned - which is usually far easier to do than make it from its elements - then we have a source of energy data that we can use to calculate its H_f^o. This is because the products of combustion are nearly always compounds whose values of H_f^o are known or can be measured by direct means. The combustion of sucrose in an atmosphere of pure oxygen proceeds by the following equation:

$$C_{12}H_{22}O_{11}(s) + 12\ O_2(g) \text{ -----> } 12\ CO_2(g) + 11\ H_2O(l)$$

If the standard enthalpy change for this reaction can be measured, and if we can look up the values of H_f^o for three of the four chemicals in the equation, then we can use the enthalpies of formation equation to find the H_f^o of the remaining substance, sucrose.

The above equation is a combustion equation and data tables would have H_c^o values (standard enthalpy of combustion). The combustion of sucrose has a value of -5639.7 kJ/mol :

$H_c^o = -5639.7$ kJ/mol $\qquad$ $H_c^o = [(12)CO_2(g) + (11)H_2O(l)] - [C_{12}H_{22}O_{11}(s) + (12)O_2(g)]$

All values are in kJ/moles.

$$-5639.7 = [(12)\text{-}393.5 + (11)\text{-}285.8] - [C_{12}H_{22}O_{11}(s) + (12)0]$$
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The value of H_c^o for glycine is -973.49 kJ/mole. Using this information and the values of H_f^o calculate the H_f^o for glycine.

Solution

For this problem, enthalpy of formation equation becomes

$\Delta H^o = [(8)CO_2(g) + (10)H_2O(l) + (2)N_2(g)]-[(4)C_2H_5NO_2(s) + (9)O_2(g)]$

No, I didn't forget the H_c^o. The first term ΔH^o, is obtained from the standard heat of combustion of glycine. Since the chemical equation for this reaction is for the combustion of *four* moles of glycine, we have to multiple $H^o_{combustion}$ by four.

$H^o = 4 \text{ mol} \times -973.49 \text{ kJ/mol} = -3894.0 \text{ kJ}$

Now we can substitute into enthalpy of formation equation the correct values.

$$-3894.0 \text{ kJ} = [(8)-393.5 + (10)-285.8 + (2)0] - [(4)C_2H_5NO_2 + (9)0]$$
$$-3894.0 = -3148.0 -2858 -[(4)C_2H_5NO_2]$$
$$2112.0 = -[(4)C_2H_5NO_2]$$

Therefore by rearranging we get

$$H_f^o \text{ for } C_2H_5NO_2 = \frac{-2112.0 \text{ kJ/mole}}{4 \text{ moles}} = -528.0 \text{ kJ/mole}$$

Thus, the standard heat of formation of glycine is -528.0 kJ/mol, and we have seen how we can determine this quantity without making glycine directly from its elements.