Practice Questions Section 2.4

Hess's Law

1. Given the following equations:

   $4\ NH_{3\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(l)}$ $\Delta H° = -1170$ kJ

   $4\ NH_{3\,(g)} + 3\ O_{2\,(g)} \rightarrow 2\ N_{2\,(g)} + 6\ H_2O_{(l)}$ $\Delta H° = -1530$ kJ

   Using these two equations, determine the heat of formation, $\Delta H°_f$, for nitrogen monoxide, NO.

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Hess's Law Answers

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Solution

Recall that a formation reaction describes the formation of **one mole** of the compound **from its elements**. Thus, the formation reaction for nitrogen monoxide, NO, is:

$\frac{1}{2}\ N_{2\,(g)} + \frac{1}{2}\ O_{2\,(g)} \rightarrow NO_{(g)}$

Next, examine the two reactions given to you. Some things you will want to notice:

Both reactions contain ammonia, $NH_{3\,(g)}$, which does not appear in the formation reaction. It must get cancelled out when we add the equations together.

Similarly, $H_2O_{(l)}$ must also be cancelled out since it does not appear in our desired equation.

Oxygen also appears in both equations, as well as in our desired reaction. It will be of no help to us in decided which, if any, of the equations need to be reversed.

NO only appears in one of the original equations (equation 1), and it also appears in our desired equation. In Equation 1, NO is on the product side of the equation, which is where we want it to appear in our final equation. So, we don't want to reverse equation 1.

N_2 appears only in Reaction 2 and in our desired equation. In the formation reaction, N_2 appears on the reactant side of the equation, but it appears on the product side of Reaction 2. Therefore we must reverse the entire equation.

IMPORTANT - when you reverse the equation, the sign in front of ΔH will change.

This gives us the following. Reaction 1 is unchanged; reaction 2 has been reversed:

$4\ NH_{3\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(l)}$ $\Delta H° = -1170$ kJ

$2\ N_{2\,(g)} + 6\ H_2O_{(l)} \rightarrow 4\ NH_{3\,(g)} + 3\ O_{2\,(g)}$ $\Delta H° = +1530$ kJ

Next, look carefully at the coefficients in the balanced equation. Remember, we want both NH_3 and H_2O to get cancelled out, as they do not appear in the final equation. We find that if we now add together our two equations, both NH_3 and H_2O will indeed cancel out (4 NH_3 on the reactant side cancels out the 4 NH_3 on the product side; H_2O also cancels out as 6 moles appear on both sides of the equation).

Notice O_2 - the number of moles are not the same on the two sides of the equation. This will leave us with a net of 2 moles of O_2 on the reactant side of the equation.

Once we can add up the reactants and products, we can then add the $\Delta H°$ values:

$4\ NH_{3\ (g)}$	$+\ 5\ O_{2\ (g)}$	$\rightarrow$	$4\ NO_{\ (g)}$	$+$	$6\ H_2O_{\ (l)}$	$\Delta H° = -1170$ kJ
$2\ N_{2\ (g)}$	$+\ 6\ H_2O_{\ (l)}$	$\rightarrow$	$4\ NH_{3\ (g)}$	$+$	$3\ O_{2\ (g)}$	$\Delta H° = +1530$ kJ
$2\ N_{2\ (g)}$	$+\ 2\ O_{2\ (g)}$	$\rightarrow$	$4\ NO_{\ (g)}$			$\Delta H° = +360$ kJ

We are almost done, but not quite. A heat of formation reaction calls for the production of one mole of the compound, NO. Our equation above produces 4 moles. Therefore, we need to divide everything in the equation by 4, including ΔH°.

Thus, our final answer is: $\frac{1}{2}\ N_{2\ (g)} + \frac{1}{2}\ O_{2\ (g)} \rightarrow 4\ NO_{\ (g)}$ $\Delta H° = +90$ kJ