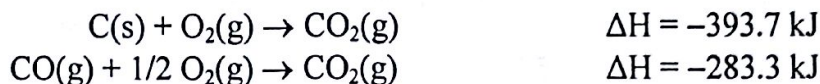
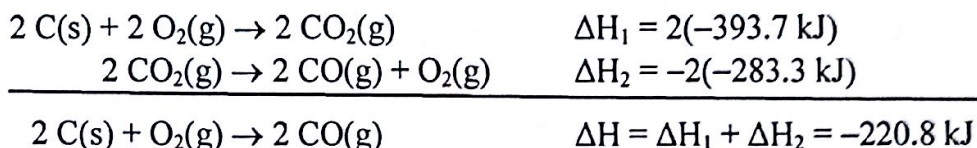


Hess's Law

69. Information given:

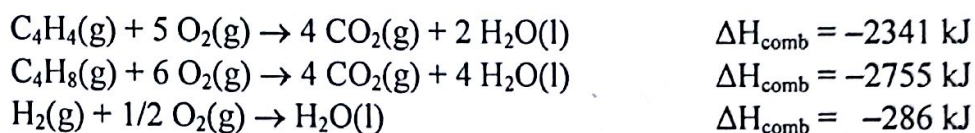


Using Hess's law:



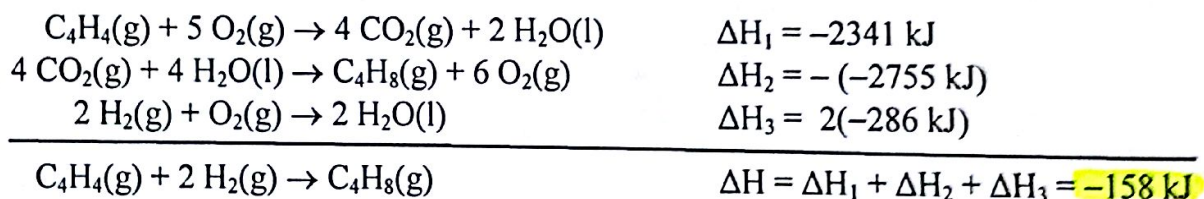
Note: When an equation is reversed, the sign on ΔH is reversed. When the coefficients in a balanced equation are multiplied by an integer, then the value of ΔH is multiplied by the same integer.

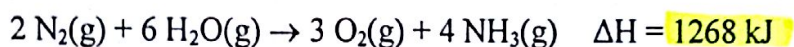
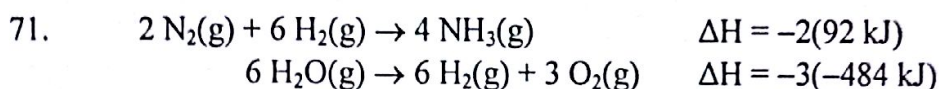
70. Given:



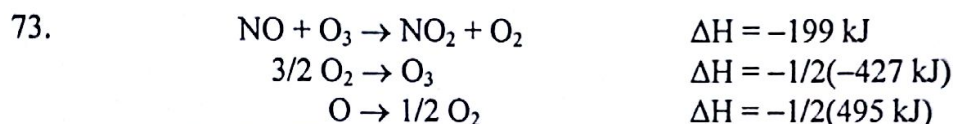
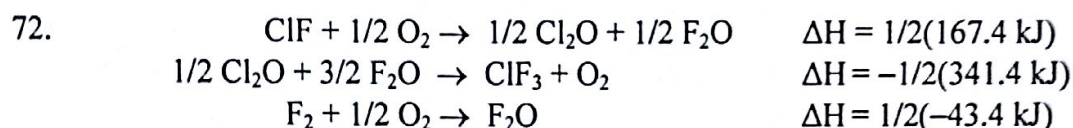
By convention, $\text{H}_2\text{O(l)}$ is produced when enthalpies of combustion are given, and because per-mole quantities are given, the combustion reaction refers to 1 mole of that quantity reacting with $\text{O}_2\text{(g)}$.

Using Hess's law to solve:

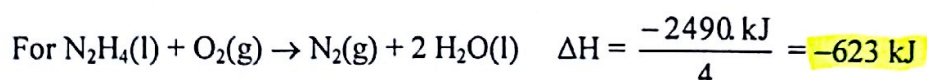
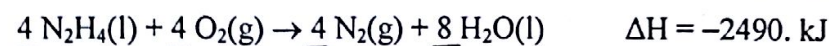
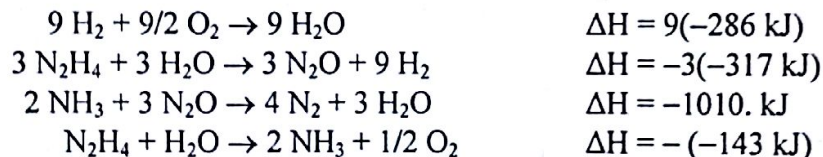




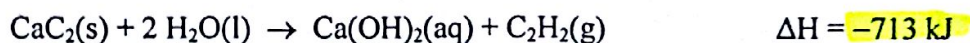
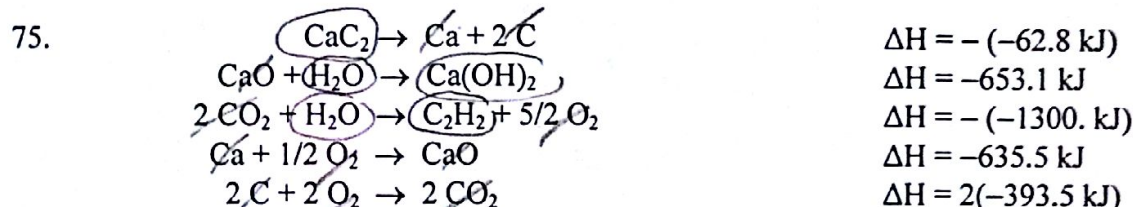
No, because the reaction is very endothermic (requires a lot of heat to react), it would not be a practical way of making ammonia because of the high energy costs required.

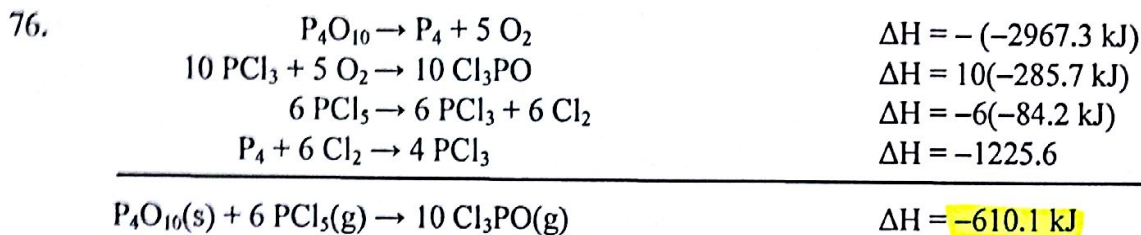


74. We want ΔH for $\text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$. It will be easier to calculate ΔH for the combustion of four moles of N_2H_4 because we will avoid fractions.



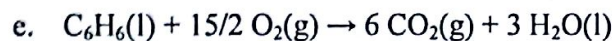
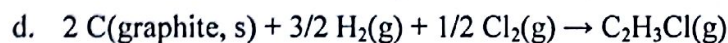
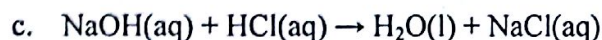
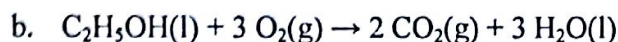
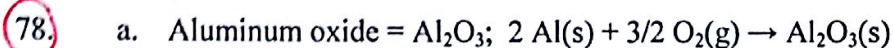
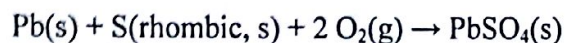
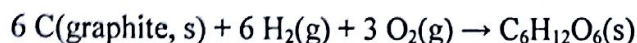
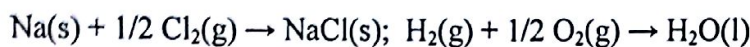
Note: By the significant figure rules, we could report this answer to four significant figures. However, because the ΔH values given in the problem are only known to ± 1 kJ, our final answer will at best be ± 1 kJ.



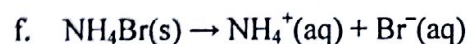


Standard Enthalpies of Formation

77. The change in enthalpy that accompanies the formation of 1 mole of a compound from its elements, with all substances in their standard states, is the standard enthalpy of formation for a compound. The reactions that refer to ΔH_f° are:



Note: ΔH_{comb} values assume 1 mole of compound combusted.



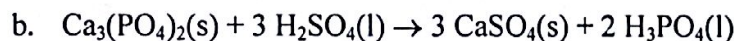
79. In general, $\Delta H^\circ = \sum n_p \Delta H_{f, \text{products}}^\circ - \sum n_r \Delta H_{f, \text{reactants}}^\circ$, and all elements in their standard state have $\Delta H_f^\circ = 0$ by definition.

a. The balanced equation is $2 NH_3(g) + 3 O_2(g) + 2 CH_4(g) \rightarrow 2 HCN(g) + 6 H_2O(g)$.

$$\Delta H^\circ = (2 \text{ mol HCN} \times \Delta H_{f, \text{HCN}}^\circ + 6 \text{ mol H}_2\text{O}(g) \times \Delta H_{f, \text{H}_2\text{O}}^\circ)$$

$$- (2 \text{ mol NH}_3 \times \Delta H_{f, \text{NH}_3}^\circ + 2 \text{ mol CH}_4 \times \Delta H_{f, \text{CH}_4}^\circ)$$

$$\Delta H^\circ = [2(135.1) + 6(-242)] - [2(-46) + 2(-75)] = -940. \text{ kJ}$$



$$\Delta H^\circ = \left[3 \text{ mol CaSO}_4(\text{s}) \left(\frac{-1433 \text{ kJ}}{\text{mol}} \right) + 2 \text{ mol H}_3\text{PO}_4(\text{l}) \left(\frac{-1267 \text{ kJ}}{\text{mol}} \right) \right] \\ - \left[1 \text{ mol Ca}_3(\text{PO}_4)_2(\text{s}) \left(\frac{-4126 \text{ kJ}}{\text{mol}} \right) + 3 \text{ mol H}_2\text{SO}_4(\text{l}) \left(\frac{-814 \text{ kJ}}{\text{mol}} \right) \right]$$

$$\Delta H^\circ = -6833 \text{ kJ} - (-6568 \text{ kJ}) = -265 \text{ kJ}$$

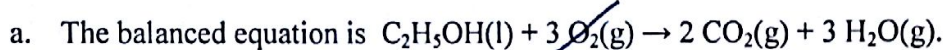


$$\Delta H^\circ = (1 \text{ mol NH}_4\text{Cl} \times \Delta H_{\text{f}, \text{NH}_4\text{Cl}}^\circ) - (1 \text{ mol NH}_3 \times \Delta H_{\text{f}, \text{NH}_3}^\circ + 1 \text{ mol HCl} \times \Delta H_{\text{f}, \text{HCl}}^\circ)$$

$$\Delta H^\circ = \left[1 \text{ mol} \left(\frac{-314 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-46 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-92 \text{ kJ}}{\text{mol}} \right) \right]$$

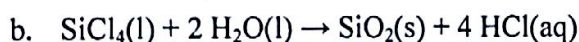
$$\Delta H^\circ = -314 \text{ kJ} + 138 \text{ kJ} = -176 \text{ kJ}$$

80.



$$\Delta H^\circ = \left[2 \text{ mol} \left(\frac{-393.5 \text{ kJ}}{\text{mol}} \right) + 3 \text{ mol} \left(\frac{-242 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-278 \text{ kJ}}{\text{mol}} \right) \right]$$

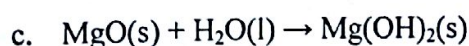
$$\Delta H^\circ = -1513 \text{ kJ} - (-278 \text{ kJ}) = -1235 \text{ kJ}$$



Because $\text{HCl}(\text{aq})$ is $\text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$, $\Delta H_{\text{f}}^\circ = 0 - 167 = -167 \text{ kJ/mol}$.

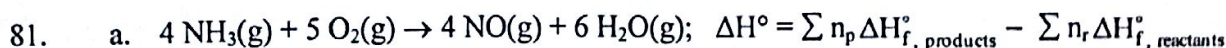
$$\Delta H^\circ = \left[4 \text{ mol} \left(\frac{-167 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-911 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-687 \text{ kJ}}{\text{mol}} \right) + 2 \text{ mol} \left(\frac{-286 \text{ kJ}}{\text{mol}} \right) \right]$$

$$\Delta H^\circ = -1579 \text{ kJ} - (-1259 \text{ kJ}) = -320. \text{ kJ}$$

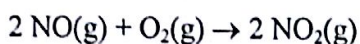


$$\Delta H^\circ = \left[1 \text{ mol} \left(\frac{-925 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-602 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-286 \text{ kJ}}{\text{mol}} \right) \right]$$

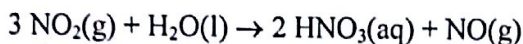
$$\Delta H^\circ = -925 \text{ kJ} - (-888 \text{ kJ}) = -37 \text{ kJ}$$



$$\Delta H^\circ = \left[4 \text{ mol} \left(\frac{90. \text{ kJ}}{\text{mol}} \right) + 6 \text{ mol} \left(\frac{-242 \text{ kJ}}{\text{mol}} \right) \right] - \left[4 \text{ mol} \left(\frac{-46 \text{ kJ}}{\text{mol}} \right) \right] = -908 \text{ kJ}$$



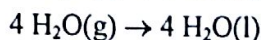
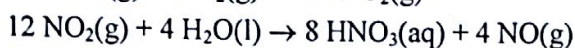
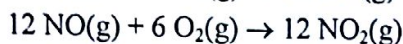
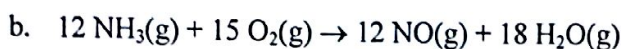
$$\Delta H^\circ = \left[2 \text{ mol} \left(\frac{34 \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \text{ mol} \left(\frac{90. \text{ kJ}}{\text{mol}} \right) \right] = -112 \text{ kJ}$$



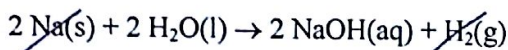
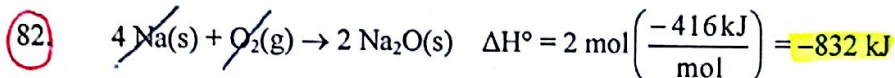
$$\Delta H^\circ = \left[2 \text{ mol} \left(\frac{-207 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{90. \text{ kJ}}{\text{mol}} \right) \right] - \left[3 \text{ mol} \left(\frac{34 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-286 \text{ kJ}}{\text{mol}} \right) \right]$$

-140. kJ

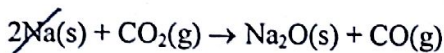
Note: All ΔH_f° values are assumed ± 1 kJ.



The overall reaction is exothermic because each step is exothermic.

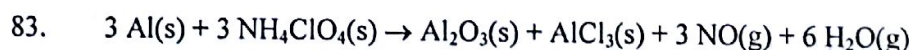


$$\Delta H^\circ = \left[2 \text{ mol} \left(\frac{-470. \text{ kJ}}{\text{mol}} \right) \right] - \left[2 \text{ mol} \left(\frac{-286 \text{ kJ}}{\text{mol}} \right) \right] = -368 \text{ kJ}$$



$$\Delta H^\circ = \left[1 \text{ mol} \left(\frac{-416 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-110.5 \text{ kJ}}{\text{mol}} \right) \right] - \left[1 \text{ mol} \left(\frac{-393.5 \text{ kJ}}{\text{mol}} \right) \right] = -133 \text{ kJ}$$

In Reactions 2 and 3, sodium metal reacts with the "extinguishing agent." Both reactions are exothermic, and each reaction produces a flammable gas, H_2 and CO , respectively.



$$\Delta H^\circ = \left[6 \text{ mol} \left(\frac{-242 \text{ kJ}}{\text{mol}} \right) + 3 \text{ mol} \left(\frac{90. \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-704 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left(\frac{-1676 \text{ kJ}}{\text{mol}} \right) \right]$$

$$- \left[3 \text{ mol} \left(\frac{-295 \text{ kJ}}{\text{mol}} \right) \right] = -2677 \text{ kJ}$$