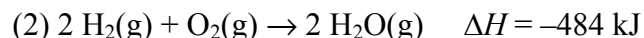# Section 5.4: Hess's Law
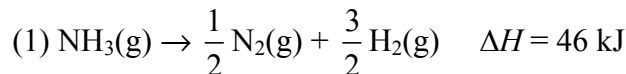
Tutorial 1 Practice, page 317
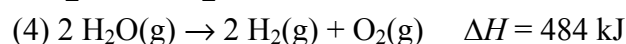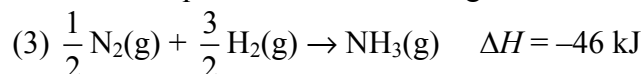
1. (a) Solution:

Step 1: Label the equations.

The desired equation is: $2 N_2(g) + 6 H_2O(g) \rightarrow 3 O_2(g) + 4 NH_3(g)$

(1) $NH_3(g) \rightarrow \dfrac{1}{2} N_2(g) + \dfrac{3}{2} H_2(g)$ $\Delta H = 46$ kJ

(2) $2 H_2(g) + O_2(g) \rightarrow 2 H_2O(g)$ $\Delta H = -484$ kJ

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation. Reverse (1) and (2) so that reactants and products are on the same side as in the desired equation. Reverse the sign of ΔH.

(3) $\dfrac{1}{2} N_2(g) + \dfrac{3}{2} H_2(g) \rightarrow NH_3(g)$ $\Delta H = -46$ kJ

(4) $2 H_2O(g) \rightarrow 2 H_2(g) + O_2(g)$ $\Delta H = 484$ kJ

Step 3: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.
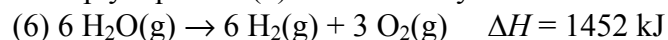
Equation (3) has $\dfrac{1}{2}$ mol of nitrogen gas on the reactant side, while the desired equation has

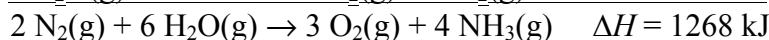2 mol. Multiply equation (3) and its ΔH by the factor 4:

(5) $2 N_2(g) + 6 H_2(g) \rightarrow 4 NH_3(g)$ $\Delta H = -184$ kJ
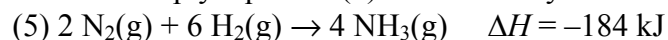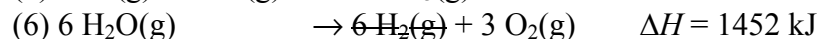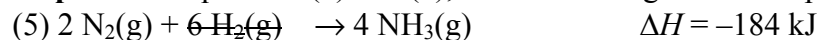
Equation (4) has 2 mol of water on the reactant side, while the desired equation has 6 mol. Multiply equation (4) and its ΔH by the factor 3:

(6) $6 H_2O(g) \rightarrow 6 H_2(g) + 3 O_2(g)$ $\Delta H = 1452$ kJ

Step 4: Add equations (5) and (6), and their changes in enthalpies.

(5) $2 N_2(g) + \cancel{6 H_2(g)} \rightarrow 4 NH_3(g)$ $\Delta H = -184$ kJ

(6) $\underline{6 H_2O(g) \rightarrow \cancel{6 H_2(g)} + 3 O_2(g) \qquad \Delta H = 1452 \text{ kJ}}$

   $2 N_2(g) + 6 H_2O(g) \rightarrow 3 O_2(g) + 4 NH_3(g)$ $\Delta H = 1268$ kJ

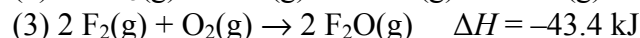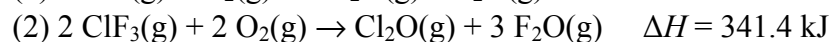**Statement:** The ΔH for the reaction is 1268 kJ.

(b) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Enthalpy, H (kJ)." Since ΔH is positive, the products will have greater enthalpy than the reactants. Point the arrow from reactants upward to products. Write the reactants on the lower left-hand side of the graph and the products on the upper right-hand side of the graph. Include the ΔH of the reaction. An enthalpy diagram of the reaction is:
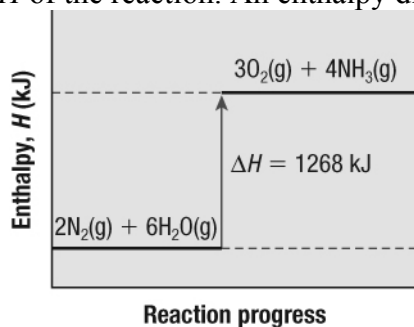
2. (a) Solution:
Step 1: Label the equations.
The desired equation is: $ClF(g) + F_2(g) \rightarrow ClF_3(g)$
(1) $2\ ClF(g) + O_2(g) \rightarrow Cl_2O(g) + F_2O(g)$ $\Delta H = 167.4$ kJ
(2) $2\ ClF_3(g) + 2\ O_2(g) \rightarrow Cl_2O(g) + 3\ F_2O(g)$ $\Delta H = 341.4$ kJ
(3) $2\ F_2(g) + O_2(g) \rightarrow 2\ F_2O(g)$ $\Delta H = -43.4$ kJ
Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.
Equation (1) and (3) contain 2 mol of $ClF(g)$ and 2 mol of $F_2(g)$ as reactants, so use them as written:
(1) $2\ ClF(g) + O_2(g) \rightarrow Cl_2O(g) + F_2O(g)$ $\Delta H = 167.4$ kJ
(3) $2\ F_2(g) + O_2(g) \rightarrow 2\ F_2O(g)$ $\Delta H = -43.4$ kJ
Step 3: Reverse (2) so that the product is on the same side as in the desired equation. Reverse the sign of ΔH.
(4) $Cl_2O(g) + 3\ F_2O(g) \rightarrow 2\ ClF_3(g) + 2\ O_2(g)$ $\Delta H = -341.4$ kJ
Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.
Equation (1) has 2 mol of $ClF(g)$ on the reactant side, while the desired equation has 1 mol.

Multiply equation (1) and its ΔH by the factor $\dfrac{1}{2}$:

(5) $ClF(g) + \dfrac{1}{2} O_2(g) \rightarrow \dfrac{1}{2} Cl_2O(g) + \dfrac{1}{2} F_2O(g)$ $\Delta H = 83.7$ kJ

Equation (3) has 2 mol of $F_2(g)$ on the reactant side, while the desired equation has 1 mol.

Multiply equation (3) and its ΔH by the factor $\dfrac{1}{2}$:

(6) $F_2(g) + \dfrac{1}{2} O_2(g) \rightarrow F_2O(g)$ $\Delta H = -21.7$ kJ

Equation (4) has 2 mol of $ClF_3(g)$ on the product side, while the desired equation has 1 mol.

Multiply equation (4) and its ΔH by the factor $\dfrac{1}{2}$:

(7) $\dfrac{1}{2} Cl_2O(g) + \dfrac{3}{2} F_2O(g) \rightarrow ClF_3(g) + O_2(g)$ $\Delta H = -170.7$ kJ

Step 5: Add equations (5), (6), and (7), and their changes in enthalpies.

(5) $ClF(g) + \dfrac{1}{2} \cancel{O_2(g)}$ $\rightarrow \dfrac{1}{2}\cancel{Cl_2O(g)} + \dfrac{1}{2}\cancel{F_2O(g)}$ $\Delta H =$ 83.7 kJ

(6) $F_2(g) + \dfrac{1}{2}\cancel{O_2(g)}$ $\rightarrow \cancel{F_2O(g)}$ $\Delta H =$ −21.7 kJ

(7) $\dfrac{1}{2}\cancel{Cl_2O(g)} + \dfrac{3}{2}\cancel{F_2O(g)} \rightarrow ClF_3(g) + \cancel{O_2(g)}$ $\Delta H = -170.7$ kJ

$ClF(g) + F_2(g)$ $\rightarrow ClF_3(g)$ $\Delta H = -108.7$ kJ
Statement: The ΔH for the reaction is −108.7 kJ.
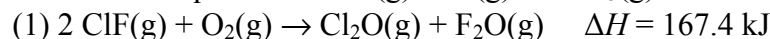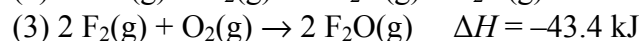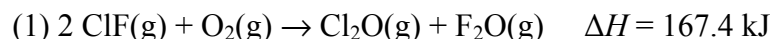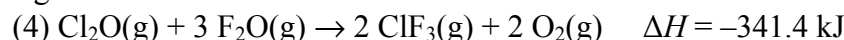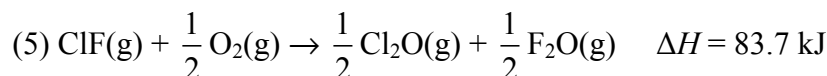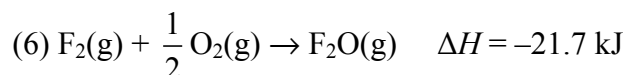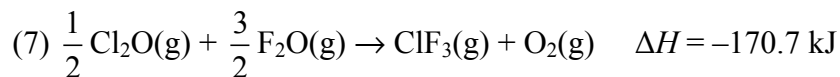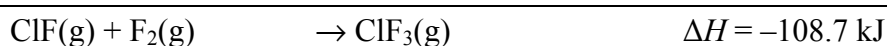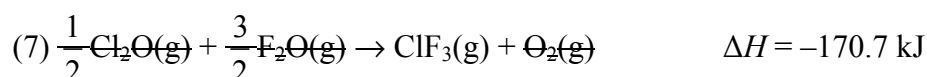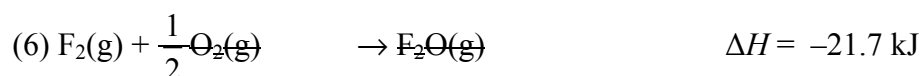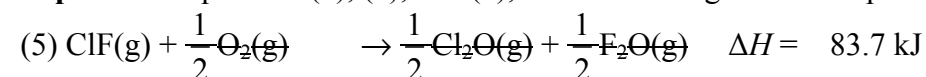
Chapter 5: Thermochemistry

(b) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Enthalpy, *H* (kJ)." Since Δ*H* is negative, the products will have less enthalpy than the reactants. Point the arrow from reactants downward to products. Write the reactants on the upper left-hand side of the graph and the products on the lower right-hand side of the graph. Include the Δ*H* of the reaction. An enthalpy diagram of the reaction is:
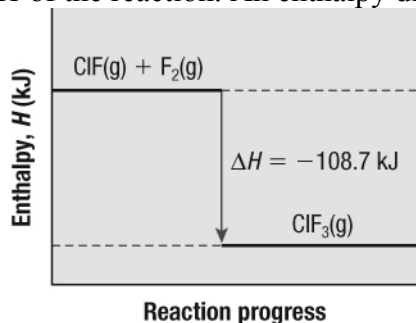
3. Solution:
Step 1: Label the equations.

The desired equation is: $Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$

(1) $C(graphite) + \dfrac{1}{2} O_2(g) \rightarrow CO(g) \quad \Delta H = -110.5$ kJ

(2) $C(graphite) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5$ kJ

(3) $2\ Fe(s) + \dfrac{3}{2} O_2(g) \rightarrow Fe_2O_3(s) \quad \Delta H = -824.2$ kJ

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.

Equation (2) contains $CO_2(g)$ as the product, so use it as written:

(2) $C(graphite) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5$ kJ

Step 3: Reverse (1) and (3) so that reactants and products are on the same side as in the desired equation. Reverse the sign of Δ*H*.

(4) $CO(g) \rightarrow C(graphite) + \dfrac{1}{2} O_2(g) \quad \Delta H = 110.5$ kJ

(5) $Fe_2O_3(s) \rightarrow 2\ Fe(s) + \dfrac{3}{2} O_2(g) \quad \Delta H = 824.2$ kJ

Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Equation (4) has 1 mol of $CO(g)$ on the reactant side, while the desired equation has 3 mol. Multiply equation (4) and its Δ*H* by the factor 3:
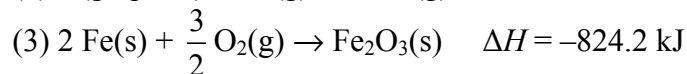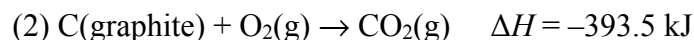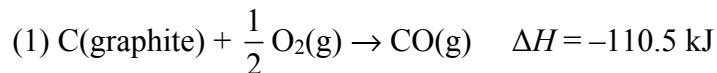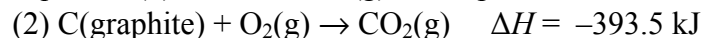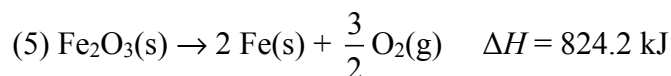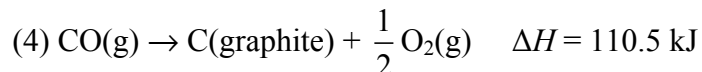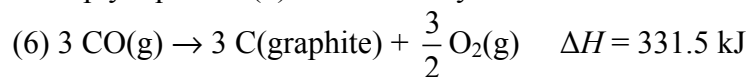
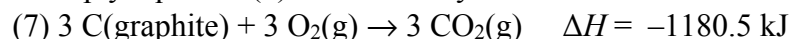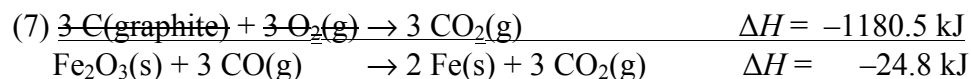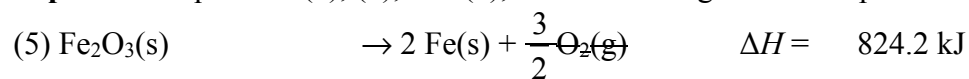(6) $3\ CO(g) \rightarrow 3\ C(graphite) + \dfrac{3}{2} O_2(g) \quad \Delta H = 331.5$ kJ

Equation (2) has 1 mol of $CO_2(g)$ on the product side, while the desired equation has 3 mol. Multiply equation (2) and its Δ*H* by the factor 3:

(7) $3\ C(graphite) + 3\ O_2(g) \rightarrow 3\ CO_2(g) \quad \Delta H = -1180.5$ kJ

Step 5: Add equations (5), (6), and (7), and their changes in enthalpies.

(5) $Fe_2O_3(s)$ $\rightarrow 2\ Fe(s) + \dfrac{3}{2}\ \cancel{O_2(g)}$ $\Delta H =$ 824.2 kJ

(6) $3\ CO(g)$ $\rightarrow \cancel{3\ C(graphite)} + \dfrac{3}{2}\ \cancel{O_2(g)}$ $\Delta H =$ 331.5 kJ

(7) $\cancel{3\ C(graphite)} + \cancel{3\ O_2(g)} \rightarrow 3\ CO_2(g)$ $\Delta H = -1180.5$ kJ

$Fe_2O_3(s) + 3\ CO(g)$ $\rightarrow 2\ Fe(s) + 3\ CO_2(g)$ $\Delta H =$ −24.8 kJ

Statement: The ΔH for the reaction is −24.8 kJ.

Section 5.4 Questions, page 318

1. (a) Answers may vary. It should be a rewording of Hess's law which states: The enthalpy change for the conversion of reactants to products is the same whether the conversion occurs in one step or several steps.
(b) In a correct equation for a reverse reaction, the magnitude of ΔH stays the same but the sign is reversed (from positive to negative or from negative to positive).
(c) When you multiply the coefficients of a balanced equation by a factor, multiply the value of ΔH by the same factor. You do not have to change the sign.
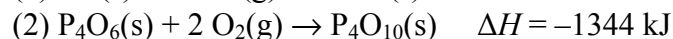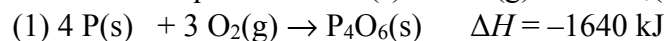**2.** Answers may vary. Sample answer: The law of conservation of energy states that energy cannot be created or destroyed, however, it can be converted from one form to another, and transferred between a system and its surroundings. Hess's law supports this theory as the overall energy change of a chemical process equals the sum of the energy changes of the individual steps. This indicates that the overall change in energy of a process is constant—energy cannot be created or destroyed but transferred from step to step.
3. The characteristic of enthalpy change that is the basis of Hess's Law is the flow of thermal energy. If the direction of the reaction is reversed, the direction of the thermal energy flow must also be reversed. That is, the sign of the ΔH value, which indicates the flow of thermal energy, is reversed.
4. (a) Solution:
Step 1: Label the equations.
The desired equation is: $4\ P(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$
(1) $4\ P(s)\ \ + 3\ O_2(g) \rightarrow P_4O_6(s)$ $\Delta H = -1640$ kJ
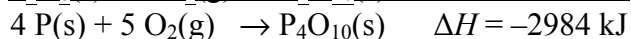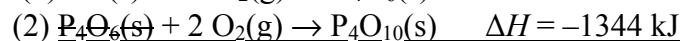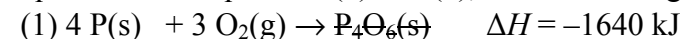(2) $P_4O_6(s) + 2\ O_2(g) \rightarrow P_4O_{10}(s)$ $\Delta H = -1344$ kJ
Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.
Equation (1) contains P(s) as the reactant, so use it as written.
Equation (2) contains $P_4O_{10}(s)$ as the product, so use it as written.
Step 3: The coefficients of the entities in the equations equal to the coefficients in the desired equation. Add equations (1) and (2), and their changes in enthalpies.
(1) $4\ P(s)\ \ + 3\ O_2(g) \rightarrow \cancel{P_4O_6(s)}$ $\Delta H = -1640$ kJ
(2) $\cancel{P_4O_6(s)} + 2\ O_2(g) \rightarrow P_4O_{10}(s)$ $\Delta H = -1344$ kJ

$4\ P(s) + 5\ O_2(g)\ \ \rightarrow P_4O_{10}(s)$ $\Delta H = -2984$ kJ

Statement: The enthalpy change for the combustion of phosphorus is −2984 kJ.

(b) The reaction for the combustion of phosphorus in part (a) is for 4 mol of phosphorus. To determine the molar enthalpy of combustion, divide the enthalpy change by 4 mol.

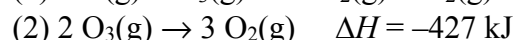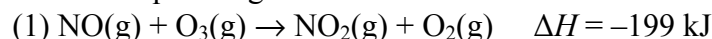$$\frac{-2984 \text{ kJ}}{4 \text{ mol}} = -746 \text{ kJ/mol}$$

So, the molar enthalpy of combustion for phosphorus is –746 kJ/mol.

5. Solution:

Step 1: Label the equations.

The desired equation is: $2 \text{ NO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ NO}_2\text{(g)}$

The first equation given is not needed.

(1) $\text{NO(g)} + \text{O}_3\text{(g)} \rightarrow \text{NO}_2\text{(g)} + \text{O}_2\text{(g)}$ $\Delta H = -199 \text{ kJ}$

(2) $2 \text{ O}_3\text{(g)} \rightarrow 3 \text{ O}_2\text{(g)}$ $\Delta H = -427 \text{ kJ}$

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.

Equation (1) contains NO(g) as the reactant, so use it as written.

Step 3: Reverse (2) so that reactants and products are on the same side as in the desired equation. Reverse the sign of ΔH.

(3) $3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ O}_3\text{(g)}$ $\Delta H = 427 \text{ kJ}$
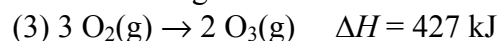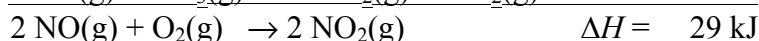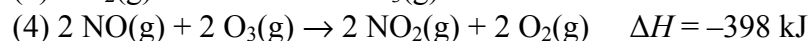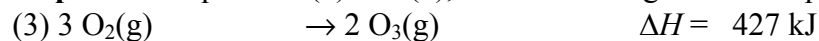
Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Equation (1) has 1 mol of NO(g) on the reactant side, while the desired equation has 2 mol. Multiply equation (1) and its ΔH by the factor 2:

(4) $2 \text{ NO(g)} + 2 \text{ O}_3\text{(g)} \rightarrow 2 \text{ NO}_2\text{(g)} + 2 \text{ O}_2\text{(g)}$ $\Delta H = -398 \text{ kJ}$

Step 5: Add equations (3) and (4), and their changes in enthalpies.

(3) $3 \text{ O}_2\text{(g)}$ $\rightarrow 2 \text{ O}_3\text{(g)}$ $\Delta H = $ 427 kJ

(4) $\underline{2 \text{ NO(g)} + 2 \text{ O}_3\text{(g)} \rightarrow 2 \text{ NO}_2\text{(g)} + 2 \text{ O}_2\text{(g)}}$ $\underline{\Delta H = -398 \text{ kJ}}$

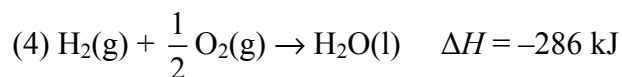 $2 \text{ NO(g)} + \text{O}_2\text{(g)}$ $\rightarrow 2 \text{ NO}_2\text{(g)}$ $\Delta H = $ 29 kJ
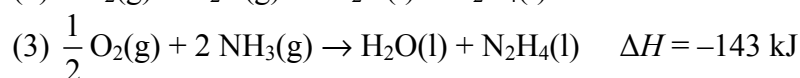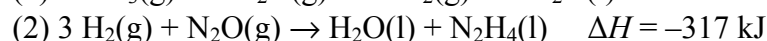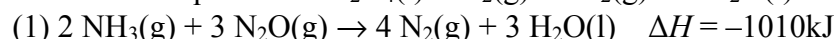
Statement: The enthalpy change of the desired reaction is 29 kJ.
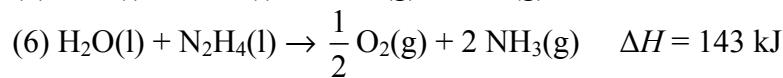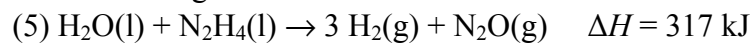
6. Solution:

Step 1: Label the equations.

The desired equation is: $\text{N}_2\text{H}_4\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{N}_2\text{(g)} + 2 \text{ H}_2\text{O(l)}$

(1) $2 \text{ NH}_3\text{(g)} + 3 \text{ N}_2\text{O(g)} \rightarrow 4 \text{ N}_2\text{(g)} + 3 \text{ H}_2\text{O(l)}$ $\Delta H = -1010 \text{kJ}$

(2) $3 \text{ H}_2\text{(g)} + \text{N}_2\text{O(g)} \rightarrow \text{H}_2\text{O(l)} + \text{N}_2\text{H}_4\text{(l)}$ $\Delta H = -317 \text{ kJ}$

(3) $\frac{1}{2} \text{O}_2\text{(g)} + 2 \text{ NH}_3\text{(g)} \rightarrow \text{H}_2\text{O(l)} + \text{N}_2\text{H}_4\text{(l)}$ $\Delta H = -143 \text{ kJ}$

(4) $\text{H}_2\text{(g)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O(l)}$ $\Delta H = -286 \text{ kJ}$

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.
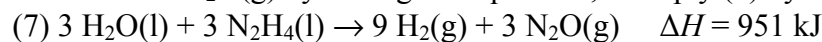
Equation (1) has $\text{N}_2\text{(g)}$ and $\text{H}_2\text{O(l)}$ as the product, so use it as written.

Chapter 5: Thermochemistry

Step 3: Reverse (2) and (3) so that $N_2H_4(l)$ is on the same side as in the desired equation. Reverse the sign of ΔH.

(5) $H_2O(l) + N_2H_4(l) \rightarrow 3\ H_2(g) + N_2O(g) \quad \Delta H = 317$ kJ

(6) $H_2O(l) + N_2H_4(l) \rightarrow \dfrac{1}{2}\ O_2(g) + 2\ NH_3(g) \quad \Delta H = 143$ kJ

Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.
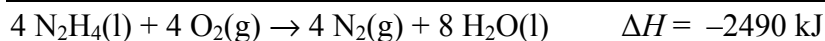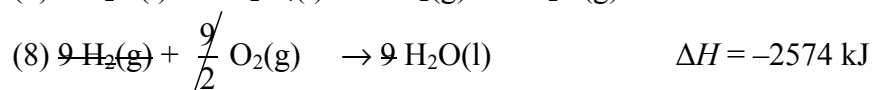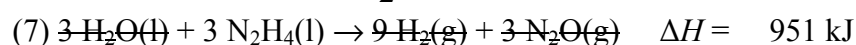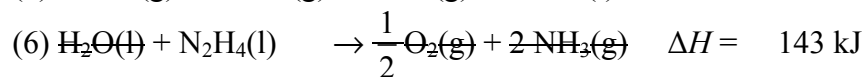To eliminate $N_2O(g)$ by adding the equations, multiply (5) by a factor of 3:
(7) $3\ H_2O(l) + 3\ N_2H_4(l) \rightarrow 9\ H_2(g) + 3\ N_2O(g) \quad \Delta H = 951$ kJ
To eliminate $H_2(g)$ by adding the equations, multiply (4) by a factor of 9:

(8) $9\ H_2(g) + \dfrac{9}{2}\ O_2(g) \rightarrow 9\ H_2O(l) \quad \Delta H = -2574$ kJ
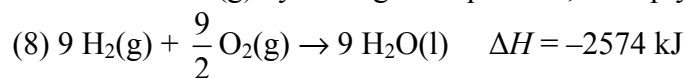
Step 5: Add equations (1), (5), (7), and (8), and their changes in enthalpies.

(1) $\cancel{2\ NH_3(g)} + \cancel{3\ N_2O(g)} \rightarrow 4\ N_2(g) + \cancel{3\ H_2O(l)} \quad \Delta H = -1010$ kJ

(6) $\cancel{H_2O(l)} + N_2H_4(l) \quad \rightarrow \dfrac{1}{2}\cancel{O_2(g)} + \cancel{2\ NH_3(g)} \quad \Delta H = \quad 143$ kJ

(7) $\cancel{3\ H_2O(l)} + 3\ N_2H_4(l) \rightarrow \cancel{9\ H_2(g)} + \cancel{3\ N_2O(g)} \quad \Delta H = \quad 951$ kJ

(8) $\cancel{9\ H_2(g)} + \cancel{\dfrac{9}{2}}\ O_2(g) \quad \rightarrow \cancel{9}\ H_2O(l) \qquad\qquad \Delta H = -2574$ kJ

$\overline{\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad}$

$\quad 4\ N_2H_4(l) + 4\ O_2(g) \rightarrow 4\ N_2(g) + 8\ H_2O(l) \qquad \Delta H = \ -2490$ kJ
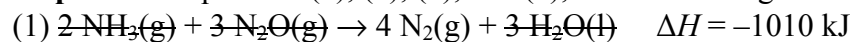
This enthalpy change is for 4 mol of $N_2H_4(l)$. For 1 mol of $N_2H_4(l)$ in the desired equation, divide -2490 kJ by 4.
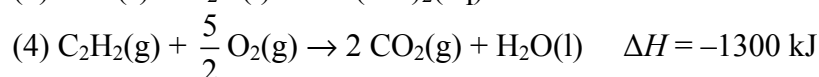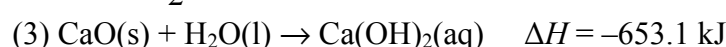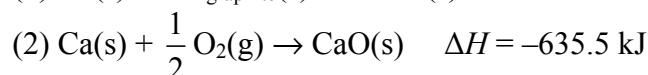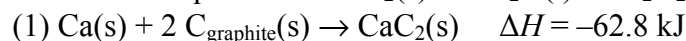
$\dfrac{-2490 \text{ kJ}}{4 \text{ mol}} = -623 \text{ kJ/mol}$

Statement: For the combustion of liquid hydrazine in oxygen gas to form nitrogen gas and liquid water, the enthalpy change is -623 kJ/mol.
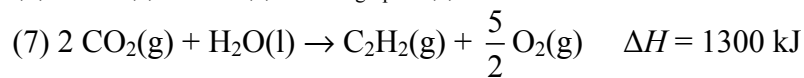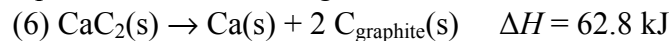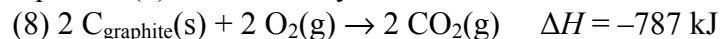
7. Solution:

Step 1: Label the equations.
The desired equation is: $CaC_2(s) + 2\ H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$
(1) $Ca(s) + 2\ C_{graphite}(s) \rightarrow CaC_2(s) \quad \Delta H = -62.8$ kJ
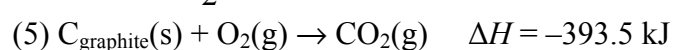
(2) $Ca(s) + \dfrac{1}{2}\ O_2(g) \rightarrow CaO(s) \quad \Delta H = -635.5$ kJ

(3) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) \quad \Delta H = -653.1$ kJ

(4) $C_2H_2(g) + \dfrac{5}{2}\ O_2(g) \rightarrow 2\ CO_2(g) + H_2O(l) \quad \Delta H = -1300$ kJ

(5) $C_{graphite}(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5$ kJ

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.
Equation (3) contains $Ca(OH)_2(aq)$ as the product, so use it as written.

Step 3: Reverse (1) and (4) so that $CaC_2(s)$ and $C_2H_2(g)$ are on the same side as in the desired equation. Reverse the sign of ΔH.

(6) $CaC_2(s) \rightarrow Ca(s) + 2\ C_{graphite}(s)$ $\Delta H = 62.8$ kJ

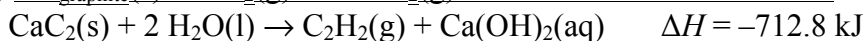(7) $2\ CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + \dfrac{5}{2} O_2(g)$ $\Delta H = 1300$ kJ
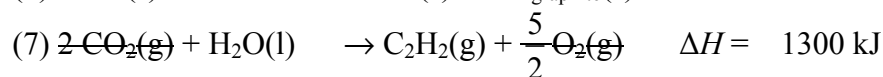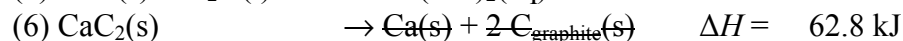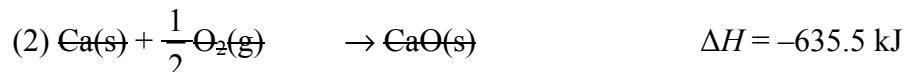
Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.

Equation (5) has 1 mol of $C_{graphite}(s)$ on the reactant side, while Equation (1) has 2 mol. Multiply equation (5) and its ΔH by the factor 2:

(8) $2\ C_{graphite}(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g)$ $\Delta H = -787$ kJ

Step 5: Add equations (2), (3), (6), (7), and (8), and their changes in enthalpies.

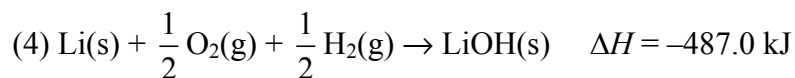(2) $\cancel{Ca(s)} + \cancel{\dfrac{1}{2} O_2(g)}$ $\rightarrow \cancel{CaO(s)}$ $\Delta H = -635.5$ kJ

(3) $\cancel{CaO(s)} + H_2O(l)$ $\rightarrow Ca(OH)_2(aq)$ $\Delta H = -653.1$ kJ

(6) $CaC_2(s)$ $\rightarrow \cancel{Ca(s)} + \cancel{2\ C_{graphite}(s)}$ $\Delta H = \quad 62.8$ kJ

(7) $\cancel{2\ CO_2(g)} + H_2O(l)$ $\rightarrow C_2H_2(g) + \cancel{\dfrac{5}{2} O_2(g)}$ $\Delta H = \quad 1300$ kJ

(8) $\cancel{2\ C_{graphite}(s)} + \cancel{2\ O_2(g)} \rightarrow \cancel{2\ CO_2(g)}$ $\Delta H = \quad -787$ kJ

$\quad CaC_2(s) + 2\ H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$ $\Delta H = -712.8$ kJ
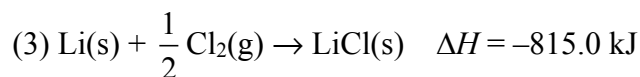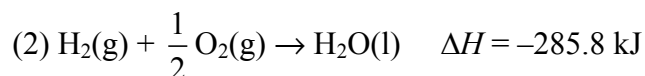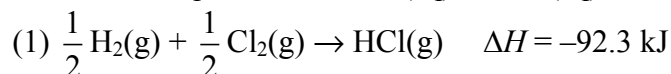
Statement: The enthalpy change for the reaction of solid calcium carbide with liquid water to produce ethyne is -712.8 kJ.

8. Solution:

Step 1: Label the equations.

The desired equation is: $LiOH(aq) + HCl\ (aq) \rightarrow LiCl(aq) + H_2O(l)$
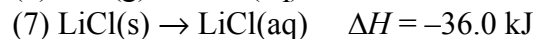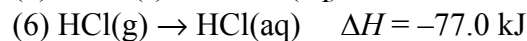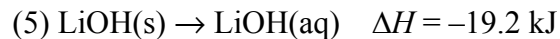
(1) $\dfrac{1}{2} H_2(g) + \dfrac{1}{2} Cl_2(g) \rightarrow HCl(g)$ $\Delta H = -92.3$ kJ

(2) $H_2(g) + \dfrac{1}{2} O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285.8$ kJ

(3) $Li(s) + \dfrac{1}{2} Cl_2(g) \rightarrow LiCl(s)$ $\Delta H = -815.0$ kJ

(4) $Li(s) + \dfrac{1}{2} O_2(g) + \dfrac{1}{2} H_2(g) \rightarrow LiOH(s)$ $\Delta H = -487.0$ kJ

(5) $LiOH(s) \rightarrow LiOH(aq)$ $\Delta H = -19.2$ kJ
(6) $HCl(g) \rightarrow HCl(aq)$ $\Delta H = -77.0$ kJ
(7) $LiCl(s) \rightarrow LiCl(aq)$ $\Delta H = -36.0$ kJ

Step 2: Identify the equations that have reactants or products on the same side of the arrow as in the desired equation.

Equation (2) contains $H_2O(l)$ as the product, so use it as written.

Equation (7) contains $LiCl(aq)$ as the product, so use it as written.

Step 3: Reverse (5) and (6) so that LiOH(aq) and HCl (aq) are on the same side as in the desired equation. Reverse the sign of ΔH.
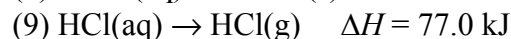
(8) $LiOH(aq) \rightarrow LiOH(s) \quad \Delta H = 19.2 \text{ kJ}$

(9) $HCl(aq) \rightarrow HCl(g) \quad \Delta H = 77.0 \text{ kJ}$

Step 4: Check the coefficients of all entities in the equations to see if any must be adjusted by a factor so as to be equal to the coefficients in the desired equation.
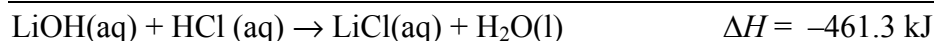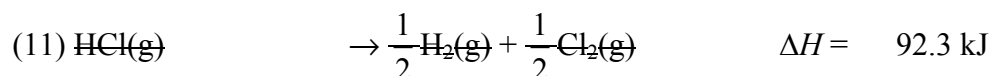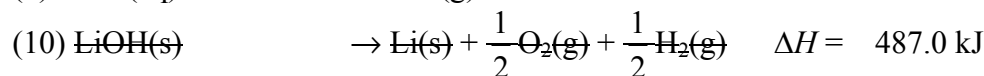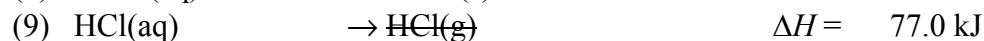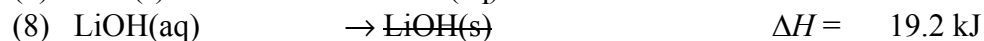
To eliminate LiOH(s) by adding the equations, reverse (4) and the sign of ΔH.

(10) $LiOH(s) \rightarrow Li(s) + \dfrac{1}{2} O_2(g) + \dfrac{1}{2} H_2(g) \quad \Delta H = 487.0 \text{ kJ}$
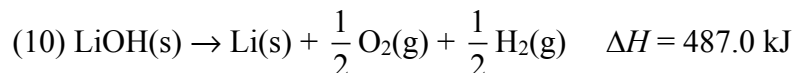
To eliminate HCl(g) by adding the equations, reverse (1) and the sign of ΔH.

(11) $HCl(g) \rightarrow \dfrac{1}{2} H_2(g) + \dfrac{1}{2} Cl_2(g) \quad \Delta H = 92.3 \text{ kJ}$
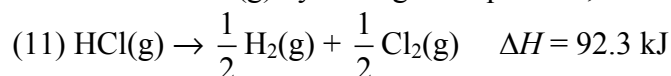
Step 5: Add equations (2), (3), (7), (8), (9), (10), and (11), and their changes in enthalpies.

(2) $\cancel{H_2(g) + \dfrac{1}{2} O_2(g)} \rightarrow H_2O(l) \qquad\qquad \Delta H = -285.8 \text{ kJ}$

(3) $\cancel{Li(s) + \dfrac{1}{2} Cl_2(g)} \rightarrow \cancel{LiCl(s)} \qquad\qquad \Delta H = -815.0 \text{ kJ}$

(7) $\cancel{LiCl(s)} \rightarrow LiCl(aq) \qquad\qquad \Delta H = \ -36.0 \text{ kJ}$

(8) $LiOH(aq) \rightarrow \cancel{LiOH(s)} \qquad\qquad \Delta H = \ \ 19.2 \text{ kJ}$

(9) $HCl(aq) \rightarrow \cancel{HCl(g)} \qquad\qquad \Delta H = \ \ 77.0 \text{ kJ}$

(10) $\cancel{LiOH(s)} \rightarrow \cancel{Li(s) + \dfrac{1}{2} O_2(g) + \dfrac{1}{2} H_2(g)} \quad \Delta H = \ 487.0 \text{ kJ}$

(11) $\cancel{HCl(g)} \rightarrow \cancel{\dfrac{1}{2} H_2(g) + \dfrac{1}{2} Cl_2(g)} \qquad\qquad \Delta H = \ \ 92.3 \text{ kJ}$

$LiOH(aq) + HCl (aq) \rightarrow LiCl(aq) + H_2O(l) \qquad \Delta H = -461.3 \text{ kJ}$
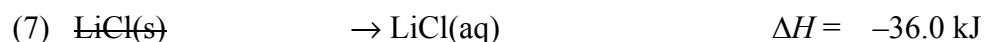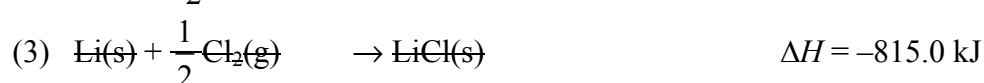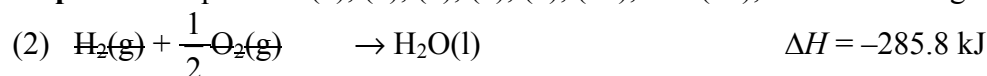
Statement: The enthalpy change of neutralization for 1 mol of aqueous lithium hydroxide is -461.3 kJ.