Section 5.2: Calorimetry and Enthalpy
Tutorial 1 Practice, page 297
1. Given: $V_{H_2O(l)} = 6.0 \text{ mL}$; $T_{initial} = 25 \text{ °C}$; $T_{final} = 75 \text{ °C}$; $c_{H_2O(l)} = 4.18 \text{ J/(g·°C)}$; $d_{H_2O(l)} = 1.00 \text{ g/mL}$

Required: thermal energy required, q

Analysis: $q = mc\Delta T$

Solution:

Step 1: Determine the mass of water, m.

$m = Vd$

$= 6.0 \text{ mL} \times \dfrac{1.00 \text{ g}}{1 \text{ mL}}$

$m = 6.0 \text{ g}$

Step 2: Determine the change in temperature, ΔT.

$\Delta T = T_{final} - T_{initial}$

$= 75 \text{ °C} - 25 \text{ °C}$

$\Delta T = 50 \text{ °C}$

Step 3: Calculate the quantity of thermal energy, q.

$q = mc\Delta T$

$= (6.0 \text{ g})\left(\dfrac{4.18 \text{ J}}{\text{g} \cdot \text{°C}}\right)(50 \text{ °C})$

$q = 1.3 \text{ kJ}$

Statement: The quantity of thermal energy required is 1.3 kJ.

2. Given: $m_{ethylene\ glycol} = 4.0 \text{ kg}$; $c_{ethylene\ glycol} = 3.5 \text{ J/(g·°C)}$; $q = 250 \text{ kJ}$

Required: change in temperature, ΔT

Analysis: $q = mc\Delta T$

Solution:

Step 1: Convert quantities to the appropriate units.

$m = 4.0 \text{ kg} = 4.0 \times 10^3 \text{ g}$

$q = 250 \text{ kJ} = 250 \times 10^3 \text{ J}$

Step 2: Rearrange $q = mc\Delta T$ to solve for ΔT.

$q = mc\Delta T$

$\Delta T = \dfrac{q}{mc}$

$= \dfrac{250 \times 10^3 \text{ J}}{(4.0 \times 10^3 \text{ g})\left(\dfrac{3.5 \text{ J}}{\text{g} \cdot \text{°C}}\right)}$

$\Delta T = 18 \text{ °C}$

Statement: The temperature change of the solution was 18 °C.

3. Given: $V_{HCl(aq)} = 50.0 \text{ mL}$; $V_{NaOH(aq)} = 75.0 \text{ mL}$; $T_{initial} = 20.2 \text{ °C}$; $T_{final} = 25.6 \text{ °C}$

Required: quantity of energy transferred, q

Analysis: $q = mc\Delta T$

Solution:

Step 1: Determine the total volume of HCl(aq) and NaOH(aq), V.

$V = V_{HCl(aq)} + V_{NaOH(aq)}$

$\quad = 50.0 \text{ mL} + 75.0 \text{ mL}$

$V = 125.0 \text{ mL}$

Step 2: Determine the total mass of two solutions , m.

Since the solution contains dilute HCl(aq) and NaOH(aq), the density and heat capacity are assumed to be the same as those for water.

$m = Vd$

$\quad = 125.0 \text{ mL} \times \dfrac{1.00 \text{ g}}{1 \text{ mL}}$

$m = 125.0 \text{ g}$

Step 3: Determine the change in temperature, ΔT.

$\Delta T = T_{final} - T_{initial}$

$\quad = 25.6 \text{ °C} - 20.2 \text{ °C}$

$\Delta T = 5.4 \text{ °C}$

Step 4: Calculate the quantity of energy transferred, q.

$q = mc\Delta T$

$\quad = (125.0 \text{ g}) \left(\dfrac{4.18 \text{ J}}{\text{g} \cdot \text{°C}} \right) (5.4 \text{ °C})$

$q = 2800 \text{ J}$

Statement: The quantity of energy transferred is 2800 J, or 2.8 kJ. Since the temperature of the surroundings (liquid water in the calorimeter) increased, the reaction was exothermic.

Mini Investigation: Thermal Energy Transfer in a Coffee-Cup Calorimeter, page 297

Answers may vary. Sample answers:

mass of water in calorimeter = 100.00 g

mass of aluminum block = 89.80 g

temperature of aluminum block = temperature of hot water bath = 80.0 °C

final temperature of water in calorimeter = 27.8 °C

initial temperature of tap water = 19.8 °C

A. Given: $m_{H_2O(l)} = 100.00 \text{ g}$; $T_{initial} = 19.8 \text{ °C}$; $T_{final} = 27.8 \text{ °C}$

Required: quantity of thermal energy gained by the water, q

Analysis: $q = mc\Delta T$

Solution:
Step 1: Determine the change in temperature, ΔT.
$$\Delta T = T_{final} - T_{initial}$$
$$= 27.8\ °C - 19.8\ °C$$

$$\Delta T = 8.0\ °C$$
Step 2: Calculate the quantity of thermal energy gained by the water, q.
$$q = mc\Delta T$$

$$= (100.00\ \cancel{g})\left(\frac{4.18\ J}{\cancel{g} \cdot \cancel{°C}}\right)(8.0\ \cancel{°C})$$

$$q = 3300\ J$$
Statement: The quantity of thermal energy gained by the water is 3300 J, or 3.3 kJ.
B. Given: $m_{Al(s)} = 89.80\ g$; $T_{initial} = 80.0\ °C$; $T_{final} = 27.8\ °C$; $c_{Al(s)} = 0.900\ J/(g \cdot °C)$

Required: quantity of thermal energy transferred to or from aluminum, q
Analysis: $q = mc\Delta T$
Solution:
Step 1: Determine the change in temperature, ΔT.
$$\Delta T = T_{final} - T_{initial}$$
$$= 27.8\ °C - 80.0\ °C$$

$$\Delta T = -52.2\ °C$$
Step 2: Calculate the quantity of thermal energy transferred to or from aluminum, q.
$$q = mc\Delta T$$

$$= (89.80\ \cancel{g})\left(\frac{0.900\ J}{\cancel{g} \cdot \cancel{°C}}\right)(-52.2\ \cancel{°C})$$

$$q = -4220\ J$$
Statement: Since the value of q is negative, energy is transferred from aluminum to water, and the quantity transferred is 4220 J, or 4.22 kJ.
C. Answers may vary. Sample answer: There is a difference of about 0.9 kJ in the answers to A and B. The difference could be due to assumptions or experimental errors.
D. Answers may vary. Sample answer: Experimental errors include thermal energy lost to the tongs, to the air on transfer from hot to cold water, to the thermometer, to the calorimeter, or to the air above the water–aluminum mixture. The investigation could be improved by taking into account the temperature changes and thus energy transferred to the thermometer, tongs, and calorimeter. Other improvements would be increasing the insulation of calorimeter and faster transfer from hot to cold water.

Tutorial 2 Practice, page 301

1. Given: $m_{H_2O(l)} = 50.0 \text{ g}$; $\Delta H_{vap} = 44.0 \text{ kJ/mol}$

Required: enthalpy change, ΔH

Analysis: $\Delta H = n\Delta H_{vap}$

Solution:

Step 1: Calculate the amount of water in 50.0 g, $n_{H_2O(l)}$.

$$M_{H_2O(l)} = 18.02 \text{ g/mol}$$

$$n_{H_2O(l)} = \frac{m_{H_2O(l)}}{M_{H_2O(l)}}$$

$$= \frac{50.0 \text{ g}}{18.02 \text{ g/mol}}$$

$$n_{H_2O(l)} = 2.7747 \text{ mol (2 extra digits carried)}$$

Step 2: Solve for the change in enthalpy, ΔH.

$$\Delta H = n_{H_2O(l)}\Delta H_{vap}$$

$$= (2.7747 \text{ mol})\left(\frac{44.0 \text{ kJ}}{1 \text{ mol}}\right)$$

$$\Delta H = 122 \text{ kJ}$$

Statement: The enthalpy change expected when 50.0 g of water vaporizes is 122 kJ.

2. Given: $m_{Na_2SO_4 \cdot 10H_2O(l)} = 2.50 \text{ kg}$; $\Delta H_r = -78.0 \text{ kJ/mol}$

Required: enthalpy change, ΔH

Analysis: $\Delta H = n_{Na_2SO_4 \cdot 10H_2O(l)}\Delta H_r$

Solution:

Step 1: Calculate the amount of salt in 2.50 kg, $n_{Na_2SO_4 \cdot 10H_2O(l)}$.

$$M_{Na_2SO_4 \cdot 10H_2O(l)} = 322.24 \text{ g/mol}$$

$$n_{Na_2SO_4 \cdot 10H_2O(l)} = \frac{m_{Na_2SO_4 \cdot 10H_2O(l)}}{M_{Na_2SO_4 \cdot 10H_2O(l)}}$$

$$= \frac{2.50 \times 10^3 \text{ g}}{322.24 \text{ g/mol}}$$

$$n_{Na_2SO_4 \cdot 10H_2O(l)} = 7.7582 \text{ mol (2 extra digits carried)}$$

Step 2: Solve for the change in enthalpy, ΔH.

$$\Delta H = n_{Na_2SO_4 \cdot 10H_2O(l)}\Delta H_r$$

$$= (7.7582 \text{ mol})\left(\frac{-78.0 \text{ kJ}}{1 \text{ mol}}\right)$$

$$\Delta H = -605 \text{ kJ}$$

Statement: The change in enthalpy when 2.50 kg of Glauber's salt enters the solid state is −605 kJ.

3. Given: $\Delta H = 100.0$ kJ ; $\Delta H_{vap} = 15.7$ kJ/mol

Required: mass of propane, $m_{C_3H_8(l)}$

Analysis: $\Delta H = n\Delta H_{vap}$

Solution:

Step 1: Rearrange $\Delta H = n\Delta H_{vap}$ to solve for the amount of propane, $n_{C_3H_8(l)}$.

$$\Delta H = n_{C_3H_8(l)}\Delta H_{vap}$$

$$n_{C_3H_8(l)} = \frac{\Delta H}{\Delta H_{vap}}$$

$$= \frac{100.0 \text{ k\!\!\!/J}}{15.7 \text{ k\!\!\!/J}/\text{mol}}$$

$$n_{C_3H_8(l)} = 6.3694 \text{ mol (2 extra digits carried)}$$

Step 2: Convert amount of propane into mass, $m_{C_3H_8(l)}$.

$$M_{C_3H_8(l)} = 44.11 \text{ g/mol}$$

$$n_{C_3H_8(l)} = \frac{m_{C_3H_8(l)}}{M_{C_3H_8(l)}}$$

$$m_{C_3H_8(l)} = n_{C_3H_8(l)}M_{C_3H_8(l)}$$

$$= (6.3694 \text{ m\!\!\!/ol})\left(\frac{44.11 \text{ g}}{1 \text{ m\!\!\!/ol}}\right)$$

$$m_{C_3H_8(l)} = 281 \text{ g}$$

Statement: The mass of propane that would vaporize as a result of absorbing 100.0 kJ of energy is 281 g.

4. Given: $m_{NaOH(s)} = 4.00$ g; $V_{H_2O(l)} = 100.0$ mL; $\Delta T = 10.6$ °C

Required: molar enthalpy of dissolution, ΔH_{sol}

Analysis: $q = mc\Delta T$; $\Delta H = n\Delta H_{sol}$

Solution:

Step 1: Determine the mass of water, m.

$$m = Vd$$

$$= 100.0 \text{ m\!\!\!/L} \times \frac{1.00 \text{ g}}{1 \text{ m\!\!\!/L}}$$

$$m = 100 \text{ g}$$

Step 2: Determine the quantity of thermal energy absorbed by the water, q.

$$q = mc\Delta T$$

$$= (100 \text{ g\!\!\!/})\left(\frac{4.18 \text{ J}}{\text{g\!\!\!/} \cdot °\!\!\!/C}\right)(10.6 \text{ °\!\!\!/C})$$

$$q = 4430.8 \text{ J (2 extra digits carried)}$$

This is the quantity of energy released, ΔH, when 4.00 g of NaOH(s) is dissolved.

Step 3: Convert mass of NaOH(s) into amount, $n_{NaOH(s)}$.

$M_{NaOH(s)} = 40.00$ g/mol

$$n_{NaOH(s)} = \frac{m_{NaOH(s)}}{M_{NaOH(s)}}$$

$$= \frac{4.00 \; \cancel{g}}{40.00 \; \cancel{g}/mol}$$

$n_{NaOH(s)} = 0.100$ mol

Step 4: Rearrange $\Delta H = n\Delta H_{sol}$ to solve for the molar enthalpy of dissolution, ΔH_{sol}.
Since energy is released, ΔH for the reaction is negative; $\Delta H = -4430.8$ J.

$$\Delta H = n_{NaOH(s)}\Delta H_{sol}$$

$$\Delta H_{sol} = \frac{\Delta H}{n_{NaOH(s)}}$$
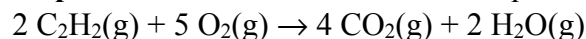
$$= \frac{-4430.8 \text{ J}}{0.100 \text{ mol}}$$

$\Delta H_{sol} = -44\;300$ J/mol

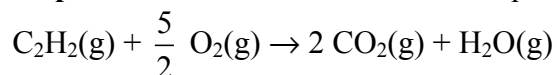**Statement:** The molar enthalpy of dissolution of sodium hydroxide is -44.3×10^3 J/mol, or -44.3 kJ/mol.

Tutorial 3 Practice, page 304

1. (a) Solution:
Step 1: Write the balanced chemical equation without the energy term.
2 $C_2H_2(g)$ + 5 $O_2(g) \rightarrow$ 4 $CO_2(g)$ + 2 $H_2O(g)$
Step 2: Write the balanced chemical equation for the combustion of 1 mol $C_2H_2(g)$.

$C_2H_2(g) + \dfrac{5}{2}$ $O_2(g) \rightarrow$ 2 $CO_2(g)$ + $H_2O(g)$

Step 3: Since the enthalpy change has a negative value, the combustion reaction of ethyne is exothermic. Since an exothermic reaction releases energy, place the energy term on the right side of the equation.
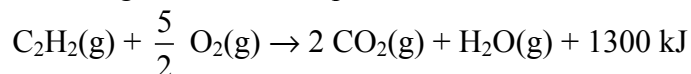
$C_2H_2(g) + \dfrac{5}{2}$ $O_2(g) \rightarrow$ 2 $CO_2(g)$ + $H_2O(g)$ + 1300 kJ

(b) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Potential energy, *E*p (kJ)." Since the reaction is exothermic, the products will have lower potential energy 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 ΔH of the reaction. This process is combustion, so ΔH is ΔH_c.

Potential Energy Changes during the Combustion Reaction
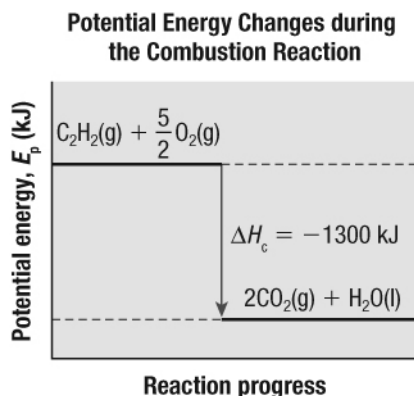
2. (a) Since the temperature of the surroundings decreases, the reaction is endothermic.
(b) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Potential energy, *E*p (kJ)." Since the reaction is endothermic, the products will have greater potential energy 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. This process is dissolution, so ΔH is ΔH_{sol}.

Potential Energy Changes during the Dissolution Reaction
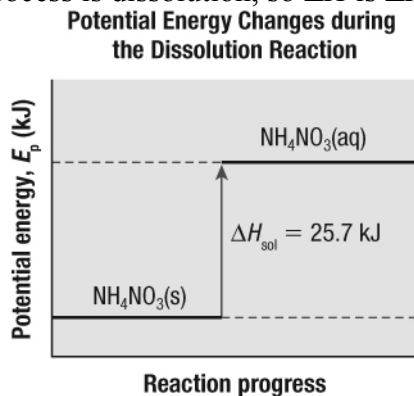
(c) A thermochemical equation for the reaction is:
$NH_4NO_3(s) + 25.7 \text{ kJ} \xrightarrow{\text{H}_2\text{O}} NH_4NO_3(aq)$
3. (a) The corresponding thermochemical equation for the reaction depicted by the potential energy diagram is: $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$ $\Delta H_{vap} = 38.6 \text{ kJ}$
(b) The corresponding thermochemical equation for the reaction depicted by the potential energy diagram is: $C_8H_{18}(l) + \dfrac{25}{2}O_2(g) \rightarrow 8\ CO_2(g) + 9\ H_2O(l)$ $\Delta H_c = -5074 \text{ kJ}$

4. Given: $[Ag^+(aq)] = [Cl^-(aq)] = 0.100$ mol/L ; $m_{solution} = 100.0$ g; $T_{initial} = 22.6$ °C;
$T_{final} = 23.4$ °C; $c = 4.18$ J/(g·°C)
Required: molar enthalpy of reaction, ΔH_r
Analysis: $q = mc\Delta T$; $\Delta H = n\Delta H_r$
Solution:
Step 1: Determine the change in temperature of the solution, ΔT.
$\Delta T = T_{final} - T_{initial}$
$\quad = 23.4$ °C $- 22.6$ °C
$\Delta T = 0.8$ °C
Step 2: Calculate the quantity of thermal energy for the reaction, q.
$q = mc\Delta T$
$\quad = (100.0 \ \cancel{g}) \left(\dfrac{4.18 \text{ J}}{\cancel{g} \cdot \cancel{°C}} \right) (0.8 \ \cancel{°C})$
$q = 334$ J (2 extra digits carried)
This is the quantity of energy released, ΔH. Since there is an increase in temperature, the reaction is exothermic, and ΔH is negative; $\Delta H = -334$ J.
Step 3: Determine the amount of AgCl(s) formed in the 100 mL solution, $n_{AgCl(s)}$.

$n_{AgCl(s)} = n_{Ag^+(aq)}$

$\quad = 50.0 \ \cancel{mL} \times \dfrac{0.100 \text{ mol}}{1000 \ \cancel{mL}}$

$n_{AgCl(s)} = 0.005$ mol
Step 4: Rearrange $\Delta H = n\Delta H_r$ to solve for the enthalpy of formation for 1 mol AgCl(s),
ΔH_r.
$\Delta H = n_{AgCl(s)}\Delta H_r$

$\Delta H_r = \dfrac{\Delta H}{n_{AgCl(s)}}$

$\quad = \dfrac{-334 \text{ J}}{0.005 \text{ mol}}$

$\Delta H_r = -70 \ 000$ J/mol
Statement: The thermal energy that accompanies the formation of AgCl(s) is
$-70 \ 000$ J/mol or -70 kJ/mol of AgCl(s) formed.

Section 5.2 Questions, page 306

1. Specific heat capacity is the quantity of thermal energy required to raise the temperature of 1 g of a substance by 1 °C. The specific heat capacity of water, 4.18 J/(g·°C), is higher than that of aluminum, 0.900 J/(g·°C). When the same mass of aluminum and water were heated to the same temperature, the water would have more thermal energy and take longer to dissipate the excess energy, whereas aluminum would have less thermal energy and take less time to cool down. Therefore, the aluminum foil can be safely removed using your fingers, but touching the inside of the potato results in a serious burn.

2. The burning of gasoline into carbon dioxide and water is a combustion reaction, which is an exothermic process. The system releases energy and the surroundings absorb energy. Therefore, the sign of q_{system} is negative (energy released), the sign for ΔH is also negative (exothermic process), and the sign of $q_{surroundings}$ is positive (energy absorbed).

3. (a) From the balanced equation, the enthalpy change of –891 kJ is for the formation of 2 moles of water. Calculate the enthalpy change, ΔH, for 1 mole:

$$\Delta H = \frac{-891 \text{ kJ}}{2}$$

$\Delta H = -446$ kJ

So, the enthalpy change for each mole of water formed is –446 kJ.

(b) From the balanced equation, the enthalpy change of –891 kJ is for the formation of 1 mole of carbon dioxide. So, the enthalpy change for each mole of carbon dioxide formed is also –891 kJ.

(c) From the balanced equation, the enthalpy change of –891 kJ is for 2 moles of oxygen reacted. Calculate the enthalpy change, ΔH, for 1 mole:

$$\Delta H = \frac{-891 \text{ kJ}}{2}$$

$\Delta H = -446$ kJ

So, the enthalpy change for each mole of oxygen reacted is –446 kJ.

4. Given: $m_{NH_4Cl(s)} = 20.0$ g; $\Delta H_{sol} = +14.8$ kJ/mol; $V_{H_2O(l)} = 125$ mL; $\Delta T_{initial} = 20.0$ °C

Required: final temperature, ΔT_{final}

Analysis: $\Delta H = n\Delta H_{sol}$; $q = mc\Delta T$

Solution:

Step 1: Convert mass of $NH_4Cl(s)$ into amount, $n_{NH_4Cl(s)}$.

$M_{NH_4Cl(s)} = 53.50$ g/mol

$$n_{NH_4Cl(s)} = \frac{m_{NH_4Cl(s)}}{M_{NH_4Cl(s)}}$$

$$= \frac{20.0 \text{ g}}{53.50 \text{ g/mol}}$$

$n_{NH_4Cl(s)} = 0.373\ 83$ mol (2 extra digits carried)

Step 2: Calculate the enthalpy change for the reaction, ΔH.

$$\Delta H = n_{\text{NH}_4\text{Cl(s)}} \Delta H_{\text{sol}}$$

$$= (0.373\ 83\ \cancel{\text{mol}}) \left(\frac{14.8\ \text{kJ}}{1\ \cancel{\text{mol}}} \right)$$

$\Delta H = 5.5327\ \text{kJ}$ (2 extra digits carried)

Since ΔH is positive, thermal energy is absorbed from the surroundings. So, the change in thermal energy of the surrounding water, q, is negative; $q = -5.5327\ \text{kJ}$.

Step 3: Determine the mass of water, m.

$$m = Vd$$

$$= 125\ \cancel{\text{mL}} \times \frac{1.00\ \text{g}}{1\ \cancel{\text{mL}}}$$

$m = 125\ \text{g}$

Step 4: Rearrange $q = mc\Delta T$ to solve for the temperature change, ΔT.

$$q = mc\Delta T$$

$$\Delta T = \frac{q}{mc}$$

$$= \frac{-5.5327 \times 10^3\ \text{J}}{(125\ \cancel{\text{g}})(4.18\ \text{J} \cdot \cancel{\text{g}}^{-1} \cdot {}^\circ\text{C}^{-1})}$$

$\Delta T = -10.6\ {}^\circ\text{C}$

Step 5: Calculate the final temperature, ΔT_{final}.

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

$$T_{\text{final}} = \Delta T + T_{\text{initial}}$$

$$= -10.6\ {}^\circ\text{C} + 20.0\ {}^\circ\text{C}$$

$T_{\text{final}} = 9.4\ {}^\circ\text{C}$

Statement: The final temperature observed when 20.0 g of ammonium chloride is added to 125 mL water at 20.0 °C is 9.4 °C.

5. (a) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Potential energy, E_p (kJ)." Since the reaction is endothermic, the products will have greater potential energy 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. This process is formation, so ΔH is ΔH_f.

**Potential Energy Changes during the Formation Reaction**
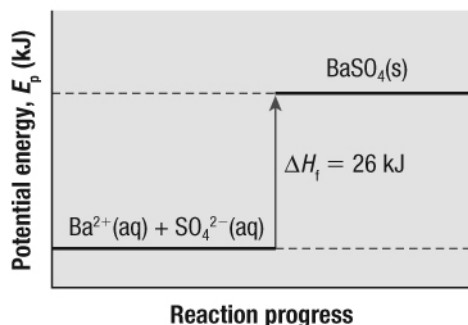
(b) Draw the reaction coordinates with the *x*-axis labelled "Reaction progress" and the *y*-axis labelled "Potential energy, *E*p (kJ)." Since the reaction is exothermic, the products will have lower potential energy than the reactants. Point the arrow from reactants downward to products. Write the reactants on the upper left-hand side of the graph and the products on the lower right-hand side of the graph. Include the ΔH of the reaction. This process is combustion, so ΔH is ΔH_c.

Potential Energy Changes during the Combustion Reaction
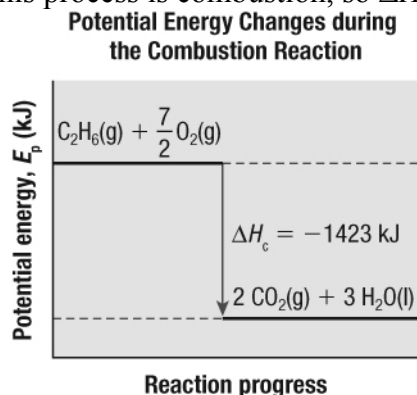
6. (a) A thermochemical equation for the combustion for propane is:

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)\quad \Delta H = -2220$ kJ
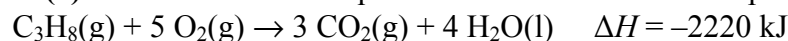
(b) A thermochemical equation for the separation of atoms in the chlorine molecule is:

$Cl_2(g) \rightarrow 2\ Cl(g)\quad \Delta H = 243$ kJ
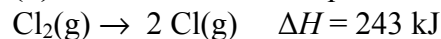
(c) A thermochemical equation for the formation of iron(III) oxide from its elements is:

$2\ Fe(s) + \dfrac{3}{2}\ O_2(g) \rightarrow Fe_2O_3(s)\quad \Delta H = -824$ kJ
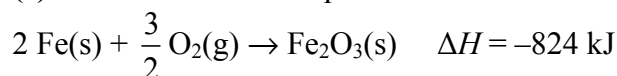
(d) A thermochemical equation for the decomposition of hydrogen chloride into its elements is:

$HCl(g) \rightarrow \dfrac{1}{2}\ H_2(g) + \dfrac{1}{2}\ Cl_2(g)\quad \Delta H = 93$ kJ
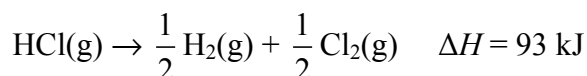
7. (a) A thermochemical equation for the formation of 1 mol of copper(II) chloride from its elements is:

$Cu(s) + Cl_2(g) \rightarrow CuCl_2(s) + 220.1$ kJ
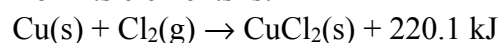
(b) A thermochemical equation for the conversion of graphite to diamond is:

$C_{graphite}(s) + 2.0\ \text{kJ} \rightarrow C_{diamond}(s)$
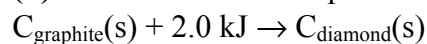
(c) A thermochemical equation for the decomposition of silver chloride into its elements is:

$AgCl(s) + 127.1\ \text{kJ} \rightarrow Ag(s) + \dfrac{1}{2}\ Cl_2(g)$