CHAPTER 7 IN REVIEW

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CONCEPTS

The Nature of Energy and Thermodynamics (7.2, 7.3)

- Energy, which is measured in the SI unit of joules (J), is the capacity to do work.
- Work is the result of a force acting through a distance.
- Many different kinds of energy exist, including kinetic energy, thermal energy, potential energy, and chemical energy, a type of potential energy associated with the relative positions of electrons and nuclei in atoms and molecules.
- The first law of thermodynamics states that energy can be converted from one form to another, but the total amount of energy is always conserved.
- The internal energy (*E*) of a system is the sum of all of its kinetic and potential energy. Internal energy is a state function, which means that it depends only on the state of the system and not on the pathway by which it got to that state.
- A chemical system exchanges energy with its surroundings through heat (the transfer of thermal energy caused by a temperature difference) or work. The total change in internal energy is the sum of these two quantities.

Heat and Work (7.4)

- We quantify heat with the equation $q = m \times C_s \times \Delta T$. In this expression, C_s is the specific heat capacity, the amount of heat required to change the temperature of 1 g of the substance by 1 °C. Compared to most substances, water has a very high heat capacity—it takes a lot of heat to change its temperature.
- The type of work most characteristic of chemical reactions is pressure–volume work, which occurs when a gas expands against an external pressure. Pressure–volume work can be quantified with the equation $w = -P_{ext}\Delta V$.
- The change in internal energy (ΔE) that occurs during a chemical reaction is the sum of the heat (q) exchanged and the work (w) done: $\Delta E = q + w$.

Enthalpy (7.6)

- The heat evolved in a chemical reaction occurring at constant pressure is the change in enthalpy (ΔH) for the reaction. Like internal energy, enthalpy is a state function.
- An endothermic reaction has a positive enthalpy of reaction; an exothermic reaction has a negative enthalpy of reaction.
- We can use the enthalpy of reaction to stoichiometrically determine the heat evolved when a specific amount of reactant reacts.

Calorimetry (7.5, 7.7)

- Calorimetry is a method of measuring ΔE or ΔH for a reaction.
- In bomb calorimetry, the reaction is carried out under conditions of constant volume, so $\Delta E = q_v$. We can therefore use the temperature change of the calorimeter to calculate ΔE for the reaction.
- When a reaction takes place at constant pressure, energy may be released both as heat and as work.
- In coffee-cup calorimetry, a reaction is carried out under atmospheric pressure in a solution, so $q = \Delta H$. We use the temperature change of the solution to calculate ΔH for the reaction.

Calculating ΔH_{rxn} (7.8, 7.9)

- We can calculate the enthalpy of reaction (ΔH_{rxn}) from known thermochemical data using the following relationships: (a) when a reaction is multiplied by a factor, ΔH_{rxn} is multiplied by the same factor; (b) when a reaction is reversed, ΔH_{rxn} changes sign; and (c) if a chemical reaction can be expressed as a sum of two or more steps, ΔH_{rxn} is the sum of the ΔH's for the individual steps (Hess's law). We can use these relationships to determine the enthalpy change of an unknown reaction from reactions with known enthalpy changes.
- A second method to calculate ΔH_{rxn} from known thermochemical data involves using tabulated standard enthalpies of formation for the reactants and products of the reaction. These are usually tabulated for substances in their standard states, and the enthalpy of reaction is called the standard enthalpy of reaction (ΔH_{rxn}°). For any reaction, we obtain ΔH_{rxn}° by subtracting the sum of the enthalpies of formation of the reactants multiplied by their stoichiometric coefficients from the sum of the enthalpies of formation of the products multiplied by their stoichiometric coefficients.

Environmental Problems Associated with Fossil Fuel Use (7.10)

- Fossil fuels are nonrenewable fuels; once humans consume them, they cannot be replaced.
- At current rates of consumption, natural gas and petroleum reserves will be depleted in 50–100 years.
- In addition to their limited supply, the products of the combustion of fossil fuels—directly or indirectly formed—contribute to environmental problems including air pollution, acid rain, and global climate change, which involves an increase in Earth's average temperature caused by CO_2 emission.

EQUATIONS AND RELATIONSHIPS

Kinetic Energy (7.2)

$$KE = \frac{1}{2}mv^2$$

Change in Internal Energy (ΔE) of a Chemical System (7.3)

$$\Delta E = E_{\text{products}} - E_{\text{reactants}}$$

Energy Flow between System and Surroundings (7.3)

$$\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$$

Relationship between Internal Energy (ΔE), Heat (q), and Work (w) (7.3)

$$\Delta E = q + w$$

Relationship between Heat (q), Temperature (T), and Heat Capacity (C) (7.4)

$$q = C \times \Delta T$$

Relationship between Heat (q), Mass (m), Temperature (T), and Specific Heat Capacity of a Substance (C_s) (7.4)

$$q = m \times C_s \times \Delta T$$

Relationship between Work (w), Force (F), and Distance (D) (7.4)

$$w = F \times D$$

Relationship between Work (w), Pressure (P), and Change in Volume (ΔV) (7.4)

$$w = -P_{\text{ext}}\,\Delta V$$

Change in Internal Energy (ΔE) of System at Constant Volume (7.5)

$$\Delta E = q_v$$

Heat of a Bomb Calorimeter (q_{cal}) (7.5)

$$q_{\text{cal}} = C_{\text{cal}} \times \Delta T$$

Heat Exchange between a Calorimeter and a Reaction (7.5)

$$q_{\text{cal}} = -q_{\text{rxn}}$$

Relationship between Enthalpy (ΔH), Internal Energy (ΔE), Pressure (P), and Volume (V) (7.6)

$$\Delta H = \Delta E + P\Delta V$$

$$\Delta H = q_p$$

Relationship between Enthalpy of a Reaction ($\Delta H^\circ_{\text{rxn}}$) and the Heats of Formation (ΔH°_f) (7.9)

$$\Delta H^\circ_{\text{rxn}} = \sum n_p \Delta H^\circ_f\,(\text{products}) - \sum n_r \Delta H^\circ_f\,(\text{reactants})$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Analyze types of energy and convert between energy units (7.2)	**Exercises** 33–36
Analyze changes in internal energy in terms of heat and work (7.3)	Example 7.1 **For Practice** 7.1 **Exercises** 37–44
Determine heat from temperature changes (7.4)	Example 7.2 **For Practice** 7.2 **For More Practice** 7.2 **Exercises** 45–50
Calculate quantities in thermal energy transfer (7.4)	Example 7.3 **For Practice** 7.3 **Exercises** 65–70
Analyze processes involving pressure–volume work (7.4)	Example 7.4 **For Practice** 7.4 **For More Practice** 7.4 **Exercises** 51–54
Analyze energy changes for combustion reactions inside a bomb calorimeter (7.5)	Example 7.5 **For Practice** 7.5 **For More Practice** 7.5 **Exercises** 73–74
Predict endothermic and exothermic processes (7.6)	Example 7.6 **For Practice** 7.6 **Exercises** 55–58
Perform stoichiometric calculations involving the enthalpy of reaction (7.6)	Example 7.7 **For Practice** 7.7 **For More Practice** 7.7 **Exercises** 59–64
Analyze enthalpy changes for reactions in a coffee-cup calorimeter (7.7)	Example 7.8 **For Practice** 7.8 **Exercises** 71–72, 75–76
Analyze how changes in chemical reactions affect the enthalpy of reaction (7.8)	Example 7.9 **For Practice** 7.9 **For More Practice** 7.9 **Exercises** 77–82
Determine the standard enthalpy change for a reaction using standard enthalpies of formation (7.9)	Examples 7.10, 7.11, 7.12 **For Practice** 7.10, 7.11, 7.12 **Exercises** 83–92
Analyze the effects of energy use on the environment (7.10)	Example 7.13 **For Practice** 7.13 **Exercises** 93–96

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. What is thermochemistry? Why is it important?

2. What is energy? What is work? List some examples of each.

3. What is kinetic energy? What is potential energy? List some examples of each.

4. State the law of conservation of energy. How does it relate to energy exchanges between a thermodynamic system and its surroundings?

5. What is the SI unit of energy? List some other common units of energy.

6. State the first law of thermodynamics. What are its implications?

7. A friend claims to have constructed a machine that creates electricity but requires no energy input. Explain why you should be suspicious of your friend's claim.

8. What is a state function? List some examples of state functions.

9. What is internal energy? Is internal energy a state function?

10. If energy flows out of a chemical system and into the surroundings, what is the sign of ΔE_{system}?

11. If the internal energy of the products of a reaction is higher than the internal energy of the reactants, what is the sign of ΔE for the reaction? In which direction does energy flow?

12. What is heat? Explain the difference between heat and temperature.

13. How is the change in internal energy of a system related to heat and work?

14. Explain how the sum of heat and work can be a state function, even though heat and work are themselves not state functions.

15. What is heat capacity? Explain the difference between heat capacity and specific heat capacity.

16. Explain how the high specific heat capacity of water can affect the weather in coastal regions.

17. If two objects, A and B, of different temperature come into direct contact, what is the relationship between the heat lost by one object and the heat gained by the other? What is the relationship between the temperature changes of the two objects? (Assume that the two objects do not lose any heat to anything else.)

18. What is pressure–volume work? How is it calculated?

19. What is calorimetry? Explain the difference between a coffee-cup calorimeter and a bomb calorimeter. What is each designed to measure?

20. What is the change in enthalpy (ΔH) for a chemical reaction? How is ΔH different from ΔE?

21. Explain the difference between an exothermic and an endothermic reaction. Give the sign of ΔH for each type of reaction.

22. From a molecular viewpoint, where does the energy emitted in an exothermic chemical reaction come from? Why does the reaction mixture undergo an increase in temperature even though energy is emitted?

23. From a molecular viewpoint, where does the energy absorbed in an endothermic chemical reaction go? Why does the reaction mixture undergo a decrease in temperature even though energy is absorbed?

24. Is the change in enthalpy for a reaction an extensive property? Explain the relationship between ΔH for a reaction and the amounts of reactants and products that undergo reaction.

25. Explain how the value of ΔH for a reaction changes upon each operation.

a. multiplying the reaction by a factor
b. reversing the reaction

Why do these relationships hold?

26. What is Hess's law? Why is it useful?

27. What is a standard state? What is the standard enthalpy change for a reaction?

28. What is the standard enthalpy of formation for a compound? For a pure element in its standard state?

29. How do you calculate $\Delta H^\circ_{\text{rxn}}$ from tabulated standard enthalpies of formation?

30. What are the main sources of the energy consumed in the United States?

31. What are the main environmental problems associated with fossil fuel use?

32. Explain global climate change. What causes global warming? What is the evidence that global warming is occurring?

PROBLEMS BY TOPIC

Energy Units

33. Convert between energy units.
MISSED THIS? *Read Section 7.2*

a. 534 kWh to J
b. 215 kJ to Cal
c. 567 Cal to J
d. 2.85×10^3 J to cal

34. Convert between energy units.

a. 231 cal to kJ
b. 132×10^4 kJ to kcal
c. 4.99×10^3 kJ to kWh
d. 2.88×10^4 J to Cal

35. Suppose that a person eats 2387 Calories per day. Convert this amount of energy into each unit.
MISSED THIS? *Read Section 7.2*

a. J
b. kJ
c. kWh

36. A particular frost-free refrigerator uses about 745 kWh of electrical energy per year. Express this amount of energy in each unit.

a. J
b. kJ
c. Cal

Internal Energy, Heat, and Work

37. Which statement is true of the internal energy of a system and its surroundings during an energy exchange with a negative ΔE_{sys}? **MISSED THIS?** *Read Section 7.3; Watch KCV 7.3*

a. The internal energy of the system increases and the internal energy of the surroundings decreases.

b. The internal energy of both the system and the surroundings increases.

c. The internal energy of both the system and the surroundings decreases.

d. The internal energy of the system decreases and the internal energy of the surroundings increases.

38. During an energy exchange, a chemical system absorbs energy from its surroundings. What is the sign of ΔE_{sys} for this process? Explain.

39. Identify each energy exchange as primarily heat or work and determine whether the sign of ΔE is positive or negative for the system.
MISSED THIS? *Read Section 7.3; Watch KCV 7.3*

a. Sweat evaporates from skin, cooling the skin. (The evaporating sweat is the system.)

b. A balloon expands against an external pressure. (The contents of the balloon is the system.)

c. An aqueous chemical reaction mixture is warmed with an external flame. (The reaction mixture is the system.)

40. Identify each energy exchange as primarily heat or work and determine whether the sign of ΔE is positive or negative for the system.

a. A rolling billiard ball collides with another billiard ball. The first billiard ball (defined as the system) stops rolling after the collision.

b. A book falls to the floor. (The book is the system.)

c. A father pushes his daughter on a swing. (The daughter and the swing are the system.)

41. A system releases 622 kJ of heat and does 105 kJ of work on the surroundings. What is the change in internal energy of the system? **MISSED THIS?** *Read Section 7.3*

42. A system absorbs 196 kJ of heat, and the surroundings do 117 kJ of work on the system. What is the change in internal energy of the system?

43. The gas in a piston (defined as the system) warms and absorbs 655 J of heat. The expansion performs 344 J of work on the surroundings. What is the change in internal energy for the system? **MISSED THIS?** *Read Section 7.3*

44. The air in an inflated balloon (defined as the system) warms over a toaster and absorbs 115 J of heat. As it expands, it does 77 kJ of work. What is the change in internal energy for the system?

Heat, Heat Capacity, and Work

45. We pack two identical coolers for a picnic, placing 24 12-ounce soft drinks and five pounds of ice in each. However, the drinks that we put into cooler A were refrigerated for several hours before they were packed in the cooler, while the drinks that we put into cooler B were at room temperature. When we open the two coolers three hours later, most of the ice in cooler A is still present, while nearly all of the ice in cooler B has melted. Explain this difference.
MISSED THIS? *Read Section 7.4; Watch KCV 7.4*

46. A kilogram of aluminum metal and a kilogram of water are each warmed to 75 °C and placed in two identical insulated containers. One hour later, the two containers are opened, and the temperature of each substance is measured. The aluminum has cooled to 35 °C, while the water has cooled only to 66 °C. Explain this difference.

47. How much heat is required to warm 1.50 L of water from 25.0 °C to 100.0 °C? (Assume a density of 1.0 g/mL for the water.) **MISSED THIS?** *Read Section 7.3; Watch KCV 7.3, IWE 7.2*

48. How much heat is required to warm 1.50 kg of sand from 25.0 °C to 100.0 °C?

49. Suppose that 25 g of each substance is initially at 27.0 °C. What is the final temperature of each substance upon absorbing 2.35 kJ of heat? **MISSED THIS?** *Read Section 7.3; Watch KCV 7.3, IWE 7.2*

a. gold **b.** silver **c.** aluminum **d.** water

50. An unknown mass of each substance, initially at 23.0 °C, absorbs 1.95×10^3 J of heat. The final temperature is recorded. Find the mass of each substance.

a. Pyrex glass ($T_f = 55.4$ °C)

b. sand ($T_f = 62.1$ °C)

c. ethanol ($T_f = 44.2$ °C)

d. water ($T_f = 32.4$ °C)

51. How much work (in J) is required to expand the volume of a pump from 0.0 L to 2.5 L against an external pressure of 1.1 atm? **MISSED THIS?** *Read Section 7.4*

52. The average human lung expands by about 0.50 L during each breath. If this expansion occurs against an external pressure of 1.0 atm, how much work (in J) is done during the expansion?

53. The air within a piston equipped with a cylinder absorbs 565 J of heat and expands from an initial volume of 0.10 L to a final volume of 0.85 L against an external pressure of 1.0 atm. What is the change in internal energy of the air within the piston? **MISSED THIS?** *Read Section 7.4*

54. A gas is compressed from an initial volume of 5.55 L to a final volume of 1.22 L by an external pressure of 1.00 atm. During the compression the gas releases 124 J of heat. What is the change in internal energy of the gas?

Enthalpy and Thermochemical Stoichiometry

55. When 1 mol of a fuel burns at constant pressure, it produces 3452 kJ of heat and does 11 kJ of work. What are ΔE and ΔH for the combustion of the fuel? **MISSED THIS?** *Read Section 7.6*

56. The change in internal energy for the combustion of 1.0 mol of octane at a pressure of 1.0 atm is 5084.3 kJ. If the change in enthalpy is 5074.1 kJ, how much work is done during the combustion?

57. Determine whether each process is exothermic or endothermic and indicate the sign of ΔH.
MISSED THIS? *Read Section 7.6; Watch KCV 7.6*

a. natural gas burning on a stove

b. isopropyl alcohol evaporating from skin

c. water condensing from steam

58. Determine whether each process is exothermic or endothermic and indicate the sign of ΔH.

a. dry ice evaporating

b. a sparkler burning

c. the reaction that occurs in a chemical cold pack used to ice athletic injuries

59. Consider the thermochemical equation for the combustion of acetone (C_3H_6O), the main ingredient in nail polish remover:

$$C_3H_6O(l) + 4\,O_2(g) \longrightarrow 3\,CO_2(g) + 3\,H_2O(g) \qquad \Delta H^\circ_{\mathrm{rxn}} = -1790\ \mathrm{kJ}$$

If a bottle of nail polish remover contains 177 mL of acetone, how much heat is released by its complete combustion? The density of acetone is 0.788 g/mL.
MISSED THIS? *Read Section 7.6; Watch IWE 7.7*

60. What mass of natural gas (CH_4) must burn to emit 267 kJ of heat?

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g) \qquad \Delta H^\circ_{\mathrm{rxn}} = -802.3\ \mathrm{kJ}$$

61. Nitromethane (CH_3NO_2) burns in air to produce significant amounts of heat.

$$2\,CH_3NO_2(l) + {}^{3}\!/_{2}\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(l) + N_2(g) \qquad \Delta H^\circ_{\mathrm{rxn}} = -1418\ \mathrm{kJ}$$

How much heat is produced by the complete reaction of 5.56 kg of nitromethane?
MISSED THIS? *Read Section 7.6; Watch IWE 7.7*

62. Titanium reacts with iodine to form titanium(III) iodide, emitting heat.

$$2\,Ti(s) + 3\,I_2(g) \longrightarrow 2\,TiI_3(s) \qquad \Delta H^\circ_{\mathrm{rxn}} = -839\ \mathrm{kJ}$$

Determine the masses of titanium and iodine that react if 1.55×10^3 kJ of heat is emitted by the reaction.

63. The propane fuel (C_3H_8) used in gas barbeques burns according to the thermochemical equation:

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g) \qquad \Delta H^\circ_{\mathrm{rxn}} = -2044\ \mathrm{kJ}$$

If a pork roast must absorb 1.6×10^3 kJ to fully cook, and if only 10% of the heat produced by the barbeque is actually absorbed by the roast, what mass of CO_2 is emitted into the atmosphere during the grilling of the pork roast?
MISSED THIS? *Read Section 7.6; Watch IWE 7.7*

64. Charcoal is primarily carbon. Determine the mass of CO_2 produced by burning enough carbon (in the form of charcoal) to produce 5.00×10^2 kJ of heat.

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H^\circ_{\mathrm{rxn}} = -393.5\ \mathrm{kJ}$$

Thermal Energy Transfer

65. A silver block, initially at 58.5 °C, is submerged into 100.0 g of water at 24.8 °C, in an insulated container. The final temperature of the mixture upon reaching thermal equilibrium is 26.2 °C. What is the mass of the silver block?
MISSED THIS? *Read Section 7.4; Watch KCV 7.4, IWE 7.3*

66. A 32.5-g iron rod, initially at 22.7 °C, is submerged into an unknown mass of water at 63.2 °C, in an insulated container. The final temperature of the mixture upon reaching thermal equilibrium is 59.5 °C. What is the mass of the water?

67. A 31.1-g wafer of pure gold, initially at 69.3 °C, is submerged into 64.2 g of water at 27.8 °C in an insulated container. What is the final temperature of both substances at thermal equilibrium?
MISSED THIS? *Read Section 7.4; Watch KCV 7.4, IWE 7.3*

68. A 2.85-g lead weight, initially at 10.3 °C, is submerged in 7.55 g of water at 52.3 °C in an insulated container. What is the final temperature of both substances at thermal equilibrium?

69. Two substances, A and B, initially at different temperatures, come into contact and reach thermal equilibrium. The mass of substance A is 6.15 g and its initial temperature is 20.5 °C. The mass of substance B is 25.2 g and its initial temperature is 52.7 °C. The final temperature of both substances at thermal equilibrium is 46.7 °C. If the specific heat capacity of substance B is 1.17 J/g·°C, what is the specific heat capacity of substance A?
MISSED THIS? *Read Section 7.4; Watch KCV 7.4, IWE 7.3*

70. A 2.74-g sample of a substance suspected of being pure gold is warmed to 72.1 °C and submerged into 15.2 g of water initially at 24.7 °C. The final temperature of the mixture is 26.3 °C. What is the heat capacity of the unknown substance? Could the substance be pure gold?

Calorimetry

71. Exactly 1.5 g of a fuel burns under conditions of constant pressure and then again under conditions of constant volume. In measurement A the reaction produces 25.9 kJ of heat, and in measurement B the reaction produces 23.3 kJ of heat. Which measurement (A or B) corresponds to conditions of constant pressure? Which one corresponds to conditions of constant volume? Explain.
MISSED THIS? *Read Section 7.6*

72. In order to obtain the largest possible amount of heat from a chemical reaction in which there is a large increase in the number of moles of gas, should you carry out the reaction under conditions of constant volume or constant pressure? Explain.

73. When 0.514 g of biphenyl ($C_{12}H_{10}$) undergoes combustion in a bomb calorimeter, the temperature rises from 25.8 °C to 29.4 °C. Find ΔE_{rxn} for the combustion of biphenyl in kJ/mol biphenyl. The heat capacity of the bomb calorimeter, determined in a separate experiment, is 5.86 kJ/°C.
MISSED THIS? *Read Section 7.5; Watch IWE 7.5*

74. Mothballs are composed primarily of the hydrocarbon naphthalene ($C_{10}H_8$). When 1.025 g of naphthalene burns in a bomb calorimeter, the temperature rises from 24.25 °C to 32.33 °C. Find ΔE_{rxn} for the combustion of naphthalene. The heat capacity of the bomb calorimeter, determined in a separate experiment, is 5.11 kJ/°C.

75. Zinc metal reacts with hydrochloric acid according to the balanced equation:

$$Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

When 0.103 g of Zn(*s*) is combined with enough HCl to make 50.0 mL of solution in a coffee-cup calorimeter, all of the zinc reacts, raising the temperature of the solution from 22.5 °C to 23.7 °C. Find ΔH_{rxn} for this reaction as written. (Use 1.0 g/mL for the density of the solution and 4.18 J/g·°C as the specific heat capacity.)
MISSED THIS? *Read Section 7.7; Watch IWE 7.8*

76. Instant cold packs used to ice athletic injuries on the field contain ammonium nitrate and water separated by a thin plastic divider. When the divider is broken, the ammonium nitrate dissolves according to the endothermic reaction:

$$NH_4NO_3(s) \longrightarrow NH_4^+(aq) + NO_3^-(aq)$$

In order to measure the enthalpy change for this reaction, 1.25 g of NH_4NO_3 is dissolved in enough water to make 25.0 mL of solution. The initial temperature is 25.8 °C and the final temperature (after the solid dissolves) is 21.9 °C. Calculate the change in enthalpy for the reaction in kJ. (Use 1.0 g/mL as the density of the solution and 4.18 J/g·°C as the specific heat capacity.)

Quantitative Relationships Involving Δ*H* and Hess's Law

77. For each generic reaction, determine the value of ΔH_2 in terms of ΔH_1. **MISSED THIS?** *Read Section 7.8*

a. $A + B \longrightarrow 2\,C \qquad \Delta H_1$
$2\,C \longrightarrow A + B \qquad \Delta H_2 = ?$

b. $A + {}^{1}\!/_{2}\,B \longrightarrow C \qquad \Delta H_1$
$2\,A + B \longrightarrow 2\,C \qquad \Delta H_2 = ?$

c. $A \longrightarrow B + 2\,C \qquad \Delta H_1$
${}^{1}\!/_{2}\,B + C \longrightarrow {}^{1}\!/_{2}\,A \qquad \Delta H_2 = ?$

78. Consider the generic reaction:

$$\text{A} + 2\,\text{B} \longrightarrow \text{C} + 3\,\text{D} \qquad \Delta H = 155\text{ kJ}$$

Determine the value of ΔH for each related reaction.

a. $3\,\text{A} + 6\,\text{B} \longrightarrow 3\,\text{C} + 9\,\text{D}$ **b.** $\text{C} + 3\,\text{D} \longrightarrow \text{A} + 2\,\text{B}$

c. $\frac{1}{2}\text{C} + \frac{3}{2}\text{D} \longrightarrow \frac{1}{2}\text{A} + \text{B}$

79. Calculate ΔH_{rxn} for the reaction:

$$Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g)$$

Use the following reactions and given ΔH's:

$2\,Fe(s) + \frac{3}{2}O_2(g) \longrightarrow Fe_2O_3(s) \quad \Delta H = -824.2\text{ kJ}$

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -282.7\text{ kJ}$

MISSED THIS? *Read Section 7.8; Watch IWE 7.9*

80. Calculate ΔH_{rxn} for the reaction:

$$CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$$

Use the following reactions and given ΔH's:

$Ca(s) + CO_2(g) + \frac{1}{2}O_2(g) \longrightarrow CaCO_3(s) \quad \Delta H = -812.8\text{ kJ}$

$2\,Ca(s) + O_2(g) \longrightarrow 2\,CaO(s) \quad \Delta H = -1269.8\text{ kJ}$

81. Calculate ΔH_{rxn} for the reaction:

$$5\,C(s) + 6\,H_2(g) \longrightarrow C_5H_{12}(l)$$

Use the following reactions and given ΔH's:

$C_5H_{12}(l) + 8\,O_2(g) \longrightarrow 5\,CO_2(g) + 6\,H_2O(g) \quad \Delta H = -3244.8\text{ kJ}$

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5\text{ kJ}$

$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \quad \Delta H = -483.5\text{ kJ}$

MISSED THIS? *Read Section 7.8; Watch IWE 7.9*

82. Calculate ΔH_{rxn} for the reaction:

$$CH_4(g) + 4\,Cl_2(g) \longrightarrow CCl_4(g) + 4\,HCl(g)$$

Use the following reactions and given ΔH's:

$C(s) + 2\,H_2(g) \longrightarrow CH_4(g) \quad \Delta H = -74.4\text{ kJ}$

$C(s) + 2\,Cl_2(g) \longrightarrow CCl_4(g) \quad \Delta H = -95.7\text{ kJ}$

$H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g) \quad \Delta H = -92.3\text{ kJ}$

Enthalpies of Formation and ΔH

83. Write an equation for the formation of each compound from its elements in their standard states, and find ΔH_f° for each in Appendix IIB. **MISSED THIS?** *Read Section 7.9*

a. $NH_3(g)$ **b.** $CO_2(g)$ **c.** $Fe_2O_3(s)$ **d.** $CH_4(g)$

84. Write an equation for the formation of each compound from its elements in their standard states, and find $\Delta H_{\text{rxn}}^\circ$ for each in Appendix IIB.

a. $NO_2(g)$ b. $MgCO_3(s)$ c. $C_2H_4(g)$ **d.** $CH_3OH(l)$

85. Hydrazine (N_2H_4) is a fuel used by some spacecraft. It is normally oxidized by N_2O_4 according to the equation:

$$N_2H_4(l) + N_2O_4(g) \longrightarrow 2\,N_2O(g) + 2\,H_2O(g)$$

Calculate $\Delta H_{\text{rxn}}^\circ$ for this reaction using standard enthalpies of formation.

MISSED THIS? *Read Section 7.9; Watch KCV 7.9, IWE 7.11*

86. Pentane (C_5H_{12}) is a component of gasoline that burns according to the following balanced equation:

$$C_5H_{12}(l) + 8\,O_2(g) \longrightarrow 5\,CO_2(g) + 6\,H_2O(g)$$

Calculate $\Delta H_{\text{rxn}}^\circ$ for this reaction using standard enthalpies of formation. (The standard enthalpy of formation of liquid pentane is -146.8 kJ/mol.)

87. Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for each reaction.

MISSED THIS? *Read Section 7.9; Watch KCV 7.9, IWE 7.11*

a. $C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$

b. $CO(g) + H_2O(g) \longrightarrow H_2(g) + CO_2(g)$

c. $3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(aq) + NO(g)$

d. $Cr_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Cr(s) + 3\,CO_2(g)$

88. Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for each reaction.

a. $2\,H_2S(g) + 3\,O_2(g) \longrightarrow 2\,H_2O(l) + 2\,SO_2(g)$

b. $SO_2(g) + \frac{1}{2}O_2(g) \longrightarrow SO_3(g)$

c. $C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)$

d. $N_2O_4(g) + 4\,H_2(g) \longrightarrow N_2(g) + 4\,H_2O(g)$

89. During photosynthesis, plants use energy from sunlight to form glucose ($C_6H_{12}O_6$) and oxygen from carbon dioxide and water. Write a balanced equation for photosynthesis and calculate $\Delta H_{\text{rxn}}^\circ$.

MISSED THIS? *Read Section 7.9; Watch KCV 7.9, IWE 7.11*

90. Ethanol (C_2H_5OH) can be made from the fermentation of crops and has been used as a fuel additive to gasoline. Write a balanced equation for the combustion of ethanol and calculate $\Delta H_{\text{rxn}}^\circ$.

91. Top fuel dragsters and funny cars burn nitromethane as fuel according to the balanced combustion equation:

$$2\,CH_3NO_2(l) + \tfrac{3}{2}O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(l) + N_2(g)$$

$$\Delta H_{\text{rxn}}^\circ = -1418\text{ kJ}$$

Calculate the standard enthalpy of formation (ΔH_f°) for nitromethane.

MISSED THIS? *Read Section 7.9; Watch KCV 7.9, IWE 7.11*

92. The explosive nitroglycerin ($C_3H_5N_3O_9$) decomposes rapidly upon ignition or sudden impact according to the balanced equation:

$$4\,C_3H_5N_3O_9(l) \longrightarrow 12\,CO_2(g) + 10\,H_2O(g) + 6\,N_2(g) + O_2(g)$$

$$\Delta H_{\text{rxn}}^\circ = -5678\text{ kJ}$$

Calculate the standard enthalpy of formation (ΔH_f°) for nitroglycerin.

Energy Use and the Environment

93. Determine the mass of CO_2 produced by burning enough of each fuel to produce 1.00×10^2 kJ of heat. Which fuel contributes least to global warming per kJ of heat produced?

MISSED THIS? *Read Sections 7.4, 7.6; KCV 7.4, 7.6, IWE 7.2, 7.7*

a. $CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$

$$\Delta H_{\text{rxn}}^\circ = -802.3\text{ kJ}$$

b. $C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g)$

$$\Delta H_{\text{rxn}}^\circ = -2043\text{ kJ}$$

c. $C_8H_{18}(l) + \frac{25}{2}O_2(g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(g)$

$$\Delta H_{\text{rxn}}^\circ = -5074.1\text{ kJ}$$

94. Methanol (CH_3OH) has been suggested as a fuel to replace gasoline. Write a balanced equation for the combustion of methanol, find $\Delta H_{\text{rxn}}^\circ$, and determine the mass of carbon dioxide emitted per kJ of heat produced. Use the information from the previous exercise to calculate the same quantity for octane, C_8H_{18}. How does methanol compare to octane with respect to global warming?

95. The citizens of the world burn the fossil fuel equivalent of 7×10^{12} kg of petroleum per year. Assume that all of this petroleum is in the form of octane (C_8H_{18}) and calculate how much CO_2 (in kg) the world produces from fossil fuel combustion per year. (*Hint:* Begin by writing a balanced equation for the combustion of octane.) If the atmosphere currently contains approximately 3×10^{15} kg of CO_2, how long will it take for the world's fossil fuel combustion to double the amount of atmospheric carbon dioxide?

MISSED THIS? *Read Sections 7.4, 7.6; KCV 7.4, 7.6, IWE 7.2, 7.7*

96. In a sunny location, sunlight has a power density of about 1 kW/m^2. Photovoltaic solar cells can convert this power into electricity with 15% efficiency. If a typical home uses 385 kWh of electricity per month, how many square meters of solar cells are required to meet its energy requirements? Assume that electricity can be generated from the sunlight for 8 hours per day.

▲ What area of solar cells do you need to power a home?

CUMULATIVE PROBLEMS

97. The kinetic energy of a rolling billiard ball is given by KE $= \frac{1}{2}mv^2$. Suppose a 0.17-kg billiard ball is rolling down a pool table with an initial speed of 4.5 m/s. As it travels, it loses some of its energy as heat. The ball slows down to 3.8 m/s and then collides head-on with a second billiard ball of equal mass. The first billiard ball completely stops, and the second one rolls away with a velocity of 3.8 m/s. Assume the first billiard ball is the system. Calculate w, q, and ΔE for the process.

98. A 100-W lightbulb is placed in a cylinder equipped with a moveable piston. The lightbulb is turned on for 0.015 hour, and the assembly expands from an initial volume of 0.85 L to a final volume of 5.88 L against an external pressure of 1.0 atm. Use the wattage of the lightbulb and the time it is on to calculate ΔE in joules (assume that the cylinder and lightbulb assembly is the system and assume two significant figures). Calculate w and q.

99. Evaporating sweat cools the body because evaporation is an endothermic process:

$$H_2O(l) \longrightarrow H_2O(g) \qquad \Delta H^\circ_{rxn} = +44.01\text{ kJ}$$

Estimate the mass of water that must evaporate from the skin to cool the body by 0.50 °C. Assume a body mass of 95 kg and assume that the specific heat capacity of the body is 4.0 J/g·°C.

100. LP gas burns according to the exothermic reaction:

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g)$$

$$\Delta H^\circ_{rxn} = -2044\text{ kJ}$$

What mass of LP gas is necessary to heat 1.5 L of water from room temperature (25.0 °C) to boiling (100.0 °C)? Assume that during heating, 15% of the heat emitted by the LP gas combustion goes to heat the water. The rest is lost as heat to the surroundings.

101. Use standard enthalpies of formation to calculate the standard change in enthalpy for the melting of ice. (The ΔH°_f for $H_2O(s)$ is -291.8 kJ/mol.) Use this value to calculate the mass of ice required to cool 355 mL of a beverage from room temperature (25.0 °C) to 0.0 °C. Assume that the specific heat capacity and density of the beverage are the same as those of water.

102. Dry ice is solid carbon dioxide. Instead of melting, solid carbon dioxide sublimes according to the equation:

$$CO_2(s) \longrightarrow CO_2(g)$$

When dry ice is added to warm water, heat from the water causes the dry ice to sublime more quickly. The evaporating carbon dioxide produces a dense fog often used to create special effects. In a simple dry ice fog machine, dry ice is added to warm water in a Styrofoam cooler. The dry ice produces fog until it evaporates away, or until the water gets too cold to sublime the dry ice quickly enough. Suppose that a small Styrofoam cooler holds 15.0 L of water heated to 85 °C. Use standard enthalpies of formation to calculate the change in enthalpy for dry ice sublimation, and calculate the mass of dry ice that should be added to the water so that the dry ice completely sublimes away when the water reaches 25 °C. Assume no heat loss to the surroundings. (The ΔH°_f for $CO_2(s)$ is -427.4 kJ/mol.)

▲ When carbon dioxide sublimes, the gaseous CO_2 is cold enough to cause water vapor in the air to condense, forming fog.

103. A 25.5-g aluminum block is warmed to 65.4 °C and plunged into an insulated beaker containing 55.2 g water initially at 22.2 °C. The aluminum and the water are allowed to come to thermal equilibrium. Assuming that no heat is lost, what is the final temperature of the water and aluminum?

104. If 50.0 mL of ethanol (density = 0.789 g/mL) initially at 7.0 °C is mixed with 50.0 mL of water (density = 1.0 g/mL) initially at 28.4 °C in an insulated beaker, and assuming that no heat is lost, what is the final temperature of the mixture?

105. Palmitic acid ($C_{16}H_{32}O_2$) is a dietary fat found in beef and butter. The caloric content of palmitic acid is typical of fats in general. Write a balanced equation for the complete combustion of palmitic acid and calculate the standard enthalpy of combustion. What is the caloric content of palmitic acid in Cal/g? Do the same calculation for table sugar (sucrose, $C_{12}H_{22}O_{11}$). Which dietary substance (sugar or fat) contains more Calories per gram? The standard enthalpy of formation of palmitic acid is –208 kJ/mol and that of sucrose is –2226.1 kJ/mol. [Use $H_2O(l)$ in the balanced chemical equations because the metabolism of these compounds produces liquid water.]

106. Hydrogen and methanol have both been proposed as alternatives to hydrocarbon fuels. Write balanced reactions for the complete combustion of hydrogen and methanol and use standard enthalpies of formation to calculate the amount of heat released per kilogram of the fuel. Which fuel contains the most energy in the least mass? How does the energy of these fuels compare to that of octane (C_8H_{18})?

107. Derive a relationship between ΔH and ΔE for a process in which the temperature of a fixed amount of an ideal gas changes.

108. Under certain nonstandard conditions, oxidation by $O_2(g)$ of 1 mol of $SO_2(g)$ to $SO_3(g)$ absorbs 89.5 kJ. The enthalpy of formation of $SO_3(g)$ is –204.2 kJ under these conditions. Find the enthalpy of formation of $SO_2(g)$.

109. One tablespoon of peanut butter has a mass of 16 g. It is combusted in a calorimeter whose heat capacity is 120.0 kJ/°C. The temperature of the calorimeter rises from 22.2 °C to 25.4 °C. Find the food caloric content of peanut butter.

110. A mixture of 2.0 mol of $H_2(g)$ and 1.0 mol of $O_2(g)$ is placed in a sealed evacuated container made of a perfect insulating material at 25 °C. The mixture is ignited with a spark and reacts to form liquid water. Determine the temperature of the water.

111. A 20.0-L volume of an ideal gas in a cylinder with a piston is at a pressure of 3.0 atm. Enough weight is suddenly removed from the piston to lower the external pressure to 1.5 atm. The gas then expands at constant temperature until its pressure is 1.5 atm. Find ΔE, ΔH, q, and w for this change in state.

112. When 10.00 g of phosphorus is burned in $O_2(g)$ to form $P_4O_{10}(s)$, enough heat is generated to raise the temperature of 2950 g of water from 18.0 °C to 38.0 °C. Calculate the enthalpy of formation of $P_4O_{10}(s)$ under these conditions.

113. The ΔH for the oxidation of sulfur in the gas phase to SO_3 is –204 kJ/mol and for the oxidation of SO_2 to SO_3 is 89.5 kJ/mol. Find the enthalpy of formation of SO_2 under these conditions.

114. The ΔH_f° of $TiI_3(s)$ is –328 kJ/mol and the ΔH° for the reaction $2\,Ti(s) + 3\,I_2(g) \longrightarrow 2\,TiI_3(s)$ is –839 kJ. Calculate the ΔH of sublimation of $I_2(s)$, which is a solid at 25 °C.

115. A gaseous fuel mixture contains 25.3% methane (CH_4), 38.2% ethane (C_2H_6), and the rest propane (C_3H_8) by volume. When the fuel mixture contained in a 1.55 L tank, stored at 755 mmHg and 298 K, undergoes complete combustion, how much heat is emitted? (Assume that the water produced by the combustion is in the gaseous state.)

116. A gaseous fuel mixture stored at 745 mmHg and 298 K contains only methane (CH_4) and propane (C_3H_8). When 11.7 L of this fuel mixture is burned, it produces 769 kJ of heat. What is the mole fraction of methane in the mixture? (Assume that the water produced by the combustion is in the gaseous state.)

117. A copper cube with an edge measuring 1.55 cm and an aluminum cube with an edge measuring 1.62 cm are both heated to 55.0 °C and submerged in 100.0 mL of water at 22.2 °C. What is the final temperature of the water when equilibrium is reached? (Assume a density of 0.998 g/mL for water.)

118. A pure gold ring and a pure silver ring have a total mass of 14.9 g. The two rings are heated to 62.0 °C and dropped into 15.0 mL of water at 23.5 °C. When equilibrium is reached, the temperature of the water is 25.0 °C. What is the mass of each ring? (Assume a density of 0.998 g/mL for water.)

CHALLENGE PROBLEMS

119. A typical frostless refrigerator uses 655 kWh of energy per year in the form of electricity. Suppose that all of this electricity is generated at a power plant that burns coal containing 3.2% sulfur by mass and that all of the sulfur is emitted as SO_2 when the coal is burned. If all of the SO_2 goes on to react with rainwater to form H_2SO_4, what mass of H_2SO_4 does the annual operation of the refrigerator produce? (*Hint:* Assume that the remaining percentage of the coal is carbon and begin by calculating $\Delta H^\circ_{\text{rxn}}$ for the combustion of carbon.)

120. A large sport utility vehicle has a mass of 2.5×10^3 kg. Calculate the mass of CO_2 emitted into the atmosphere upon accelerating the SUV from 0.0 mph to 65.0 mph. Assume that the required energy comes from the combustion of octane with 30% efficiency. (*Hint:* Use $KE = {}^1\!/_2\, mv^2$ to calculate the kinetic energy required for the acceleration.)

121. Combustion of natural gas (primarily methane) occurs in most household heaters. The heat given off in this reaction is used to raise the temperature of the air in the house. Assuming that all the energy given off in the reaction goes to heating up only the air in the house, determine the mass of methane required to heat the air in a house by 10.0 °C. Assume that the house dimensions are 30.0 m × 30.0 m × 3.0 m, specific heat capacity of air is 30 J/K·mol, and 1.00 mol of air occupies 22.4 L for all temperatures concerned.

122. When backpacking in the wilderness, hikers often boil water to sterilize it for drinking. Suppose that you are planning a backpacking trip and will need to boil 35 L of water for your group. What volume of fuel should you bring? Assume that the fuel has an average formula of C_7H_{16}, 15% of the heat generated from combustion goes to heat the water (the rest is lost to the surroundings), the density of the fuel is 0.78 g/mL, the initial temperature of the water is 25.0 °C, and the standard enthalpy of formation of C_7H_{16} is –224.4 kJ/mol.

123. An ice cube of mass 9.0 g is added to a cup of coffee. The coffee's initial temperature is 90.0 °C and the cup contains 120.0 g of liquid. Assume the specific heat capacity of the coffee is the same as that of water. The heat of fusion of ice (the heat associated with ice melting) is 6.0 kJ/mol. Find the temperature of the coffee after the ice melts.

124. Find ΔH, ΔE, q, and w for the freezing of water at –10.0 °C. The specific heat capacity of ice is 2.04 J/g·°C and its heat of fusion (the quantity of heat associated with melting) is –332 J/g.

125. Starting from the relationship between temperature and kinetic energy for an ideal gas, find the value of the molar heat capacity of an ideal gas when its temperature is changed at constant volume. Find its molar heat capacity when its temperature is changed at constant pressure.

126. An amount of an ideal gas expands from 12.0 L to 24.0 L at a constant pressure of 1.0 atm. Then the gas is cooled at a constant volume of 24.0 L back to its original temperature. Then it contracts back to its original volume. Find the total heat flow for the entire process.

127. The heat of vaporization of water at 373 K is 40.7 kJ/mol. Find q, w, ΔE, and ΔH for the evaporation of 454 g of water at this temperature at 1 atm.

128. Find ΔE, ΔH, q, and w for the change in state of 1.0 mol $H_2O(l)$ at 80 °C to $H_2O(g)$ at 110 °C. The heat capacity of $H_2O(l)$ $= 75.3$ J/mol K, the heat capacity of $H_2O(g) = 25.0$ J/mol K, and the heat of vaporization of H_2O is 40.7×10^3 J/mol at 100 °C.

129. The heat of combustion of liquid octane (C_8H_{18}) to carbon dioxide and liquid water at 298 K is -1303 kJ/mol. Find ΔE for this reaction.

130. Find ΔH for the combustion of ethanol (C_2H_6O) to carbon dioxide and liquid water from the following data. The heat capacity of the bomb calorimeter is 34.65 kJ/K, and the combustion of 1.765 g of ethanol raises the temperature of the calorimeter from 294.33 K to 295.84 K.

CONCEPTUAL PROBLEMS

131. Which statement is true of the internal energy of the system and its surroundings following a process in which $\Delta E_{sys} = +65$ kJ? Explain.

a. The system and the surroundings both lose 65 kJ of energy.
b. The system and the surroundings both gain 65 kJ of energy.
c. The system loses 65 kJ of energy and the surroundings gain 65 kJ of energy.
d. The system gains 65 kJ of energy and the surroundings lose 65 kJ of energy.

132. The internal energy of an ideal gas depends only on its temperature. Which statement is true of an isothermal (constant-temperature) expansion of an ideal gas against a constant external pressure? Explain.

a. ΔE is positive. **b.** w is positive.
c. q is positive. **d.** ΔE is negative.

133. Which expression describes the heat evolved in a chemical reaction when the reaction is carried out at constant pressure? Explain.

a. $\Delta E - w$ **b.** ΔE **c.** $\Delta E - q$

134. Two identical new refrigerators are plugged in for the first time. Refrigerator A is empty (except for air) and refrigerator B is filled with jugs of water. The compressors of both refrigerators immediately turn on and begin cooling the interiors of the refrigerators. After two hours, the compressor of refrigerator A turns off, while the compressor of refrigerator B continues to run. The next day, the compressor of refrigerator A can be heard turning on and off every few minutes, while the compressor of refrigerator B turns off and on every hour or so (and stays on longer each time). Explain these observations.

135. A 1-kg cylinder of aluminum and 1-kg jug of water, both at room temperature, are put into a refrigerator. After 1 hour, the temperature of each object is measured. One of the objects is much cooler than the other. Which one is cooler and why?

136. Two substances A and B, initially at different temperatures, are thermally isolated from their surroundings and allowed to come into thermal contact. The mass of substance A is twice the mass of substance B, but the specific heat capacity of substance B is four times the specific heat capacity of substance A. Which substance will undergo a larger change in temperature?

137. When 1 mol of a gas burns at constant pressure, it produces 2418 J of heat and does 5 J of work. Determine ΔE, ΔH, q, and w for the process.

138. In an exothermic reaction, the reactants lose energy, and the reaction feels hot to the touch. Explain why the reaction feels hot even though the reactants are losing energy. Where does the energy come from?

139. Which statement is true of a reaction in which ΔV is positive? Explain.

a. $\Delta H = \Delta E$ **b.** $\Delta H > \Delta E$ **c.** $\Delta H < \Delta E$

QUESTIONS FOR GROUP WORK

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

140. Have each group member write a problem involving the transfer of heat from one material in Table 7.4 to another material in the table. Working as a group, solve each problem. The group member who wrote each problem may act as the group facilitator when the group is working on his or her problem. What do all of your problems have in common? How do they differ?

141. Classify each process as endothermic or exothermic. What is the sign of ΔH for each process? Explain your answers.

a. gasoline burning in an engine
b. steam condensing on a mirror
c. water boiling in a pot

Provide at least two additional examples of exothermic processes and two additional examples of endothermic processes. Have each member of your group provide an example.

142. A propane tank on a home barbeque contains 10.4×10^3 g of propane.

a. Write the balanced chemical reaction for the combustion of gaseous propane (C_3H_8) to form water vapor and gaseous carbon dioxide.
b. Use the value for ΔH_{rxn} provided in the text to calculate the total amount of heat produced when the entire contents of the tank of propane is burned.
c. What mass of water could be warmed from 25 °C to 100 °C with this much heat?

143. Consider the decomposition of liquid hydrogen peroxide (H_2O_2) to form water and oxygen.

a. What is the heat of formation for hydrogen peroxide?
b. What is the heat of formation for liquid water?
c. What is the heat of formation for gaseous oxygen? Why?
d. Write the balanced chemical equations that correspond to the ΔH values you looked up for parts a, b, and c.
e. Write the balanced chemical equation for the decomposition of hydrogen peroxide to form water and oxygen. (Write the equation such that the coefficient on oxygen is 1.)
f. What is the heat of reaction for the process in part e?
g. Draw a scale diagram of this reaction (1 cm = 100 kJ) that shows the relative energies of reactants (on the left), products (on the right), and the elements in their most stable states (in the middle). Label all the energies you know.

DATA INTERPRETATION AND ANALYSIS

Heat Content of Fuels

144. The heating value of combustible fuels is evaluated based on the quantities known as the higher heating value (HHV) and the lower heating value (LHV). The HHV has a higher absolute value and assumes that the water produced in the combustion reaction is formed in the liquid state. The LHV has a lower absolute value and assumes that the water produced in the combustion reaction is formed in the gaseous state. The LHV is therefore the sum of the HHV (which is negative) and the heat of vaporization of water for the number of moles of water formed in the reaction (which is positive). The table lists the enthalpy of combustion—which is equivalent to the HHV—for several closely related hydrocarbons.

Hydrocarbon	ΔH_{comb} (kJ/mol)
$CH_4(g)$	−890
$C_2H_6(g)$	−1560
$C_3H_8(g)$	−2219
$C_4H_{10}(g)$	−2877
$C_5H_{12}(l)$	−3509
$C_6H_{14}(l)$	−4163
$C_7H_{16}(l)$	−4817
$C_8H_{18}(l)$	−5470

Use the information in the table to answer the following questions.

a. Write two balanced equations for the combustion of C_3H_8; one assuming the formation of liquid water and the other assuming the formation of gaseous water.

b. Given that the heat of vaporization of water is 44.0 kJ/mol, what is ΔH_{rxn} for each reaction in part a? Which quantity is the HHV? The LHV?

c. When propane is used to cook in an outdoor grill, is the amount of heat released the HHV or the LHV? What amount of heat is released upon combustion of 1.00 kg of propane in an outdoor grill?

d. For each CH_2 unit added to a hydrocarbon, what is the average increase in the absolute value of ΔH_{comb}?

ANSWERS TO CONCEPTUAL CONNECTIONS

System and Surroundings

7.1 **(a)** When ΔE_{sys} is negative, energy flows out of the system and into the surroundings. The energy increase in the surroundings must exactly match the decrease in the system.

Heat and Work

7.2 **(b)** Rolling a metal cylinder up a ramp requires that a force be applied to the cylinder over a distance and is therefore work.

The Heat Capacity of Water

7.3 **(a)** Bring the water into your sleeping bag; it has the higher heat capacity and will therefore release more heat as it cools.

Thermal Energy Transfer

7.4 **(c)** The specific heat capacity of substance B is twice that of A, but since the mass of B is half that of A, the quantity $m \times C_s$ is identical for both substances so that the final temperature is exactly midway between the two initial temperatures.

Pressure–Volume Work

7.5 **(d)** The work done is given by $w = -P_{ext} \cdot \Delta V$, which is equal to $-1.00 \text{ atm} (2.00 \text{ L} - 1.00 \text{ L}) = -1.00 \text{ L} \cdot \text{atm}$. Since $1 \text{ L} \cdot \text{atm} = 101.3 \text{ J}$, the work done (rounded to the correct number of significant figures) is -101 J.

The Difference between ΔH and ΔE

7.6 **(c)** ΔH represents only the heat exchanged; therefore, $\Delta H = -2658$ kJ. ΔE represents the heat *and work* exchanged; therefore, $\Delta E = -2661$ kJ. The signs of both ΔH and ΔE are negative because heat and work are flowing out of the system and into the surroundings. Notice that the values of ΔH and ΔE are similar in magnitude, as is the case in many chemical reactions.

Exothermic and Endothermic Reactions

7.7 **(b)** The temperature falls. An endothermic reaction absorbs heat from the surroundings. Since the flask is part of the immediate surroundings, the temperature of the flask falls.

Thermochemical Equations

7.8 **(c)** From the thermochemical equation, you can get the ratio

$$2 \text{ mol A} : -51.0 \text{ J} \qquad 6 \,\cancel{\text{mol A}} \left(\frac{-51.0 \text{ J}}{2 \,\cancel{\text{mol A}}} \right) = -153 \text{ J}$$

Constant-Pressure versus Constant-Volume Calorimetry

7.9 **(a)** The value of q_{rxn} with the greater magnitude (−12.5 kJ) must have come from the bomb calorimeter. Recall that $\Delta E_{rxn} = q_{rxn} + w_{rxn}$. In a bomb calorimeter, the energy change that occurs in the course of the reaction all takes the form of heat (q). In a coffee-cup calorimeter, the amount of energy released as heat may be smaller because some of the energy may be used to do work (w).

Relationships Involving ΔH_{rxn}

7.10 **(d)** When a chemical equation is multiplied by some factor, ΔH_{rxn} must be multiplied by the same factor and when a chemical equation is reversed, ΔH_{rxn} changes sign. So $\Delta H_{rxn} = -2(122 \text{ J}) = -144 \text{ J}$.