

## CHAPTER 10

# Energy

### CHAPTER ANSWERS

1. heat
2. potential
3. kinetic
4. The total energy of the universe is constant. Energy cannot be created or destroyed, but can only be converted from one form to another.
5. state
6. Ball A initially possesses potential energy by virtue of its position at the top of the hill. As Ball A rolls down the hill, its potential energy is converted to kinetic energy and frictional (heat) energy. When Ball A reaches the bottom of the hill and hits Ball B, it transfers its kinetic energy to Ball B. Ball A then has only the potential energy corresponding to its new position.
7. *Temperature* is a measure of the random motions of the particles in a substance, that is, temperature is a measure of the average kinetic energies of the particles. *Heat* is the energy that flows because of a temperature difference.
8. The hot tea is at a higher temperature, which means the particles in the hot tea have higher average kinetic energies. When the tea spills on the skin, energy flows from the hot tea to the skin, until the tea and skin are at the same temperature. This sudden inflow of energy causes the burn.
9. The thermal energy of an object represents the random motions of the particles of matter that constitute the object.
10. Temperature is the concept by which we express the thermal energy contained in a sample. We cannot measure the motions of the particles/kinetic energy in a sample of matter directly. We know, however, that if two objects are at different temperatures, the one with the higher temperature has molecules that have higher average kinetic energies than those in the object at the lower temperature.
11. The system is the part of the universe upon which we want to focus attention. In a chemical reaction, the system represents the reactants and products of the chemical reaction.
12. When the chemical system evolves energy, the energy evolved from the reacting chemicals is transferred to the surroundings.
13. An endothermic reaction is one that absorbs energy; therefore the products will have a higher potential energy than did the reactants.
14. exactly equal to
15. Thermodynamics is the study of energy and energy transfers. The first law of thermodynamics is the same as the law of conservation of energy, which is usually worded as "the energy of the universe is constant."
16. internal

17. changed  
18. losing  
19. positive  
20. gaining  
21. The calorie represents the amount of energy required to warm one gram of water by one Celsius degree. The "Calorie" or "nutritional" calorie represents 1000 calories (one kilocalorie). The Joule is the SI unit of energy, and 1 calorie is equivalent to 4.184 Joules. The Joule has a specific SI definition as the energy required to exert a force of 1 Newton over a distance of 1 m, but we tend to use the experimental definition above in terms of heating water.

22.

a.  $\frac{1 \text{ J}}{4.184 \text{ cal}}$

b.  $\frac{4.184 \text{ cal}}{1 \text{ J}}$

c.  $\frac{1 \text{ kcal}}{1000 \text{ cal}}$

d.  $\frac{1000 \text{ J}}{1 \text{ kJ}}$

23.  $2 \times 2.75 \text{ kcal} = 5.50 \text{ kcal}$  for twice the increase in temperature

24.  $6540 \text{ J} = 6.54 \text{ kJ}$  for ten times more water

25.

a.  $75.2 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 315 \text{ kJ} = 3.15 \times 10^3 \text{ J}$

b.  $75.2 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 315 \text{ J} = 0.315 \text{ kJ}$

c.  $1.41 \times 10^3 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 5.90 \times 10^3 \text{ J} = 5.90 \text{ kJ}$

d.  $1.41 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 5.90 \text{ kJ} = 5.90 \times 10^3 \text{ J}$

26.

a.  $7518 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 7.518 \text{ kcal}$

b.  $7.518 \times 10^3 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 7.518 \text{ kcal}$

c.  $1 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 0.001 \text{ kcal}$

$$d. \quad 655,200 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 655.2 \text{ kcal}$$

27.

$$a. \quad 491.4 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 117.4 \text{ kcal}$$

$$b. \quad 24.22 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 5.789 \text{ kcal}$$

$$c. \quad 81.01 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 19.36 \text{ kcal}$$

$$d. \quad 111.5 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 26.65 \text{ kcal}$$

28.

$$a. \quad 243,000 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 243 \text{ kJ to 3 significant figures}$$

$$b. \quad 4.184 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.004184 \text{ kJ}$$

$$c. \quad 0.251 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.000251 \text{ kJ}$$

$$d. \quad 450.3 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.4503 \text{ kJ}$$

29.

$$a. \quad 76.52 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.320 \text{ kJ}$$

$$b. \quad 7.824 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 7824 \text{ J}$$

$$c. \quad 52.99 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 2.22 \times 10^5 \text{ J}$$

$$d. \quad 221.4 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 5.29 \times 10^{-2} \text{ kcal}$$

30.

$$a. \quad 89.74 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 21.45 \text{ kcal}$$

$$b. \quad 1.756 \times 10^4 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 17.56 \text{ kJ}$$

$$c. \quad 1.756 \times 10^4 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 4.20 \text{ kcal}$$

$$d. \quad 1.00 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 239 \text{ cal}$$

$$31. \quad Q = s \times m \times \Delta T$$

$$69,500 \text{ J} = s \times (1012 \text{ g}) \times (11.4 \text{ }^\circ\text{C})$$

$$s = 6.02 \text{ J/g }^\circ\text{C}$$

$$32. \quad Q = s \times m \times \Delta T$$

$$\text{Temperature increase} = 41.5 - 27.2 = 14.3 \text{ }^\circ\text{C}$$

$$Q = (0.89 \text{ J/g }^\circ\text{C}) \times (29.2 \text{ g}) \times (14.3 \text{ }^\circ\text{C}) = 3.7 \times 10^2 \text{ J} = 89 \text{ cal}$$

$$33. \quad Q = s \times m \times \Delta T$$

Specific heat capacity of iron is  $0.45 \text{ J/g }^\circ\text{C}$

$$595 \text{ J} = (0.45 \text{ J/g }^\circ\text{C}) \times m \times (20 \text{ }^\circ\text{C})$$

$$m = 66 \text{ g}$$

$$34. \quad Q = s \times m \times \Delta T$$

Specific heat capacity of mercury is  $0.14 \text{ J/g }^\circ\text{C}$

$$100. \text{ J} = (0.14 \text{ J/g }^\circ\text{C}) \times 25 \text{ g} \times \Delta T$$

$$\Delta T = 28.6 \text{ }^\circ\text{C} = 29 \text{ }^\circ\text{C}$$

$$35. \quad Q = s \times m \times \Delta T$$

Specific heat capacity of gold is  $0.13 \text{ J/g }^\circ\text{C}$

$$Q = (0.13 \text{ J/g }^\circ\text{C}) \times (55.5 \text{ g}) \times (25 \text{ }^\circ\text{C}) = 180 \text{ J} = 1.8 \times 10^2 \text{ J}$$

$$36. \quad 0.24 \frac{\text{J}}{\text{g }^\circ\text{C}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 0.057 \text{ cal/g }^\circ\text{C}$$

$$37. \quad Q = s \times m \times \Delta T$$

specific heat capacities: Hg,  $0.14 \text{ J/g }^\circ\text{C}$ ; Fe,  $0.45 \text{ J/g }^\circ\text{C}$ ; C,  $0.71 \text{ J/g }^\circ\text{C}$

$$\text{for Hg: } Q = (0.14 \text{ J/g }^\circ\text{C}) \times (10.0 \text{ g}) \times (25 \text{ }^\circ\text{C}) = 35 \text{ J}$$

$$\text{for Fe: } Q = (0.45 \text{ J/g }^\circ\text{C}) \times (10.0 \text{ g}) \times (25 \text{ }^\circ\text{C}) = 113 \text{ J}$$

$$\text{for C: } Q = (0.71 \text{ J/g }^\circ\text{C}) \times (10.0 \text{ g}) \times (25 \text{ }^\circ\text{C}) = 178 \text{ J}$$

$$38. \quad Q = s \times m \times \Delta T$$

$$133 \text{ J} = s \times (5.00 \text{ g}) \times (55.1 - 25.2 \text{ }^\circ\text{C})$$

$$s = 0.89 \text{ J/g }^\circ\text{C}$$

The specific heat is  $0.89 \text{ J/g }^\circ\text{C}$ , so the element is most likely aluminum.

39. The enthalpy change is the heat that flows on a molar basis for a reaction at constant pressure.

40. A calorimeter is an insulated device in which reactions are performed and temperature changes measured, enabling the calculation of heat flows. (See Figure 10.6.)

41.

- a. exothermic (The sign of the enthalpy change is negative.)  
 b. released (The sign of the enthalpy change is negative.)

c. 
$$-285.8 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = -68.31 \text{ kcal/mol}$$

42.

- a. molar mass of S = 32.07 g

$$1.00 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.0312 \text{ mol S}$$

From the balanced chemical equation, combustion of 0.0312 mol S would produce 0.0312 mol SO<sub>2</sub>

$$0.0312 \text{ mol} \times \frac{-296 \text{ kJ}}{\text{mol}} = -9.23 \text{ kJ}$$

- b. From the balanced chemical equation, combustion of 0.0312 mol S would produce 0.0312 mol SO<sub>2</sub>

$$0.501 \text{ mol} \times \frac{-296 \text{ kJ}}{\text{mol}} = -148 \text{ kJ}$$

- c. The enthalpy change would be the same in magnitude, but opposite in sign = +296 kJ/mol.

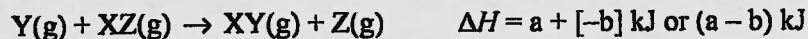
43.

- a. The equation as given is for 1 mol of Hg: +90.7 J.  
 b. The equation as given is for ½ mol O<sub>2</sub>; therefore 2 × +90.7 = +181.4 J would be required for 1 mol O<sub>2</sub>.  
 c. The required equation is *double the reverse* of the given equation; therefore the heat evolved would be 2 × -90.7 = -181.4 J.

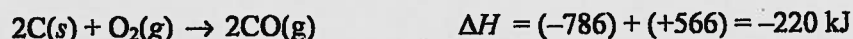
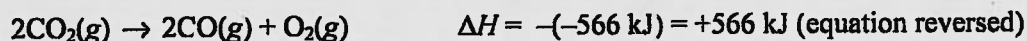
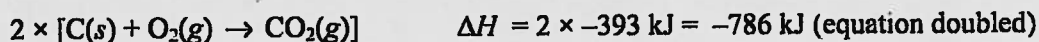
44. It is important when dealing with enthalpy changes to make sure you understand the *language* in which the number representing  $\Delta H$  is given. Sometimes the  $\Delta H$  value is given "for the equation as written," including all the stoichiometric coefficients of the balanced equation. At other times, the  $\Delta H$  value may be given on a per-mole basis for one specific reactant or product. In this question, the  $\Delta H$  value is given for the reaction as written, in which one mole of methane reacts with two moles of dioxygen to produce one mole of carbon dioxide and two moles of water.  $\Delta H$  would be -890 kJ for each mole of methane reaction, but the question asks about the quantities of heat evolved when one mole of dioxygen reacts or when one mole of water is produced.

- a. -445 kJ/mol of water (because 2 moles of water are formed in the reaction as given)  
 b. -445 kJ/mol of dioxygen (because 2 moles of dioxygen react in the equation as given)

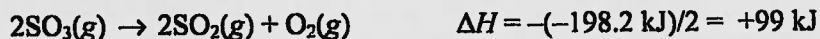
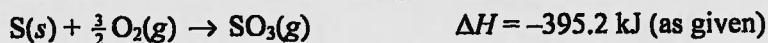
45. If the second equation is *reversed* and then added to the first equation, the desired equation can be generated:



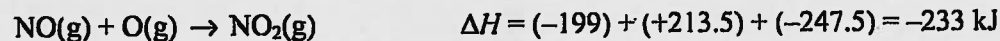
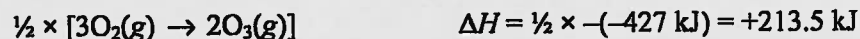
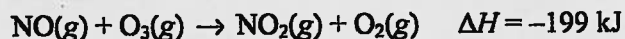
46. The desired equation,  $2\text{C(s)} + \text{O}_2(\text{g}) \rightarrow 2\text{CO(g)}$ , can be generated by taking twice the first equation and adding it to the reverse of the second equation:



47. The desired equation,  $\text{S(s)} + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$ , can be generated by reversing the second equation and dividing the equation by 2 and then adding this to the first equation:



48. The desired equation can be generated as follows:



49. The energy is converted to the energy of motion of the car and to friction as the car's tires interact with the road. Once the potential energy has been dispersed, it cannot be reused.
50. Once everything in the universe is at the same temperature, no further thermodynamic work can be done. Even though the total energy of the universe will be the same, the energy will have been dispersed evenly, making it effectively useless.
51. Petroleum is especially useful because it is a concentrated, easy to transport and use source of energy.
52. Concentrated sources of energy, such as petroleum, are being used to disperse the energy they contain, thus making it unavailable for further use.
53. These sources of energy originally came from living plants and animals that used their metabolic processes to store energy.
54. Petroleum consists mainly of hydrocarbons, which are molecules containing chains of carbon atoms with hydrogen atoms attached to the chains. The fractions are based on the number of carbon atoms in the chains. For example, gasoline is a mixture of hydrocarbons with five to ten carbon atoms in the chains whereas asphalt is a mixture of hydrocarbons with 25 or more carbon atoms in the chains. Different fractions have different physical properties and uses, but all can be combusted to produce energy. (See Table 10.3.)

55. Natural gas consists primarily of methane with small amounts of ethane, propane and butane. It is generally found in association with petroleum deposits.
56. Tetraethyl lead was used as an additive for gasoline to promote smoother running of engines. It is no longer widely used because of concerns about the lead being released to the environment as the leaded gasoline is burned.
57. Coal matures through four stages: lignite, sub-bituminous, bituminous, and anthracite. The four types of coal differ in the ratio of carbon to the other elements. Anthracite has the highest fraction of carbon in it and, when burned, releases a larger amount of heat for a given mass.
58. The greenhouse effect is a warming effect due to the presence of gases in the atmosphere that absorb infrared radiation that has reached the earth from the sun and do not allow it to pass back into space. A limited greenhouse effect is desirable because it moderates the temperature changes in the atmosphere that would otherwise be more drastic between daytime when the sun is shining and nighttime. Having too high a concentration of greenhouse gases, however, will elevate the temperature of the earth too much, affecting climate, crops, the polar ice caps, temperature of the oceans, and so on. Carbon dioxide produced by combustion reactions is our greatest concern as a greenhouse gas.
59. A driving force is some factor that tends to make a process occur.
60. If a proposed reaction involves either or both of those phenomena, the reaction will tend to be favorable.
61. an increase in entropy
62. Formation of a solid precipitate represents a concentration of matter.
63. Entropy is a measure of the randomness or disorder in a system. The entropy of the universe increases because the natural tendency is for things to become more disordered.
64. The molecules in liquid water are moving around freely and are therefore more "disordered" than when the molecules are held rigidly in a solid lattice in ice. The entropy increases during melting.
65. As the steam is cooled from 150 °C to 100 °C, the molecules of vapor gradually slow down as they lose kinetic energy. At 100 °C, the steam condenses into liquid water, and the temperature remains at 100 °C until all the steam has condensed. As the liquid water cools, the molecules in the liquid move more and more slowly as they lose kinetic energy. At 0 °C, the liquid water freezes.
- 66.
- $462.4 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 110.5 \text{ kcal}$
  - $18.28 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 4.369 \text{ kcal}$
  - $1.014 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 0.2424 \text{ kcal}$
  - $190.5 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 45.53 \text{ kcal}$

67.

$$a. \quad 45.62 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 190.9 \text{ kJ}$$

$$b. \quad 72.94 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 17.43 \text{ kcal}$$

$$c. \quad 2.751 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 657.5 \text{ cal}$$

$$d. \quad 5.721 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 2.394 \times 10^4 \text{ cal}$$

$$68. \quad \text{Temperature increase} = 75.0 - 22.3 = 52.7^\circ\text{C}$$

$$145 \text{ g} \times 4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 52.7^\circ\text{C} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 7641.5 \text{ cal} = 7.65 \text{ kcal}$$

$$69. \quad \text{Specific heat of silver} = 0.24 \text{ J/g}^\circ\text{C}; 1.25 \text{ kJ} = 1250 \text{ J}$$

$$\text{Temperature increase} = 15.2 - 12.0 = 3.2^\circ\text{C}$$

$$Q = s \times m \times \Delta T$$

$$1250 \text{ J} = (0.24 \text{ J/g}^\circ\text{C}) \times (\text{mass of silver}) \times (3.2^\circ\text{C})$$

$$\text{mass of silver} = 1627 \text{ g} = 1.6 \times 10^3 \text{ g}$$

70. The specific heat capacity of iron is  $0.45 \text{ J/g}^\circ\text{C}$ . Then 50. joules is the heat that is applied to the sample of iron and must equal the product of the mass of iron, the specific heat capacity of the iron, and the temperature change undergone by the iron (which is what we want).

$$Q = s \times m \times \Delta T$$

$$50. \text{ J} = (0.45 \text{ J/g}^\circ\text{C})(10. \text{ g})(\Delta T)$$

$$\Delta T = 11^\circ\text{C}$$

$$71. \quad 0.13 \frac{\text{J}}{\text{g}^\circ\text{C}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 0.31 \frac{\text{cal}}{\text{g}^\circ\text{C}}$$

$$72. \quad 2.5 \text{ kg water} = 2500 \text{ g}$$

$$\text{Temperature change} = 55.0 - 18.5 = 36.5^\circ\text{C}$$

$$Q = s \times m \times \Delta T$$

$$Q = 4.184 \text{ J/g}^\circ\text{C} \times 2500 \text{ g} \times 36.5^\circ\text{C} = 3.8 \times 10^5 \text{ J}$$

73. For a given mass of substance, the substance with the *smallest* specific heat capacity (gold,  $0.13 \text{ J/g}^\circ\text{C}$ ) will undergo the *largest* increase in temperature. Conversely, the substance with the largest specific heat capacity (water,  $4.184 \text{ J/g}^\circ\text{C}$ ) will undergo the smallest increase in temperature.



74. Let  $T_f$  represent the final temperature reached by the system.

$$\text{For the hot water, heat lost} = 50.0 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (100. - T_f ^\circ\text{C})$$

$$\text{For the cold water, heat gained} = 50.0 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (T_f - 25 ^\circ\text{C})$$

The heat lost by the hot water must *equal* the heat gained by the cold water; therefore

$$50.0 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (100. - T_f ^\circ\text{C}) = 50.0 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (T_f - 25 ^\circ\text{C})$$

$$\text{Solving this equation for } T_f \text{ gives } T_f = 62.5 ^\circ\text{C} = 63 ^\circ\text{C}$$

75. Let  $T_f$  be the final temperature reached.

$$\text{Heat gained by water} = 75 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (T_f - 20 ^\circ\text{C})$$

$$\text{Heat lost by iron} = 25.0 \text{ g} \times 0.45 \text{ J/g } ^\circ\text{C} \times (85 - T_f ^\circ\text{C})$$

The heat lost by the iron must equal the heat gained by the water.

$$75 \text{ g} \times 4.184 \text{ J/g } ^\circ\text{C} \times (T_f - 20 ^\circ\text{C}) = 25.0 \text{ g} \times 0.45 \text{ J/g } ^\circ\text{C} \times (85 - T_f ^\circ\text{C})$$

$$\text{Solving for } T_f \text{ gives } T_f = 22.3 ^\circ\text{C} = 22 ^\circ\text{C}$$

76. 9.0 J (It requires twice as much heat to warm a sample of twice the mass over the same temperature interval.)
77. For any substance,  $Q = s \times m \times \Delta T$ . The basic calculation for each of the substances is the same: Heat required =  $150. \text{ g} \times (\text{specific heat capacity}) \times 11.2 ^\circ\text{C}$

Substance	Specific Heat Capacity	Heat Required
water (l)	4.184 J/g °C	$7.03 \times 10^3 \text{ J}$
water (s)	2.03 J/g °C	$3.41 \times 10^3 \text{ J}$
water (g)	2.0 J/g °C	$3.4 \times 10^3 \text{ J}$
aluminum	0.89 J/g °C	$1.5 \times 10^3 \text{ J}$
iron	0.45 J/g °C	$7.6 \times 10^2 \text{ J}$
mercury	0.14 J/g °C	$2.4 \times 10^2 \text{ J}$
carbon	0.71 J/g °C	$1.2 \times 10^3 \text{ J}$
silver	0.24 J/g °C	$4.0 \times 10^2 \text{ J}$
gold	0.13 J/g °C	$2.2 \times 10^2 \text{ J}$

78. Because, for any substance,  $Q = s \times m \times \Delta T$ , we can solve this equation for the temperature change,  $\Delta T$ . The results are tabulated:

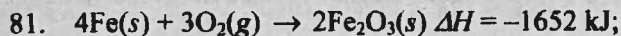
Substance	Specific Heat Capacity	Temperature Change
water (l)	4.184 J/g °C	23.9 °C
water (s)	2.03 J/g °C	49.3 °C
water (g)	2.0 J/g °C	50. °C
aluminum	0.89 J/g °C	$1.1 \times 10^2 \text{ } ^\circ\text{C}$
iron	0.45 J/g °C	$2.2 \times 10^2 \text{ } ^\circ\text{C}$
mercury	0.14 J/g °C	$7.1 \times 10^2 \text{ } ^\circ\text{C}$
carbon	0.71 J/g °C	$1.4 \times 10^2 \text{ } ^\circ\text{C}$
silver	0.24 J/g °C	$4.2 \times 10^2 \text{ } ^\circ\text{C}$
gold	0.13 J/g °C	$7.7 \times 10^2 \text{ } ^\circ\text{C}$

79.

- a.  $\Delta E = q + w = -47 \text{ kJ} + 88 \text{ kJ} = 41 \text{ kJ}$   
 b.  $\Delta E = 82 + 47 = 129 \text{ kJ}$   
 c.  $\Delta E = 47 + 0 = 47 \text{ kJ}$   
 d. When the surroundings deliver work to the system,  $w > 0$ . This is the case for a and b.

80.

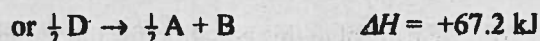
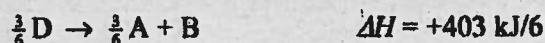
- a. The combustion of gasoline releases heat, so this is an exothermic process.  
 b.  $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l)$ ; Heat is released when water vapor condenses, so this is an exothermic process.  
 c. To convert a solid to a gas, heat must be absorbed, so this is an endothermic process.  
 d. Heat must be added (absorbed) in order to break a bond, so this is an endothermic process.



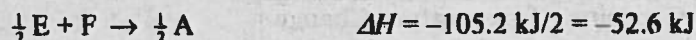
Note that 1652 kJ of heat are released when 4 mol Fe reacts with 3 mol  $\text{O}_2$  to produce 2 mol  $\text{Fe}_2\text{O}_3$ .

- a.  $4.00 \text{ mol Fe} \times \frac{-1652 \text{ kJ}}{4 \text{ mol Fe}} = -1652 \text{ kJ}$   
 b.  $1.00 \text{ mol Fe}_2\text{O}_3 \times \frac{-1652 \text{ kJ}}{2 \text{ mol Fe}_2\text{O}_3} = -826 \text{ kJ}$   
 c.  $1.00 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{-1652 \text{ kJ}}{4 \text{ mol Fe}} = -7.39 \text{ kJ}$   
 d.  $10.0 \text{ g Fe} = 0.179 \text{ mol Fe}$ ;  $2.00 \text{ g O}_2 = 0.0625 \text{ mol O}_2$ ;  $\text{O}_2$  is the limiting reactant.  
 $0.0625 \text{ mol O}_2 \times \frac{-1652 \text{ kJ}}{3 \text{ mol O}_2} = -34.4 \text{ kJ heat released.}$

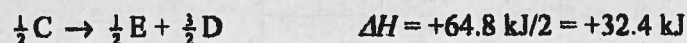
82. Reversing the first equation and dividing by 6 we get



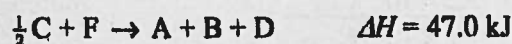
Dividing the second equation by 2 we get



Dividing the third equation by 2 we get



Adding these equations together we get



83. During exercise, the body generates about 5500 kJ/hr, or about 11,000 kJ in 2 h. Assuming all of this heat is lost through evaporation, we can calculate the volume of the perspiration. Water has a heat of vaporization of 40.7 kJ/mol. So,

$$11,000 \text{ kJ} \times \frac{1 \text{ mol H}_2\text{O}}{40.7 \text{ kJ}} = 270 \text{ mol H}_2\text{O}$$

$$270 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 4900 \text{ g H}_2\text{O}$$

Assuming a density of 1 g/mL for water, 4900 g of H<sub>2</sub>O would occupy a volume of 4900 mL or 4.9 L.

84. If 400 kcal is burned per hour walking 4.0 mph, then the total amount of heat burned while walking at 4.0 mph will be given by

$$Q = 400 \text{ kcal/hr} \times t$$

where  $t$  is the amount of time spent walking (in hours). One gram of fat is consumed for every 7.7 kcal of heat. Therefore, one pound of fat requires

$$7.7 \text{ kcal/g} \times 454 \text{ g/lb} = 3500 \text{ kcal/lb}$$

Since we want to lose one pound,  $Q = 3500 \text{ kcal}$  and

$$t = \frac{Q}{400 \text{ kcal/hr}} = \frac{3500 \text{ kcal}}{400 \text{ kcal/hr}} = 8.75 \text{ hr} = 9 \text{ hr}$$