

Section Review Answers

Student Textbook, page 261

- Students should ask the following questions: What definition of efficiency are you using? What is your definition of “useful energy”? What factors are included in your definition of “energy used”?
- Adding steps to remove polluting compounds from effluent such as coal smoke require an input of energy. Therefore, they reduce the overall efficiency of the process.
- A simple experiment involves heating water in a beaker. Students will need to determine the flow rate of the gas, and hence the quantity of gas burned. They should compare ideal enthalpy change for the burning to the amount of heat actually used to increase the temperature of the water, say, by 30°C. An experiment like this would give the efficiency of a laboratory burner in performing the specific task of heating water in a beaker. The efficiency definition is based on the ratio between the actual energy used to heat the water as compared to the ideal quantity of heat released by burning the gas.
- Under ordinary conditions, the enthalpy of combustion of methane is $\Delta H^\circ_{\text{comb}} = -802 \text{ kJ/mol}$. However, if the gaseous water produced is condensed, further heat is released, as shown in the following equations:

$$\text{CH}_{4(g)} + 2\text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2\text{H}_2\text{O}_{(g)} \quad \Delta H^\circ_{\text{comb}} = -802 \text{ kJ}$$

$$2\text{H}_2\text{O}_{(g)} \rightarrow 2\text{H}_2\text{O}_{(\ell)} \quad \Delta H^\circ_{\text{comb}} = 2(-44 \text{ kJ})$$

$$\text{CH}_{4(g)} + 2\text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2\text{H}_2\text{O}_{(g)} \quad \Delta H^\circ_{\text{comb}} = -890 \text{ kJ}$$
- Students defending nuclear energy over coal will probably mention the greenhouse effect and acid precipitation. Students defending coal will probably suggest that dealing with radioactive waste products and warm water pollution are more significant problems.
- The manufacturer is probably using a definition that compares electrical energy used by the device to thermal energy used directly to heat the water in the kettle.
 - $$\text{efficiency} = \frac{\text{heat absorbed to boil water}}{\text{electrical energy used to produce heat}} \times 100\%$$
 - Students should outline an experiment that monitors the electrical energy used to raise a measured volume of water in the kettle to 100°C, comparing it to the calculated enthalpy change based on the specific heat capacity of water. Safety precautions should include care using electricity and water.
- Students will probably suggest that using methane hydrates is less efficient than using methane from natural gas, since it currently is more difficult, and thus requires more energy, to extract methane from methane hydrates.
- Adding the three equations together, without manipulation, gives the following equation:

$$\text{H}_2\text{O}_{(g)} \rightarrow \text{H}_2(g) + \frac{1}{2} \text{O}_{2(g)}$$
 - The enthalpy change for the reaction is:

$$318 \text{ kJ} - 249 \text{ kJ} + 173 \text{ kJ} = 242 \text{ kJ}$$
 The enthalpy of formation of gaseous water is -242 kJ/mol .

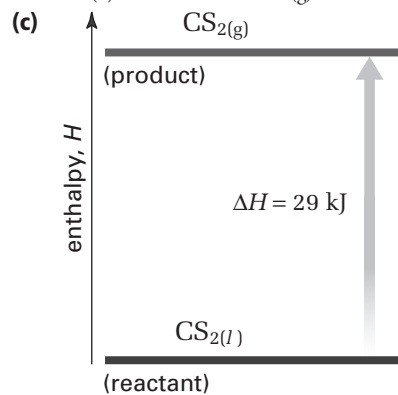
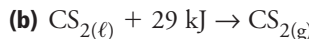
Chapter 5 Review Answers

Knowledge/Understanding

- There is a direct relationship between the system and the surroundings as summarized by the equation:

$$E_{\text{universe}} = E_{\text{system}} + E_{\text{surroundings}}$$
 The system is the immediate part of the universe being studied or observed (e.g., reactants in a test tube). The surroundings are everything else in the universe. Any change in the energy of a system affects the energy of the surroundings but the effects are directly observed on its immediate surroundings rather than the wider surroundings. For example, if a reaction occurs between the reactants in the test tube, and the reaction releases energy, that energy will be transferred as heat to increase the energy of the test tube itself, the surrounding air, and the test tube support.

- The reaction is endothermic. $\Delta H^\circ_{\text{vap}} = +29 \text{ kJ/mol}$.



3. If the enthalpy change for the original equation is ΔH , then the enthalpy change for the manipulated equation will be $-3 \times \Delta H$.
4. A good calorimeter should isolate the system from the surroundings. Two nested polystyrene coffee cups with a lid make a good constant pressure calorimeter because polystyrene is an excellent insulator and the layer of air between the cups further helps to insulate the container. The container is not particularly strong, however this is not a problem because the calorimeter need not withstand high pressures.
5. (a) $\text{Li}_{(s)} + \frac{1}{2} \text{Cl}_{2(g)} \rightarrow \text{LiCl}_{(s)}$
 (b) $\text{C}_2\text{H}_4(g) + \text{H}_2\text{O}(g) \rightarrow \text{C}_2\text{H}_5\text{OH}(\ell)$
 (c) $\text{NH}_3(g) + \text{HNO}_3(\ell) \rightarrow \text{NH}_4\text{NO}_3(s)$
6. In its standard state, iodine is a solid. For 1 mol of $\text{I}_{2(s)}$ to sublime and form $\text{I}_{2(g)}$, it needs to absorb 21 kJ of energy.

Inquiry

7. Step 1: Determine ΔH for the combustion carbon in excess oxygen to form carbon dioxide, per equation (1), shown below.
- $$(1) \text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} \quad \Delta H_1$$
- Step 2: Determine ΔH for the combustion of $\text{CO}_{(g)}$ in excess oxygen to form $\text{CO}_{2(g)}$, per equation (2), shown below.
- $$(2) \text{CO}_{(g)} + \frac{1}{2} \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} \quad \Delta H_2$$
- Step 3: Add equation (1) and the reverse of equation (2) to obtain the target equation: the formation of carbon monoxide. Perform the same operations on the enthalpies of reaction.
- $$[(1) - (2)] \text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{(g)}$$
- $$(\Delta H_1 - \Delta H_2) = \Delta H_f(\text{CO}_{(g)})$$
8. (a) $100.0^\circ\text{C} + \frac{20.0^\circ\text{C}}{2} = 60.0^\circ\text{C}$
 (b) $(50.0 \text{ g})(20.0^\circ\text{C}) + (30.0 \text{ g})(100.0^\circ\text{C}) = 80.0 \text{ g} \times 50.0^\circ\text{C}$
9. $\Delta H^\circ_{\text{rxn}}$
 $= 12\Delta H^\circ_f(\text{CO}_{2(g)}) + 11\Delta H^\circ_f(\text{H}_2\text{O}(\ell))$
 $- [\Delta H^\circ_f(\text{C}_{12}\text{H}_{22}\text{O}_{11(s)}) + 12\Delta H^\circ_f(\text{O}_{2(g)})]$
 $\Delta H^\circ_f(\text{C}_{12}\text{H}_{22}\text{O}_{11(s)})$
 $= -\Delta H^\circ_{\text{rxn}} + 12\Delta H^\circ_f(\text{CO}_{2(g)}) + 11\Delta H^\circ_f(\text{H}_2\text{O}(\ell))$
 $- 12\Delta H^\circ_f(\text{O}_{2(g)})$
 $= 5641 \text{ kJ/mol} + 12(-393.5 \text{ kJ/mol})$
 $+ 11(-285.8 \text{ kJ/mol}) - 12(0 \text{ kJ/mol})$
 $= -2225 \text{ kJ/mol}$

10. Use $Q = m \times c \times \Delta T$, where
 $m = 2000 \text{ g}$ (2000 mL of water, $d = 1.00 \text{ g/ml}$)
 $c = 4.184 \text{ J/g}\cdot^\circ\text{C}$
 $\Delta T = 39.6^\circ\text{C} - 22.3^\circ\text{C} = 17.3^\circ\text{C}$
 $Q_{\text{surroundings}} = 2000 \text{ g} \times 4.184 \text{ J/g}\cdot^\circ\text{C} \times 17.3^\circ\text{C}$
 $= 1.45 \times 10^5 \text{ J} = 1.45 \times 10^2 \text{ kJ}$
 $\Delta H_{\text{reaction}} = Q_{\text{system}} = -Q_{\text{surroundings}} = -145 \text{ kJ}$
 Since 10.0 g of acetic acid was used, in terms of kJ/g acetic acid, the enthalpy of combustion of acetic acid is:
 -14.5 kJ/g
 Molar mass of acetic acid = 60.06 g/mol
 $n \text{ mol acetic acid} = 10.0 \text{ g} \div 60.06 \text{ g/mol} = 0.167 \text{ mol}$
 $\Delta H_{\text{comb}}(\text{acetic acid}) = -145 \text{ kJ} / 0.167 \text{ mol}$
 $= 1.24 \times 10^3 \text{ kJ/mol}$
11. Equation (1) should be reversed and multiplied by 3. Equation (2) should be left as is. Equation (3) should be reversed. When these operations are carried out, the resulting equations add up to the target equation. Therefore:
 $\Delta H_{\text{reaction}} = -3(-92.3 \text{ kJ/mol}) - 74.8 \text{ kJ/mol}$
 $- (-134.5 \text{ kJ/mol})$
 $= +336.6 \text{ kJ/mol}$
12. $\Delta H^\circ_{\text{rxn}}$
 $= 12\Delta H^\circ_f(\text{CO}_{2(g)}) + 3\Delta H^\circ_f(\text{H}_2\text{O}(\ell))]$
 $- [\Delta H^\circ_f((\text{CH}_2\text{OH})_{2(s)}) + \frac{5}{2}\Delta H^\circ_f(\text{O}_{2(g)})]$
 $\Delta H^\circ_f((\text{CH}_2\text{OH})_{2(s)})$
 $= -\Delta H^\circ_{\text{rxn}} + 2\Delta H^\circ_f(\text{CO}_{2(g)})$
 $+ 3\Delta H^\circ_f(\text{H}_2\text{O}(\ell)) - \frac{5}{2}\Delta H^\circ_f(\text{O}_{2(g)})$
 $= -(-1178 \text{ kJ/mol}) + 2(-393.5 \text{ kJ/mol})$
 $+ 3(-285.8 \text{ kJ/mol}) - \frac{5}{2}(0 \text{ kJ/mol})$
 $= -466 \text{ kJ/mol}$
13. Molar mass of $\text{H}_2\text{S}_{(g)} = 34.06 \text{ g/mol}$
 $n \text{ mol H}_2\text{S}_{(g)} = \frac{15.0 \text{ g}}{34.09 \text{ g/mol}}$
 $= 0.440 \text{ mol}$
 $\Delta H = 0.440 \text{ mol H}_2\text{S}_{(g)} \times -519 \text{ kJ/mol}$
 $= -228 \text{ kJ}$
 Therefore, 228 kJ of energy is released.

14. (a) $\text{H}_{2(\text{g})} + \text{O}_{2(\text{g})} \rightarrow \text{H}_2\text{O}_{2(\ell)}$
 (b) Divide equation (1) by 2 and reverse its direction. Leave equation (2) as is. The equations now add up to the target equation (shown in part (a)). Therefore, you can calculate the enthalpy of formation of liquid hydrogen peroxide as follows:
- $$\begin{aligned}\Delta H^\circ_f &= -\frac{1}{2}(-196 \text{ kJ/mol}) - 286 \text{ kJ/mol} \\ &= -188 \text{ kJ/mol}\end{aligned}$$
15. Divide equation (1) by two. Divide equation (2) by two and reverse its direction. Reverse the direction of equation (3). The equations now add up to the target equation. Therefore, you can calculate the enthalpy change of the target equation as follows:
- $$\begin{aligned}\Delta H^\circ &= \frac{1}{2}(270 \text{ kJ}) - \frac{1}{2}(-92 \text{ kJ}) - (-75 \text{ kJ}) \\ &= 256 \text{ kJ}\end{aligned}$$
16. (a) $\Delta H^\circ_{\text{comb}} = 5(-241.8 \text{ kJ/mol}) + 4(-393.5 \text{ kJ/mol}) - (-126 \text{ kJ/mol}) = -2657 \text{ kJ/mol}$
 (b) $\text{C}_2\text{H}_{6(\text{g})} + 3.5\text{O}_{2(\text{g})} \rightarrow 3\text{H}_2\text{O}_{(\text{g})} + 2\text{CO}_{2(\text{g})}$

$$\begin{aligned}\Delta H^\circ_{\text{comb}} &= 3(-241.8 \text{ kJ/mol}) + 2(-393.5 \text{ kJ/mol}) \\ &\quad - (-84.0 \text{ kJ/mol}) \\ &= -1428.4 \text{ kJ/mol}\end{aligned}$$

 (c) The sample contains 3.0 g C_2H_6 and 7.0 g C_4H_{10} .
 For C_2H_6 :

$$\Delta H^\circ = \frac{3.0 \text{ g}}{30.08 \text{ g/mol}} \times -1428.4 \text{ kJ/mol} = 1.4 \times 10^2 \text{ kJ}$$

 For C_4H_{10} :

$$\Delta H^\circ = \frac{7.0 \text{ g}}{58.14 \text{ g/mol}} \times -2657 \text{ kJ/mol} = -3.2 \times 10^2 \text{ kJ}$$

 Therefore, the total amount of energy released is $4.6 \times 10^2 \text{ kJ}$.
17. You may want to suggest to students that they search on the Internet for ideas—numerous methods can be found online. Use the keywords pop can (or soda can, for U.S. sites), calorimeter, and food. Students should use the pop can to contain water, and design a device to suspend the cheese or cracker under the pop can (a paper clip works). The heat produced by burning the food raises the temperature of the water in the pop can. A second, larger can provides support for the food suspension device and channels hot air towards the pop can. Ensure students include safety precautions that recognize the dangers of fire and working with sharp metal. Note that this is a constant pressure calorimeter, not a bomb-type calorimeter.

18. Accept all reasonable procedures. Students will probably use a measured mass of ethanol in the lamp to heat water and determine the enthalpy change based on the temperature change of the water. Students should state that they would assume all heat from the lamp heats the water, not the surrounding air or the water container. They will need to assume that the fuel is pure ethanol. They will also need to assume that the only reaction that takes place is complete combustion (although this is probably not the case).

Communication

19. There are two ways to determine the enthalpy change of the reaction:
- The student could find reactions that have equations adding up to the target equations. If the enthalpies of reaction of those reactions can be determined in the laboratory, the student can use them to determine the enthalpy of reaction of the target equation.
 - The student could use the standard molar enthalpies of formation for the substances in the target equation to determine the enthalpy of reaction for the target equation. (The student could find the values in a table such as the one found in Appendix E).
20. (a) Since ΔH° for the equation is negative, ΔH°_f for C_2H_2 must have either a positive value or a negative value that is smaller than ΔH°_f for C_2H_4 . (ΔH°_f for $\text{H}_{2(\text{g})}$ is 0 kJ/mol.)
 (b) The combustion of 1 mol C_2H_2 produces 1 mol H_2O and 1 mol CO_2 . The combustion of 1 mol C_2H_4 produces 2 mol H_2O and 1 mol CO_2 . There is a difference of 175 kJ between the enthalph of formation of C_2H_2 and C_2H_4 , however, the extra 1 mol of water produced by the combustion of C_2H_4 means a difference of -285.9 kJ/mol . Therefore, per mol of fuel, C_2H_4 provides more energy.

Making Connections

21. (a) When the aerosol is sprayed into the car, the liquid evaporates. When the liquid evaporates, it absorbs heat from the air, and the overall temperature in the car drops.

