

$$= \frac{\frac{8.57 \text{ kJ}}{^{\circ}\text{C}} \times 5.28^{\circ}\text{C}}{1.14 \text{ g}}$$

$$h_c = 39.7 \text{ kJ/g}$$

The exothermic specific heat of combustion of stearic acid is 39.7 kJ/g.

- (b) The energy release from burning a gram of fat is more than double that of a gram of sucrose. The specific heat of fat is  $39.7/16.5 = 2.4$  times as great as the specific heat of sucrose.

### Making Connections

- Student answers will vary widely, but will indicate clearly that fats and oils are very high in energy, while carbohydrates and proteins are lower.
- (a) The energy characteristic of gasoline that we use is its exothermic heat of combustion — converting the gasoline to heat and mechanical energy to move the vehicle.  
(b) Things like the speed with which the gasoline burns, the degree of pollution it produces, and the toxicity might also be important characteristics.

## 12.3 HEATS OF REACTION

### PRACTICE

(Page 585)

### Understanding Concepts

- Calorimetry is used to determine heats of reaction.
- The term *specific* for a scientific quantity means “per unit of mass.”
- Specific heat of reaction has units such as kJ/g. Molar heat of reaction has units such as kJ/mol.
- Molar mass will convert specific heat of reaction to molar heat of reaction.
- Combustion reactions are exothermic, such as the burning of gasoline, propane, natural gas, and candle wax.
- (a)  $h_c = 49.90 \text{ kJ/g}$   
 $M = 26.04 \text{ g/mol}$   
 $\Delta H_c = ?$  (combustion of acetylene,  $\text{C}_2\text{H}_{2(g)}$ )

$$\Delta H_c = \frac{49.90 \text{ kJ}}{1 \text{ g}} \times \frac{26.04 \text{ g}}{1 \text{ mol}}$$

$$\Delta H_c = 1299 \text{ kJ/mol}$$

The exothermic molar heat of combustion of acetylene is 1299 kJ/mol.

- (b)  $\text{C}_2\text{H}_{2(g)} + 3/2 \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + \text{H}_2\text{O}_{(g)} + 1299 \text{ kJ}$   
 or  $2 \text{C}_2\text{H}_{2(g)} + 3 \text{O}_{2(g)} \rightarrow 2 \text{CO}_{2(g)} + 2 \text{H}_2\text{O}_{(g)} + 2598 \text{ kJ}$
7. (a)  $h_r = 10.9 \text{ kJ/g}$

$$M = 12.01 \text{ g/mol}$$

$$\Delta H_r = ? \quad (\text{reaction of carbon, } \text{C}_{(s)})$$

$$\Delta H_r = \frac{10.9 \text{ kJ}}{1 \text{ g}} \times \frac{12.01 \text{ g}}{1 \text{ mol}}$$

$$\Delta H_r = 131 \text{ kJ/mol}$$

The exothermic molar heat of reaction of carbon is 131 kJ/mol.

- (b)  $\text{C}_{(s)} + \text{H}_2\text{O}_{(g)} \rightarrow \text{CO}_{(g)} + \text{H}_{2(g)} + 131 \text{ kJ}$

8. (a)  $\text{CO}_{(\text{g})} + 1/2 \text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{g})} + 283 \text{ kJ}$   
 (b)  $\Delta H_{\text{c}} = \frac{566 \text{ kJ}}{2 \text{ mol}}$  (combustion of carbon monoxide,  $\text{CO}_{(\text{g})}$ )  
 $\Delta H_{\text{c}} = 283 \text{ kJ/mol}$

The exothermic molar heat of combustion of carbon monoxide is 283 kJ/mol.

9. (a)  $\text{CH}_3\text{OH}_{(\text{l})} + 3/2 \text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})} + 638.0 \text{ kJ}$   
 or  $2 \text{CH}_3\text{OH}_{(\text{l})} + 3 \text{O}_{2(\text{g})} \rightarrow 2 \text{CO}_{2(\text{g})} + 4 \text{H}_2\text{O}_{(\text{g})} + 1276 \text{ kJ}$   
 b)  $\text{C}_{(\text{s})} + 1/4 \text{S}_{8(\text{s})} + 89.0 \text{ kJ} \rightarrow \text{CS}_{2(\text{l})}$   
 or  $4 \text{C}_{(\text{s})} + \text{S}_{8(\text{s})} + 356 \text{ kJ} \rightarrow 4 \text{CS}_{2(\text{l})}$   
 (c)  $\text{ZnS}_{(\text{s})} + 3/2 \text{O}_{2(\text{g})} \rightarrow \text{ZnO}_{(\text{s})} + \text{SO}_{2(\text{g})} + 441.3 \text{ kJ}$   
 or  $2 \text{ZnS}_{(\text{s})} + 3 \text{O}_{2(\text{g})} \rightarrow 2 \text{ZnO}_{(\text{s})} + 2 \text{SO}_{2(\text{g})} + 882.6 \text{ kJ}$   
 (d)  $\text{Fe}_2\text{O}_{3(\text{s})} + 824.2 \text{ kJ} \rightarrow 2 \text{Fe}_{(\text{s})} + 3/2 \text{O}_{2(\text{g})}$   
 or  $2 \text{Fe}_2\text{O}_{3(\text{s})} + 1648 \text{ kJ} \rightarrow 4 \text{Fe}_{(\text{s})} + 3 \text{O}_{2(\text{g})}$

10.  $h_{\text{c}} = 48.70 \text{ kJ/g}$

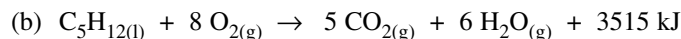
$M = 72.17 \text{ g/mol}$

$\Delta H_{\text{c}} = ?$  (combustion of pentane,  $\text{C}_5\text{H}_{12(\text{l})}$ )

$$\Delta H_{\text{c}} = \frac{48.70 \text{ kJ}}{1 \text{ g}} \times \frac{72.17 \text{ g}}{1 \text{ mol}}$$

$$\Delta H_{\text{c}} = 3515 \text{ kJ/mol}$$

The exothermic molar heat of combustion of pentane is 3515 kJ/mol.



## Reflecting

11. Exothermic and endothermic can be applied as terms to any situation involving energy change. They can apply to substances warming or cooling, or phase changes like ice melting, or nuclear changes.

## PRACTICE

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## Understanding Concepts

12. Bond energy refers to the energy change when a bond is formed, or broken. For example, when bonds in methane and oxygen molecules are broken, new bonds can form between carbon and oxygen atoms, and between hydrogen and oxygen atoms. The energy released from the formation of the new bonds in this case is much more than the energy absorbed to break the original bonds, so this change is exothermic, overall.
13. (a)  $\text{H}_{2(\text{g})} + \text{Br}_{2(\text{g})} \rightarrow 2 \text{HBr}_{(\text{g})}$   
 (b) H—H and Br—Br covalent bonds must be broken, and H—Br bonds must form.  
 (c) The reaction must have lost energy, so it is exothermic.  
 (d) The energy released from the formation of the new bonds in this case is more than the energy absorbed to break the original bonds.
14. (a) The surroundings (the water, its container ...) decrease in temperature.  
 (b)  $\text{NaCl}_{(\text{s})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$   
 (c) Ionic bonds are broken in this dissolving process.  
 (d) Dipole-dipole and hydrogen bonds form to produce the aqueous ions.  
 (e) The energy released from the formation of the new bonds in this case is less than the energy absorbed to break the original bonds.

## Reflecting

15. Effort must be expended to pull a magnet off a refrigerator door. This can be thought of as the breaking of a bond (in this case, magnetic), requiring a loss of energy for the surroundings (your hand) and a gain by the system (the magnet and refrigerator). When the magnet attaches to the door, energy leaves the system and transfers to your hand. ( It “pulls” your hand.)

## SECTION 12.3 QUESTIONS

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### Understanding Concepts

1. (a)  $h_c = 44.72 \text{ kJ/g}$

$$M = 86.20 \text{ g/mol}$$

$$\Delta H_c = ? \text{ (combustion of hexane, } C_6H_{14(l)})$$

$$\Delta H_c = \frac{44.72 \text{ kJ}}{1 \text{ g}} \times \frac{86.20 \text{ g}}{1 \text{ mol}}$$

$$\Delta H_c = 3855 \text{ kJ/mol}$$

The exothermic molar heat of combustion of hexane is 3855 kJ/mol.

- (b)  $C_6H_{14(l)} + 19/2 O_{2(g)} \rightarrow 6 CO_{2(g)} + 7 H_2O_{(g)} + 3855 \text{ kJ}$
- (c)  $2 C_6H_{14(l)} + 19 O_{2(g)} \rightarrow 12 CO_{2(g)} + 14 H_2O_{(g)} + 7710 \text{ kJ}$
- (d) The energy released from the formation of the new bonds in this combustion reaction is much greater than the energy absorbed to break the original bonds, so this chemical change is exothermic, overall.
2. (a)  $C_3H_{8(g)} + 5 O_{2(g)} \rightarrow 3 CO_{2(g)} + 4 H_2O_{(g)} + 2.04 \text{ MJ}$
- (b)  $1/2 N_{2(g)} + 1/2 O_{2(g)} + 90.2 \text{ kJ} \rightarrow NO_{(g)}$
- (c)  $C_2H_5OH_{(l)} + 3 O_{2(g)} \rightarrow 2 CO_{2(g)} + 3 H_2O_{(g)} + 1.28 \text{ MJ}$
3. (a)  $\Delta H_c = \frac{483.6 \text{ kJ}}{2 \text{ mol}}$
- $$\Delta H_c = 241.8 \text{ kJ/mol}$$
- The molar heat of combustion of hydrogen,  $H_{2(g)}$ , is 241.8 kJ/mol.
- (b)  $\Delta H_c = \frac{1134.4 \text{ kJ}}{4 \text{ mol}}$
- $$\Delta H_c = 283.6 \text{ kJ/mol}$$
- The molar heat of combustion of ammonia,  $NH_{3(g)}$ , is 283.6 kJ/mol.
4. (a)  $H_{2(g)} + I_{2(s)} \rightarrow 2 HI_{(g)}$
- H—H and I—I covalent bonds must be broken, and H—I bonds form.
- Note:* If the physical state of the iodine is considered to be that of the pure solid at SATP, then students can correctly assume that the (London force) intermolecular bonds must be broken as well.
- (b) The energy released from the formation of the new bonds in this case is less than the energy absorbed to break the original bonds, because the reaction is endothermic.
5. (a) Vaporization of propane is endothermic. Energy transfers from the surroundings to the propane.
- (b)  $C_3H_{8(l)} + \text{energy} \rightarrow C_3H_{8(g)}$

### Question

- (c) What is the specific heat of vaporization of propane?

### Experimental Design (many designs are possible — this is a suggestion)

The mass of propane in a small tank (independent variable) is allowed to vaporize while the tank is surrounded with a fixed mass (controlled variable) of water. The temperature change of the water (dependent variable) is measured and the endothermic specific heat of vaporization of propane calculated. Heat transfer from the calorimeter materials is ignored (assumed negligible).

### Materials

- small propane tank and valve (for propane torch or camping stove)
- styrofoam camping cooler and cover
- water
- large measuring cup
- thermometer
- long-handled wooden spoon for stirring
- electronic or triple-beam lab balance

### Procedure

1. This procedure releases a flammable gas that is denser than air. It **MUST** be done in a fume hood, well away from any source of spark or flame. Wear lab safety eye protection.
2. Place a valve on the propane tank, and measure and record the initial mass to the nearest 0.1 g.
3. Place the propane tank in the Styrofoam cooler.
4. Add measured volumes of water to the cooler until the propane tank is completely submerged. The propane tank may have to be taped or wired to the cooler bottom to keep it submerged.
5. Record the total measured volume of water to the nearest 10 mL.
6. Allow water and propane to sit, while stirring occasionally, until they are both assumed to be at the same initial temperature.
7. Measure and record the initial temperature to the nearest 0.1°C.
8. Cut a hole in the cooler cover large enough for the spoon to be inserted.
9. Allow the propane to vaporize and escape by opening the tank valve. Cover the cooler, and stir the water continuously while the propane is vaporizing.
10. When all the propane has vaporized, measure and record the final temperature of the water.
11. Dry the propane tank/valve assembly, and measure and record its final mass to the nearest 0.1 g.
12. Dispose of the empty propane tank at an approved facility.

6. (a) **Evidence** (values from research will vary with source used)

$\Delta H_c$  values for: (in MJ/mol)

methane (1 C atom)	0.89
propane (3 C atoms)	2.22
pentane (5 C atoms)	3.53
heptane (7 C atoms)	4.81
decane (10 C atoms)	6.73

- (b) **Analysis**

Dividing the molar heat of combustion by the number of carbons in the alkane gives values for:  
(in MJ/(mol · C atom))

methane (1 C atom)	$0.89/1 = 0.89$
propane (3 C atoms)	$2.22/3 = 0.740$
pentane (5 C atoms)	$3.53/5 = 0.706$
heptane (7 C atoms)	$4.81/7 = 0.687$
decane (10 C atoms)	$6.73/10 = 0.673$

There appears to be a clear relationship between the number of carbon atoms and the molar heat of combustion: the greater the number of carbon atoms in the alkane, the lower the molar heat of combustion per carbon atom. The relationship seems to show that the values get closer together (there is less change from one alkane to the next) as the number of carbons increase. This can probably be explained by considering the other alkane atoms — the hydrogens. The ratio H/C is 4.00/1 for methane, 2.67/1 for propane, and 2.20/1 for decane.

The mathematics of series shows that for very large alkanes this ratio must approach 2/1 as a series limit. (When  $n$  is very large,  $2n+2/n \cong 2$ .)

### Making Connections



$$n_{\text{C}} = 1.0 \text{ MJ} \times \frac{1 \text{ mol}}{0.395 \text{ MJ}}$$

$$n_{\text{C}} = 2.53 \text{ mol}$$

$$n_{\text{CO}_2} = 2.53 \text{ mol} \times \frac{1}{1}$$

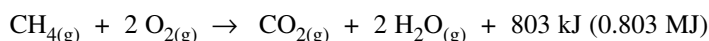
$$n_{\text{CO}_2} = 2.53 \text{ mol}$$

or

$$n_{\text{CO}_2} = 1.0 \text{ MJ} \times \frac{1 \cancel{\text{mol C}}}{0.395 \text{ MJ}} \times \frac{1 \text{ mol CO}_2}{1 \cancel{\text{mol C}}}$$

$$n_{\text{CO}_2} = 2.53 \text{ mol}$$

Releasing 1.0 MJ of energy from burning coal produces 2.53 mol of carbon dioxide.



$$n_{\text{CH}_4} = 1.0 \text{ MJ} \times \frac{1 \text{ mol}}{0.803 \text{ MJ}}$$

$$n_{\text{CH}_4} = 1.25 \text{ mol}$$

$$n_{\text{CO}_2} = 1.25 \text{ mol} \times \frac{1}{1}$$

$$n_{\text{CO}_2} = 1.25 \text{ mol}$$

or

$$n_{\text{CO}_2} = 1.0 \text{ MJ} \times \frac{1 \cancel{\text{mol CH}_4}}{0.803 \text{ MJ}} \times \frac{1 \text{ mol CO}_2}{1 \cancel{\text{mol CH}_4}}$$

$$n_{\text{CO}_2} = 1.25 \text{ mol}$$

Releasing 1.0 MJ of energy from burning methane produces 1.25 mol of carbon dioxide.

- (b) Methane releases half the amount of carbon dioxide that coal does, for equal energy production. If evaluated on that basis, methane would be considered the better fuel.
8. Information about the specific heats of combustion would be more useful, because ordinarily, fuel tank capacities and vehicle mileages are calculated in terms of mass and volume units, not of amount units (moles).

## 12.4 OUR USE OF FOSSIL FUELS

### PRACTICE

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#### Understanding Concepts

- Since 1950 the use of coal has greatly decreased, and the use of oil and gas has greatly increased in Canada.
- Much of Canada's electricity is produced by burning fossil fuels. Conservation of electricity makes more fossil fuels available for petrochemicals.
- The U.K. has a lower living standard, a much warmer climate, and much shorter transport and travel distances than Canada.
- Products produced from ethylene include pipe, tile, fabrics, antifreeze, and all sorts of plastic products.
- |                          |     |
|--------------------------|-----|
| (a) Economic perspective | con |
| (b) Ecological           | con |
| (c) Social               | con |
| (d) Economic             | pro |
| (e) Ecological           | pro |
| (f) Social               | pro |